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World of Chemistry

Zumdahl/Zumdahl/DeCoste



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University of Illinois

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McDougal Littell

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Dedication

For Olivia, Madeline, and Michael
Jessica, Joshua, Moriah, David, Lucia, and Alexander
Who bring love, laughter, and joy
to all of us

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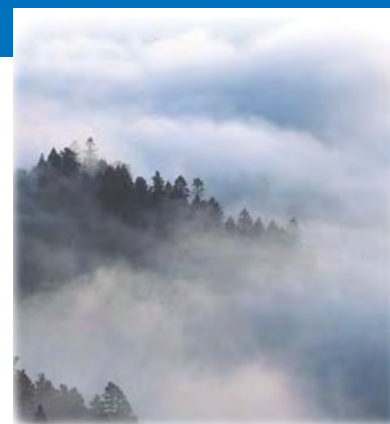
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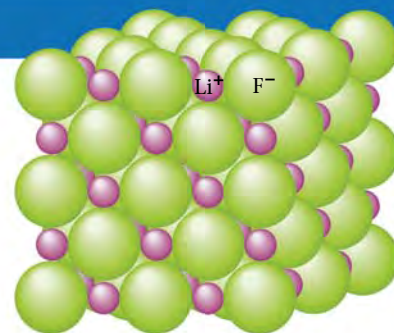


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Reading Chemistry

Chemistry textbooks are written differently from nontechnical textbooks. With this in mind, be aware that reading five pages in a chemistry textbook will probably take much more time than reading five pages in an English or a history textbook.

If you want to understand this chemistry text, prepare to spend a great deal of time reading each section within a chapter. If you flip through this book, you will notice many examples, explanations, diagrams, charts, symbols, and photos to read, analyze, and interpret. You should read the text in each section and incorporate these visuals in your reading. You will quickly find that these visuals are very useful in helping you understand the subject matter.

Use these suggestions to become an efficient and effective reader of chemistry.

1. Preview your textbook.

To become familiar with the design and structure of your textbook, take a look at these key features:

- Information about the authors, page vi
- Tables of Contents and list of features, pages vii–xxi
- Appendices, pages A1–A12
- Glossary, pages A49–A57
- Index, pages A61–A83
- Periodic Table, inside back cover

2. Plan when and where you read.

Although you might find that relaxing on the couch at night is a comfortable and convenient place to do some of your homework, it will not be a good place to read about a chemistry concept.

- Try to read your chemistry section during daylight hours, or in a well-lit area.
- Sit in a straight-backed chair, such as the chairs in the classroom, a kitchen chair, or a desk chair.
- Read in a quiet atmosphere; turn off your television, radio, or earphones.
- Remember that reading chemistry requires active participation. Make sure that you have a pencil, notebook, and any other necessary items to help you study.
- Allow plenty of time to read through each section. Even if a section is short, it still requires deep concentration and focus.

3. Preview each chapter.

When you start a new chapter, be sure to preview it before you begin reading a section within the chapter. Taking time to get an overview of the concepts within the chapter will help you become a more efficient learner.

- Read the title of the chapter and the “Looking Ahead” box. Ask yourself what you have learned previously that may apply to these new concepts.
- Read the first paragraph of the chapter. This paragraph will give you a quick preview of what will be covered and why this chapter material is important.
- Read each section title and the corresponding objectives.
- Glance at the highlighted features, photos, graphs, charts, and symbols in the chapter, to get a sense of the content.
- Read the Chapter Review found at the end of each chapter. This indicates what is most important and what material you will most likely have to master.

4. Read each section prior to class.

Make sure that you read a section before it is presented in class. By previewing the section, you will be better prepared for class.

- Note the section title and objectives in your notebook.
- Make a list of questions you have about the section and of any concepts you need to clarify.
- Write down any vocabulary, symbols, or structures that you have difficulty interpreting.
- Try to work through the example problems, and note any steps that confuse you.
- Attempt to answer the Review Questions if they appear at the end of a section.

5. Reread each section after class.

To reinforce the concepts explained during class, read the section again.

- Highlight in your notes the concepts your teacher emphasized during class.
- Make flashcards to help you memorize structures, symbols, and vocabulary. Review your flashcards daily.
- Work through each example problem, and complete the Practice Problem Exercises if they appear at the end of an example.
- Write a summary of the section. Putting the material in your own words will help you learn this new information.



Writing Chemistry

Up to this point in your studies, you have learned and practiced many different types and styles of writing. In this class, you will be expected to use technical writing. Technical writing differs from essay, report, letter, and creative writing in many ways. Technical writing is factual, precise, clear, and free of bias and personal opinion. Sentences should be specific and to the point.

For instance, if you were asked to write about the current weather conditions in your area for a technical course, your description should use the technical writing style shown in Example 1, rather than the creative style shown in Example 2.

EXAMPLE 1

Technical Writing Sample

YES *The temperature is 75 °F. It is sunny with minimal cloud coverage. The wind is blowing north to northwest at 15 miles per hour.*

Note that the writer in Example 1 states factual information regarding the temperature and wind speed, includes units of measure, and writes without bias and personal opinion.

EXAMPLE 2

Creative Writing Sample

NO *It is a beautiful spring day! The sun is shining without a cloud in the sky, the air is light and breezy, and the temperature is simply perfect for a picnic in the park.*

Although the writer in Example 2 paints a visual picture of the weather conditions, it would not be an appropriate description in a technical course. There are too many adjectives, there are not enough detailed facts, and the writing contains personal opinion about the weather.

When writing a response to a laboratory question, get right to the point. There is no need for extra words. For instance, you may be asked: “What evidence did you observe that indicates a reaction occurred?”

A good technical response would be:

YES *A yellow precipitate formed in the container.*

The wordy response below would be inappropriate:

NO *After we observed the reaction for a few minutes, we saw something yellow starting to form at the bottom of the glass. We think that is how we know that a reaction occurred.*

To strengthen your skills in technical writing for this chemistry course, as well as for other technical courses, follow these helpful hints.

- Write in third person. Do not use I, we, our, your, my, or us.
- Be specific. Do not ramble to make a point.
- Use the correct verb tense. Use present tense when writing research papers and lab reports and when analyzing data. Use past tense to describe objectives and experimental results.
- Make sure that responses to questions are organized, logical, and precise. A reader should be able to clearly follow your train of thought.
- Avoid stating your opinion. Most of your descriptions and observations should be based on things that can be measured. Keep in mind that your writing should be based on facts, not on the opinion of the observer.
- Write numbers as numerals when they are greater than 10 or when they are linked with a measurement. For example: 230, six, 6 cm.
- Avoid beginning sentences with a number, unless the number is part of a chemical name.
- Include units of measure as part of your data. For example: 50 °F, not 50°; 0.025 M, not 0.025.
- Use scientific notation when appropriate.
- Use the correct number of significant figures when reporting your data.
- Use lowercase for chemical names, unless they are at the beginning of a sentence.
- Make sure that tables, graphs, and diagrams used to help explain a response are completely and accurately labeled.
- Describe a laboratory procedure in detail so that another person can repeat the experiment.
- Make sure that a summary of an experiment includes the meaning of your results and how they are the basis for your conclusions. Incorporate possible sources of error in your description.
- Cite all references.

A vertical photograph on the left side of the page shows a woman with blonde hair, wearing a red bikini, surfing on a large, curling blue wave. The water is a vibrant blue, and the scene is captured from a low angle, emphasizing the height and power of the wave.

Testing in Chemistry

The key to earning a successful grade on a chemistry test is to be actively involved in learning the content and to be well-prepared for every class. Chemistry is a cumulative subject, meaning that each section builds on the one before it. Preparation cannot begin the night before the test, but must be ongoing from the first day of class to the last day of class.

Designate a notebook for your chemistry notes and homework. Review your class notes with the corresponding section in the textbook, every day. If you are having trouble remembering new vocabulary words, chemical structures, or formulas, create a set of flashcards. Study your flashcards as often as possible.

Keep in mind that if you find yourself “cramming” for a chemistry test, it is likely that your test grade will reflect a lack of effort. Many of the concepts that you will be learning in this chemistry course will be unfamiliar and challenging. Embrace this challenge by organizing your approach to studying for this course.

Follow these test preparation tips to help you achieve a successful test score.

Before a Chemistry Test

- Read the course outline to know when your tests are scheduled.
- Review your class notes daily, and review the entire section after you have completed it in class. Do not put off reviewing material until right before test time; review needs to be ongoing.
- As you review, ask yourself questions about the material that is difficult for you.
- Create and answer your own test questions about what the teacher emphasized in class.
- Create and study your flashcards.
- Rework any homework problems that you solved incorrectly.
- Find out from your teacher what topics will be covered on the test, the style of the test, and whether a formula reference sheet will be provided.
- Get a good night's sleep before the test. Eat a good breakfast the day of the test so that you can think, process, and recall quickly and accurately.
- Remember to take pencils, erasers, a calculator, and a watch.

During a Chemistry Test

- Relax!
- Read the test directions before you start answering the test questions.
- Look through the test so that you know how many test questions there are, the style of the test items, and what references you have available. This will help determine how much time you should spend on each question.
- If a test question is too hard or too time consuming, skip it and return to it later. Sometimes you may be reminded about a concept or have a thought triggered by another test item. Circle any test items that you skip so that when you go back over the test, you will quickly know which items are not completed.

- Try to save time at the end of the testing period to review and check your answers. Verify that you have completed all parts of each question and that your answers are written clearly.

After a Chemistry Test

- Correct any wrong or incomplete answers.
- If necessary, make changes to your study habits or to your organization to prepare better for the next test.
- Do not give up! Chemistry concepts and applications can be tough to grasp, but with steadfast effort, you will be successful.

Strategies for Various Styles of Test Items

When taking a standardized test, you may encounter various styles of test items, such as multiple choice, gridded response, short answer, and extended response. Each style of question assesses your knowledge and understanding of chemistry, but each is graded differently.

Multiple Choice

A multiple choice test item consists of a test question and four answer choices, such as the following:

A compound contains 16% carbon and 84% sulfur by mass. What is the empirical formula of this compound?

A CS_2 **B** C_2S_2 **C** CS **D** C_2S

To answer a multiple choice question:

- Read the entire question slowly before considering any of the answer choices. Take note of key words that indicate what you are being asked to do.
- Try to determine an answer to the question before looking at the answer choices. Be sure to read all of the answer choices before selecting an answer.
- If you are not sure of the answer, try methods such as eliminating answer choices or working backward from the answer choices provided to make an educated guess.
- Completely fill in the correct answer choice on the answer sheet. Verify that the test item number matches the item number on the answer sheet.



Short Response

Short response test items are structured questions designed to test comprehension and understanding. In a multiple choice or gridded response test item, you simply solve a problem and select or write your answer. In a short response question, however, you must show your work and describe your reasoning process in order to receive full credit.

- Before responding to a short response test item, take time to thoroughly read the question. You may find it helpful to underline key words that signal what should be included in your answer.
- Use the margin of your test page to write a brief outline of your response.
- Use your technical writing skills to write an organized and logical description or explanation.
- Verify that any needed units of measure are included in your answer and that your response contains complete sentences that are free of spelling and grammatical errors.

Extended Response

Extended response test items are usually questions with multiple parts that require a high level of thinking. Because the extended response questions are longer and more involved, your answer may be scored using several points and may involve partial credit.

The following is a sample scoring rubric (method) used to score an answer to an extended response test item.

A Sample Scoring Rubric Method

4 points:	Student shows a thorough understanding of the concept, correctly answers the question, and provides a complete, detailed explanation.
3 points:	Student shows most of the work and provides an explanation but has a minor error OR student shows all work and arrives at a correct response, but does not provide an explanation.
2 points:	Student shows work but makes major errors resulting in an incorrect response or explanation. Explanation is missing, incorrect, or incomplete.
1 point:	Student shows some work and has an incorrect response without an explanation OR student does not follow directions.

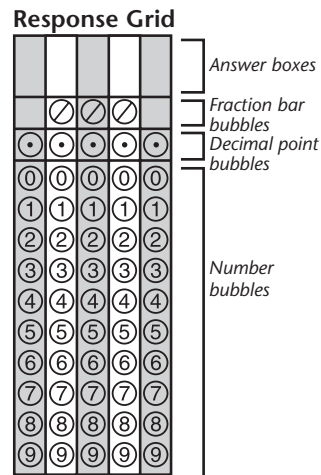
To answer this type of test item, follow the same suggestions given for short response test items. You must also pay attention to the number of parts in the question. Underline key words or phrases in the question, such as *list*, *explain*, *solve*, *compare*, *contrast*, *design*, *identify*, and so on to help you include all aspects of the question in your response.

Gridded Response

When a test item requires you to place your answer in a grid, you must answer the test item correctly *and* fill out accurately the grid on your answer sheet, or the item will be marked as incorrect. The answer to a gridded response test item can be only a positive numerical value, such as a positive integer, a fraction, an improper fraction, or a decimal.

After you work the test item and determine the answer, you must write your answer in the answer boxes at the top of the grid. You must then fill in the corresponding bubble under each box.

- Write your answer with the first digit in the far left answer box, or with the last digit in the far right answer box.
- Write only one digit or symbol in each answer box.
- Do not leave a blank in the middle of an answer.
- Include the decimal point or the fraction bar if it is part of the answer.
- Fill in only one bubble for each answer box that has a number or symbol. Bubbles should be filled in by making a solid black mark.
- Do not grid a mixed number. You must first convert the mixed number to an improper fraction.



EXAMPLES

Whole Number = 50

5	0			
/	/	/	/	/
•	•	•	•	•
0	0	0	0	0
1	1	1	1	1
2	2	2	2	2
3	3	3	3	3
4	4	4	4	4
•	5	5	5	5
6	6	6	6	6
7	7	7	7	7
8	8	8	8	8
9	9	9	9	9

OR

			5	0
/	/	/	/	/
•	•	•	•	•
0	0	0	0	•
1	1	1	1	1
2	2	2	2	2
3	3	3	3	3
4	4	4	4	4
5	5	5	•	5
6	6	6	6	6
7	7	7	7	7
8	8	8	8	8
9	9	9	9	9

Decimal = 5.06

5	.	0	6	
/	/	/	/	/
•	•	•	•	•
0	0	•	0	0
1	1	1	1	1
2	2	2	2	2
3	3	3	3	3
4	4	4	4	4
•	5	5	5	5
6	6	6	•	6
7	7	7	7	7
8	8	8	8	8
9	9	9	9	9

OR

	5	.	0	6
/	/	/	/	/
•	•	•	•	•
0	0	0	•	0
1	1	1	1	1
2	2	2	2	2
3	3	3	3	3
4	4	4	4	4
5	•	5	5	5
6	6	6	6	•
7	7	7	7	7
8	8	8	8	8
9	9	9	9	9

Fractions = $\frac{16}{5} = 3.2$

1	6	/	5	
/	•	/	/	/
•	•	•	•	•
0	0	0	0	0
•	1	1	1	1
2	2	2	2	2
3	3	3	3	3
4	4	4	4	4
5	5	5	•	5
6	•	6	6	6
7	7	7	7	7
8	8	8	8	8
9	9	9	9	9

OR

		3	.	2
/	/	/	/	/
•	•	•	•	•
0	0	0	0	0
1	1	1	1	1
2	2	2	2	•
3	3	•	3	3
4	4	4	4	4
5	5	5	5	5
6	6	6	6	6
7	7	7	7	7
8	8	8	8	8
9	9	9	9	9

Mixed Number = $12 \frac{1}{5} = \frac{61}{5}$

6	1	/	5	
/	/	•	/	/
•	•	•	•	•
0	0	0	0	0
1	1	•	1	1
2	2	2	2	2
3	3	3	3	3
4	4	4	4	4
5	5	5	5	•
6	•	6	6	6
7	7	7	7	7
8	8	8	8	8
9	9	9	9	9

NOT

1	2	1	/	5
/	/	/	•	/
•	•	•	•	•
0	0	0	0	0
•	1	•	1	1
2	•	2	2	2
3	3	3	3	3
4	4	4	4	4
5	5	5	5	•
6	6	6	6	6
7	7	7	7	7
8	8	8	8	8
9	9	9	9	9

Chemistry: An Introduction

LOOKING AHEAD

1.1 The Science of Chemistry

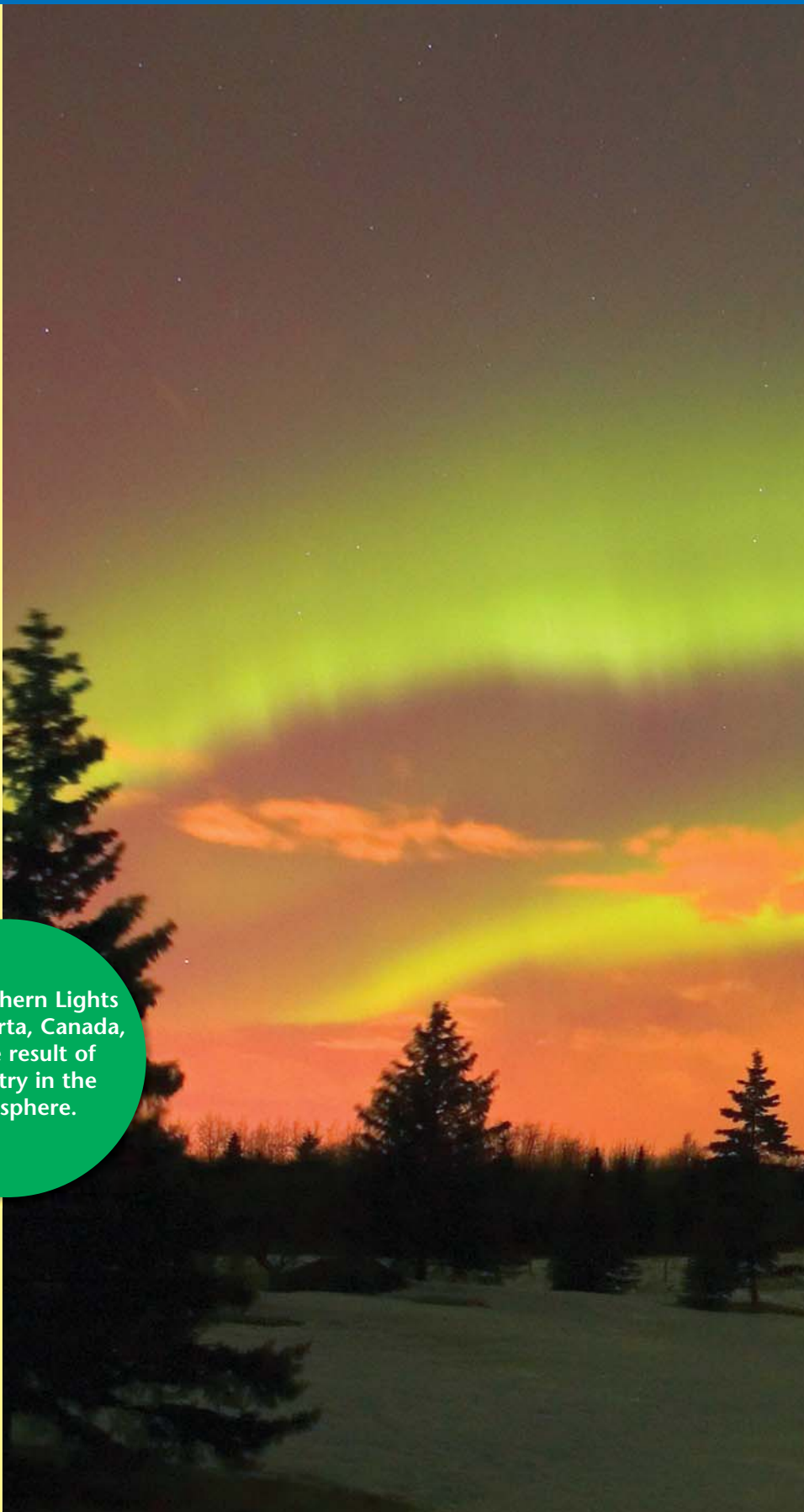
- A. The Importance of Learning Chemistry
 - Why Study Chemistry?
 - A Real-World Chemist
 - Real-World Chemistry
 - Learning Chemistry
- B. What Is Chemistry?
 - Macroscopic and Microscopic Worlds

1.2 Using Science to Solve Problems

- A. Solving Everyday Problems
- B. Using Scientific Thinking to Solve a Problem
- C. The Scientific Method
 - Observations Are Not Theories
 - Theories Do Not Become Laws

1.3 Learning Chemistry

- A. Strategies for Learning Chemistry
 - Solving Problems: Try and Try Again
 - Learning Chemistry: It's Your Job



The Northern Lights over Alberta, Canada, are the result of chemistry in the atmosphere.

IN YOUR LIFE

Did you ever see a fireworks display on the Fourth of July and wonder how it's possible to produce those beautiful, intricate designs in the air? Have you read about dinosaurs—how they ruled the earth for millions of years and then suddenly disappeared?

Although the extinction happened 65 million years ago and may seem unimportant, could the same thing happen to us? Have you ever wondered why an ice cube (pure water) floats in a glass of water (also pure water)? Did you know that the “lead” in your pencil is made of the same substance (carbon) as the diamond in an engagement ring? Did you ever wonder how a corn plant or a palm tree grows seemingly by magic, or why leaves turn beautiful colors in autumn? Do you know how the battery works to start your car or run your calculator? Surely some of these things and many others in the world around you have intrigued you. The fact is that we can explain all of these things in convincing ways by using the models of chemistry and the related physical and life sciences.



WHAT DO YOU KNOW?



Prereading Questions

1. List the first three things you think of when you hear the word *chemistry*.
2. List the first three things you think of when you hear the word *chemist*.
3. Consider a time when you solved a problem in your life. What was the problem? Explain how you went about solving the problem.

SECTION 1.1

The Science of Chemistry

Key Term

- Chemistry

Objectives

- To understand the importance of learning chemistry
- To define chemistry

A. The Importance of Learning Chemistry

A knowledge of chemistry is useful to almost everyone—chemistry occurs all around us all of the time, and an understanding of chemistry is useful to doctors, lawyers, mechanics, businesspeople, firefighters, and poets, among others. Chemistry is important—there is no doubt about that. It lies at the heart of our efforts to produce new materials that make our lives safer and easier, to produce new sources of energy that are abundant and nonpolluting, and to understand and control the many diseases that threaten us and our food supplies. Even if your future career does not require the daily use of chemical principles, your life will be greatly influenced by chemistry.



Why Study Chemistry?

It is clear that the applications of chemistry have enriched our lives in many ways. However, these applications have brought with them some problems. As a society we tend to be shortsighted; we concentrate too much on present benefits without considering the long-range implications of our actions. As a result, we have damaged the environment. We need to pursue a “greener” approach to the production and use of chemicals. To accomplish this goal, we need to learn more about the chemical nature of the earth.

A Real-World Chemist

One of the “hottest” fields in the chemical sciences is environmental chemistry—an area that involves studying our environmental ills and finding creative ways to address them. For example, meet Bart Eklund, who works in the atmospheric chemistry field for Radian Corporation in Austin, Texas. Eklund’s interest in a career in environmental science started with chemistry and ecology courses he took in college. The diverse nature of environmental problems has allowed Eklund to pursue his interest in several fields at the same time.

Bart Eklund’s career demonstrates how chemists help solve our environmental problems. It is how we use our chemical knowledge that makes all the difference.

Image not available for electronic use. Please refer to the image in the textbook.

Eklund's Career Highlights

- Spending a winter month doing air sampling in the Grand Tetons, where he also met his wife and learned to ski;
- Driving sampling pipes by hand into the rocky ground of Death Valley Monument in California;
- Getting to work in and see Alaska, Yosemite Park, Niagara Falls, Hong Kong, the People's Republic of China, Mesa Verde, New York City, and dozens of other interesting places.

Real-World Chemistry

An example that shows how technical knowledge can be a “double-edged sword” is the case of chlorofluorocarbons (CFCs). When the compound called Freon-12 was first synthesized, it was hailed as a near-miracle substance. Because of its noncorrosive nature and its unusual ability to resist decomposition, Freon-12 seemed ideal for refrigeration and air-conditioning systems and many other applications. For years everything seemed fine—the CFCs actually replaced more dangerous materials, such as the ammonia used earlier in refrigeration systems. The CFCs were definitely viewed as “good guys.” But then a problem was discovered—the ozone in the upper atmosphere that protects us from the high-energy radiation of the sun began to decline. What was happening to cause the destruction of the vital ozone?



A chemist in the laboratory

CELEBRITY CHEMICAL

Helium (He)

Have you ever heard recordings of deep-sea divers' voices? They sound a lot like Donald Duck. It's the helium gas mixed with the oxygen gas in the “air” that the divers are breathing that produces the effect. Normal air is about 8/10 nitrogen and 2/10 oxygen. This mixture doesn't work for diving because too much nitrogen dissolves in the blood at the high pressures under the ocean. When the diver returns to the surface, this nitrogen can form bubbles leading to the “bends”—a terribly painful condition that can be fatal. Helium does not dissolve in the blood easily, so it does not lead to the bends readily.

Helium is a very interesting substance. You are no doubt familiar with it because you have had balloons filled with helium. If you let go of a helium-filled balloon, it floats up and away from you. That's because helium is lighter than air.

The helium we use for deep-sea diving and filling balloons is separated from natural gas as it is brought to the surface. This helium was formed in the depths of the earth by radioactive decay and remains trapped with the natural gas deposits.

Helium has many uses. For example, it is used to fill blimps and weather balloons. In addition, because of its very low boiling point ($-452\text{ }^{\circ}\text{F}$), it is used as an extreme coolant in scientific experiments.

For a “light-weight,” helium is pretty important.

Spider-Man balloon filled with helium



Much to everyone's amazement, the culprits turned out to be the CFCs. Large quantities of CFCs had leaked and drifted into the upper atmosphere, where they reacted to destroy ozone molecules. This process allowed more harmful radiation to reach the earth's surface (see **Figure 1.1**). Thus a substance that possessed many advantages in earthbound applications turned against us in the atmosphere. Who could have guessed it would end this way?

The good news is that the U.S. chemical industry is leading the way to find environmentally safe alternatives to CFCs, and the levels of CFCs in the atmosphere are already dropping.

The story of the CFCs shows that we can respond relatively quickly to a serious environmental problem if we decide to do so. Also, it is important to understand that chemical manufacturers have a new attitude about the environment—they are now among the leaders in finding ways to address our environmental problems.

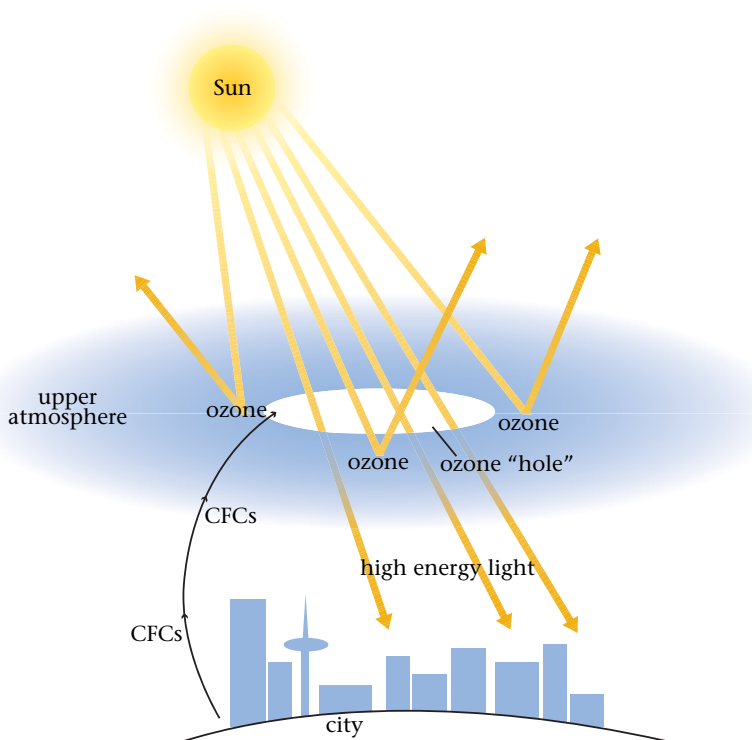


Figure 1.1
Destruction of ozone molecules in the upper atmosphere occurs when CFCs drift up from populated areas.

Learning Chemistry

Learning chemistry is both interesting and important. A chemistry course can do more than simply help you learn the principles of chemistry, however. A major byproduct of your study of chemistry is that you will become a better problem solver. One reason that chemistry has the reputation of being “tough” is that it often deals with rather complicated systems that require some effort to figure out. Although this complexity may seem to be a disadvantage at first, you can turn it to your advantage if you have the right attitude.

Although learning chemistry is often not easy, it's never impossible. In fact, anyone who is interested, patient, and willing to work can learn chemistry. In this book, we will try very hard to help you understand what chemistry is and how it works, and to point out how chemistry applies to the things going on in your life.

Our sincere hope is that this text will motivate you to learn chemistry, make its concepts understandable to you, and demonstrate how interesting and vital the study of chemistry is.

B. What Is Chemistry?

Chemistry can be defined as the science that deals with the materials of the universe and the changes that these materials undergo.

Chemists are involved in activities as diverse as examining the fundamental particles of matter, looking for molecules in space, making new materials of all types, using bacteria to produce such chemicals as insulin, and finding new methods for early detection of disease.

Chemistry is often called the central science—and with good reason. Most of the phenomena that occur in the world around us involve chemical changes—changes in which one or more substances become different substances. Here are some examples of chemical changes:

- Wood burns in air, forming water, carbon dioxide, and other substances.
- A plant grows by assembling simple substances into more complex substances.
- The steel in a car rusts.
- Eggs, flour, sugar, and baking powder are mixed and baked to yield a cake.
- The definition of the term *chemistry* is learned and stored in the brain.
- Emissions from a power plant lead to the formation of acid rain.

Active Reading Question

What types of things does a real-world chemist do?

Macroscopic and Microscopic Worlds

As we proceed, you will see how the concepts of chemistry allow us to understand the nature of these and other changes. To understand these processes and the many others that occur around us, chemists take a special view of things. Chemists “look inside” ordinary objects to observe how the fundamental components are behaving.

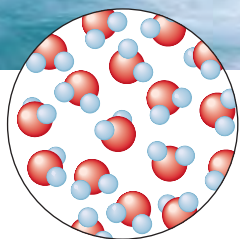
To understand how this approach works, consider a tree. When we view a tree from a distance, we see the tree as a whole. The trunk, the branches, and the leaves all blend together to give the tree. We call this overall view of the tree the *macroscopic* picture.

As we get closer to the tree, we begin to see the detail—pieces of bark, individual leaves, large and tiny branches, and so on. Now imagine that we examine a single leaf. We see veins, variation in color, surface irregularities, and more. Our curiosity is whetted. What lies inside the leaf? What causes it to change from a bud in the spring to a green leaf in the summer and

Chemistry

The science of materials and the changes that these materials undergo



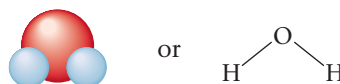


The macroscopic view of water (the mountain stream) and the microscopic view (the individual water molecules)

then to a red or golden color in the fall? To answer these questions, we need a microscope. As we examine the leaf under a microscope, we see cells and motion. Because we don't "live" in this *microscopic* world, the commonplace leaf becomes fascinating and mysterious.

When we speak of "motion" in the macroscopic world, we refer to the swaying of the tree and the rustling of the leaves. In the microscopic world, "motion" refers to the cells acting as tiny machines that absorb energy from the sun and nutrients from the air and the soil. We are now in the microscopic world, but as chemists, we want to go even further. What are the building blocks of the cells and what are the components of the water that contains the dissolved nutrients?

Think about water, a very familiar substance. In the macroscopic world, it flows and splashes over rocks in mountain streams and freezes on ponds in the winter. What is the microscopic nature of water? As you may know already, water is composed of tiny molecules that we can represent as



Here H represents a hydrogen atom and O represents an oxygen atom. We often write this molecule as H_2O because it contains two hydrogens (H) and one oxygen (O).

This is the microscopic world of the chemist—a world of molecules and atoms. This is the world we will explore in this book. One of our main goals is to connect the macroscopic world in which you live to the microscopic world that makes it all work. We think you will enjoy the trip!

Active Reading Question

What are some examples of chemistry that you see in everyday life?

SECTION 1.1

REVIEW QUESTIONS

- 1 How do CFC's illustrate that technical advances can be a "double-edged" sword?
- 2 Why is chemistry often called a central science?
- 3 What is the difference between a *microscopic* picture and a *macroscopic* picture?
- 4 What is chemistry?
- 5 Name three activities that a chemist may do.



Using Science to Solve Problems

Objectives

- To understand scientific thinking
- To illustrate scientific thinking
- To describe the method scientists use to study nature

Key Terms

- Scientific method
- Theory
- Natural law
- Measurement

A. Solving Everyday Problems

One of the most important things we do in everyday life is solve problems. In fact, most of the decisions you make each day can be described as solving problems.

It's 8:30 A.M. on Friday. Which is the best way to drive to school to avoid traffic congestion?

You have two tests on Monday. Should you divide your study time equally or allot more time to one than to the other?

Your car stalls at a busy intersection and your little brother is with you. What should you do next?


These are everyday problems of the type we all face. What process do we use to solve them? You may not have thought about it before, but there are several steps that almost everyone uses to solve problems:

1. Recognize the problem and state it clearly. Some information becomes known, or something happens that requires action. In science we call this step **making an observation**.
2. Propose possible solutions to the problem or possible explanations for the observation. In scientific language, suggesting such a possibility is called **formulating a hypothesis**.
3. Decide which of the solutions is the best, or decide whether the explanation proposed is reasonable. To do this, we search our memory for any pertinent information or we seek new information. In science we call searching for new information **performing an experiment**.

As we will see, citizens as well as scientists use these same procedures to study what happens in the world around us. The important point here is that scientific thinking can help you in all parts of your life. It's worthwhile to learn how to think scientifically—whether you want to be a scientist, an auto mechanic, a doctor, a politician, or a poet!

B. Using Scientific Thinking to Solve a Problem

To illustrate how science helps us solve problems, consider a true story about two people, David and Susan (not their real names). Several years ago David and Susan were healthy 40-year-olds living in California, where David was serving in the Air Force. Gradually Susan became quite ill, showing flulike symptoms including nausea and severe muscle pains. Even her personality changed: she became uncharacteristically grumpy. She seemed like a totally different person from the healthy, happy woman of a few



months earlier. Following her doctor's orders, she rested and drank plenty of fluids, including large quantities of orange juice from her favorite mug, part of a 200-piece set of pottery dishes recently purchased in Italy. However, she just became sicker, developing extreme abdominal cramps and severe anemia.

During this time David also became ill and exhibited symptoms much like Susan's: weight loss, excruciating pain in his back and arms, and uncharacteristic fits of temper. The disease became so serious that he retired early from the Air Force and the couple moved to Seattle. For a short time their health improved, but after they unpacked all their belongings (including those pottery dishes), their health began to deteriorate again. Susan's body became so sensitive that she could not tolerate the weight of a blanket. She was near death. What was wrong? The doctors didn't know, but one of them suggested that she might have porphyria, a rare blood disease.

Desperate, David began to search the medical literature himself. One day while he was reading about porphyria, a statement jumped off the page: "Lead poisoning can sometimes be confused with porphyria." Could the problem be lead poisoning?

We have described a very serious problem with life-or-death implications. What should David do next? Overlooking for a moment the obvious response of calling the couple's doctor immediately to discuss the possibility of lead poisoning, could David solve the problem by scientific thinking? Let's use the three steps of the scientific approach to attack the problem one part at a time. This is important: usually we solve complex problems by breaking them down into manageable parts. We can then assemble the solution to the overall problem from the answers we have found "from the parts."



CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Land Mine Buzzers

An estimated 100 million plastic land mines are scattered throughout the earth on former battlefields. Every day, these hidden mines kill or maim 60 people. Finding these mines is very difficult—they were designed to resist detection. Scientists are now enthusiastic about a new way to identify the mines—with honeybees. Previous work has shown that bees foraging in chemically contaminated areas carry these substances back to their hives. The hope is that bees searching for food in mined areas will bring back traces of the explosives from "leaky" mines, alerting people to nearby danger. Scientists also plan to train bees to seek out explosives by associating the scents of the explosive compounds with food. The researchers will keep track of the

bees' movements by fitting electronic identification tags (about the size of a grain of rice) on the bees' backs.

If this idea works, it would be an inexpensive, safe method for characterizing a minefield. These "mine buzzers" would be much safer than the current practice of prodding the soil with pokers.



What Is the Disease?			
Observation	Hypothesis	Experiment	Results
David and Susan are ill with specific systems.	The disease is lead poisoning.	Look up symptoms of lead poisoning.	Symptoms match almost exactly.

Is It Lead Poisoning?			
Observation	Hypothesis	Experiment	Results
Lead poisoning results from high levels of lead in the blood.	David and Susan have high levels of lead in their blood.	Perform blood analysis.	High levels of lead in both people's blood.

Where Is the Lead Coming From?			
Observation	Hypothesis	Experiment	Results
There is lead in the couple's blood.	The lead is in their food and drink when they purchase it.	Determine whether anyone else shopping at the store has symptoms (no one does).	Moving to a new area (and new store) did not solve the problem.
The food they bought is free of lead.	The dishes they use at home are the source of the lead.	Determine whether the dishes contain lead.	Analysis of the pottery showed that lead was present in the glaze.
Lead is present in the dishes, so they could be a source of lead poisoning.	The lead is being leached into their food.	Place a beverage (such as orange juice) in one of the cups. Analyze the beverage for lead.	High levels of lead in the drink, so the dishes are the source of the lead poisoning.

After many applications of the scientific method, the problem is solved. We can summarize the answer to the problem (David and Susan's illness) as follows: the Italian pottery they used for everyday dishes contained a lead glaze that contaminated their food and drink with lead. This lead accumulated in their bodies to the point at which it interfered seriously with normal functions and produced severe symptoms. This overall explanation, which summarizes the hypotheses that agree with the experimental results, is called a *theory* in science. This explanation accounts for the results of all of the experiments performed.*

We could continue to use the scientific method to study other aspects of this problem, such as the following:

What types of food or drink leach the most lead from the dishes?

Do all pottery dishes with lead glazes produce lead poisoning?

As we answer questions by using the scientific method, other questions naturally arise. By repeating the three steps again and again, we can come to understand a given phenomenon thoroughly.



*"David" and "Susan" recovered from their lead poisoning and are now publicizing the dangers of using lead-glazed pottery. This happy outcome is the answer to the third part of their overall problem, "Can the disease be cured?" They simply stopped using that pottery for food and beverages.

C. The Scientific Method

Science is a framework for gaining and organizing knowledge. Science is not simply a set of facts but also a plan of action—a *procedure* for processing and understanding certain types of information. Although scientific thinking is useful in all aspects of life, in this text we will use it to understand how the natural world operates. The process that lies at the center of scientific inquiry is called the **scientific method**, which consists of the following steps:

Scientific method

Systematic process for studying nature that involves observations, hypotheses, and experiments

Measurement

A quantitative observation

Steps in the Scientific Method

1. State the problem and collect data (make observations).

Observations may be *qualitative* (the sky is blue; water is a liquid) or *quantitative* (water boils at 100 °C; a certain chemistry book weighs 4.5 pounds). A qualitative observation does not involve a number. A quantitative observation is called a **measurement** and does involve a number (and a unit, such as pounds or inches). We will discuss measurements in detail in Chapter 5.

2. Formulate hypotheses.

A hypothesis is a *possible* explanation for the observation.

3. Perform experiments.

An experiment is something we do to test the hypothesis. We gather new information that allows us to decide whether the hypothesis is supported by the new information we have learned from the experiment. Experiments always produce new observations, and these observations bring us back to the beginning of the process.

Information

Quantitative observations involve a number. Qualitative observations do not.

To explain the behavior of a given part of nature, we repeat these steps many times. Gradually, we gather the knowledge necessary to understand what is going on.

When we have a set of hypotheses that agrees with our various observations, we assemble them into a theory that is often called a *model*. A **theory** (model) is a set of tested hypotheses that gives an overall explanation of some part of nature (see **Figure 1.2**).

Observations Are Not Theories

It is important to distinguish between observations and theories.

An observation is something that is witnessed and can be recorded. A theory is an interpretation—a possible explanation of why nature behaves in a particular way.

Theories inevitably change as more information becomes available. For example, the motions of the sun and stars have remained virtually the same over the thousands of years during which humans have been observing them, but our explanations—our theories—have changed greatly since ancient times.

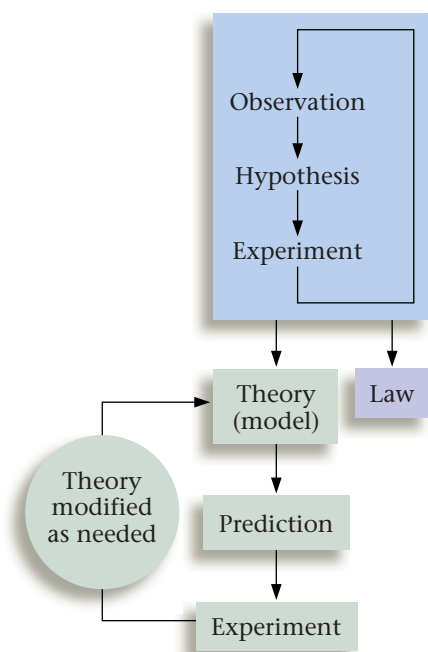


Figure 1.2
The various parts of the scientific method

The point is that we don't stop asking questions just because we have devised a theory that seems to account satisfactorily for some aspect of natural behavior. We continue doing experiments to refine our theories. We do this by using the theory to make a prediction and then performing an experiment (making a new observation) to see whether the results match this prediction.

Always remember that theories (models) are human inventions. They represent our attempts to explain observed natural behavior in terms of our human experiences. We must continue to do experiments and refine our theories to be consistent with new knowledge if we hope to approach a more nearly complete understanding of nature.

Theories Do Not Become Laws

As we observe nature, we often see that the same observation applies to many different systems. For example, studies of innumerable chemical changes have shown that the total mass of the materials involved is the same before and after the change. We often summarize such generally observed behavior into a statement called a **natural law**. The observation that the total mass of materials is not affected by a chemical change in those materials is called the law of conservation of mass.

You must recognize the difference between a law and a theory. A law is a summary of observed (measurable) behavior, whereas a theory is an explanation of behavior.

Natural law

Statement that summarizes generally observed behavior

A law tells what happens; a theory (model) is our attempt to explain why it happens.

CRITICAL THINKING ?

What if everyone in the government used the scientific method to analyze and solve society's problems, and politics were never involved in the solutions?

Write a paragraph explaining how this approach would be different from the present situation, and whether it would be better or worse.

In this section, we have described the scientific method (which is summarized in **Figure 1.2**) as it might ideally be applied. However, it is important to remember that science does not always progress smoothly and efficiently. Scientists are human. They have prejudices, they misinterpret data, they can become emotionally attached to their theories and thus lose objectivity, and they may play politics. Science is affected by profit motives, budgets, fads, wars, and religious beliefs.

Galileo, for example, was forced to deny his astronomical observations in the face of strong religious resistance. Lavoisier, the father of modern chemistry, was beheaded because of his politics. And great progress in the chemistry of nitrogen fertilizers resulted from the desire to produce explosives to fight wars. The progress of science is often slowed more by the frailties of humans and their institutions than by the limitations of scientific measuring devices. The scientific method is only as effective as the humans using it. It does not automatically lead to progress.

Active Reading Question

In what ways is using a scientific approach to solving a problem similar to approaches you have used in solving problems in everyday life? In what ways is it different?

SECTION 1.2 REVIEW QUESTIONS

1 Describe how you would set up an experiment to test the relationship between doing chemistry homework and the final grade in the course.

Apply a scientific approach to solving this problem. Label each of your steps appropriately.

2 What is the difference between an observation and a theory? Give an example of each (different from the ones in your book!)

3 What are the steps in the scientific method?

4 Which is an interpretation—a law or a theory?



Learning Chemistry

Objective

- To develop successful strategies for learning chemistry

A. Strategies for Learning Chemistry

Chemistry courses have a universal reputation for being difficult. There are some good reasons for this. For example, as we have already discussed, we live in the macroscopic world; but to understand this world, we need to think in terms of the unfamiliar microscopic world. In addition, the language of chemistry is unfamiliar in the beginning; many terms and definitions need to be memorized. As with any language, *you must know the vocabulary* before you can communicate effectively. We will try to help you by pointing out those things that need to be memorized.

But memorization is only the beginning. Don't stop there or your experience with chemistry will be frustrating. Memorization is not the same as understanding. We can teach a parrot to say anything we can say, but the parrot does not understand the meaning of its words. You need to memorize facts, but you also need to strive to understand the fundamental ideas—the models of chemistry. Be willing to do some thinking, and learn to trust yourself to figure things out. To solve a typical chemistry problem, you must sort through the given information and decide what is really crucial.

Solving Problems: Try and Try Again

It is important to realize that chemical systems tend to be complicated—there are typically many components—and we must make approximations in describing them. Therefore, trial and error play a major role in solving chemical problems. In tackling a complicated system, a practicing chemist really does not expect to be right the first time he or she analyzes the problem. The usual practice is to make several simplifying assumptions and then give it a try. If the answer obtained doesn't make sense, the chemist adjusts the assumptions, using feedback from the first attempt, and tries again.

The point is this: in dealing with chemical systems, do not expect to understand everything immediately. In fact, it is typical (even for an experienced chemist) *not* to understand at first. Make an attempt to solve the problem and then analyze the feedback. *It is not a disaster to make a mistake as long as you learn from it.*

The only way to develop your confidence as a problem solver is to practice solving problems. To help you, this book contains many examples worked out in detail. Follow these examples carefully, making sure you understand each step. These examples are followed by a Practice Problem that you should try on your own. Use the practice problems to test whether you are understanding the material as you go along.



Top Ten Reasons for Taking Chemistry

- Do experiments using fire and chemicals.
- Learn to answer chemistry questions on game shows.
- Learn to solve problems.
- See cool demonstrations in class.
- Learn to use and dispose of household chemicals safely.
- Wear fashionable lab aprons and goggles.
- Learn to think like a scientist.
- Impress your friends by using words like *molarity* and *stoichiometry*.
- Find a practical use for the math you've been learning.
- Meet the most interesting people—the people who teach it!

CHEMISTRY IN YOUR WORLD

Consumer Connection

Learn Chemistry—Prepare for Life

What is the purpose of education?

Some people seem to equate education with the storage of facts in the brain. These people apparently believe that education simply means memorizing the answers to all of life's present and future problems. Although this idea is clearly unreasonable, many students behave as though it were their guiding principle. These students want to memorize lists of facts and to reproduce them on tests. They regard as unfair any exam questions that require some original thought or some processing of information. Indeed, it might be tempting to reduce education to a simple filling up with facts because that approach can produce short-term satisfaction for both student and teacher. Of course, storing facts in the brain *is* important. You cannot function without knowing that red means stop, electricity is dangerous, ice is slippery, and so on.

However, mere recall of abstract information, without the ability to process it, makes you little better than a talking encyclopedia. Successful people have several common characteristics. They have a thorough knowledge of their fields, the ability to recognize and solve problems, and the ability to communicate effectively. Typically, they also have a high level of motivation.

How does studying chemistry help you achieve these characteristics? The fact that chemical systems are complicated is really a blessing, though one that

is well disguised. Studying chemistry will not by itself make you a good problem solver; but it can help you develop a positive, aggressive attitude toward problem solving, and it can help boost your confidence. Learning to “think like a chemist” can be valuable to anyone in any field. People who are trained as chemists often excel not only in chemical research and production but also in personnel, marketing, sales, development, finance, and management. The point is that much of what you learn in this course can be applied to any career. So be careful not to take too narrow a view of this course. Look beyond short-term frustration to long-term benefits.

It may not be easy to learn to be a good problem solver, but it's well worth the effort.



Learning chemistry is both fun and important.



SECTION 1.3

REVIEW QUESTIONS

- 1 What resources are available to help you learn chemistry?
- 2 What characteristics do former chemistry students find most important to their success? Which of these are easiest for you? Which are most difficult?
- 3 Why does trial and error play a major role in problem solving?



Key Terms

1.1 Chemistry

1.2 Scientific method

Theory

Natural law

Measurement

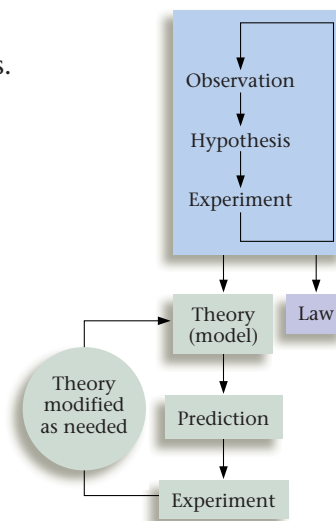
Key Ideas

1.1 The Science of Chemistry

- Chemistry is important to everyone because chemistry occurs all around us in our daily lives.
- Chemistry is the science that deals with the materials of the universe and the changes that these materials undergo.
- Chemistry “looks inside” ordinary objects to study how their components behave.

1.2 Using Science to Solve Problems

- Scientific thinking helps us solve all types of problems that we confront in our lives.
- Scientific thinking involves observations that enable us to clearly define both a problem and the construction and evaluation of possible explanations or solutions to the problem.
- The scientific method is a procedure for processing the information that flows from the world around us in which we
 - Make observations.
 - Formulate hypotheses.
 - Perform experiments.



- Models represent our attempt to understand the world around us.
 - Models are not the same as “reality.”
 - Elementary models are based on the properties of atoms and molecules.

1.3

1.3 Learning Chemistry

- Understanding chemistry takes hard work and patience.
- As you learn chemistry, you should be able to understand, explain, and predict phenomena in the macroscopic world by using models based in the microscopic world.
- Understanding is different from memorizing.
- It is okay to make mistakes as long as you learn from them.

All exercises with **blue numbers** have answers in the back of this book.

1.1 The Science of Chemistry

A. The Importance of Learning Chemistry

- The first few paragraphs of this book asked you if you had ever wondered how and why various things in everyday life behave the way they do. Surely there are many other such phenomena that have intrigued you. List five of these and present them in class for discussion with your instructor and fellow students.
- Chemistry is used in many professions, and a basic understanding of chemistry is of great importance. Suggest two ways that each of the following professionals might make use of chemistry in their jobs.

a. physician	e. photographer
b. lawyer	f. farmer
c. pharmacist	g. nurse
d. artist	
- This section presents several ways in which our day-to-day lives have been enriched by chemistry. List three materials or processes involving chemistry that you believe have contributed to such an enrichment, and explain your choices.
- The text admits that there has been a “dark side” to our use of chemicals and chemical processes, and uses the example of chlorofluorocarbons (CFCs) to explain this. List three additional improper or unfortunate uses of chemicals or chemical processes, and explain your reasoning.

B. What Is Chemistry?

- This textbook provides a specific definition of chemistry: the study of the materials of which the universe is made and the transformations that these materials undergo. Obviously, such a general definition has to be very broad and nonspecific. From your point of view at this time, how would *you* define chemistry? In your mind, what are chemicals? What do chemists do?
- We use chemical reactions in our everyday lives, not just in the science laboratory. Give at least five examples of chemical transformations that you make use of in your daily activities. Identify the “chemical” in each of your examples, and how you recognize that a chemical change has taken place.

1.2 Using Science to Solve Problems

A. Solving Everyday Problems

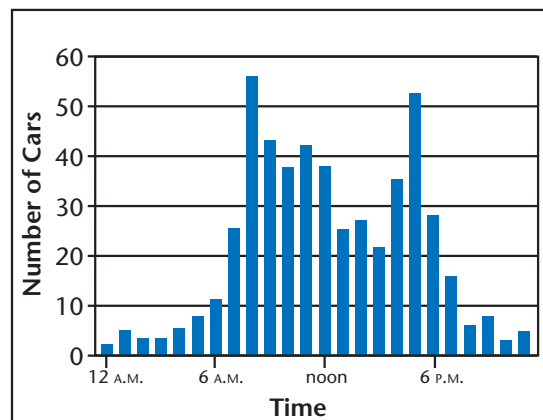
- Discuss two situations in which you have analyzed a problem such as those presented in this section. What hypotheses did you suggest? How did you test those hypotheses?

B. Using Scientific Thinking to Solve a Problem

- Consider the lead poisoning case given in this section. Discuss how David and Susan analyzed the situation and arrived at the theory that the lead glaze on the pottery was responsible for their symptoms.

C. The Scientific Method

- What are the three operations involved in applying the scientific method? How does the scientific method help us understand our observations of nature?
- As part of a science project, you study traffic patterns in your city at an intersection in the middle of downtown. You set up a device that counts the cars passing through this intersection for a 24-hour period during a weekday. The graph of hourly traffic looks like this.



- At what time(s) does the highest number of cars pass through the intersection?
- At what time(s) does the lowest number of cars pass through the intersection?
- Briefly describe the trend in numbers of cars over the course of the day.
- Provide a hypothesis explaining the trend in numbers of cars over the course of the day.
- Provide a possible experiment that could test your hypothesis.

- 11.** Which of the following are quantitative observations and which are qualitative observations?
- My waist size is 31 inches.
 - My eyes are blue.
 - My right index finger is $\frac{1}{4}$ inch longer than my left.
 - The leaves of most trees are green in summer.
 - An apple is more than 95% water.
 - Chemistry is an easy subject.
 - My score on the last chemistry exam was 90%.

12. What is the difference between a hypothesis and a theory? How are the two similar? How do they differ?

13. What is a natural law? Give examples of such laws. How does a law differ from a theory?

14. For the following passage, label each statement as “observation,” “theory,” or “law.”

Every time we let go of something heavy, it falls to the ground. For example, when an apple breaks from a branch of a tree, it falls to the ground. This happens because apples are attracted to earth.

15. Discuss several political, social, or personal considerations that might affect a scientist’s evaluation of a theory. Give examples of how such external forces have influenced scientists in the past. Discuss methods by which such bias might be excluded from future scientific investigations.

1.3 Learning Chemistry

A. Strategies for Learning Chemistry

16. In some academic subjects, it may be possible to receive a good grade primarily by memorizing facts. Why is chemistry not one of these subjects?

17. Why is the ability to solve problems important in the study of chemistry? Why is it that the method used to attack a problem is as important as the answer itself?

18. Students approaching the study of chemistry must learn certain basic facts (such as the names and symbols of the most common elements), but it is much more important that they learn to think critically and to go beyond the specific examples discussed in class or in the textbook. Explain how learning to do this might be helpful in any career, even one far removed from chemistry.

1 Chemists are involved in activities such as

- A producing new materials.
- B producing new sources of energy.
- C understanding and controlling various diseases.
- D All of the above

2 Which of the following is another term for quantitative observation?

- A theory
- B measurement
- C hypothesis
- D natural law

Use the information in the following paragraph for questions 3–5.

Ronika and Thomas are eating raisins and drinking ginger ale. Thomas accidentally drops a raisin into his ginger ale. Both he and Ronika notice that the raisin falls to the bottom of the glass. Soon, the raisin rises to the surface of the ginger ale, and then sinks. Within a couple of minutes, it rises and sinks again. Ronika asks “I wonder why that happened?” Thomas says, “I don’t know, but let’s see if it works in water.” Ronika fills a glass with water and drops a raisin into the glass. After a few minutes, Thomas says, “No, it doesn’t go up and down in the water.” Ronika closely observes the raisins in the two cups and states, “Look, there are bubbles on the raisin in the ginger ale but not on the raisin in the water.” Thomas says “It must be the bubbles that make the raisin rise.” Ronika asks, “OK, but then why does it sink again?”

3 Which of the following items from the passage is an *observation*?

- A Accidentally dropping the raisin into the ginger ale
- B Claiming that the bubbles make the raisin rise
- C Asking “I wonder why that happened?”
- D Noticing the raisin sinks in the ginger ale

4 Which of the following from the passage is a *theory*?

- A Asking “I wonder why that happened?”
- B Noticing that the raisin does not rise and sink in the water
- C Claiming that the bubbles make the raisin rise
- D Wondering why the raisin sinks after it rises to the surface of the ginger ale

5 Which of the following from the passage is an *experiment*?

- A Claiming that the bubbles make the raisin rise
- B Dropping a raisin into water
- C Asking “I wonder why that happened?”
- D Wondering why the raisin sinks after it rises to the surface of the ginger ale

6 What does a chemist mean by a “microscopic view” of the world, and why is taking such a view important to a chemist?

For questions 7–10, label each of the following statements as true or false. For the statements that you label false, provide an explanation.

7 We can prove a theory to be correct by performing the right experiment.

8 If a theory is accepted for a long time, it eventually becomes a law.

9 A law is a summary of observed behavior.

10 A theory is an attempt to explain what happens in an experiment.

LOOKING AHEAD

2.1 The Nature of Matter


- A. The Particulate Nature of Matter
 - The Atomic Nature of Matter
- B. Elements and Compounds
 - Compounds
 - Elements
- C. The States of Matter

2.2 Properties of Matter

- A. Physical and Chemical Properties and Changes

2.3 Classifying Matter

- A. Mixtures and Pure Substances
 - Mixtures
 - Pure Substances
 - Homogeneous and Heterogeneous Mixtures
- B. Separation of Mixtures



Chemical reactions cause the leaves to change color in the fall.

IN YOUR LIFE

As you look around you, you must wonder about the properties of matter. How do plants grow and why are they green? Why is the sun hot? Why does soup get hot in a microwave oven? Why does wood burn whereas rocks do not? What is a flame? How does soap work? Why does soda fizz when you open the bottle? When iron rusts, what's happening? And why doesn't aluminum rust? How does a cold pack for an athletic injury, which is stored for weeks or months at room temperature, suddenly get cold when you need it? How does a hair permanent work?

The answers to these and endless other questions lie in the domain of chemistry. In this chapter we begin to explore the nature of matter: how it is organized and how and why it changes. We will also consider the energy that accompanies these changes.

Why does soda fizz when you open the bottle?



WHAT DO YOU KNOW?



Prereading Questions

1. How do atoms differ from molecules?
2. Give an example of each of these: a solid, a liquid, a gas.
3. Describe a chemical change (chemical reaction) you have witnessed outside of school. How do you know that it is a chemical change?

SECTION 2.1

The Nature of Matter

Key Terms

- Matter
- Atoms
- Compounds
- Molecule
- Elements
- Solid
- Liquid
- Gas

Objectives

- To learn about the composition of matter
- To learn the difference between elements and compounds
- To define the three states of matter

A. The Particulate Nature of Matter

Matter, the “stuff” of which the universe is composed, has two characteristics: it has mass and it occupies space. Matter comes in a great variety of forms: the stars, the air that you are breathing, the gasoline that you put in your car, the chair on which you are sitting, the turkey in the sandwich you may have eaten for lunch, the tissues in your brain that allow you to read and understand this sentence, and so on.

As we look around our world, we are impressed by the great diversity of matter. Given the many forms and types of matter, it seems difficult to believe that all matter is composed of a small number of fundamental particles. It is surprising that the fundamental building blocks in chocolate cake are very similar to the components of air.

The Atomic Nature of Matter

How do we know that matter is composed of the tiny particles we call atoms? After all, they are far too small to be seen with the naked eye. It turns out that after literally thousands of years of speculation, we can finally “see” the atoms that are present in matter. In recent years scientists have developed a device called a scanning tunneling microscope (STM) that, although it works quite differently from an optical microscope, can produce images of atoms.

For example, look at the penny shown in **Figure 2.1**. The small objects represent tiny copper atoms. When chemists look at other metals with powerful microscopes, they see atoms in these substances as well. You can see an example of another metal in **Figure 2.2**.

With ultra-high magnification, objects appear more similar. All objects are made up of small particles called atoms.

All matter consists of tiny particles called atoms. But when you look at objects such as nails or pennies, you don’t see these particles. Why not? The atoms are very tiny and can be seen only with a powerful magnifying instrument. You may have encountered the same concept in your life when you looked at a beach from a distance. The sand looks uniform—you can’t see the separate particles. As you get close, however, the individual grains of sand become apparent. Observe the Impressionist painting by Seurat shown in the photo on the following page. From a distance the scene looks normal. Only when you stand very close to the painting do you see that it is composed of tiny dots of paint.

So, the conclusion is that although objects in the macroscopic world typically look quite continuous and uniform, they are really particulate in nature—they are made of atoms. We can “see” them with powerful electronic microscopes.

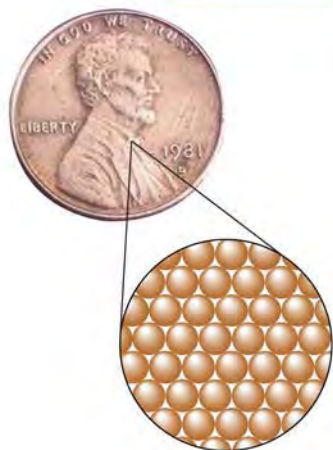


Figure 2.1

The surface of the copper penny is made of copper atoms represented as they would be seen through the lens of a very powerful electronic microscope.

Image not available for electronic use. Please refer to the image in the textbook.

Figure 2.2

A scanning tunneling microscope image of nickel metal. Each peak represents a nickel atom.

Active Reading Question

How do we know that all matter is made of atoms?

B. Elements and Compounds

We have just considered the most important idea in chemistry: matter is composed of tiny particles we call **atoms**. If all matter is made up of tiny particles called atoms, are all atoms alike? That is, is copper metal made of the same kind of atoms as gold? The answer is no. Copper atoms and gold atoms are different. Scientists have learned that all matter is composed from about 100 different types of atoms. For example, air is mostly gaseous oxygen and nitrogen. The nitrogen atoms are different from the oxygen atoms, which in turn are different from copper atoms, which differ from gold atoms.

You can think of the matter in the universe like the words in a book. If you break all the words in this book apart into their component letters, you will end up with “large piles” of only 26 letters. The English alphabet allows you to construct thousands of words from just 26 letters. Similarly, all the matter in the universe is constructed by putting approximately 100 types of atoms together in various ways. The different types of matter are like the different words in a book. When we separate all of the universe into its atoms, we find approximately 100 different atoms. We call these 100 types of atoms the *elements* of the universe.

To illustrate this idea, consider the letters A, D, and M. Using these letters you can make many words, such as MAD, DAM, DAD, and MADAM (can you think of others?). Each word represents something very different. Thus, with only three letters, you can represent several unique things or ideas.

Atom

Fundamental unit of which elements are made



Sand on a beach looks uniform from a distance, but up close the irregular sand grains are visible.



The painting “Sunday Afternoon on the Island of La Grande Jatte” by Georges Seurat illustrates the Impressionist style.



Top Ten Elements in the Universe

Element	Percent (by atoms)
Hydrogen	73.9
Helium	24.0
Oxygen	1.1
Carbon	0.46
Neon	0.13
Iron	0.11
Nitrogen	0.097
Silicon	0.065
Magnesium	0.058
Sulfur	0.044

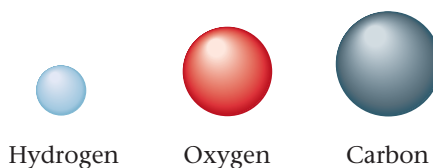
CRITICAL THINKING ?

The scanning tunneling microscope allows us to “see” atoms.

What if you were sent back in time before the invention of the scanning tunneling microscope? Write a paragraph giving evidence to support the theory that all matter is made of atoms and molecules.

Compounds

In much the same way that we can use a few letters to make thousands of words, we can use a few types of atoms to construct all matter. For example, consider the atoms hydrogen, oxygen, and carbon:







Notice that we represent atoms by using spheres. We get this idea from the highly magnified pictures of metals that show the atoms. Notice in Figure 2.1 that atoms look like spheres.

We can combine the hydrogen, oxygen, and carbon atoms in a variety of ways. Just as letters combine to form different words, atoms combine to form different compounds. **Compounds** are substances made by bonding atoms together in specific ways. These substances contain two or more different types of atoms bound together in a particular way. A specific compound consists of the same particles throughout. Table 2.1 shows some examples of atoms combined into compounds.

Table 2.1


Some Common Compounds

Atom Combinations	Name	Characteristics
	carbon monoxide	Carbon monoxide is a poisonous gas.
	carbon dioxide	You breathe out carbon dioxide as a waste material, and plants use carbon dioxide to make oxygen.
	water	Water is the most important liquid on Earth.
	hydrogen peroxide	Hydrogen peroxide is used to disinfect cuts and bleach hair.

Consider a glass of water. If you could magically travel inside the water and examine its individual parts, you would see particles consisting of two hydrogen atoms bonded to an oxygen atom:



We call this particle a molecule. A **molecule** is made up of atoms that are “stuck” together. A glass of water, for example, contains a huge number of molecules packed closely together (see **Figure 2.3**).

Carbon dioxide is another example of a compound. For example, “dry ice”—solid carbon dioxide—contains molecules of the type  packed together as shown in **Figure 2.4**.

Molecule

A collection of atoms bonded together that behave as a unit

Notice that the particles (molecules) in water are all the same. Likewise, all the molecules in dry ice are the same. However, the molecules in water differ from the molecules in dry ice. Water and carbon dioxide are different compounds.

Figure 2.3

A glass of water contains millions of tiny water molecules packed closely together.

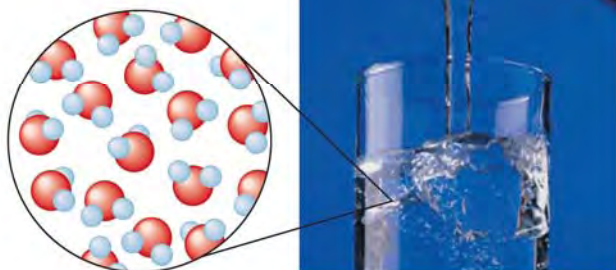
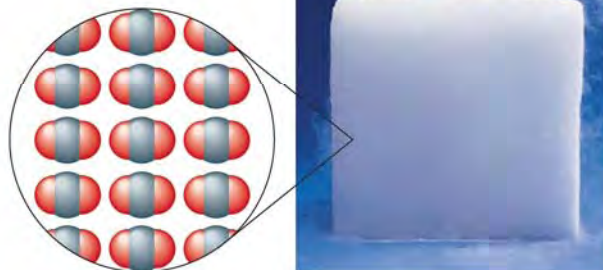




Figure 2.4

Dry ice contains molecules packed closely together.



Elements

Just as hydrogen, oxygen, and carbon can form the compounds carbon dioxide and water, atoms of the same type can also combine with one another to form molecules. For example, hydrogen atoms can pair up , as can oxygen atoms . For reasons we will consider later, carbon atoms form much larger groups, leading to substances such as diamond, graphite, and buckminsterfullerene.

Because pure hydrogen and oxygen each contains only one type of atom, the substances are called elemental substances or, more commonly, **elements**. Elements are substances that contain only one type of atom. For example, pure gold contains only gold atoms, elemental copper contains only copper atoms, and so on. Thus an element contains only one kind of atom; a sample of iron contains many atoms, but they are all iron atoms. Samples of certain pure elements do contain molecules; for example, hydrogen gas contains H—H (usually written H₂) molecules, and oxygen gas contains O—O (O₂) molecules. However, any pure sample of an element contains only atoms of that element, *never* any atoms of any other element.

Figure 2.5 shows examples of elements.

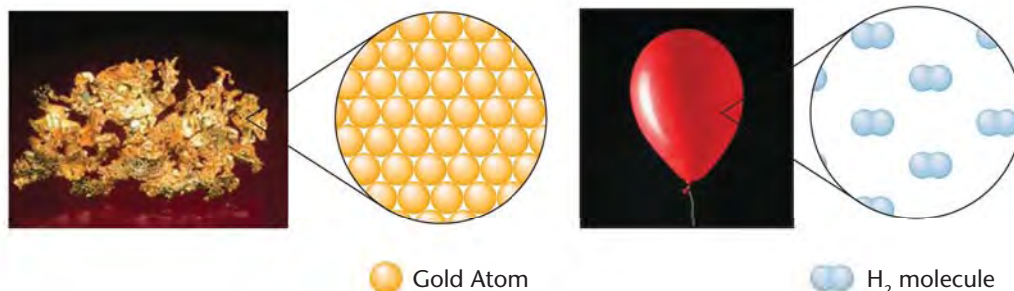


Figure 2.5

Gold and hydrogen are both elements. Gold consists of gold atoms packed together as a solid. Hydrogen is an element that is composed of molecules of hydrogen, not single atoms.

Elements

Substances containing only one type of atom

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Carbon—Element of Many Forms

Did you know that the “lead” in your pencil and the diamond in an engagement ring are made of exactly the same thing—elemental carbon? It may come as a surprise to learn that the black, slippery material that makes up pencil “lead,” which is called **graphite**, has the same composition as the brilliant, hard gemstone called **diamond**. Although both substances are made of carbon atoms, their properties vary dramatically because the carbon atoms are arranged differently in the two substances.

A third form of elemental carbon also exists—**buckminsterfullerene** (see **Figure 2.6**). This strange name comes from a famous industrial designer, Buckminster Fuller, who popularized geodesic domes. The fundamental component of

buckminsterfullerene is a C_{60} molecule that has a framework like a geodesic dome. The C_{60} molecule also resembles a soccer ball with its interconnecting hexagons (six-sided polygons) and pentagons (five-sided polygons). Chemists have nicknamed this form of carbon “bucky balls.” So there are three forms of elemental carbon: graphite, diamond, and buckminsterfullerene.

Because diamonds are so much more valuable than the other forms of carbon, it has long been the dream of entrepreneurs to make gem-quality diamonds from other carbon-based substances. In fact, we have been able to make artificial diamonds for about 50 years. One of the first crude diamonds was actually made by compressing peanut butter at high pressures and temperatures!

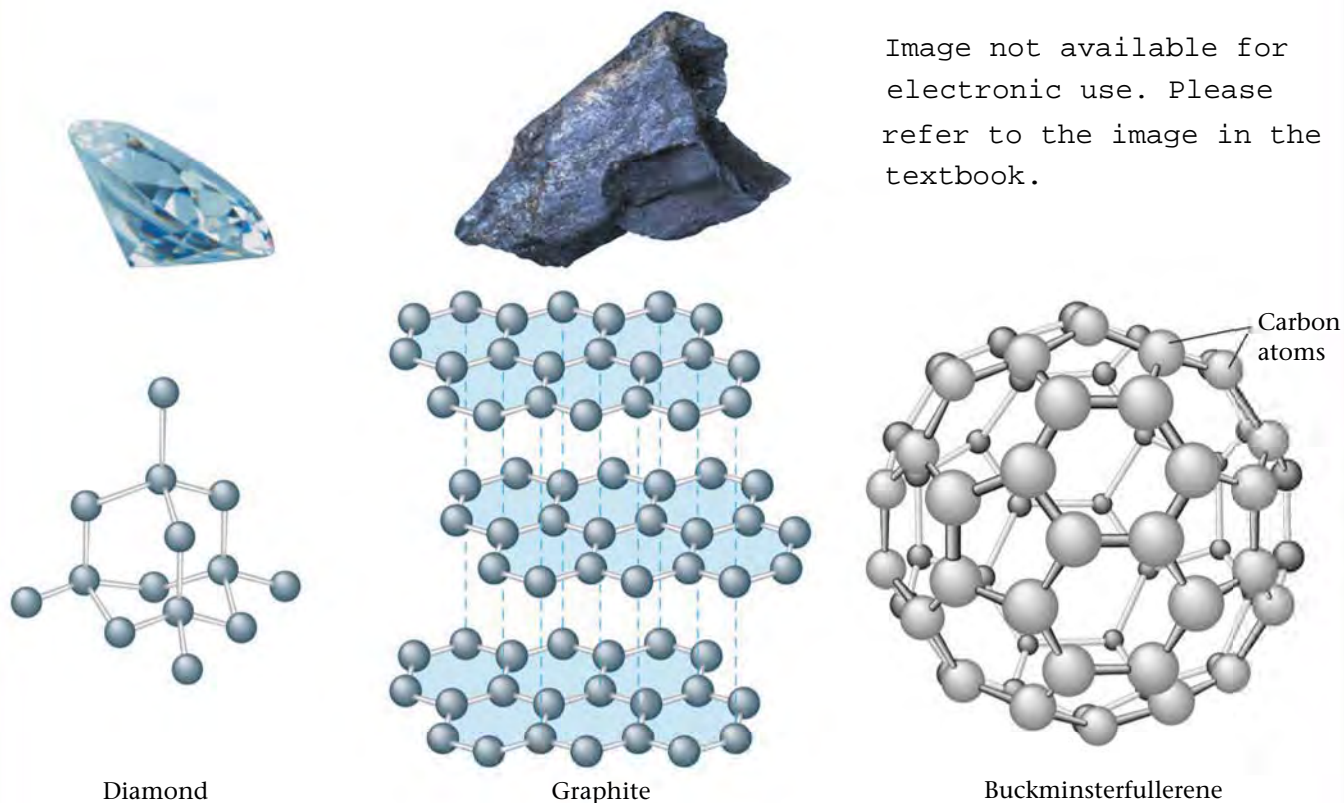




Figure 2.6

The three forms of the element carbon are diamond, graphite, and buckminsterfullerene.

A compound always contains atoms of different elements.

For example, water contains hydrogen atoms and oxygen atoms, and there are always exactly twice as many hydrogen atoms as oxygen atoms because water consists of H—O—H  molecules. A different compound, carbon dioxide, consists of CO_2  molecules and so contains carbon atoms and oxygen atoms (always in the ratio 1:2).

A compound, although it contains more than one type of atom, always has the same composition—that is, the same combination of atoms.

The properties of a compound are typically very different from those of the elements it contains. For example, the properties of water are quite different from the properties of pure hydrogen and pure oxygen.

Active Reading Question

Why are hydrogen molecules called “elements,” but water molecules are called “compounds”?

C. The States of Matter

Water is one of the most familiar substances in our world. We recognize water in three different states: solid, liquid, and gas. If we lower the temperature of liquid water, it freezes—that is, it changes to ice (solid water). On the other hand, if we heat water to its boiling point, it “disappears” into the air as a gas.

The three states of water have distinctly different properties. If a pond freezes in the winter, you can walk across it. Solid water can support your weight. Conversely, you would never try to walk across the same pond in the summertime!

We can also highlight the differing properties of liquid and solid water with food coloring. A drop of food coloring placed on an ice cube just sits there on top of the ice. In contrast, a drop of food coloring placed in liquid water spreads throughout the liquid. The fact that the food coloring spreads in all directions in the liquid water indicates that the water molecules must be moving, bouncing the “food coloring molecules” around and keeping them suspended. This property is very different from that of ice, where the food coloring does not penetrate the surface. Also, we know that gaseous water is quite different from solid and liquid water because it is invisible to the naked eye, unlike the other states of water.

Like water, all substances exist in the *three states of matter*: solid, liquid, and gas. A **solid** is rigid. It has a fixed shape and volume. A **liquid** has a definite volume but takes the shape of its container (see **Figure 2.7**). A **gas** has no fixed volume or shape. It uniformly fills any container.

Active Reading Question

What happens to food coloring when we add it to water? How do our observations support the idea that water is made up of moving molecules?



The three states of water: solid (ice), liquid (water), and gas (water vapor in the air)

Solid

Substance with a fixed shape and volume

Liquid

Substance with a definite volume that takes the shape of its container

Gas

Substance with no definite shape or volume



Figure 2.7

Liquid water takes the shape of its container.

**Full or Empty?****Materials**

- clay
- funnel
- empty soft drink bottle with cap
- water
- large safety pin

Part I: Procedure

1. Use clay to secure a funnel to an empty soft drink bottle. Make sure that the mouth of the bottle is completely covered.
2. Pour water into the funnel, and observe what happens.
3. Develop a method that allows the water to go smoothly into the bottle.
4. Develop a method that allows the water to stay in the funnel.
5. Develop a method that allows the water to periodically squirt into the bottle.

Results/Analysis (Part I)

1. List your observations from step 2.
2. Discuss each of your methods for steps 3–5, and explain why each method works.

Part II: Procedure

1. Fill a plastic soft drink bottle with water, and replace the cap tightly.
2. Using a large safety pin, poke a hole in the side of the container and observe what happens.
3. Hold the bottle over a sink, and unscrew the cap. Observe what happens.

Results/Analysis (Part II)

1. List your observations from step 2, and explain them.
2. List your observations from step 3, and explain them.
3. Relate your findings to those from Part I of this activity.

**SECTION 2.1****REVIEW QUESTIONS**

- 1 How do atoms, compounds, and elements differ?
- 2 Consider the letters in the word *chemistry*. Use them to make as many word “compounds” as is possible with 9 elements. How is an element different from a compound?
- 3 Draw an “atom” picture of a solid, a liquid, and a gas. Describe the essential differences among them.

**RESEARCH LINKS**
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Properties of Matter

Objectives

- To learn to distinguish between physical and chemical properties
- To learn to distinguish between physical and chemical changes

A. Physical and Chemical Properties and Changes

When you see a friend, you immediately respond and call him or her by name. We can recognize a friend because each person has unique characteristics or properties. The person may be thin and tall, may have brown hair and blue eyes, and so on. These characteristics are examples of **physical properties**. Substances also have physical properties. The typical physical properties of a substance include odor, color, volume, state (gas, liquid, or solid), density, melting point, and boiling point.

We can also describe a pure substance in terms of its **chemical properties**, which refer to its ability to form new substances. An example of a chemical change is wood burning in a fireplace, giving off heat and gases and leaving a residue of ashes. In this process, the wood is changed to several new substances. Other examples of chemical changes include the rusting of steel, the digestion of food, and the growth of plants. In a chemical change a given substance changes to a fundamentally different substance or substances.

Active Reading Question

Describe three physical properties of your shampoo.

EXAMPLE 2.1

Identifying Physical and Chemical Properties

Classify each of the following as a physical or a chemical property.

- The boiling point of a certain alcohol is 78 °C.
- Diamonds are very hard.
- Sugar ferments to form alcohol.
- A metal wire conducts an electric current.

Solution

Items (a), (b), and (d) are physical properties; they describe characteristics of each substance, and no change in composition occurs. A metal wire has the same composition before and after an electric current has passed through it. Item (c) is a chemical property of sugar. Fermentation of sugars involves the formation of a new substance, alcohol.

Practice Problem • Exercise 2.1

Which of the following are physical properties and which are chemical properties?

- Gallium metal melts in your hand.
- This page is white.
- The copper sheets that form the “skin” of the Statue of Liberty have acquired a greenish coating over the years.

Key Terms

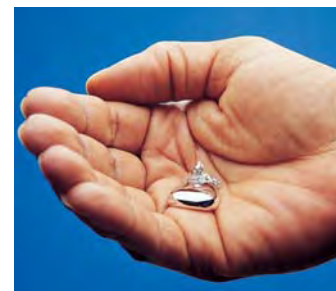
- Physical properties
- Chemical properties
- Physical changes
- Chemical change

Physical property

Characteristic of a substance that can change without the substance's becoming a different substance

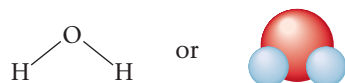
Chemical property

Characteristic that describes the ability of a substance to change to a different substance



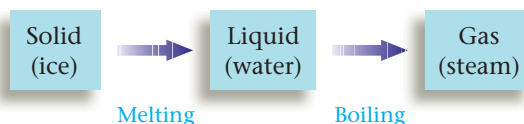
Gallium metal has such a low melting point (30 °C) that it melts from the heat of a hand.

Matter can undergo changes in both its physical and its chemical properties. To illustrate the fundamental differences between physical and chemical changes, we will consider water. As we have seen, a sample of water contains a very large number of individual units (called molecules), each made up of two atoms of hydrogen and one atom of oxygen—the familiar H_2O , which can be represented as:



Here the letters stand for atoms and the lines show attachments, called bonds, between atoms. The molecular model on the right represents water in a more three-dimensional fashion.

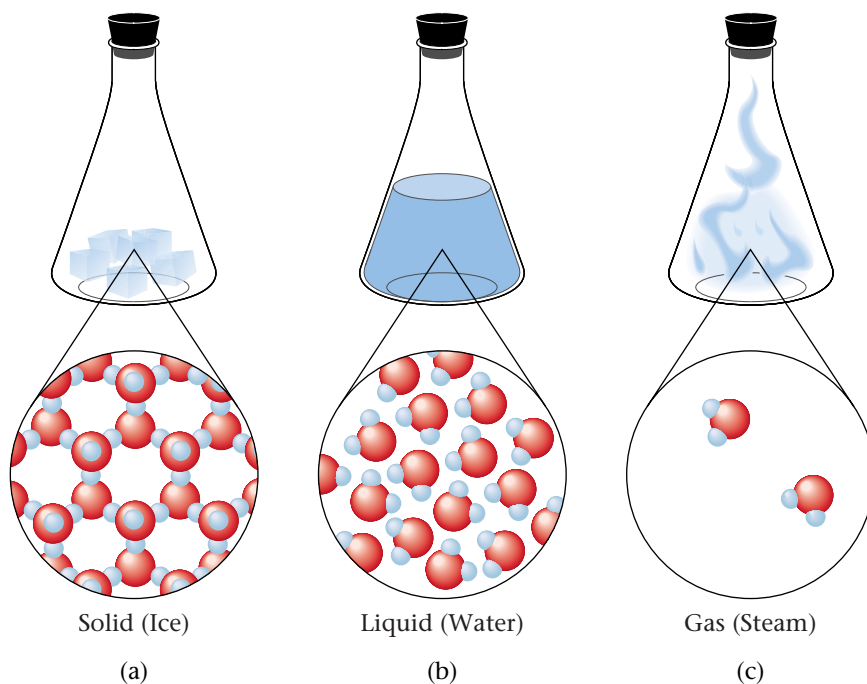
What is really occurring when water undergoes the following changes?



When ice melts, the rigid solid becomes a mobile liquid that takes the shape of its container. Continued heating brings the liquid to a boil, and the water becomes a gas or vapor that seems to disappear into the air. The changes that occur as the substance goes from solid to liquid to gas are represented in **Figure 2.8**. In ice the water molecules are locked into fixed positions. In the liquid the molecules are still very close together, but some motion is occurring; the positions of the molecules are no longer fixed as they are in ice. In the gaseous state the molecules are much farther apart and move randomly, hitting each other and the walls of the container.

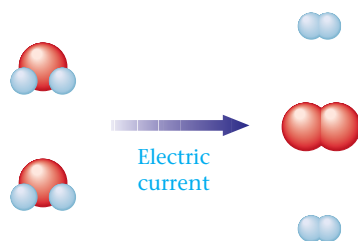
Figure 2.8

The three states of water, red spheres represent oxygen atoms and blue spheres represent hydrogen atoms. (a) Solid: The water molecules are locked into rigid positions and are close together. (b) Liquid: The water molecules are still close together but can move around to some extent. (c) Gas: The water molecules are far apart and move randomly.



The most important thing about all these changes is that the water molecules are still intact. The motions of individual molecules and the distances between them change, *but H_2O molecules are still present*. These changes of state are **physical changes** because they do not affect the composition of the substance. In each state we still have water (H_2O), not some other substance.

Now suppose that we run an electric current through water as illustrated in **Figure 2.9**. Something very different happens. The water disappears and is replaced by two new gaseous substances, hydrogen and oxygen. An electric current actually causes the water molecules to come apart—the water *decomposes* to hydrogen and oxygen. We can represent this process as follows:



This is a **chemical change** because water (consisting of H_2O molecules) has changed into different substances: hydrogen (containing H_2 molecules) and oxygen (containing O_2 molecules). In this process, the H_2O molecules have been replaced by O_2 and H_2 molecules. Let us summarize as follows.

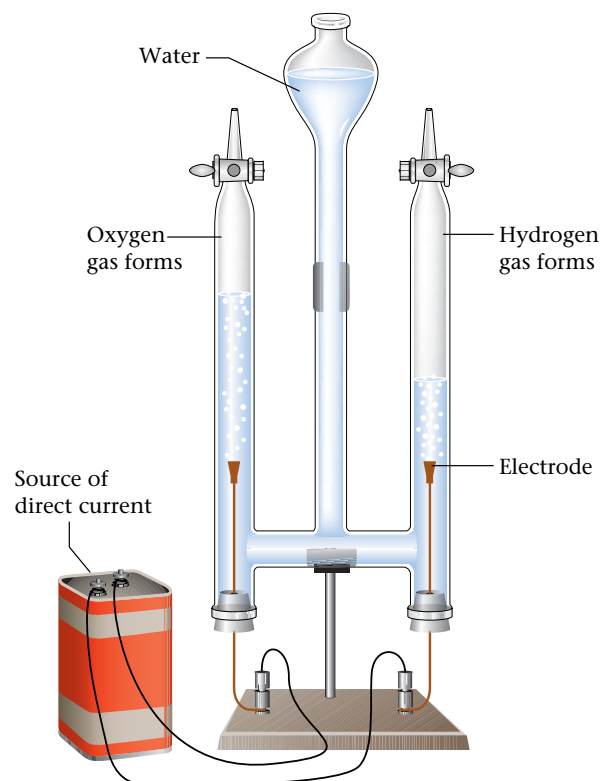


Figure 2.9
Electrolysis, the decomposition of water by an electric current, is a chemical process.

Physical and Chemical Changes

1. A *physical change* involves a change in one or more physical properties, but no change in the fundamental components that make up the substance. The most common physical changes are changes of state: solid \leftrightarrow liquid \leftrightarrow gas.
2. A *chemical change* involves a change in the fundamental components of the substance; a given substance changes into a different substance or substances. Chemical changes are called reactions: silver tarnishes by reacting with substances in the air; a plant forms a leaf by combining various substances from the air and soil; and so on.

Physical change

Change that does not affect the composition of a substance

Chemical change

Change in which a substance becomes a different substance

Active Reading Question

Why are changes of state, for example, liquid water changing to steam, not considered to be chemical changes?

EXAMPLE 2.2

Identifying Physical and Chemical Changes

Classify each of the following as a physical or a chemical change.

- Iron metal is melted.
- Iron combines with oxygen to form rust.
- Wood burns in air.
- A rock is broken into small pieces.

Solution

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Oxygen combines with the chemicals in wood to produce flames. Is a physical or chemical change taking place?

- Melted iron is just liquid iron and could cool again to the solid state. This is a physical change.
- When iron combines with oxygen, it forms a different substance, rust, that contains iron and oxygen. This is a chemical change because a different substance forms.
- Wood burns to form different substances (as we will see later, they include carbon dioxide and water). After the fire, the wood is no longer in its original form. This is a chemical change.
- When the rock is broken up, all the smaller pieces have the same composition as the whole rock. Each new piece differs from the original only in size and shape. This is a physical change.

Practice Problem • Exercise 2.2

Classify each of the following as a chemical change, a physical change, or a combination of the two.

- Milk turns sour.
- Wax is melted over a flame and then catches fire and burns.



SECTION 2.2

REVIEW QUESTIONS

- What is meant by the term *chemical property*? Provide two examples of a chemical change.
- What is meant by the term *physical property*? Provide two examples of a physical change.
- Why is the boiling of water a physical change if water changes from a liquid to a gas?
- Classify the following as physical or chemical changes.
 - Mothballs gradually vaporize in a closet.
 - Calcium chloride lowers the temperature at which water freezes and can be used to melt ice on city sidewalks and roadways.
 - An antacid tablet fizzes and releases carbon dioxide gas when it comes in contact with hydrochloric acid in the stomach.
 - A flashlight battery corrodes and leaks when stored for a long time.
 - Whole milk curdles if vinegar is added to it.
- Are all physical changes accompanied by chemical changes? Are all chemical changes accompanied by physical changes? Explain.

Classifying Matter

Objectives

- To learn to distinguish between mixtures and pure substances
- To learn two methods of separating mixtures

A. Mixtures and Pure Substances

Virtually all of the matter around us consists of mixtures of substances. For example, if you closely observe a sample of soil, you will see that it has many types of components, including tiny grains of sand and remnants of plants. The air we breathe is a mixture, too. Even the water from a drinking fountain contains many substances besides water.

Mixtures

A **mixture** can be defined as something that has variable composition. For example,

- Wood is a mixture; its composition varies greatly depending on the tree from which it originates.
- Soda is a mixture; it contains many dissolved substances, including carbon dioxide gas.
- Coffee is a mixture; it can be strong, weak, or bitter.
- Although it looks very pure, water pumped from deep in the earth is a mixture; it contains dissolved minerals and gases.
- The most common example of a mixture is the air that surrounds us. If we collect a sample of air from a mountaintop where there is no pollution, we find a mixture containing the substances shown in **Figure 2.10**.

Key Terms

- Mixture
- Alloys
- Pure substance
- Homogeneous mixture
- Solution
- Heterogeneous mixture
- Distillation
- Filtration

Mixture

A substance with variable composition

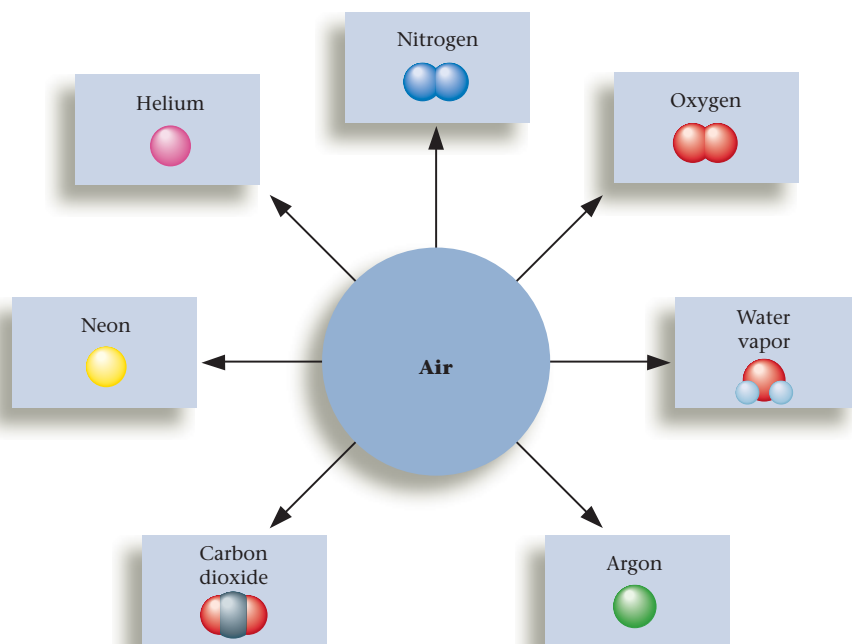
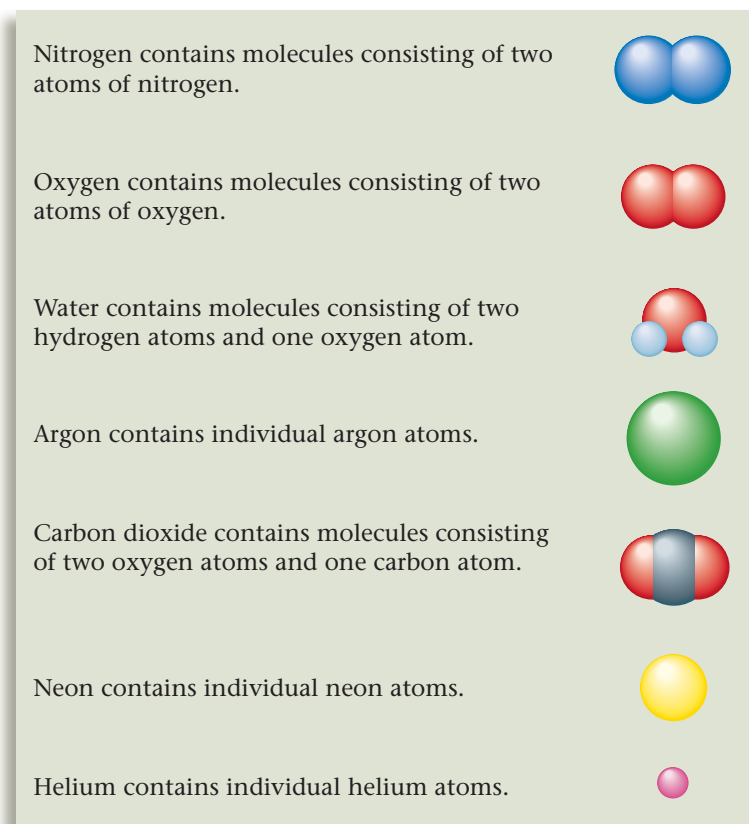


Figure 2.10

Air is composed of a variety of substances.

When we examine each of the substances in air, we find the following:



As you can see from the figure, air is a mixture of elements and compounds. Although air has a similar composition everywhere on earth, differences arise depending on where the air is collected. For example, the amount of water vapor changes greatly.

- Air collected over a desert would include very little water vapor.
- In contrast, on a humid day in Florida, air holds a relatively large amount of water.
- Also, the carbon dioxide content of air would be higher in industrial areas due to the burning of fuels.
- Likewise, other substances (pollutants) would be present due to human activities such as driving cars.

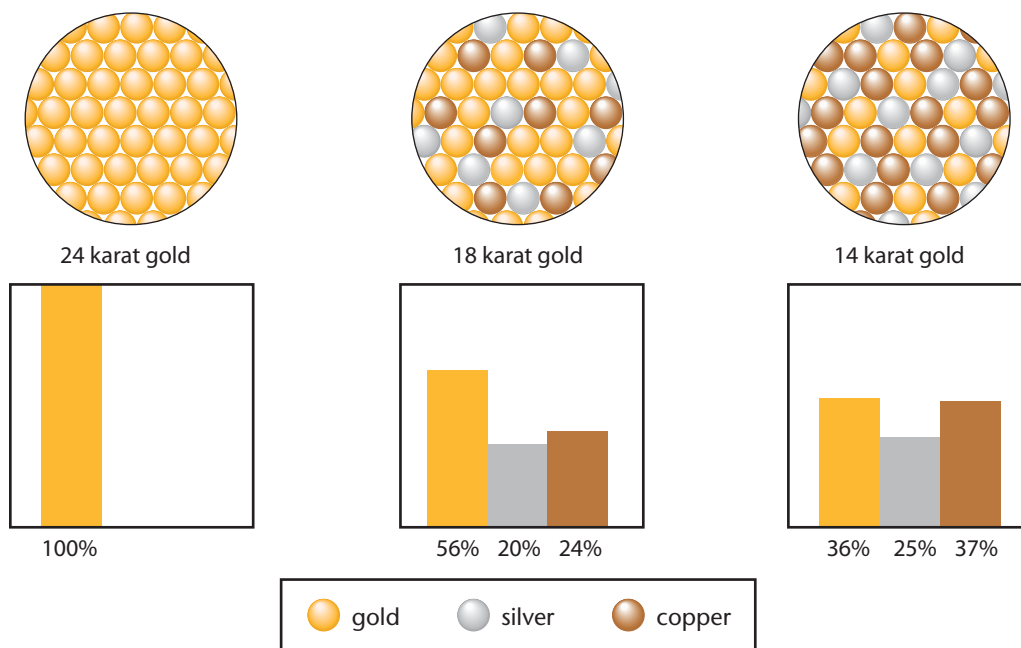
Thus air is not exactly the same everywhere. It has a somewhat different composition depending on where you get it. This variation is a characteristic of mixtures.

The composition of mixtures varies, but the composition of compounds is always the same.



Air is only one of the many mixtures we encounter in our daily lives. For example, if you wear a gold ring, you are wearing a mixture. Pure gold, like that found in the bullion stored in Fort Knox, Kentucky, is really quite soft. Because it bends easily, the pure form is not useful for jewelry. Only when we mix gold with other metals such as silver and copper does it become sturdy enough to make rings and chains.

Your ring might be 14-karat gold—that is, it contains 36 atoms of gold for every 25 atoms of silver and for every 39 atoms of copper. Or your ring might be 18-karat gold—that is, it contains 56 atoms of gold for every 20 atoms of silver and for every 24 atoms of copper. Mixtures of metals are called **alloys**. Many gold alloys exist, containing varying amounts of gold, silver, and copper atoms. Alloys are mixtures—their composition varies. They are not compounds like water. Compounds always have the same atomic composition. This is illustrated in **Figure 2.11**.



Alloy

A mixture of elements that has metallic properties

Figure 2.11

Twenty-four-karat gold is an element. It contains only gold atoms. Fourteen-karat and 18-karat gold are alloys. They contain a mixture of different atoms.

It is important to understand the difference between a compound and a mixture. A compound contains only particles of identical atomic composition. Therefore it always has the same composition. See **Figure 2.12** showing water. In contrast, a mixture is a collection of compounds and/or elements that are present in varying amounts. Thus the composition of a mixture depends on how much of each component is used when the mixture is formed.

Pure Substances

Whereas a mixture has variable composition, a **pure substance** always has the same composition. Pure substances are either elements or compounds. For example, pure water is a compound containing individual H_2O molecules. However, as we find it in nature, liquid water always contains other substances in addition to pure water—it is a mixture. This is obvious from the different tastes, smells, and colors of water samples obtained from various locations. However, if we take great pains to purify samples of water from various sources (such as oceans, lakes, rivers, and the earth's interior), we always end up with the same pure substance—water, which is made up of only H_2O molecules.

Pure water always has the same physical and chemical properties and always consists of molecules containing hydrogen and oxygen in exactly the same proportions, regardless of the original source of the water. The properties of a pure substance make it possible to identify that substance conclusively.

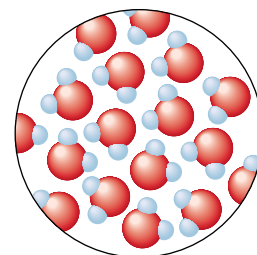


Figure 2.12

Water is a compound. All the components are the same— H_2O molecules.

Pure substance

Either a pure element or a compound

CHEMISTRY IN YOUR WORLD

Connection to History

Alchemical Cymbals

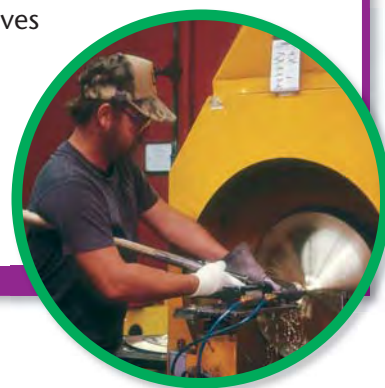
Chemistry is an ancient science dating back to at least 1000 B.C., when early “chemists” discovered how to isolate metals from ores and how to preserve bodies by using embalming fluids. The Greeks were the first to try to figure out why chemical changes occur. By 400 B.C., they had proposed a system of four elements: fire, earth, water, and air. The next 2000 years of chemical history were dominated by a pseudoscience called *alchemy*. Although many alchemists were fakes and mystics, some were serious scientists who made important discoveries.

In fact, did you know that the cymbals used by more than 60% of the rock bands in the world were invented by an alchemist? The story begins 377 years ago in Constantinople, when an alchemist named Avedis discovered an alloy that produced

better-sounding cymbals. This development was important because at that time cymbals were mainly used by armies to frighten their enemies. To honor his achievement, Avedis was given the name Zildjian, which meant “cymbal maker.”

The descendants of that alchemist now run the Avedis Zildjian Company in Norwell, Massachusetts, which manufactures 2000 of the world’s best cymbals every day. The musical world—from classical to rock—truly loves alchemical cymbals.

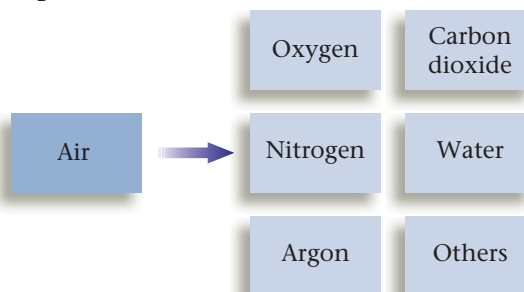
Worker at Avedis Zildjian Company making cymbals



Mixtures can be separated into pure substances: elements and/or compounds.



For example, we have seen that air can be separated into oxygen (element), nitrogen (element), water (compound), carbon dioxide (compound), argon (element), and other pure substances.



Active Reading Question

Give an example of a mixture and an example of a compound and explain how they are different.

Homogeneous and Heterogeneous Mixtures

Mixtures can be classified as either homogeneous or heterogeneous. A **homogeneous mixture** is *the same throughout*. For example, when we dissolve some salt in water and stir well, all regions of the resulting mixture have the same properties. A homogeneous mixture is also called a **solution**. Of course, different amounts of salt and water can be mixed to form various solutions, but a homogeneous mixture (a solution) does not vary in composition from one region of the solution to another (see **Figure 2.13**).

The air around you is a solution—it is a homogeneous mixture of gases. Solid solutions also exist. Brass is a homogeneous mixture of the metals copper and zinc.

A **heterogeneous mixture** contains regions that have different properties from those of other regions. For example, when we pour sand into water, the resulting mixture has one region containing water and another, very different region containing mostly sand (see **Figure 2.14**).

Homogeneous mixture

A mixture that is the same throughout

Solution

A homogeneous mixture

Heterogeneous mixture

A mixture containing regions with differing properties



Figure 2.13

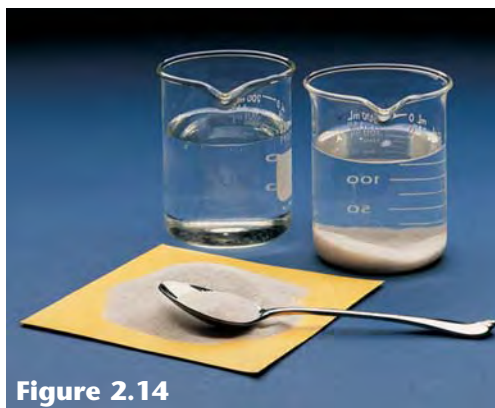


Figure 2.14

Figure 2.13

When table salt is stirred into water (left), a homogeneous mixture called a solution is formed (right).

Figure 2.14

Sand and water do not mix to form a uniform mixture. After the mixture is stirred, the sand settles back to the bottom.



HANDS-ON CHEMISTRY

• MINI-LAB •

Mysterious Mixing

Materials

- clear plastic cup
- food coloring
- water

Procedure

1. Fill a clear plastic cup halfway with water.
2. Carefully place a drop of food coloring on the surface of the water.
3. Without disturbing the water, observe the food coloring for a few minutes.
4. How can you make the food coloring mix more quickly with the water? Design an experiment and discuss it with your teacher.
5. Test the experiment.

Results/Analysis

1. Make a list of your observations from step 3.
2. The water does not appear to be moving. What do your observations tell you about the water molecules? Explain your answer.
3. Draw molecular-level pictures that explain your observations.
4. Explain the results of your experiment to make the food coloring mix more quickly with the water.

EXAMPLE 2.3

Distinguishing Between Mixtures and Pure Substances

Identify each of the following as a pure substance, a homogeneous mixture, or a heterogeneous mixture.

- a. gasoline
- b. a stream with gravel at the bottom
- c. air
- d. copper metal

Solution

- a. Gasoline is a homogeneous mixture containing many compounds.
- b. A stream with gravel on the bottom is a heterogeneous mixture.
- c. Air is a homogeneous mixture of elements and compounds.
- d. Copper metal is a pure substance (an element).

Practice Problem • Exercise 2.3

Classify each of the following as a pure substance, a homogeneous mixture, or a heterogeneous mixture.

- a. maple syrup
- b. the oxygen and helium in a scuba tank
- c. oil and vinegar salad dressing
- d. common salt (sodium chloride)

B. Separation of Mixtures

We have seen that the matter found in nature is typically a mixture of pure substances. For example, seawater is water containing dissolved minerals. We can separate the water from the minerals by boiling, which changes the water to steam (gaseous water) and leaves the minerals behind as solids. If we collect and cool the steam, it condenses to pure water. This separation process, called **distillation**, is shown in **Figure 2.15**.

Distillation

A separation process that depends on the different boiling points of the substances

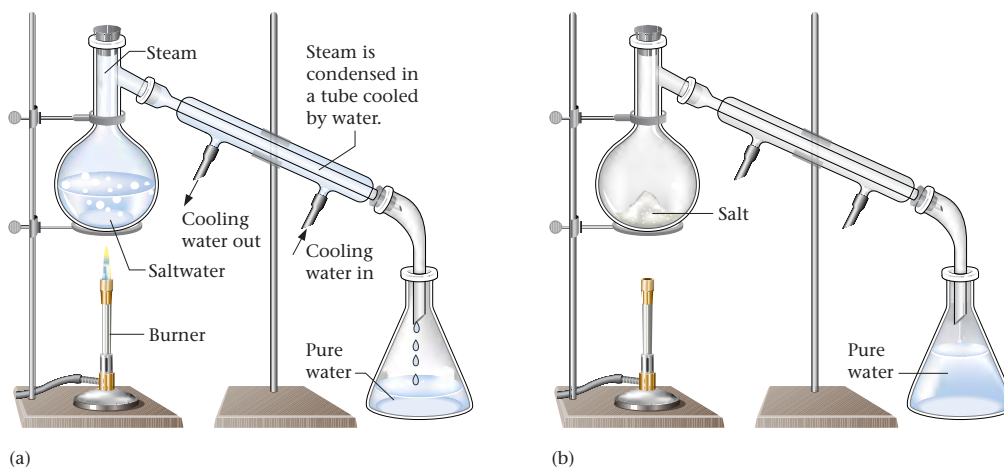


Figure 2.15

Distillation of a solution consisting of salt dissolved in water. (a) When the solution is boiled, steam is driven off. If this steam is collected and cooled, it condenses to form pure water, which drips into the collection flask. (b) After all of the water has been boiled off, the salt remains in the original flask and the water is in the collection flask.

When we carry out the distillation of saltwater, water is changed from the liquid state to the gaseous state and then back to the liquid state. These changes of state are examples of physical changes. We are separating a mixture of substances, but we are not changing the composition of the individual substances. We can represent this as shown in **Figure 2.16**.

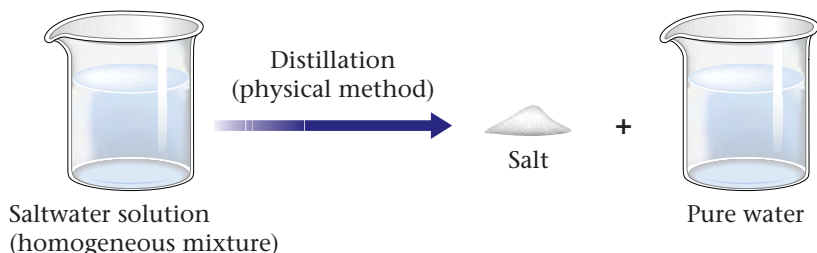


Figure 2.16

No chemical change occurs when salt water is distilled.

Suppose that we scooped up some sand with our sample of seawater. This sample is a heterogeneous mixture, because it contains an undissolved solid as well as the saltwater solution. We can separate out the sand by simple **filtration**. We pour the mixture onto a mesh, such as a filter paper, which allows the liquid to pass through and leaves the solid behind (see **Figure 2.17**). The salt can then be separated from the water by distillation. The total separation process is represented in **Figure 2.18**. All the changes involved are physical changes.

Filtration

Separation of a solid from a liquid by using filter paper

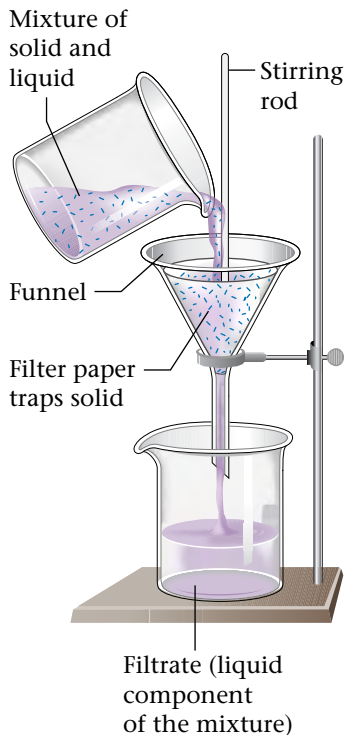
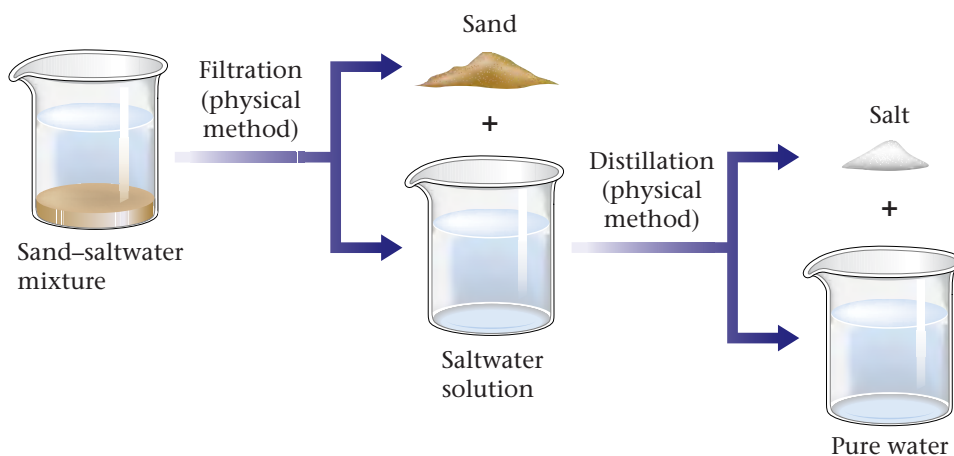


Figure 2.17

Filtration separates a liquid from a solid. The liquid passes through the filter paper, but the solid particles are trapped.

Figure 2.18

Separation of a sand-saltwater mixture



We can summarize the description of matter given in this chapter with the diagram shown in **Figure 2.19**. Note that a given sample of matter can be a pure substance (either an element or a compound) or, more commonly, a mixture (homogeneous or heterogeneous).

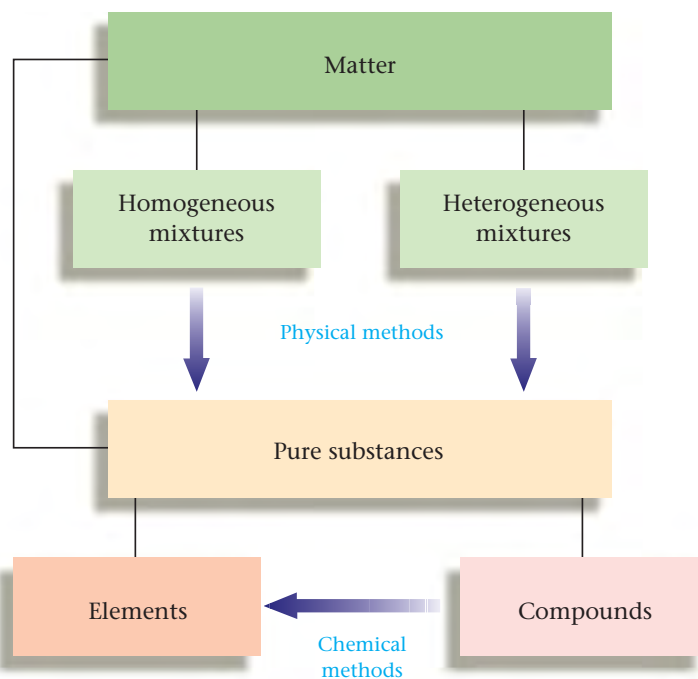
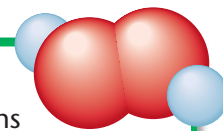


Figure 2.19

The organization of matter

We have seen that all matter exists as elements or can be broken down into elements, the most fundamental substances we have encountered up to this point. We will have more to say about the nature of elements in the next chapter.

CELEBRITY CHEMICAL

Hydrogen Peroxide (H_2O_2)

Hydrogen and oxygen atoms combine to form two different molecules: the very familiar water molecule (H_2O) and the hydrogen peroxide molecule (H_2O_2). Because these molecules have very similar compositions, you might think they should behave in very similar ways. In fact, the properties of water and hydrogen peroxide are very different.

You are already very familiar with water. We drink it, we swim in it, we cook vegetables in it, we wash with it, and so on. Water is essential for life. A few days without it and we die.

Hydrogen peroxide is very different from water. This corrosive liquid would poison us if we were foolish enough to drink it. It is most commonly used as a bleaching agent. For example, hydrogen peroxide bleaches hair by reacting with melanin, the substance responsible for the color of brown and black hair. Hydrogen peroxide changes the

composition of melanin in a way that causes it to lose its color—it turns brown hair blonde. In addition, hydrogen peroxide is used to bleach fibers, such as silk, and to bleach flour, producing the ultra-white powder that consumers demand. Small amounts of hydrogen peroxide are also added to some toothpastes as whitening agents.

One of the most common uses of hydrogen peroxide is to prevent infections in cuts. Sometime in your life when you have cut or scraped yourself, a parent or a nurse probably has applied a liquid to the wound that foamed and burned. That substance was hydrogen peroxide—a powerful *antiseptic* (killer of microorganisms).

Although it looks deceptively similar to water, hydrogen peroxide behaves very differently from water. A small change in the make-up of a molecule can produce big changes in behavior.



SECTION 2.3

REVIEW QUESTIONS

- 1 What is meant by the term *mixture*? Provide two examples of a mixture.
- 2 Why can't mixtures be classified as pure substances?
- 3 Explain the difference between a *heterogeneous mixture* and a *homogeneous mixture*. Provide two examples of each.
- 4 Identify the following as *mixtures* or as *pure substances*.
 - a. milk
 - b. the paper this book is printed on
 - c. a teaspoon of sugar
 - d. a teaspoon of sugar dissolved in a glass of water
 - e. steel
- 5 What is another name for a homogeneous mixture?
- 6 Classify the following mixtures as *homogeneous* or *heterogeneous*.
 - a. vanilla yogurt
 - b. vanilla and strawberry swirl yogurt
 - c. Italian salad dressing
 - d. kitty litter
- 7 Explain how you would separate a sugar water solution into sugar (solid) and water. Is this process a distillation or filtration?



RESEARCH LINKS
CLASSZONE.COM

Key Terms

2.1 Matter


Atoms
Compounds
Molecule
Elements
Solid
Liquid
Gas

2.2 Physical properties
Chemical properties
Physical change
Chemical change

2.3 Mixture
Alloys
Pure substance
Homogeneous mixture
Solution
Heterogeneous mixture
Distillation
Filtration

Key Ideas

2.1 The Nature of Matter

- Matter has mass and occupies space. It is composed of tiny particles called atoms.
- Matter exists in three states:
 - Solid—is a rigid substance with a definite shape
 - Liquid—has a definite volume but takes the shape of its container
 - Gas—takes the shape and volume of its container
- Elements contain only one kind of atom—elemental copper contains only copper atoms, and elemental gold contains only gold atoms.
- Compounds are substances that contain two or more kinds of atoms.
- Compounds often contain discrete molecules.
- A molecule contains atoms bound together in a particular way—an example is , the water molecule, which is written H₂O.

2.2 Properties of Matter

- Matter has both physical and chemical properties.
 - Chemical properties describe a substance's ability to change to a different substance.
 - Physical properties are the characteristics of a substance that do not involve changing to another substance.
 - ◆ Examples are shape, size, and color.
- Matter undergoes physical and chemical changes.
 - A physical change involves a change in one or more physical properties but no change in composition.
 - A chemical change transforms a substance into one or more new substances.

2.3 Classifying Matter

- Matter can be classified as a mixture or a pure substance.
 - A mixture has variable composition.
 - ◆ A homogeneous mixture has the same properties throughout.
 - ◆ A heterogeneous mixture has different properties in different parts of the mixture.
 - A pure substance always has the same composition.
- Mixtures can be separated into pure substances by various means including distillation and filtration.
- Pure substances are of two types:
 - Elements, which cannot be broken down chemically into simpler substances
 - Compounds, which can be broken down chemically into elements

All exercises with **blue numbers** have answers in the back of this book.

2.1 The Nature of Matter

A. The Particulate Nature of Matter

- How do we “know” that matter is composed of atoms?
- Why can't we see atoms with the naked eye?
- Label each of the following as an atomic element, a molecular element, or a compound.



B. Elements and Compounds

- What does it mean to say that a given compound always has the same composition?
- How do the properties of a compound, in general, compare with the properties of the elements of which it is composed? Give an example of a common compound and the elements of which it is composed to illustrate your answer.

C. The States of Matter

- Consider three 10-g samples of water: one as ice, one as a liquid, and one as vapor. How do the volumes of these three samples compare with one another? How is this difference in volume related to the physical state involved?
- In a sample of a gaseous substance, more than 99% of the overall volume of the sample is empty space. How is this fact reflected in the properties of a gaseous substance, compared with the properties of a liquid or solid substance?

2.2 Properties of Matter

A. Physical and Chemical Properties and Changes

Use the correct term (physical or chemical) to complete each of the following:

- Copper is a reddish-brown metal that is easily stretched to make wires. These characteristics are examples of _____ properties of copper.
- When copper metal is heated in concentrated nitric acid, the copper dissolves to form a deep blue solution, and a brown gas is evolved from the acid. These characteristics are examples of _____ changes.

Read the following passage to answer exercises 10 and 11.

Solutions of the substance potassium dichromate are bright orange in color. If a potassium dichromate solution is added to an acidic solution of iron(II) sulfate, the orange color of the potassium dichromate disappears, and the mixture takes on a bright green color as chromium(III) ion forms.

- From the information above, indicate one *physical* property of potassium dichromate in solution.
- From the information above, indicate one *chemical* property of potassium dichromate in solution.
- What is meant by *electrolysis*? Are the changes produced by electrolysis chemical or physical in nature? Give an example to show your reasoning.
- Classify the following as physical changes, chemical changes, physical properties, and chemical properties.
 - An iron scorches a shirt when you leave it on one spot too long.
 - The tires on your car seem to be getting low in very cold weather.
 - Your grandmother's silver tea set gets black with tarnish over time.
 - Spray-on oven cleaner converts grease in the oven into a soapy material.
 - An ordinary flashlight battery begins to leak with age and can't be recharged.
 - Acids produced by bacteria in plaque cause teeth to decay.
 - Sugar chars if overheated while making homemade candy.
 - Hydrogen peroxide fizzes when applied to a wound.
 - Dry ice “evaporates” without melting as time passes.

- State the clues that a chemical reaction has taken place.

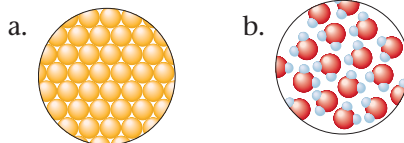


2.3 Classifying Matter

A. Mixtures and Pure Substances

15. Give three examples of heterogeneous *mixtures* and three examples of *solutions* that you might use in everyday life.
16. Classify the following as *mixtures* or as *pure substances*.
- a multivitamin tablet
 - the blue liquid in your car's windshield reservoir
 - a ham and cheese omelet
 - a diamond
17. Classify the following mixtures as *homogeneous* or *heterogeneous*.
- gasoline
 - a jar of jelly beans
 - chunky peanut butter
 - margarine
 - the paper on which this question is printed

18. Label each of the following as a heterogeneous mixture, a homogeneous mixture, or a compound.



B. Separation of Mixtures

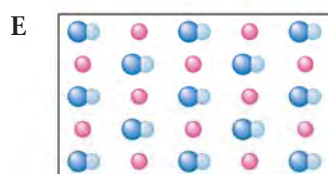
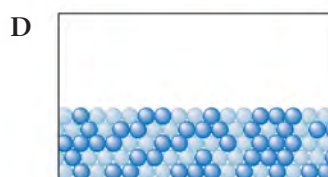
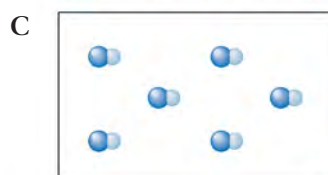
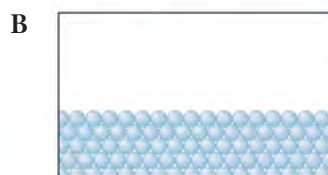
19. Describe how the process of *distillation* could be used to separate a solution into its component substances. Give an example.
20. Describe how the process of *filtration* could be used to separate a mixture into its components. Give an example.
21. In a laboratory experiment, students are asked to determine the relative amounts of benzoic acid and charcoal in a solid mixture. Benzoic acid is relatively soluble in hot water, but charcoal is not. Devise a method for separating the two components of this mixture.
22. Describe the process of distillation depicted in Figure 2.15. Does the separation of the components of a mixture by distillation represent a chemical or a physical change?

Critical Thinking

23. Pure substance X is melted, and the liquid is placed in an electrolysis apparatus like that shown in Figure 2.9. When an electric current is passed through the liquid, a brown solid forms in one chamber and a white solid forms in the other chamber. Is substance X a compound or an element?
24. If a piece of hard white blackboard chalk is heated strongly in a flame, the mass of the piece of chalk will decrease, and eventually the chalk will crumble into a fine white dust. Does this change suggest that the chalk is composed of an element or a compound?
25. During a very cold winter, the temperature may remain below the freezing point for extended periods. However, fallen snow can still disappear, even though it cannot melt. This is possible because a solid can vaporize directly, without passing through the liquid state. Is this process, called sublimation, a physical or a chemical change?
26. Discuss the similarities and differences between a liquid and a gas.
27. Write the correct term: In gaseous substances, the individual molecules are relatively (close/far apart) and are moving freely, rapidly, and randomly.
28. The fact that the substance copper(II) sulfate pentahydrate is bright blue is an example of a _____ property.
29. The fact that the substance copper(II) sulfate pentahydrate combines with ammonia in solution to form a new compound is an example of a _____ property.

- 1 Which of the following is an element?
- A air
 - B water
 - C salt
 - D helium
- 2 Which of the following describes a physical change?
- A Cookies burn in the oven.
 - B A banana ripens.
 - C Sugar dissolves in coffee.
 - D Leaves turn colors in the fall.
- 3 Which of the following is a chemical change?
- A Water condenses on a mirror.
 - B A nail rusts.
 - C A damp towel dries.
 - D Peanuts are crushed.
- 4 Consider the following substances: sodium, sugar, air, iron. How many are *pure* compounds?
- A 1
 - B 2
 - C 3
 - D 4
- 5 A homogeneous mixture is also called ____.
- A a heterogeneous mixture
 - B a pure substance
 - C a compound
 - D a solution
- 6 Which of the following processes requires chemical methods?
- A Separating a homogeneous mixture into pure substances
 - B Breaking a compound into its constituent elements
 - C Separating a heterogeneous mixture into pure substances
 - D Distilling a saltwater mixture

Use the following choices to answer questions 7–10.



- 7 Which best represents a solid element?
- 8 Which best represents a homogeneous mixture of an element and a compound?
- 9 Which best represents a gaseous compound?
- 10 Which best represents a heterogeneous mixture of two elements?
- 11 List one physical property and one chemical property of wood, labeling each.
- 12 Explain the differences among solids, liquids, and gases both macroscopically and microscopically.

Chemical Foundations: Elements, Atoms, and Ions

LOOKING AHEAD

3.1 The Elements

- A. Abundances of Elements
- B. Names and Symbols for the Elements

3.2 Atoms and Compounds

- A. Dalton's Atomic Theory
- B. Formulas of Compounds

3.3 Atomic Structure

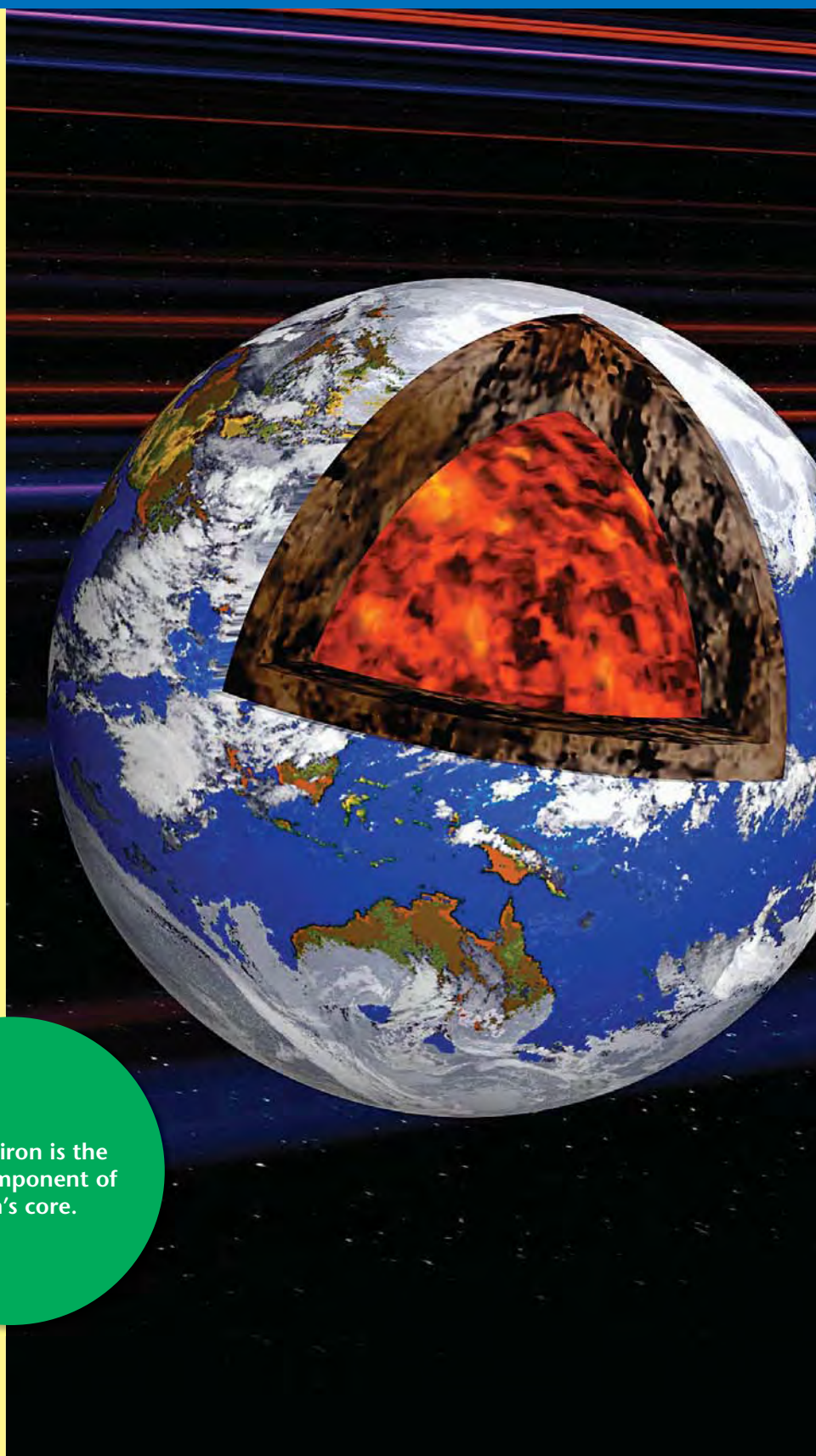
- A. The Structure of the Atom
 - Thomson's Experiment
 - The Plum Pudding Model
 - Rutherford's Experiment
- B. Introduction to the Modern Concept of Atomic Structure
- C. Isotopes

3.4 Using the Periodic Table

- A. Introduction to the Periodic Table
- B. Natural States of the Elements

3.5 Ions and Their Compounds

- A. Ions
 - Cations and Anions
 - Ion Charges and the Periodic Table
- B. Compounds that Contain Ions



Molten iron is the main component of earth's core.

IN YOUR LIFE

The chemical elements are essential to each of us in our daily lives. The most important element is carbon, which is found in virtually all of the molecules that make up the living cell. Although certain elements are present in our bodies in tiny amounts, these elements can have profound effects on our health and behavior. As we will see in this chapter, lithium can be a miracle treatment for someone with manic-depressive disease, and our cobalt levels can have a remarkable impact on behavior. Many elements in our bodies serve no useful purpose, but they are found in the food we eat, the water we drink, and the air we breathe. As a result, the human body typically contains significant amounts of elements such as aluminum, barium, strontium, uranium, and gold, which are usually deposited in the bones or liver.

Many elements have great economic significance. For example, about one billion tons of iron and 60 million tons of aluminum are produced in the world each year and are used mainly in structural materials. Iron is also important in another way that we seldom think about. Earth's core, 4,000 miles in diameter, consists mostly of molten iron. The center of earth's core, with a diameter of about 1,500 miles, may consist of solid iron.

As we will see as we progress in our study of chemistry, most of the chemical elements have an important impact on our lives.

WHAT DO YOU KNOW?



Prereading Questions

1. What elements do you think are most abundant in the human body?
2. Do you know any chemical formulas for compounds? For example, what is the chemical formula for water? What is the chemical formula for table salt? Do you know any other formulas?
3. Have you heard the terms *electrons*, *protons*, *neutrons*, or *nucleus*? What do you know about these?
4. What does the term *periodic* mean? How do you think this applies to the term *periodic table*?
5. Name at least one element that normally exists as a solid and at least one element that normally exists as a gas.

SECTION 3.1

The Elements

Key Term

- Element symbols

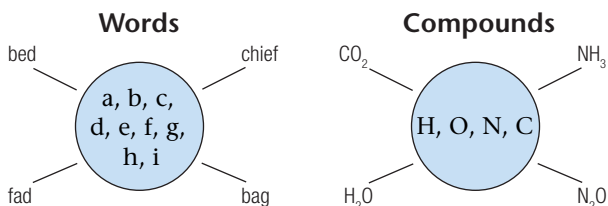
Objectives

- To learn about the relative abundances of the elements
- To learn the names of some elements
- To learn the symbols of some elements

The Greeks were the first to try to explain why chemical changes occur. By about 400 B.C. they had proposed that all matter was composed of four fundamental substances: fire, earth, water, and air.

The next 2000 years of chemical history were dominated by alchemy. Some alchemists were mystics and fakes who were obsessed with the idea of turning cheap metals into gold. However, many alchemists were sincere scientists and this period saw important events: the elements mercury, sulfur, and antimony were discovered, and alchemists learned how to prepare acids.

As we saw in Chapter 2, all of earth's materials (and those of other parts of the universe) can be broken down chemically into about 100 different elements. At first it might seem amazing that the millions of known substances are composed of so few fundamental elements. Fortunately for those trying to understand and systematize it, nature often uses a relatively small number of fundamental units to assemble even extremely complex materials. For example, proteins, a group of substances that serve the human body in almost uncountable ways, are all made by linking together a few fundamental units to form huge molecules. A nonchemical example is the English language, in which hundreds of thousands of words are constructed from only 26 letters. Compounds are made by combining atoms of the various elements, just as words are constructed from the 26 letters of the alphabet.



CHEMISTRY EXPLORERS

Robert Boyle • 1627–1691

The first scientist to recognize the importance of careful measurements was Ireland's Robert Boyle. He is best known for his pioneering work on the properties of gases, but Boyle's most important contribution to science was probably his insistence that science should be firmly grounded in experiments. For example, Boyle held no preconceived notions about how many elements there might be. His definition of the term *element* was based on experiments: a substance was an element unless it could be broken down into two or more simpler substances. For example, air could not be an element, as the Greeks believed, because it could be broken down into many substances. As Boyle's experimental definition of an element became generally accepted, the list of known elements grew and the Greek system of four elements died.



Just as you had to learn the letters of the alphabet before you could learn to read and write, you need to learn the names and symbols of the chemical elements before you can read and write chemistry.

A. Abundances of Elements

At present, about 115 different elements are known, 88 of which occur naturally. (The rest have been made in laboratories.) The elements vary greatly in abundance. In fact, only nine elements account for most of the compounds found in earth's crust. In **Table 3.1**, the elements are listed in order of abundance (mass percent) in earth's crust, oceans, and atmosphere.

Table 3.1

Distribution (Mass Percent) of the 18 Most Abundant Elements in Earth's Crust, Oceans, and Atmosphere

Element	Mass Percent	Element	Mass Percent
oxygen	49.2	titanium	0.58
silicon	25.7	chlorine	0.19
aluminum	7.50	phosphorus	0.11
iron	4.71	manganese	0.09
calcium	3.39	carbon	0.08
sodium	2.63	sulfur	0.06
potassium	2.40	barium	0.04
magnesium	1.93	nitrogen	0.03
hydrogen	0.87	fluorine	0.03
		all others	0.49

Active Reading Questions

1. What element is nearly half of earth's crust, oceans, and atmosphere?
2. What percent of the mass of the crust, ocean, and atmosphere do the top nine elements make?

Oxygen, in addition to accounting for about 20 percent of earth's atmosphere (where it occurs as O_2 molecules), is also found in virtually all rocks, sand, and soil on earth's crust. In these materials, oxygen is not present as O_2 molecules but exists in compounds that usually contain silicon and aluminum atoms.

The list of elements found in living matter is very different from that for earth's crust, as shown in the Top Ten. Oxygen, carbon, hydrogen, and nitrogen form the basis for all biologically important molecules. Some elements found in the body (called trace elements) are crucial for life, even though they are present in relatively small amounts (**Table 3.2**). For example, chromium helps the body use sugars to provide energy.

DID YOU KNOW

The number of elements changes regularly as new elements are made in particle accelerators.

Top Ten Elements in the Human Body

Element	Mass Percent
Oxygen	65.0
Carbon	18.0
Hydrogen	10.0
Nitrogen	3.0
Calcium	1.4
Phosphorus	1.0
Magnesium	0.50
Potassium	0.34
Sulfur	0.26
Sodium	0.14


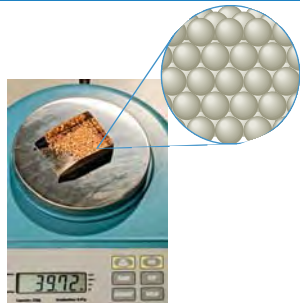
Table 3.2

Trace Elements in the Human Body

Trace Elements

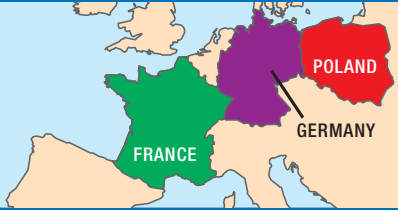
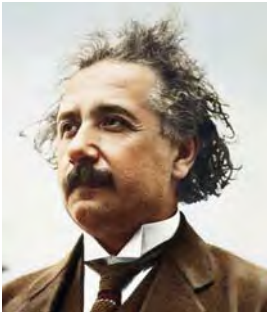


arsenic
chromium
cobalt
copper
fluorine
iodine
manganese
molybdenum
nickel
selenium
silicon
vanadium

One more general comment is important at this point. As we have seen, elements are fundamental to understanding chemistry. However, students may be confused by the many different ways that chemists use the term *element* illustrated below.

Word	Meaning	
Element	Microscopic form Single atom of that element	
Element	Macroscopic form Sample of that element large enough to weigh on a balance	
Element	Generic form When we say the human body contains the element sodium or lithium, we do not mean that free elemental sodium or lithium is present. Rather we mean that atoms of these elements are present in some form.	

B. Names and Symbols for the Elements

The names of the chemical elements have come from many sources. Sometimes the names come from descriptions of the element's properties, sometimes the name reflects the place where the element was discovered, and sometimes the name honors a famous scientist.

Sources of Names of Chemical Elements						
Greek, Latin, or German words describing a property of the element	<i>gold</i> aurum (shining down)	Latin <i>lead</i> plumbum (heavy)	<i>chlorine</i> color	Greek <i>iodine</i> color	<i>bromine</i> stench	
Place where the element was discovered	<i>americium</i> → America	<i>germanium</i> → Germany				
	<i>californium</i> → California	<i>polonium</i> → Poland				
		<i>francium</i> → France				
Famous scientists						
	Albert Einstein <i>einsteinium</i>	Alfred Nobel <i>nobelium</i>	Marie Curie <i>curium</i>			

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Trace Elements: Small but Crucial

We all know that certain chemical elements, such as calcium, carbon, nitrogen, and iron, are essential for humans to live. Many other elements, present in tiny amounts, are also essential for life.

In the Body



Chromium assists in the metabolism of sugars, cobalt is present in vitamin B₁₂, iodine is necessary for the proper functioning of the thyroid gland, manganese plays a role in maintaining proper calcium levels in the bones, and copper is involved in the production of red blood cells.

In Our Personal Environment



Exposure to trace elements in water, food, and air also affects our health. Exposure to aluminum in baked goods and cheese as well as from cookware and utensils is of concern to scientists.

In Behavior

Lithium has helped some people with bipolar disorder, a disease in which a person's behavior varies between inappropriate highs and the blackest of depressions.

Studies on inmates at a prison in Illinois have shown that trace elements in the hair of prisoners is related to behaviors. There is an inverse relationship between the levels of cobalt in the hair and the degree of violence in the prisoners.

We often use abbreviations to simplify the written word. For example, it is much easier to put MA on an envelope than to write out Massachusetts, and we often write USA instead of United States of America. Likewise, chemists have invented a set of abbreviations or **element symbols** for the chemical elements. These symbols usually consist of the first letter or the first two letters of the element names. The first letter is always capitalized, and the second is not. For example,

fluorine	F	neon	Ne
oxygen	O	silicon	Si
carbon	C		

Information

In the symbol for an element, only the first letter is capitalized.

Sometimes, however, the two letters used are not the first two letters in the name. For example,

zinc	Zn	cadmium	Cd
chlorine	Cl	platinum	Pt

The symbols for some other elements are based on the original Latin or Greek name.

<i>Current Name</i>	<i>Original Name</i>	<i>Symbol</i>
gold	aurum	Au
lead	plumbum	Pb
sodium	natrium	Na
iron	ferrum	Fe

A list of the most common elements and their symbols is given in **Table 3.3**. You can also see the elements represented on a table in the inside back cover of this text. We will explain the form of this table (which is called the periodic table) in later chapters.

Active Reading Questions

1. Give three examples of elements whose symbols are the first two letters of the name of the element.
2. Give three examples of elements with symbols in which the first letter is not the first letter of the name of the element.

Table 3.3

The Names and Symbols of the Most Common Elements*

Element	Symbol	Element	Symbol
aluminum	Al	lithium	Li
arsenic	As	mercury (hydrargyrum)	Hg
barium	Ba	neon	Ne
boron	B	nitrogen	N
bromine	Br	oxygen	O
calcium	Ca	platinum	Pt
carbon	C	potassium (kalium)	K
chromium	Cr	silicon	Si
cobalt	Co	silver (argentium)	Ag
copper (cuprum)	Cu	sodium (natrium)	Na
gold (aurum)	Au	sulfur	S
lead (plumbum)	Pb	zinc	Zn

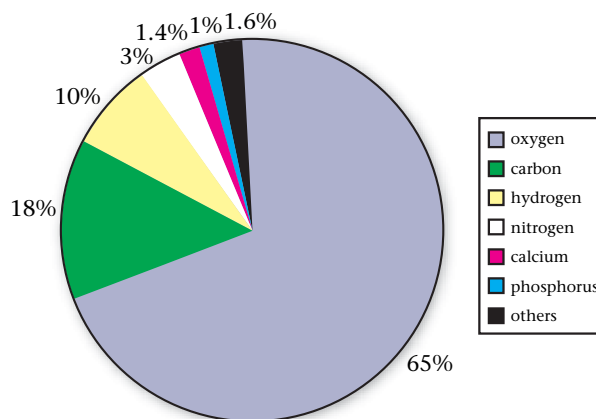
*Where appropriate, the original name is shown in parentheses so that you can see the sources of some of the symbols.



SECTION 3.1 REVIEW QUESTIONS

- Give four different chemical meanings for the word element and an example of each.
- How many of the Top Ten elements in the human body are also Top Ten elements in the earth's crust, ocean, and atmosphere? What are these elements?
- Make a list of elements that you think were named after a place.
- Give the symbol for each of the given elements.
 - Aluminum
 - Sodium
 - Mercury
 - Helium
 - Oxygen
 - Potassium
- Give the name for each of the given symbols.
 - Ca
 - P
 - S
 - Si
 - Mg
 - C

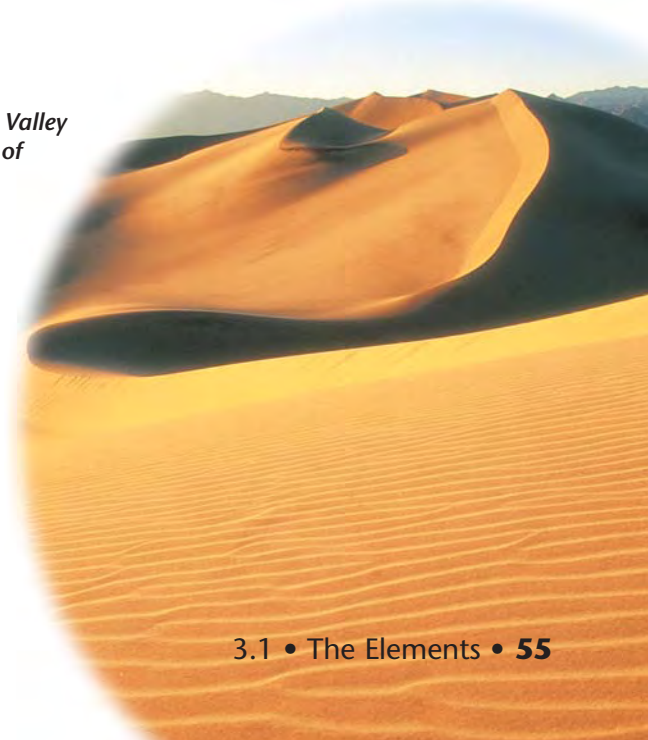
- The graph below shows the mass percentages of the five most abundant elements in the human body. Use this graph to answer the following questions.



- What is the fourth most abundant element by mass in the human body?
- Which element is almost twice as abundant by mass as hydrogen in the human body?
- What percentage of the human body is *not* made up of oxygen, carbon, or hydrogen?



The sand in these dunes at Death Valley National Monument is composed of silicon and oxygen atoms.



SECTION 3.2

Atoms and Compounds

Key Terms

- Law of constant composition
- Dalton's atomic theory
- Atoms
- Compound
- Chemical formula

Law of constant composition

A given compound always contains elements in exactly the same proportion by mass.

Objectives

- To learn about Dalton's theory of atoms
- To understand and illustrate the Law of constant composition
- To learn how a formula describes a compound's composition

A. Dalton's Atomic Theory

As scientists of the eighteenth century studied the nature of materials, several things became clear:

1. Most natural materials are mixtures of pure substances.
2. Pure substances are either elements or combinations of elements called compounds.
3. A given compound always contains the same proportions (by mass) of the elements. For example, water *always* contains 8 g of oxygen for every 1 g of hydrogen, and carbon dioxide *always* contains 2.7 g of oxygen for every 1 g of carbon. This principle became known as the **law of constant composition**. It means that a given compound always has the same composition, regardless of where it comes from.

John Dalton, an English scientist and teacher, was aware of these observations. In the early 1800s, he offered an explanation for them that became known as **Dalton's atomic theory**. The main ideas of this theory (model) can be stated as follows:

Dalton's Atomic Theory

1. Elements are made of tiny particles called **atoms**.
2. All atoms of a given element are identical.
3. The atoms of a given element are different from those of any other element.
4. Atoms of one element can combine with atoms of other elements to form compounds. A given compound always has the same relative numbers and types of atoms.
5. Atoms are indivisible in chemical processes. That is, atoms are not created or destroyed in chemical reactions. A chemical reaction simply changes the way the atoms are grouped together.

Dalton's model successfully explained important observations such as the law of constant composition. This law makes sense because if a compound always contains the same relative numbers of atoms, it will always contain the same proportions by mass of the various elements.

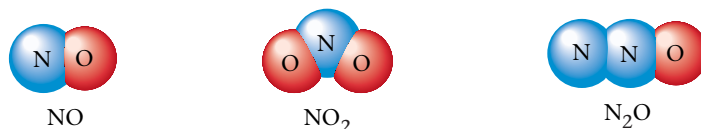
CHEMISTRY EXPLORERS

John Dalton • 1766–1844

Dalton was an English scientist who made his living as a teacher in Manchester. Although Dalton is best known for his atomic theory, he made contributions in many other areas, including meteorology (he recorded daily weather conditions for 46 years, producing a total of 200,000 data entries). A rather shy man, Dalton was colorblind to red (a special handicap for a chemist) and suffered from lead poisoning contracted from drinking substances that had been drawn through lead pipes.



Like most new ideas, Dalton's model was not accepted immediately. However, Dalton was convinced that he was right and *used his model to predict* how a given pair of elements might combine to form more than one compound. Dalton pictured compounds as collections of atoms. For example, nitrogen and oxygen might form a compound containing one atom of nitrogen and one atom of oxygen (written NO), a compound containing two atoms of nitrogen and one atom of oxygen (written N₂O), or a compound containing one atom of nitrogen and two atoms of oxygen (written NO₂).



When the existence of these substances was verified, it was a triumph for Dalton's model. Because Dalton was able to predict correctly the formation of multiple compounds between two elements, his atomic theory became widely accepted.

Active Reading Question

How does Dalton's model explain the law of constant composition?

B. Formulas of Compounds

A **compound** is a distinct substance that is composed of the atoms of two or more elements and always contains exactly the same relative masses of those elements. In light of Dalton's atomic theory, this statement simply means that a compound always contains the same relative numbers of atoms of each element. For example, water always contains two hydrogen atoms for each oxygen atom.

The types of atoms and the number of each type in each unit (molecule) of a given compound are conveniently expressed by a **chemical formula**.

In a chemical formula, the atoms are indicated by the element symbols and the number of each type of atom is indicated by a subscript, a number that appears to the right of and below the symbol for the element.

The formula for water is written H₂O, indicating that each molecule of water contains two atoms of hydrogen and one atom of oxygen

READING TIP

Here, *relative* refers to ratios.

Compound

A substance made of two or more different elements joined together in a specific way

(the subscript 1 is always understood and not written). Following are some general rules for writing formulas:

Tools for Writing Formulas

1. Each atom present is represented by its element symbol.
2. The number of each type of atom is indicated by a subscript written to the right of the element symbol.
3. When only one atom of a given type is present, the subscript 1 is not written.

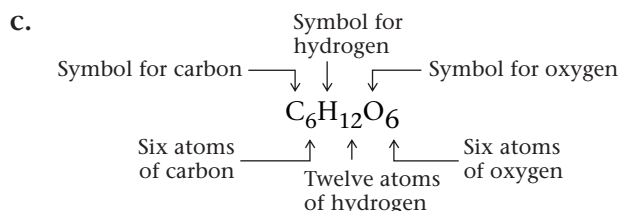
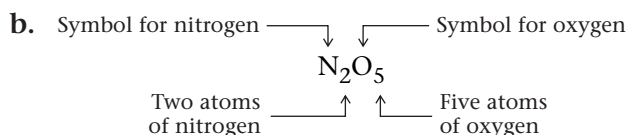
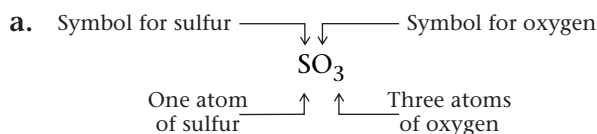
EXAMPLE 3.1

Writing Formulas of Compounds

Write the formula for each of the following compounds, listing the elements in the order given.

- a. Each molecule of a compound that has been implicated in the formation of acid rain contains one atom of sulfur and three atoms of oxygen.
- b. Each molecule of a certain compound contains two atoms of nitrogen and five atoms of oxygen.
- c. Each molecule of glucose, a type of sugar, contains six atoms of carbon, twelve atoms of hydrogen, and six atoms of oxygen.

Solution



Practice Problem • Exercise 3.1

Write the formula for each of the following compounds, listing the elements in the order given.

- a. A molecule contains four phosphorus atoms and ten oxygen atoms.
- b. A molecule contains one uranium atom and six fluorine atoms.
- c. A compound contains one aluminum atom for every three chlorine atoms.

CELEBRITY CHEMICAL Iron (Fe)

Iron is a very important element that lies at the very heart of the earth. In fact, molten iron is thought to be the main component of the earth's core. Iron is also the fourth most abundant element in the earth's crust, found mainly in compounds with oxygen.

The earliest evidence of human use of iron dates back to about 4000 B.C. and takes the form of iron beads that are thought to have come from meteors striking the earth. The first humans to obtain iron from the ores found in the earth's crust were the Hittite peoples of Asia Minor in the third millennium B.C. The way in which the Hittites made iron weapons was one of the great military secrets of the ancient world. The process became widely known only after the fall of the Hittites around 1200 B.C., leading to the "Iron Age."

Of course, the major importance of iron in the modern world relates to its presence in steel. Steel is an alloy composed mainly of iron mixed with

carbon and other metals. The principal structural material of our civilization, the annual production of steel amounts to nearly a billion tons.

Although iron is extremely important as a construction material, it is even more important to chemistry in the human body. Without the iron compounds in our systems, we would die immediately. Iron compounds in the blood absorb oxygen from the air and transport it to the tissues, where it is stored by other iron compounds. Still more iron compounds assist oxygen in reacting with the fuel from our food to provide us with the energy to live, work, and play. Iron is truly essential to our lives.



Two iron age axes found in Spain

SECTION 3.2 REVIEW QUESTIONS

- What observations led Dalton to propose his model for the atom?
- What are the main ideas of Dalton's atomic theory?
- How was Dalton able to predict the formation of multiple compounds between the elements?
- Did Dalton's model of the atom include protons, neutrons, and electrons inside the atoms? Explain.
- What information is given in a chemical formula?
- Write the formula for each of the following substances, listing the elements in the order given.
 - A molecule containing two hydrogen atoms, one sulfur atom, and four oxygen atoms
 - List the hydrogen first.
 - A molecule containing one carbon atom and four hydrogen atoms
 - List the carbon first.
 - A molecule containing one sulfur atom and five fluorine atoms
 - List the sulfur first.

SECTION 3.3

Atomic Structure

Key Terms

- Electrons
- Nuclear atom
- Nucleus
- Proton
- Neutron
- Isotopes
- Atomic number
- Mass number

Objectives

- To learn about the internal parts of an atom
- To understand Rutherford's experiment
- To describe some important features of subatomic particles
- To learn about the terms isotope, atomic number, and mass number
- To understand the use of the symbol A_ZX to describe a given atom

A. The Structure of the Atom

Dalton's atomic theory, proposed in the early 1800s, provided such a convincing explanation for the composition of compounds that it became generally accepted. Scientists came to believe that *elements consist of atoms and that compounds are a specific collection of atoms* bound together in some way. As does any new theory, Dalton's model of the atom spawned many new questions, including "What causes atoms to 'stick together' to form compounds?" and "What is an atom like?" An atom might be a tiny ball of matter that is the same throughout with no internal structure—like a ball bearing. Or the atom might be composed of parts—it might be made up of a number of subatomic particles. However, if the atom contains parts, there should be some way to break up the atom into its components.

Many scientists pondered the nature of the atom during the 1800s, but it was not until almost 1900 that convincing evidence became available to show that the atom has a number of different parts.

Active Reading Question

What are some questions that Dalton's model of the atom does not answer?

Thomson's Experiment

In the late 1890s, a British physicist named J. J. Thomson used a cathode ray tube (see **Figure 3.1**) to show that the atoms of any element can be made to emit tiny negative particles. (He knew they had a negative charge because he could show that they were repelled by the negative part of an electric field.) Thus he concluded that all types of atoms must contain these negative particles, which are now called **electrons**.

On the basis of his results, Thomson wondered what an atom must be like. Although atoms contain these tiny negative particles, he also

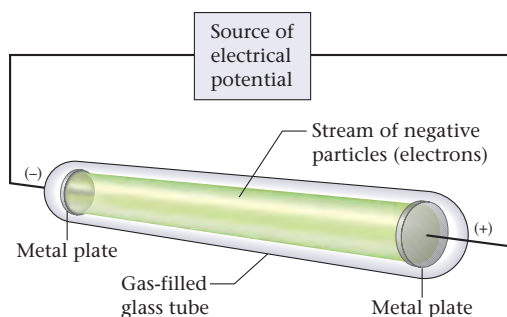
knew that whole atoms are not negatively *or* positively charged. He concluded that the atom must also contain positive particles that balance exactly the negative charge carried by the electrons, giving the atom a zero overall charge.

Electron

A negatively charged subatomic particle

Figure 3.1

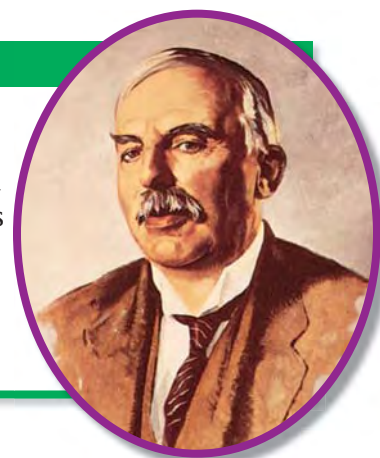
Schematic of a cathode ray tube. A stream of electrons passes between the electrodes. The fast-moving particles excite the gas in the tube, causing a glow between the plates.



CHEMISTRY EXPLORERS

Ernest Rutherford • 1871–1937

Rutherford was born on a farm in New Zealand. In 1895 he placed second in a scholarship competition to attend Cambridge University, but he was awarded the scholarship when the winner decided to stay home and get married. Rutherford was an intense, hard-driving person who became a master at designing just the right experiment to test a given idea. He was awarded the Nobel Prize in chemistry in 1908.



The Plum Pudding Model

Given J. J. Thomson's results, it was natural to wonder what the atom might look like. J. J. Thomson and William Thomson (better known as Lord Kelvin, and no relation to J. J.) are credited with proposing that the atom might be something like plum pudding (a pudding with raisins randomly distributed throughout). They reasoned that the atom might be thought of as a uniform "pudding" of positive charge with enough negative electrons scattered within to counterbalance that positive charge (see **Figure 3.2**). Thus the plum pudding model of the atom came into being.

Rutherford's Experiment

If you had taken this course in 1910, the plum pudding model would have been the only picture of the atom described. However, our ideas about the atom were changed dramatically in 1911 by a physicist named Ernest Rutherford, who learned physics in J. J. Thomson's laboratory in the late 1890s. By 1911 Rutherford had become a distinguished scientist with many important discoveries to his credit. One of his main areas of interest involved alpha particles (α particles), positively charged particles with a mass approximately 7500 times that of an electron. In studying the flight of these particles through air, Rutherford found that some of the α particles were deflected by something in the air. Puzzled by this observation, he designed an experiment that involved directing α particles toward a thin metal foil. Surrounding the foil was a detector coated with a substance that produced tiny flashes wherever it was hit by an α particle (see **Figure 3.3**). The results of the experiment were very different from those Rutherford anticipated. Although most of the α particles passed straight through the foil, some of them were deflected at large angles, as shown in Figure 3.3, and some were reflected backward.

This outcome was a great surprise to Rutherford. (He described this result as comparable to shooting a gun at a piece of paper and having the bullet bounce back.) Rutherford knew that if the plum pudding model of the atom were correct, the massive α particles would crash through the thin foil like cannonballs through paper (as shown in **Figure 3.4a**). So he expected the α particles to travel through the foil experiencing, at most, very minor deflections of their paths.

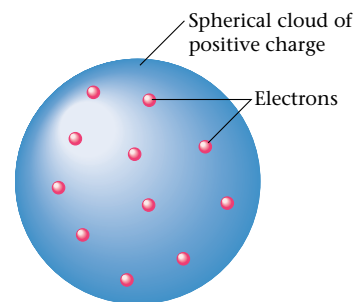


Figure 3.2

One of the early models of the atom was the plum pudding model, in which the electrons were pictured as embedded in a positively charged spherical cloud, much as raisins are distributed in an old-fashioned plum pudding.

DID YOU KNOW

If the atom were expanded to the size of a huge domed stadium, the nucleus would be only about as big as a fly at the center of the stadium.

CRITICAL THINKING ?

You have learned about three different models of the atom: Dalton's model, Thomson's model, and Rutherford's model.

What if Dalton had been correct? What would Rutherford have expected from his experiments with gold foil?

What if Thomson had been correct? What would Rutherford have expected from his experiments with gold foil?

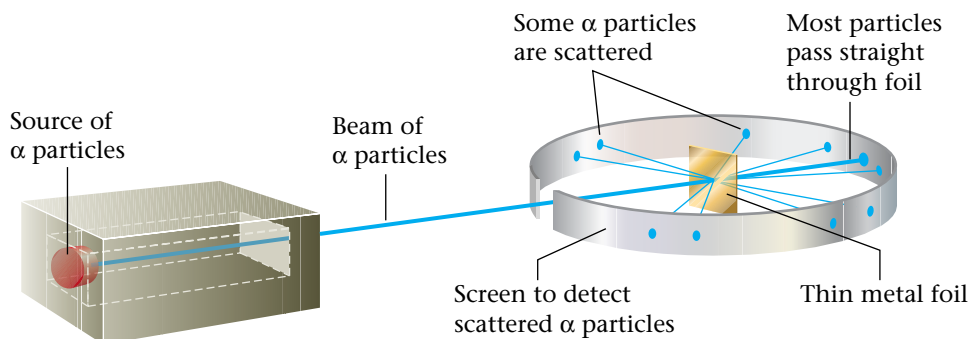


Figure 3.3

Rutherford's experiment on α particle bombardment of metal foil

From these results, Rutherford concluded that the plum pudding model for the atom could not be correct. The large deflections of the α particles could be caused only by a center of concentrated positive charge that would repel the positively charged α particles, as illustrated in **Figure 3.4b**. Most of the α particles passed directly through the foil because the atom is mostly open space. The deflected α particles were those that had a "close encounter" with the positive center of the atom, and the few reflected α particles were those that scored a "direct hit" on the positive center.

The Nuclear Atom In Rutherford's mind, these results could be explained only in terms of a **nuclear atom**—an atom with a dense center of positive charge (the **nucleus**) around which tiny electrons moved in a space that was otherwise empty.

Nuclear atom

A modern concept of the atom as having a dense center of positive charge (the nucleus) surrounded by moving electrons

Nucleus

The relatively small, dense center of positive charge in an atom

Proton

A positively charged subatomic particle located in the atomic nucleus

Active Reading Question

How does Rutherford's experiment support the existence of a positively charged, dense nucleus?

Rutherford concluded that the nucleus must have a positive charge to balance the negative charge of the electrons and that the nucleus must be small and dense. What was it made of? By 1919 Rutherford had concluded that the nucleus of an atom contained what he called protons. A **proton** has the same magnitude (size) of charge as the electron, but its charge is *positive*. We say that the proton has a charge of $1+$ and that the electron has a charge of $1-$.

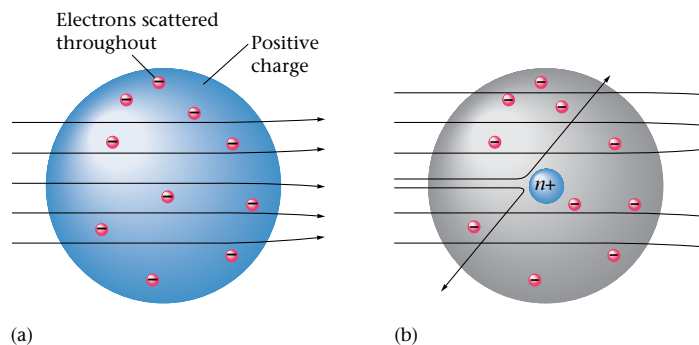


Figure 3.4

(a) The results that the metal foil experiment would have yielded if the plum pudding model had been correct (b) Actual results

Rutherford reasoned that the hydrogen atom has a single proton at its center and one electron moving through space at a relatively large distance from the proton (the hydrogen nucleus). He also reasoned that other atoms must have nuclei (the plural of *nucleus*) composed of many protons bound together in some way. In addition, in 1932 Rutherford and a coworker, James Chadwick, were able to show that most nuclei also contain a neutral particle that they named the **neutron**. A neutron is slightly more massive than a proton but has no charge.

B. Introduction to the Modern Concept of Atomic Structure

Since the time of Thomson and Rutherford, a great deal has been learned about atomic structure. The simplest view of the atom is that it consists of a tiny nucleus (about 10^{-13} cm in diameter) and electrons that move about the nucleus at an average distance of about 10^{-8} cm from it (see **Figure 3.5**). To visualize how small the nucleus is compared with the size of the atom, consider that if the nucleus were the size of a grape, the electrons would be about one *mile* away on average.

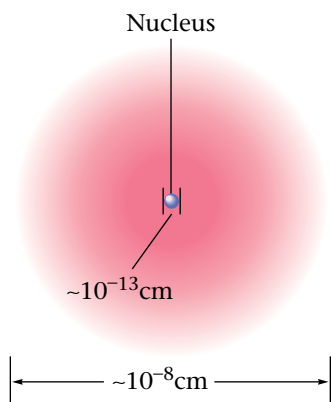


Figure 3.5

A nuclear atom viewed in cross section. (The symbol \sim means approximately.) This drawing does not show the actual scale. The nucleus is actually much smaller than the atom itself.

The nucleus contains protons, which have a positive charge equal in magnitude to the electrons' negative charge, and neutrons, which have almost the same mass as protons but no charge. The neutrons' function in the nucleus is not obvious. They may help hold the protons (which repel each other) together to form the nucleus, but we will not be concerned with that here. The relative masses and charges of the electron, proton, and neutron are shown in **Table 3.4**.

An important question arises at this point: *If all atoms are composed of these same components, why do different atoms have different chemical properties?* The answer lies in the number and arrangement of the electrons. The space in which the electrons move accounts for most of the atomic volume. The electrons are the parts of atoms that “intermingle” when atoms combine to form molecules.

Neutron

A subatomic particle with no charge located in the atomic nucleus

CRITICAL THINKING ?

The average diameter of an atom is 1.3×10^{-10} m.

What if the average diameter of an atom were 1 cm? How tall would you be?

Table 3.4

The Mass and Charge of the Electron, Proton, and Neutron

Particle	Relative Mass*	Relative Charge
electron	1	1-
proton	1836	1+
neutron	1839	none

*The electron is arbitrarily assigned a mass of 1 for comparison.

information

The *chemistry* of an atom arises from its electrons.

How Big Is an Atom?

Materials

- strip of paper 11" by 1"
- scissors

Procedure

1. Get a strip of paper 11" by 1".
2. Cut the paper in half. Discard one piece.
3. Repeat step 2 keeping track of the number of cuts until you can no longer cut the paper.

Results/Analysis

1. How many times could you cut the paper?
2. How many times would you need to cut the paper to have a piece of paper remaining that is the same width as an atom?
(average atom diameter = 1.3×10^{-10} m)

Therefore, the number of electrons a given atom possesses greatly affects the way it can interact with other atoms. As a result, atoms of different elements, which have different numbers of electrons, show different chemical behavior.

Although the atoms of different elements also differ in their numbers of protons, it is the number of electrons that really determines chemical behavior. We will discuss how this happens in later chapters.

Active Reading Question

Which of the subatomic particles are most responsible for the chemical properties of an element?

Information

All atoms of the same element have the same number of protons (the element's atomic number) and the same number of electrons.

In a free atom, the positive and negative charges always balance to yield a net zero charge.

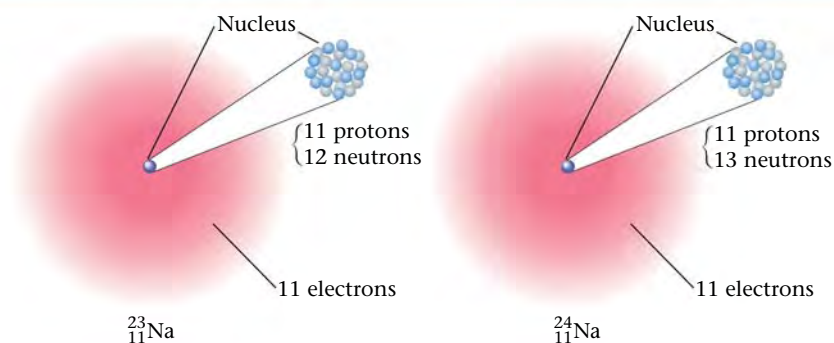
C. Isotopes

We have seen that an atom has a nucleus with a positive charge due to its protons and has electrons in the space surrounding the nucleus at relatively large distances from it.

As an example, consider a sodium atom, which has 11 protons in its nucleus. Because an atom has no overall charge, the number of electrons must equal the number of protons. Therefore, a sodium atom has 11 electrons in the space around its nucleus. It is *always* true that a sodium atom has 11 protons and 11 electrons. However, each sodium atom also has neutrons in its nucleus, and different types of sodium atoms exist that have different numbers of neutrons.

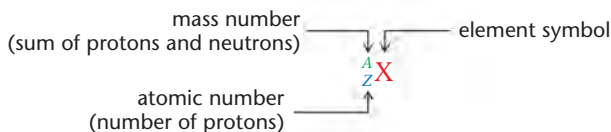
When Dalton stated his atomic theory in the early 1800s, he assumed that all of the atoms of a given element were identical. This idea persisted for over 100 years, until James Chadwick discovered that the nuclei of most atoms contain neutrons as well as protons. (This development is a good example of how a theory changes as new observations are made.) After the discovery of the neutron, Dalton's statement that all atoms of a given element are identical had to be changed to:

"All atoms of the same element contain the same number of protons and electrons, but atoms of a given element may have different numbers of neutrons."

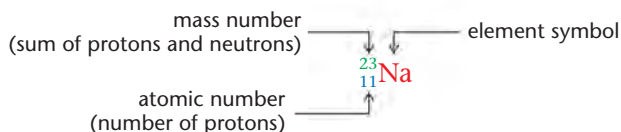
**Figure 3.6**

Two isotopes of sodium: Both have 11 protons and 11 electrons, but they differ in the number of neutrons in their nuclei.

To illustrate this idea, consider the sodium atoms depicted in **Figure 3.6**. These atoms are **isotopes**, or *atoms with the same number of protons but different numbers of neutrons*. The number of protons in a nucleus is called the atom's **atomic number**. The *sum* of the number of neutrons and the number of protons in a given nucleus is called the atom's **mass number**. To specify which of the isotopes of an element we are talking about, we use the following notation:



For example, the notation for one particular type of sodium atom is written



The particular atom represented above is called sodium-23 because it has a mass number of 23. Let's specify the number of each type of subatomic particle in sodium-23.

- Protons = atomic number = 11
- Electrons = protons = 11
- Neutrons: We need to determine the number of neutrons as follows:

$$\text{Mass number} = \text{number of protons} + \text{number of neutrons}$$

$$A = Z + \text{number of neutrons}$$

$$A - Z = \text{number of neutrons}$$

This is a general result. You can always determine the number of neutrons present in a given atom by subtracting the atomic number from the mass number. In this case ($^{23}_{11}\text{Na}$), we know that $A = 23$ and $Z = 11$. Therefore,

$$A - Z = 23 - 11 = 12 = \text{number of neutrons}$$

In summary, sodium-23 has 11 electrons, 11 protons, and 12 neutrons.

Active Reading Questions

1. How did the discovery of the neutron lead to the concept of isotopes?
2. How does the fact that isotopes exist change Dalton's model?

Isotopes

Atoms with the same number of protons but different numbers of neutrons

Atomic number

The number of protons in the nucleus of a given atom

Mass number

The total number of protons and neutrons in the nucleus of a given atom

MATH

We can solve for the number of neutrons by using algebra, subtracting Z from both sides of the equation.

$$A - Z = Z - Z + \text{number of neutrons}$$

EXAMPLE 3.2

Interpreting Symbols for Isotopes

Information

Isotopes are “top heavy”; that is, the mass number is on the top and the atomic number is on the bottom. For most isotopes, the value of A is greater than the value of Z .

In nature, elements are usually found as a mixture of isotopes. The three isotopes of carbon are:



Determine the number of each of the three types of subatomic particles in each of these carbon atoms.

Solution

The number of protons and electrons is the same in each of the isotopes and is given by the atomic number of carbon, 6. The number of neutrons can be determined by subtracting the atomic number (Z) from the mass number:

$$A - Z = \text{number of neutrons}$$

The numbers of neutrons in the three isotopes of carbon are

$${}^{12}_6\text{C}: \text{number of neutrons} = A - Z = 12 - 6 = 6$$

$${}^{13}_6\text{C}: \text{number of neutrons} = 13 - 6 = 7$$

$${}^{14}_6\text{C}: \text{number of neutrons} = 14 - 6 = 8$$

In summary,

Symbol	Number of Protons	Number of Electrons	Number of Neutrons
${}^{12}_6\text{C}$	6	6	6
${}^{13}_6\text{C}$	6	6	7
${}^{14}_6\text{C}$	6	6	8

Practice Problems • Exercise 3.2

1. Strontium-90 occurs in fallout from nuclear testing. It can accumulate in bone marrow and may cause leukemia and bone cancer. Give the number of protons, neutrons, and electrons in the atom symbolized by ${}^{90}_{38}\text{Sr}$.
2. Give the number of protons, neutrons, and electrons in the atom symbolized by ${}^{201}_{80}\text{Hg}$.

EXAMPLE 3.3

Writing Symbols for Isotopes



Magnesium burns in air to give a bright white flame.

Write the symbol for the magnesium atom (atomic number 12) with a mass number of 24. How many electrons and how many neutrons does this atom have?

Solution

The atomic number 12 means that the atom has 12 protons. The element magnesium is symbolized by Mg. The atom is represented as ${}^{24}_{12}\text{Mg}$ and is called magnesium-24. Because the atom has 12 protons, it must also have 12 electrons. The mass number gives the total number of protons and neutrons, which means that this atom has 12 neutrons ($24 - 12 = 12$).

EXAMPLE 3.4**Calculating Mass Number**

Write the symbol for the silver atom ($Z = 47$) that has 61 neutrons.

Solution

The element symbol is ${}^A_Z\text{Ag}$, where we know that $Z = 47$. We can find A from its definition, $A = Z + \text{number of neutrons}$. In this case,

$$A = 47 + 61 = 108$$

The complete symbol for this atom is ${}^{108}_{47}\text{Ag}$.

Practice Problem • Exercise 3.4

Give the symbol for the phosphorus atom ($Z = 15$) that contains 17 neutrons.

SECTION 3.3

REVIEW QUESTIONS

1 How is Thomson's model of the atom different from Dalton's model of the atom? Draw a picture of each model.

2 What caused Rutherford to propose a revised model of the atom? How is the Rutherford model different from the previous models?

3 Copy and complete the following table.

Symbol	Number of protons	Number of neutrons	Number of electrons
${}^{32}_{15}\text{P}$			
	29	36	
		21	18

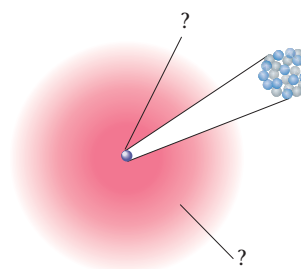
4 Write the symbol for each of the following. Determine the number of protons, neutrons, and electrons in each atom.

- A sodium atom with a mass of 24
- A carbon atom with a mass of 13
- A chlorine atom with a mass of 37

5 Write the symbol for each of the following.

- The magnesium atom that has 12 neutrons
- The bromine atom that has 45 neutrons
- The hydrogen atom that has 2 neutrons

6 Label the parts of the atom represented below and the charge of each part.



In addition, copy and complete the following table.

Atomic number of the atom	
Name of the element	
Symbol of the element	
Mass number of the atom	
Number of electrons in the atom	

7 In two isotopes of the same element, which of the following would be the same and which would be different?

- atomic number
- number of protons
- number of neutrons
- number of electrons
- element symbol

Most of the elements are metals. Metals have the following characteristic physical properties:

Physical Properties of Metals

1. Efficient conduction of heat and electricity
2. Malleability (can be hammered into thin sheets)
3. Ductility (can be pulled into wires)
4. A lustrous (shiny) appearance

Copper is a typical metal. It is lustrous (although it tarnishes readily); it is an excellent conductor of electricity (it is widely used in electrical wires); and it is readily formed into various shapes, such as pipes for water systems. Copper is one of the transition metals—the metals shown in the center of the periodic table. Iron, aluminum, and gold are other familiar elements that have metallic properties. All of the elements shown below and to the left of the heavy “stair-step” black line in Figure 3.7 are classified as metals, except for hydrogen (see **Figure 3.9**).

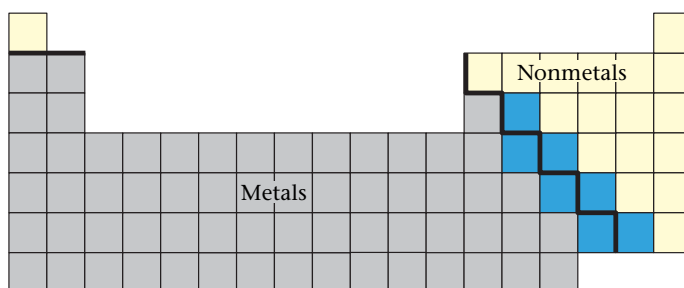


Figure 3.9
The elements classified as metals and as nonmetals



Indonesian men carry chunks of elemental sulfur in baskets.

The relatively small number of elements that appear in the upper righthand corner of the periodic table (to the right of the heavy line in Figures 3.7 and 3.9) are called **nonmetals**. Nonmetals generally lack those properties that characterize metals and show much more variation in their properties than metals do. Whereas almost all metals are solids at normal temperatures, many nonmetals (such as nitrogen, oxygen, chlorine, and neon) are gaseous, and one (bromine) is a liquid. Several nonmetals (such as carbon, phosphorus, and sulfur) are also solids.

The elements that lie close to the “stair-step” line in Figure 3.7 often show a mixture of metallic and nonmetallic properties. These elements, which are called **metalloids** or **semimetals**, include silicon, germanium, arsenic, antimony, and tellurium.

As we continue our study of chemistry, we will see that the periodic table is a valuable tool for organizing knowledge and that it helps us predict the properties we expect a given element to exhibit. We will also develop a model for atomic structure that will explain why various groups of elements have similar chemical properties.

Active Reading Question

How does the location of an element on the periodic table help you determine its properties?

EXAMPLE 3.5**Interpreting the Periodic Table**

For each of the following elements, use the periodic table in the back of the book to

- give the symbol and atomic number.
- specify whether the element is a metal or a nonmetal.
- give the named family to which the element belongs (if any).
 - a. iodine
 - b. magnesium
 - c. gold
 - d. lithium

Solution

- a. Iodine (symbol I) is element 53 (its atomic number is 53). Iodine lies to the right of the stair-step line in Figure 3.9 and so is a nonmetal. Iodine is a member of Group 7, the family of halogens.
- b. Magnesium (symbol Mg) is element 12 (atomic number 12). Magnesium is a metal and is a member of the alkaline earth metal family (Group 2).
- c. Gold (symbol Au) is element 79 (atomic number 79). Gold is a metal and is not a member of a named vertical family. It is classed as a transition metal.
- d. Lithium (symbol Li) is element 3 (atomic number 3). Lithium is a metal in the alkali metal family (Group 1).

Practice Problem • Exercise 3.5

Give the symbol and atomic number for each of the following elements. Also indicate whether each element is a metal or a nonmetal and whether it is a member of a named family.

- a. argon
- b. chlorine
- c. barium
- d. cesium

B. Natural States of the Elements

As we have noted, the matter around us consists mainly of mixtures. Most often these mixtures contain compounds, in which atoms from different elements are bound together. Most elements are quite reactive: their atoms tend to combine with those of other elements to form compounds quite readily. As a result, we do not often find elements in nature in pure form—uncombined with other elements. However, there are notable exceptions. The gold nuggets found at Sutter's Mill in California that launched the Gold Rush in 1849 are virtually pure elemental gold. Platinum and silver are also often found in nearly pure form.

**DID YOU KNOW**

A gold nugget weighing 13 lb, 7 oz, which came to be called Tom's Baby, was found by Tom Grove near Breckenridge, Colorado, on July 23, 1887.

Gold, silver, and platinum are members of a class of metals called *noble metals* because they are relatively unreactive. (The term *noble* implies a class set apart.) Other elements that appear in nature in the uncombined state are the elements in Group 8: helium, neon, argon, krypton, xenon, and radon.

Group 8

He
Ne
Ar
Kr
Xe
Rn

Because the atoms of these Group 8 elements do not combine readily with those of other elements, we call them the noble gases.

For example, helium gas is found in uncombined form in underground deposits with natural gas.

When we take a sample of air (the mixture of gases that constitute earth's atmosphere) and separate it into its components, we find several pure elements present. One of these is argon. Argon gas consists of a collection of separate argon atoms, as shown in **Figure 3.10**.

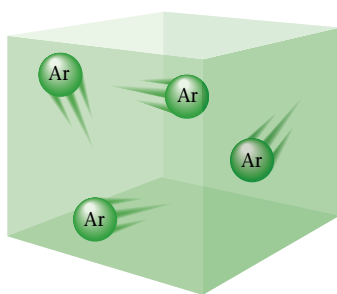


Figure 3.10

Argon gas consists of a collection of separate argon atoms.

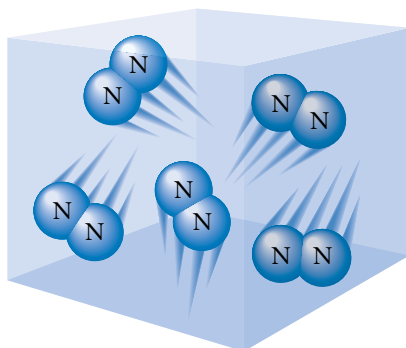
Diatomic molecule

A molecule composed of two atoms

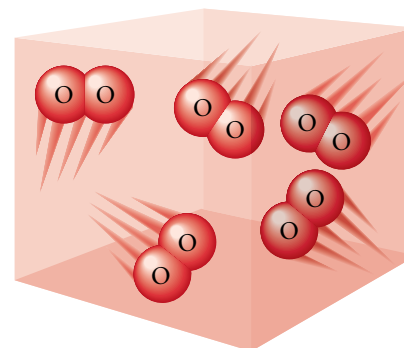
Diatomic Molecules Air also contains nitrogen gas and oxygen gas. When we examine these two gases, however, we find that they do not contain single atoms, as argon does, but instead contain **diatomic molecules**: molecules made up of *two atoms*, as represented in **Figure 3.11**. In fact, any sample of elemental oxygen gas at normal temperatures contains O_2 molecules. Likewise, nitrogen gas contains N_2 molecules.

Information

A molecule is a collection of atoms that behaves as a unit. Molecules are always electrically neutral (zero charge).



Nitrogen gas contains N_2 molecules.

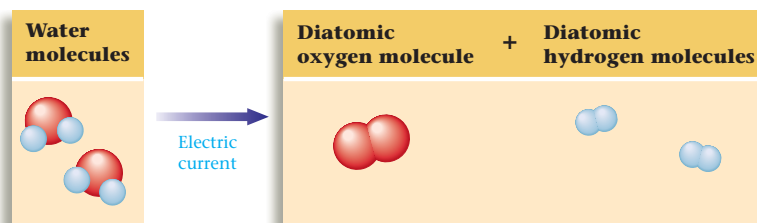


Oxygen gas contains O_2 molecules.

Figure 3.11

Gaseous nitrogen and oxygen contain diatomic (two-atom) molecules.

Hydrogen is another element that forms diatomic molecules. Although virtually all of the hydrogen found on earth is present in compounds with other elements (such as with oxygen in water), when hydrogen is prepared as a free element, it contains diatomic H_2 molecules. For example, an electric current can be used to decompose water into elemental hydrogen and oxygen containing H_2 and O_2 molecules, respectively.



Note that only the grouping of the atoms changes in this process; no atoms are created or destroyed. The same number of H atoms and O atoms must be present both before and after the process. For example, the decomposition of two H_2O molecules (containing four H atoms and two O atoms) yields one O_2 molecule (containing two O atoms) and two H_2 molecules (containing a total of four H atoms).

Several other elements, in addition to hydrogen, nitrogen, and oxygen, exist as diatomic molecules. For example, when sodium chloride is melted and subjected to an electric current, chlorine gas is produced (along with sodium metal). This chemical change is represented in **Figure 3.12**.

Chlorine gas is a pale green gas that contains Cl_2 molecules.

Chlorine is a member of Group 7, the halogen family. All of the elemental forms of the Group 7 elements contain diatomic molecules. Fluorine is a pale yellow gas containing F_2 molecules. Bromine is a brown liquid made up of Br_2 molecules. Iodine is a lustrous, purple solid that contains I_2 molecules.

Table 3.5 lists the elements that contain diatomic molecules in their pure, elemental forms.

So far we have seen that several elements are gaseous in their elemental forms at normal temperatures ($\sim 25^\circ C$). The noble gases (the Group 8 elements) contain individual atoms, whereas several other gaseous elements contain diatomic molecules (H_2 , N_2 , O_2 , F_2 , and Cl_2).

Only two elements are liquids in their elemental forms at $25^\circ C$: the nonmetal bromine (containing Br_2 molecules) and the metal mercury. The metals gallium and cesium almost qualify in this category; they are solids at $25^\circ C$, but both melt at $\sim 30^\circ C$.

Group 7

F
Cl
Br
I

DID YOU KNOW

The only elemental hydrogen found naturally on earth occurs in the exhaust gases of volcanoes.



Gallium metal has such a low melting point ($30^\circ C$) that it melts from the heat of a hand.

MATH



\sim means "approximately."



(a)



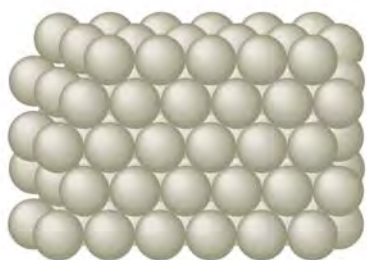
(b)

Figure 3.12

(a) Sodium chloride (common table salt) can be decomposed to the elements sodium metal and chlorine gas, shown in (b).

Table 3.5**Elements That Exist as Diatomic Molecules in Their Elemental Forms**

Element Present	Elemental State at 25 °C	Molecule
hydrogen	colorless gas	H ₂
nitrogen	colorless gas	N ₂
oxygen	pale blue gas	O ₂
fluorine	pale yellow gas	F ₂
chlorine	pale green gas	Cl ₂
bromine	reddish-brown liquid	Br ₂
iodine	lustrous, dark-purple solid	I ₂

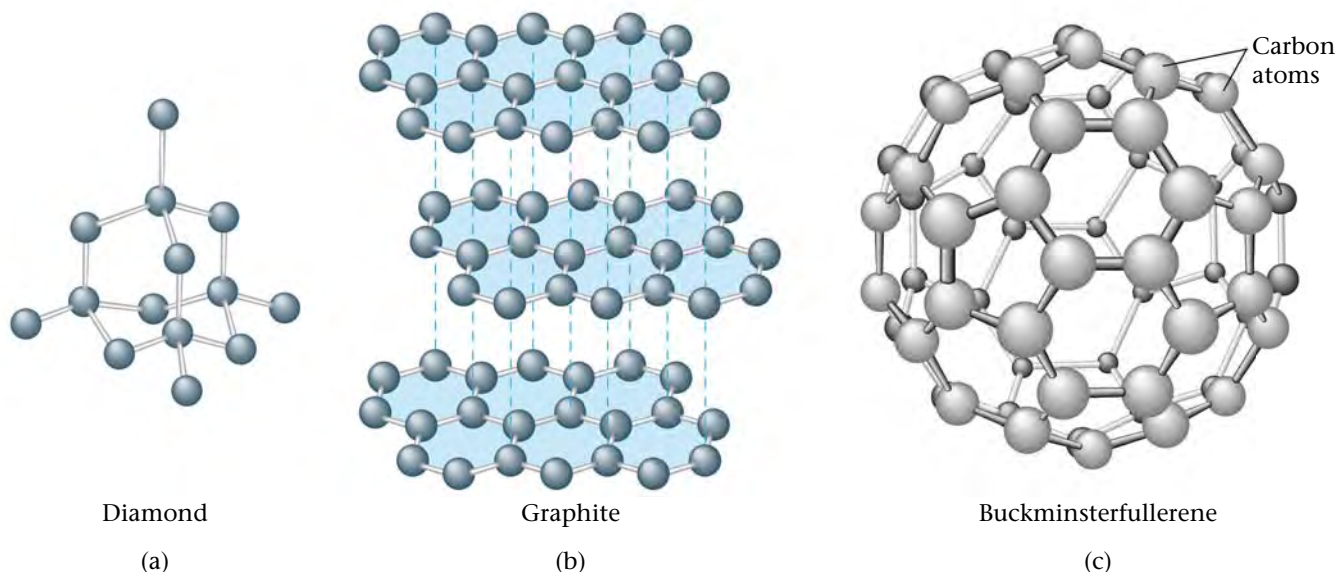
**Figure 3.13**

In solid metals, the spherical atoms are packed closely together.

Elemental Solids The other elements are solids in their elemental forms at 25 °C. The solid metals contain large numbers of atoms packed together much like marbles in a jar (see **Figure 3.13**).

The structures of solid nonmetallic elements are more varied than those of metals. In fact, different forms of the same element often occur. For example, solid carbon occurs in three forms. Different forms of a given element are called *allotropes*. The three allotropes of carbon are the familiar diamond and graphite forms and a form that has only recently been discovered—buckminsterfullerene. These elemental forms have very different properties because of their different structures (see **Figure 3.14**).

Diamond is the hardest natural substance known and is often used for industrial cutting tools. Diamonds are also valued as gemstones. Graphite, on the other hand, is a rather soft material useful for writing (pencil “lead” is really graphite) and (in the form of a powder) for lubricating locks.

**Figure 3.14**

The three solid elemental (allotropes) forms of carbon: (a) diamond, (b) graphite, and (c) buckminsterfullerene. The representations of diamond and graphite are just fragments of much larger structures that extend in all directions from the parts shown here. Buckminsterfullerene contains C₆₀ molecules, one of which is shown.

The rather odd name given to buckminsterfullerene comes from the structure of the C_{60} molecules of which it is composed. The soccer-ball-like structure contains five- and six-member rings reminiscent of the structure of geodesic domes suggested by the late industrial designer Buckminster Fuller. Other “fullerenes” containing molecules with more than 60 carbon atoms have also recently been discovered, leading to a new area of chemistry.



Active Reading Question

What does the term *allotrope* mean? Provide an example.



HANDS-ON CHEMISTRY

• MINI-LAB •

Elementary, My Dear Watson!

Materials

- index cards

Procedure

1. Make flashcards with the names of the elements in Groups 1–8 on one side and their symbols on the other side.
2. Group the elements according to
 - a. their families on the periodic table.
 - b. the similarity between their names and their symbols.

- c. the first letter of each symbol.
- d. the state of the element at 25 °C.
- e. whether the element is a metal, a nonmetal, or a metalloid.

Results/Analysis

1. Use your flashcards to learn the names and symbols for these elements.
2. Work in pairs and quiz each other with the flashcards.



SECTION 3.4

REVIEW QUESTIONS

1. To what does the term *periodic* in *periodic table* refer?
2. Why are the elements in vertical groups on the periodic table called “families”?
3. What does the “stair-step” line on the periodic table tell you?
4. What does the term *noble* mean when referring to *noble metals* and *noble gases*?
5. Which elements exist naturally as diatomic molecules? Write their formulas.
6. Which elements are gases at 25 °C? Write their formulas.
7. Copy the table shown. Then use the periodic table to complete the table.

Name	Symbol	Atomic number	Metal or nonmetal?	Family to which the element belongs
Sodium				
	F			
		4		
Calcium				
		47		
8. Which two elements are liquids at 25 °C? Write their formulas.



RESEARCH LINKS
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SECTION 3.5

Ions and Their Compounds

Key Terms

- Ion
- Cation
- Anion
- Ionic compounds

Objectives

- To describe the formation of ions from their parent atoms
- To learn to name ions
- To predict which ion a given element forms by using the periodic table
- To describe how ions combine to form neutral compounds

A. Ions

We have seen that an atom has a certain number of protons in its nucleus and an equal number of electrons in the space around the nucleus. This results in an exact balance of positive and negative charges. We say that an atom is a neutral entity—it has *zero net charge*.

We can produce a charged entity, called an **ion**, by taking a neutral atom and adding or removing one or more electrons. For example, a sodium atom ($Z = 11$) has eleven protons in its nucleus and eleven electrons outside its nucleus.

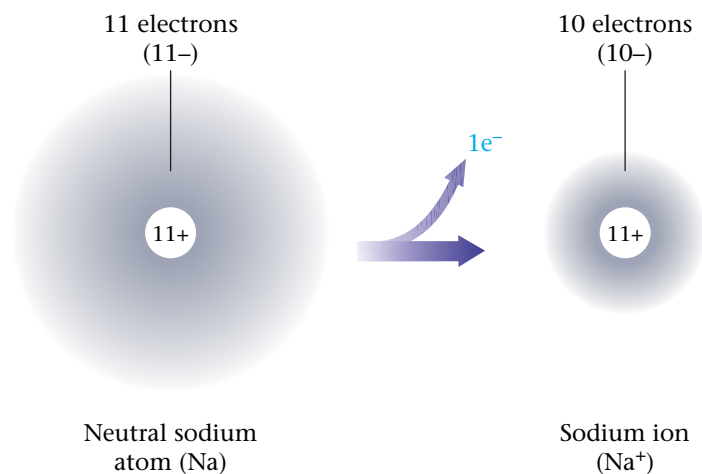
If one of the electrons is lost, there will be eleven positive charges but only ten negative charges. This gives an ion with a net positive one ($1+$) charge: $(11+) + (10-) = 1+$. We can represent this process as follows: In shorthand form,



where Na represents the neutral sodium atom, Na^+ represents the $1+$ ion formed, and e^- represents an electron.

Ion

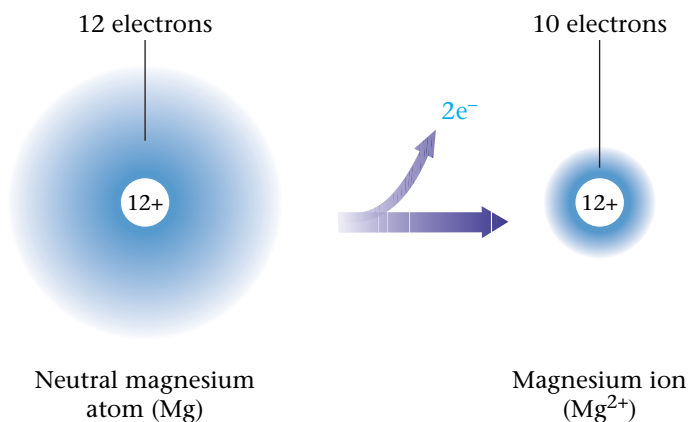
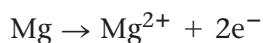
An atom or group of atoms that has a positive or negative charge



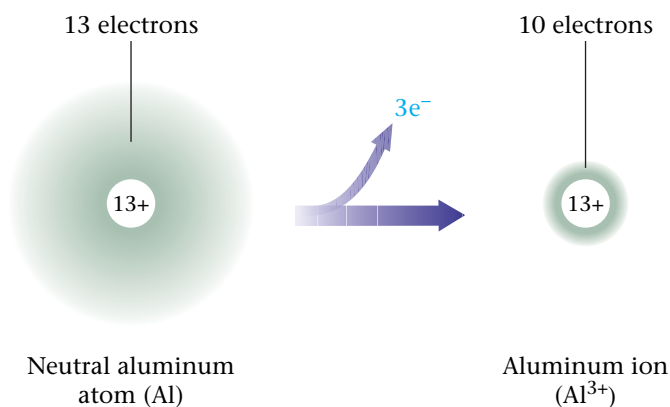
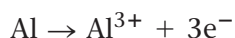
Cations and Anions

Losing electrons A positive ion, called a **cation** (pronounced *cat' eye on*), is produced when one or more electrons are *lost* from a neutral atom. We have seen that sodium loses one electron to become a 1+ cation. Some atoms lose more than one electron. For example, a magnesium atom typically loses two electrons to form a 2+ cation:

We usually represent this process as follows:



Aluminum forms a 3+ cation by losing three electrons:
That is,



A cation is named using the name of the parent atom. Thus Na^{+} is called the sodium ion (or sodium cation), Mg^{2+} is called the magnesium ion (or magnesium cation), and Al^{3+} is called the aluminum ion (or aluminum cation).

Cation

A positively charged ion

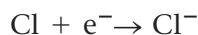
Information

Notice that when electrons are removed from an atom it becomes much smaller.

Anion

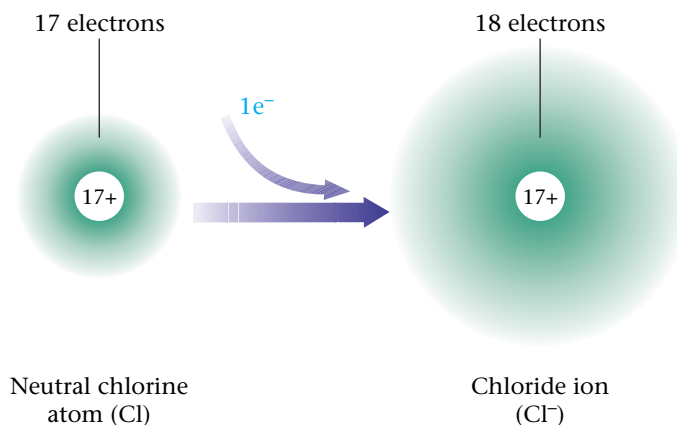
A negatively charged ion

Gaining electrons When a neutral atom *gains* electrons, an ion with a negative charge is formed. A negatively charged ion is called an **anion** (pronounced *an' ion*). An atom that gains one extra electron forms an anion with a 1⁻ charge. An example of an atom that forms a 1⁻ anion is the chlorine atom, which has seventeen protons and seventeen electrons. That is,

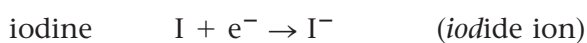
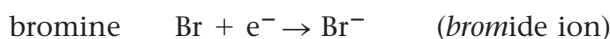
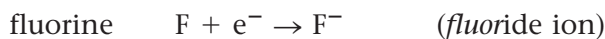


Information

When electrons are added to an atom it becomes much larger.



Note that the anion formed by chlorine has eighteen electrons but only seventeen protons, so the net charge is $(18^{-}) + (17^{+}) = 1^{-}$. Unlike a cation, which is named for the parent atom, an anion is named by taking the root name of the atom and changing the ending. For example, the Cl⁻ anion produced from the Cl (chlorine) atom is called the chloride ion (or chloride anion). Notice that the word *chloride* is obtained from the root of the atom name (*chlor-*) plus the suffix *-ide*. Other atoms that add one electron to form 1⁻ ions include the following:

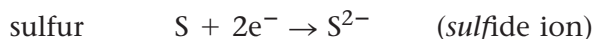
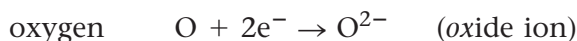


Information

The name of an anion is obtained by adding *-ide* to the root of the atom name.

Note that the name of each of these anions is obtained by adding *-ide* to the root of the atom name.

Some atoms can add two electrons to form 2⁻ anions. Examples include oxygen and sulfur:



Note that the names for these anions are derived in the same way as are those for the 1⁻ anions.

Forming ions It is important to recognize that ions are always formed by removing electrons from an atom (to form cations) or adding electrons to an atom (to form anions).

Ions are never formed by changing the number of protons in an atom's nucleus.

Miracle Coatings

Imagine a pair of plastic-lens sunglasses that are unscratchable, even if you drop them on concrete or rub them with sandpaper. Research may make such glasses possible, along with cutting tools that never need sharpening, special glass for windshields and buildings that cannot be scratched by wind-blown sand, and speakers that reproduce sound with a crispness unimagined until now. The secret of such marvels is a thin diamond coating. Diamond is so hard that virtually nothing can scratch it. A thin diamond coating on a speaker cone limits resonance and gives a remarkably pure tone.

But how do you coat something with a diamond? It is nearly impossible to melt diamond (melting point, $3500\text{ }^{\circ}\text{C}$). And even if diamond were melted, the object being coated would itself melt immediately at this temperature. Surprisingly, a diamond coating

can be applied quite easily to something even as fragile as plastic. First, the surface is bathed with a mixture of gaseous methane (CH_4) and hydrogen (H_2). Next, the methane is broken apart into its component elements by an energy source similar to that used in microwave ovens. The carbon atoms freed from the methane then form a thin diamond coating on the surface being treated.

The coating of soft, scratchable materials with a super-tough diamond layer should improve many types of consumer products in the near future.



B. Compounds That Contain Ions

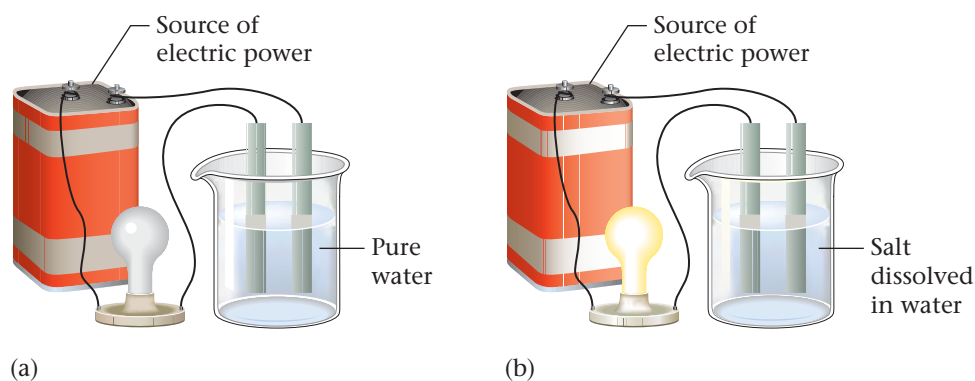
Information

Melting means that the solid, in which ions are locked into place, is changed to a liquid, in which the ions can move.

Chemists have good reasons to believe that many chemical compounds contain ions. For instance, consider some of the properties of common table salt, sodium chloride (NaCl). It must be heated to about $800\text{ }^{\circ}\text{C}$ to melt and to almost $1500\text{ }^{\circ}\text{C}$ to boil (compare to water, which boils at $100\text{ }^{\circ}\text{C}$). As a solid, salt will not conduct an electric current, but when melted it is a very good conductor. Pure water will not conduct electricity (will not allow an electric current to flow), but when salt is dissolved in water, the resulting solution readily conducts electricity (see **Figure 3.16**).

Figure 3.16

(a) Pure water does not conduct a current, so the circuit is not complete and the bulb does not light.
(b) Water containing a dissolved salt conducts electricity, and the bulb lights.



Chemists have come to realize that we can best explain these properties of sodium chloride (NaCl) by picturing it as containing Na^+ ions and Cl^- ions packed together as shown in **Figure 3.17**. Because the positive and negative charges attract each other very strongly, salt must be heated to a very high temperature ($800\text{ }^\circ\text{C}$) before it melts.

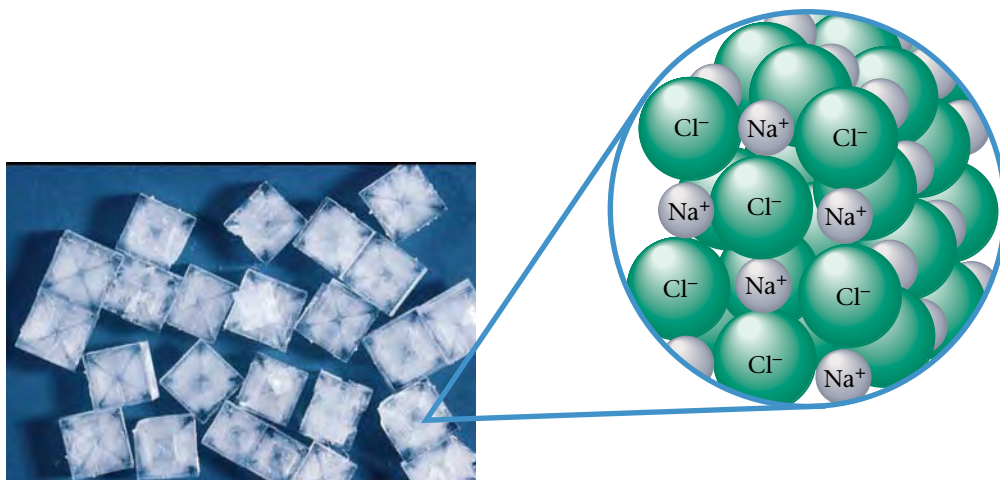


Figure 3.17

The atomic-level view shows the arrangement of sodium ions (Na^+) and chloride ions (Cl^-) in the ionic compound sodium chloride. The photo shows solid sodium chloride highly magnified.

To explore further the significance of the electrical conductivity results, we need to discuss briefly the nature of electric currents. An electric current can travel along a metal wire because *electrons are free to move* through the wire; the moving electrons carry the current. In ionic substances, the ions carry the current. Therefore, substances that contain ions can conduct an electric current *only if the ions can move*—the current travels by the movement of the charged ions. In solid NaCl, the ions are tightly held and cannot move. When the solid melts and becomes a liquid, however, the structure is disrupted and the ions can move. As a result, an electric current can travel through the melted salt.

The same reasoning applies to NaCl dissolved in water. When the solid dissolves, the ions can move around and are dispersed throughout the water, allowing it to conduct a current.

Thus we recognize substances that contain ions by their characteristic properties. They often have very high melting points, and they conduct an electric current when melted or dissolved in water.

Many substances contain ions. In fact, whenever a compound forms between a metal and a nonmetal, it can be expected to contain ions. We call these substances **ionic compounds**.

Formulas for Ionic Compounds One fact very important to remember is that *a chemical compound must have a net charge of zero*. This fact means that if a compound contains ions, then

1. Both positive ions (cations) and negative ions (anions) must be present.
2. The numbers of cations and anions must be such that the net charge is zero.

DID YOU KNOW

A substance containing ions that can move can conduct an electric current.

DID YOU KNOW

Dissolving NaCl causes the ions to be randomly dispersed in the water, allowing them to move freely. Dissolving is not the same as melting, but both processes free the ions to move.

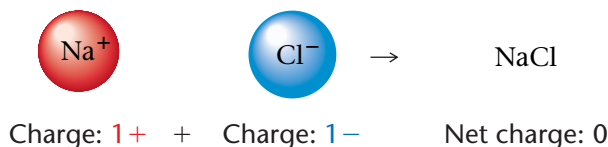
Ionic compound

A compound that results when a metal reacts with a nonmetal to form cations and anions

Information

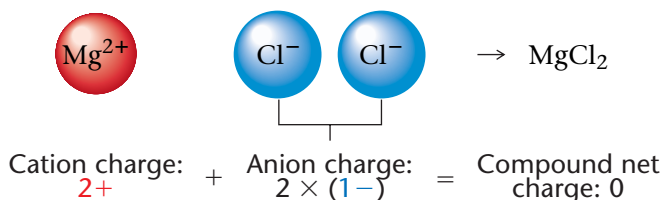
An ionic compound cannot contain only anions or only cations because the net charge of a compound must be zero.

For example, note that the formula for sodium chloride is written NaCl, indicating one of each type of these elements. This makes sense because sodium chloride contains Na^+ ions and Cl^- ions. Each sodium ion has a 1+ charge and each chloride ion has a 1- charge, so they must occur in equal numbers to give a net charge of zero.



For *any* ionic compound,
Total charge of cations + Total charge of anions = Zero net charge

Consider an ionic compound that contains the ions Mg^{2+} and Cl^- . What combination of these ions will give a net charge of zero? To balance the 2+ charge on Mg^{2+} , we will need two Cl^- ions to give a net charge of zero.



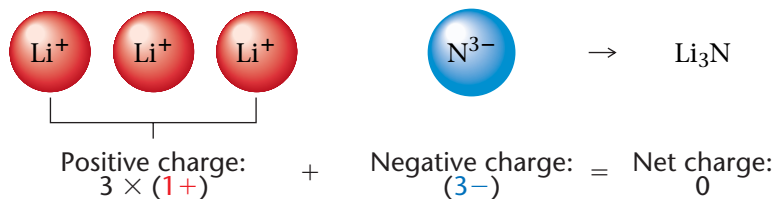
Reading Tip

The subscript 1 in a formula is not written.

This means that the formula of the compound must be MgCl_2 . Remember that subscripts are used to give the relative numbers of atoms (or ions).

Now consider an ionic compound that contains the ions Ba^{2+} and O^{2-} . What is the correct formula? These ions have charges of the same size, but opposite sign, so they must occur in equal numbers to give a net charge of zero. The formula of the compound is BaO because $(2+) + (2-) = 0$.

Similarly, the formula of a compound that contains the ions Li^+ and N^{3-} is Li_3N because three Li^+ cations are needed to balance the charge of the N^{3-} anion.



Critical Thinking

Thomson and Rutherford helped show that atoms consist of subatomic particles, two of which are charged.

What if subatomic particles had no charge? How would that affect compounds formed between metals and nonmetals?

Active Reading Question

How does knowing the ions in an ionic compound allow us to determine the chemical formula of the compound?

EXAMPLE 3.6**Writing Formulas for Ionic Compounds**

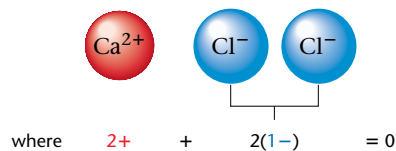
The pairs of ions contained in several ionic compounds are listed below. Give the formula for each compound.

- a. Ca^{2+} and Cl^- b. Na^+ and S^{2-} c. Ca^{2+} and P^{3-}

Solution

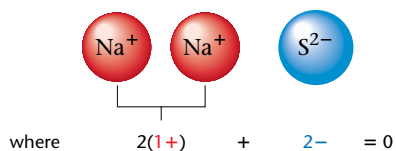
- a. Ca^{2+} has a 2+ charge, so two Cl^- ions (each with the charge 1-) will be needed.

The formula is CaCl_2 .



- b. In this case, S^{2-} , with its 2- charge, requires two Na^+ ions to produce a zero net charge.

The formula is Na_2S .



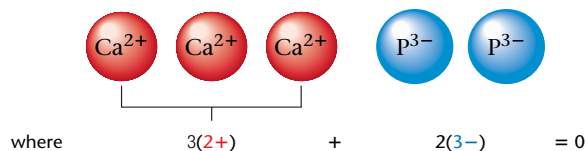
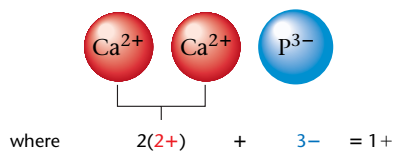
- c. We have the ions Ca^{2+} (charge 2+) and P^{3-} (charge 3-). We must figure out how many of each type is needed to balance exactly the positive and negative charges. Let's try two Ca^{2+} and one P^{3-} .

The resulting net charge is

$$2(2+) + (3-) = (4+) + (3-) = 1+$$

This doesn't work because the net charge is not zero. We can obtain the same total positive and total negative charges by having three Ca^{2+} ions and two P^{3-} ions.

The formula is Ca_3P_2 .

**Practice Problem • Exercise 3.6**

Give the formulas for the compounds that contain the following pairs of ions.

- a. K^+ and I^- b. Mg^{2+} and N^{3-} c. Al^{3+} and O^{2-}

SECTION 3.5

REVIEW QUESTIONS

- Why does an atom have a net zero charge?
- Show how fluorine forms an ion. Does it form a cation or an anion?
 - Show how lithium forms an ion. Does it form a cation or an anion?
- What are the relationships between the group number on the periodic table and the types of ions formed?
- What evidence do we have that convinces us that sodium chloride is an ionic compound?
- Give the formulas for the compounds that contain the following pairs of ions.
 - Al^{3+} and Cl^-
 - K^+ and O^{2-}
 - Ca^{2+} and S^{2-}
 - Sr^{2+} and F^-

Key Terms

3.1 Element symbols

3.2 Law of constant composition

Dalton's atomic theory

Atoms

Compound

Chemical formula

3.3 Electrons

Nuclear atom

Nucleus

Proton

Neutron

Isotopes

Atomic number

Mass number

Key Ideas

3.1 The Elements

- All of the materials in the universe can be chemically broken down into about 100 different elements.
- Nine elements account for about 98% of earth's crust, oceans, and atmosphere.
- In the human body, oxygen, carbon, hydrogen, and nitrogen are the most abundant elements.
- Each element has a name and a symbol.
 - The symbol usually consists of the first one or two letters of the element's name.
 - Sometimes the symbol is taken from the element's original Latin or Greek name.

3.2 Atoms and Compounds

- The Law of constant composition states that a given compound always contains the same proportion by mass of the elements of which it is composed.
- Dalton's atomic theory states:
 - All elements are composed of atoms.
 - All atoms of a given element are identical.
 - Atoms of different elements are different.
 - Compounds consist of the atoms of different elements.
 - Atoms are not created or destroyed in a chemical reaction.
- A compound is represented by a chemical formula in which the number and type of atoms present are shown by using the element symbols and subscripts.

3.3 Atomic Structure

- Experiments by J. J. Thomson and Ernest Rutherford showed that atoms have internal structure.
 - The nucleus, which is at the center of the atom, contains protons (positively charged) and neutrons (uncharged).
 - Electrons move around the nucleus.
 - ◆ Electrons have a small mass (1/1836 of the proton mass).
 - ◆ Electrons have a negative charge equal and opposite to that of the proton.
- Isotopes are atoms with the same number of protons but a different number of neutrons.
- A particular isotope is represented by the symbol A_ZX , in which Z represents the number of protons (atomic number) and A represents the total number of protons and neutrons (mass number) in the nucleus.

3.4 Periodic table

Groups

Alkali metals

Alkaline earth metals

Halogens

Noble gases

Transition metals

Metals

Nonmetals

Metalloids (semimetals)

Diatomic molecules

3.4 Using the Periodic Table

- The periodic table shows all of the known elements in order of increasing atomic number; the table is organized to group elements with similar properties in vertical columns.
- Most elements have metallic properties (the metals) and appear on the left side of the periodic table.
- Nonmetals appear on the right side of the periodic table.
- Metalloids are elements that have some metallic and some nonmetallic properties.

1 H	2 He																
3 Li	4 Be	5 B	6 C	7 N	8 O	9 F	10 Ne										
11 Na	12 Mg	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar										
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La*	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89 Ac†	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Uun	111 Ds	112 Rg	113 Uut	114 Uuq	115 Uup			

*Lanthanides

58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
----------	----------	----------	----------	----------	----------	----------	----------	----------	----------	----------	----------	----------	----------

†Actinides

90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr
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3.5 Ion

Cation

Anion

Ionic compounds

3.5 Ions and Their Compounds

- Atoms can form ions (species with a charge) by gaining or losing electrons.
 - Metals tend to lose one or more electrons to form positive ions called cations; these are generally named by using the name of the parent atom.
 - Nonmetals tend to gain one or more electrons to form negative ions called anions; these are named by using the root of the atom name followed by the suffix *-ide*.
- The ion that a particular atom will form can be predicted from the atom's position on the periodic table.
 - Elements in Group 1 and 2 form 1+ and 2+ ions, respectively.
 - Group 7 atoms form anions with 1- charges.
 - Group 6 atoms form anions with 2- charges.
- Ions combine to form compounds. Compounds are electrically neutral, so the sum of the charges on the anions and cations in the compound must equal zero.

All exercises with **blue numbers** have answers in the back of this book.

3.1 The Elements

A. Abundances of the Elements

- The ancient Greeks believed that all matter was composed of four fundamental substances: earth, air, fire, and water. How does this early conception of matter compare with our modern theories about matter?
- Although they were not able to transform base metals into gold, what contributions did the alchemists make to the development of chemistry?
- In addition to his important work on the properties of gases, what other valuable contributions did Robert Boyle make to the development of the study of chemistry?
- In 2006, 115 elements were known. How many of these elements occur naturally, and how many are synthesized artificially? What are the most common elements present on earth?
- Oxygen, the most abundant element on earth by mass, makes up a large percentage of the atmosphere. Where else is oxygen found? Is oxygen found more commonly as an element or in compounds?
- What are the most abundant elements found in living creatures? Are these elements also the most abundant elements found in the nonliving world?

B. Names and Symbols for the Elements

Note: Refer to the tables on the inside back cover when appropriate.

- Give the symbols and names for the elements whose chemical symbols consist of only one letter.
- In some cases, the symbol of an element does not seem to bear any relationship to the name we use for the element. Generally, the symbol for such an element is based on its name in another language. Give the symbols and names of five examples of such elements.
- Give the chemical symbol for each of the following elements.

a. neon	d. silicon
b. nickel	e. barium
c. potassium	f. silver

- Give the name or symbol, as appropriate, for each of the following elements.

Symbol	Name
Fe	
	chlorine
S	
	uranium
Ne	
K	

- Several elements have chemical symbols beginning with the letter C. For each of the following chemical symbols, give the name of the corresponding element.

a. Cu	e. Cr
b. Co	f. Cs
c. Ca	g. Cl
d. C	h. Cd

3.2 Atoms and Compounds

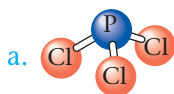
A. Dalton's Atomic Theory

- Indicate whether each of the following statements is true or false. If a statement is false, correct the statement so that it becomes true.
 - Most materials occur in nature as pure substances.
 - A given compound usually contains the same relative number of atoms of its various elements.
 - Atoms are made up of tiny particles called molecules.
- What does the law of constant composition tell us? How did Dalton's atomic theory help explain this law? Give examples.

B. Formulas of Compounds

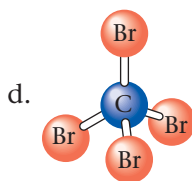
- What is a compound?
- A given compound always contains the same relative masses of its constituent elements. How is this idea related to the relative numbers of each kind of atom present?

16. Write the formula for each of the following substances, listing the elements in the order given.



List the phosphorus atom first.

- b. a molecule containing two boron atoms and six hydrogen atoms
 c. a compound containing one calcium atom for every two chlorine atoms



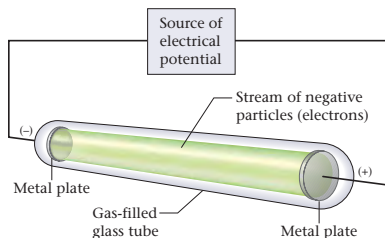
List the carbon atom first.

- e. a compound containing two iron atoms for every three oxygen atoms
 f. a molecule containing three hydrogen atoms, one phosphorus atom, and four oxygen atoms

3.3 Atomic Structure

A. The Structure of the Atom

17. Indicate whether each of the following statements is true or false. If a statement is false, correct the statement so that it becomes true.
- a. In his cathode ray tube experiments, J. J. Thomson obtained beams of different types of particles whose nature depended on which gas was contained in the tube.



- b. Thomson assumed that the atom must contain positively charged particles because isolated atoms have no overall charge.
 c. In the plum pudding model of the atom, the atom was envisioned as a sphere of negative charge in which positively charged electrons were randomly distributed.

18. Indicate whether each of the following statements is true or false. If a statement is false, correct it so that it becomes true.
- a. Rutherford's bombardment experiments with metal foil suggested that alpha particles were deflected by approaching a large, negatively charged atomic nucleus.
 b. The proton and the electron have similar masses but opposite electrical charges.
 c. Most atoms also contain neutrons, which are slightly heavier than protons but carry no charge.

B. Introduction to the Modern Concept of Atomic Structure

19. What uncharged particles are found in the nuclei of most atoms?
 20. What are the positively charged particles found in the nuclei of atoms called?
 21. Do the proton and the neutron have exactly the same mass? How do the masses of the proton and the neutron compare with the mass of the electron? Which particles make the greatest contribution to the mass of an atom? Which particles make the greatest contribution to the chemical properties of an atom?
 22. Although the nucleus of an atom is very important, it is the _____ of the atom that determine its chemical properties.

C. Isotopes

23. True or false? Atoms that have the same number of neutrons but different numbers of protons are called isotopes.
 24. True or false? The mass number of a nucleus represents the number of protons in the nucleus.
 25. How was Dalton's atomic theory modified after the discovery that several isotopes of an element may exist?
 26. Are all atoms of the same element identical? If not, how do they differ?

27. For each of the following elements, use the periodic table on the inside back cover of this book to write the element's atomic number.

- a. Ge
- b. zinc
- c. Cr
- d. tungsten
- e. Sr
- f. cobalt
- g. Be
- h. lithium

28. Write the atomic symbol A_ZX for each of the isotopes described below.

- a. $Z = 8$, number of neutrons = 9
- b. the isotope of chlorine in which $A = 37$
- c. $Z = 27$, $A = 60$
- d. number of protons = 26, number of neutrons = 31
- e. the isotope of I with a mass number of 131
- f. $Z = 3$, number of neutrons = 4

29. How many protons and neutrons are contained in the nucleus of each of the following atoms? Given that each atom is uncharged, how many electrons are present?

- a. ${}^{244}_{94}\text{Pu}$
- b. ${}^{241}_{95}\text{Am}$
- c. ${}^{227}_{89}\text{Ac}$
- d. ${}^{133}_{55}\text{Cs}$
- e. ${}^{193}_{77}\text{Ir}$
- f. ${}^{56}_{25}\text{Mn}$

30. Copy and complete the following table.

Name	Symbol	Atomic Number	Mass Number	Neutrons
sodium		11	23	
	${}^{15}_7\text{N}$			
	${}^{136}_{56}\text{Ba}$			
lithium				6
boron		5	11	

3.4 Using the Periodic Table

A. Introduction to the Periodic Table

- 31.** What property of the elements is considered when arranging the elements in the periodic table?
- 32.** In which direction on the periodic table, horizontal or vertical, are elements with similar chemical properties aligned? What are families of elements with similar chemical properties called?

33. List the characteristic physical properties that distinguish the metallic elements from the nonmetallic elements.

34. Where are the metallic elements found on the periodic table? Are there more metallic elements or nonmetallic elements?

35. Most, but not all, metallic elements are solids under ordinary laboratory conditions. Which metallic elements are *not* solids?

36. Give several examples of nonmetallic elements that occur in the gaseous state under ordinary conditions.

37. Under ordinary conditions, only a few pure elements occur as liquids. Give an example of a metallic element and a nonmetallic element that ordinarily occur as liquids.

38. What is a *metalloid*? Where are the metalloids found on the periodic table?

39. Write the number and name (if any) of the group (family) to which each of the following elements belongs.

- a. iodine
- b. Ca
- c. Na
- d. lithium
- e. Kr
- f. sodium
- g. Ne

40. Use the tables on the inside back cover of the book to give the chemical symbol, atomic number, and group number of each of the following elements and to specify whether each element is a metal, a metalloid, or a nonmetal.

- a. rubidium
- b. germanium
- c. magnesium
- d. titanium
- e. iodine

B. Natural States of the Elements

- 41.** Are most of the chemical elements found in nature in the elemental form or combined in compounds? Explain.
- 42.** Why are the elements of Group 8 called the noble or inert gas elements?

43. Give three examples of gaseous elements that exist as diatomic molecules. Give three examples of gaseous elements that exist as single atoms.
44. Most of the elements are solids at 30 °C. Give three examples of elements that are *liquids* and three examples of elements that are *gases* at 30 °C.

3.5 Ions and Their Compounds

A. Ions

45. For each of the positive ions listed in column 1, use the periodic table to find in column 2 the total number of electrons that ion contains. The same answer may be used more than once.

Column 1	Column 2
[1] Al ³⁺	[a] 2
[2] Fe ³⁺	[b] 10
[3] Mg ²⁺	[c] 21
[4] Sn ²⁺	[d] 23
[5] Co ²⁺	[e] 24
[6] Co ³⁺	[f] 25
[7] Li ⁺	[g] 36
[8] Cr ³⁺	[h] 48
[9] Rb ⁺	[i] 76
[10] Pt ²⁺	[j] 81

46. Complete the following tables.

a.

Cation	# Protons in ion	# Electrons in ion	Formula with O ²⁻
K ⁺			
Co ²⁺			
Al ³⁺			

b.

Anion	# Protons in ion	# Electrons in ion	Formula with Ca ²⁺
Cl ⁻			
S ²⁻			
P ³⁻			

47. For the following processes that show the formation of ions, use the periodic table to indicate the number of electrons and protons present in both the *ion* and the *neutral atom* from which the ion is made.

- Ca → Ca²⁺ + 2e⁻
- P + 3e⁻ → P³⁻
- Br + e⁻ → Br⁻
- Fe → Fe³⁺ + 3e⁻
- Al → Al³⁺ + 3e⁻
- N + 3e⁻ → N³⁻

48. For each of the following atomic numbers, use the periodic table to write the formula (including the charge) for the simple *ion* that the element is most likely to form.

- 53
- 38
- 55
- 88
- 9
- 13

B. Compounds That Contain Ions

49. List some properties of a substance that would lead you to believe that the substance consists of ions. How do these properties differ from those of nonionic compounds?
50. Why does an ionic compound conduct an electric current when the compound is melted but not when it is in the solid state?
51. Why must the total number of positive charges in an ionic compound equal the total number of negative charges?
52. Use the concept that a chemical compound must have a net charge of zero to predict the formula of the simplest compound that the following pairs of ions are most likely to form.
- Fe³⁺ and P³⁻
 - Fe³⁺ and S²⁻
 - Fe³⁺ and Cl⁻
 - Mg²⁺ and Cl⁻
 - Mg²⁺ and O²⁻
 - Mg²⁺ and N³⁻
 - Na⁺ and P³⁻
 - Na⁺ and S²⁻

Critical Thinking

53. What is the difference between the atomic number and the mass number of an element? Can atoms of two different elements have the same atomic number? Could they have the same mass number? Explain.

54. Which subatomic particles contribute most to the atom's mass? Which subatomic particles determine the atom's chemical properties?

55. Is it possible for the same two elements to form more than one compound? Is this possibility consistent with Dalton's atomic theory? Give an example.

56. Write the simplest formula for each of the following substances, listing the elements in the order given.

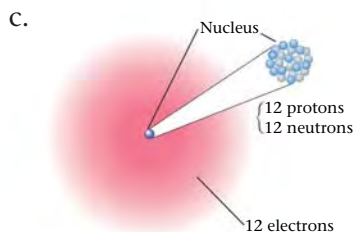
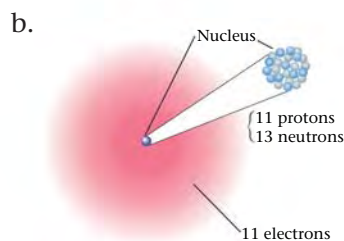
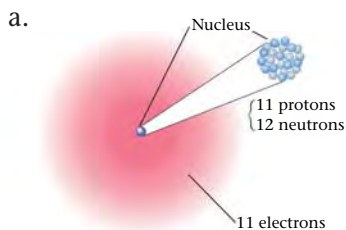
- a molecule containing one carbon atom and two oxygen atoms
- a compound containing one aluminum atom for every three chlorine atoms
- perchloric acid, which contains one hydrogen atom, one chlorine atom, and four oxygen atoms
- a molecule containing one sulfur atom and six chlorine atoms

57. List the charges of the ions formed from atoms in each of the groups.

- _____
- _____
- _____
- _____

58. An ion has a charge of $2-$ and 18 electrons. From which atom does this ion form?

59. Use the following figures to identify the element or ion. Write the symbol for each, using the A_ZX format.



60. Choose the statement that best answers this question: If all atoms are composed of the same subatomic particles, why do different atoms have different chemical properties?

- The number and arrangement of the electrons matter most because the electrons of the atoms "intermingle" when atoms combine to form molecules.
- The number and arrangement of the electrons matter most because the electrons of the atoms are located in the nucleus, and the nucleus is involved in chemical reactions.
- The number and arrangement of the protons matter most because the protons of the atoms "intermingle" when atoms combine to form molecules.
- The number and arrangement of the protons matter most because the protons of the atoms are located in the nucleus, and the nucleus is involved in chemical reactions.


LOOKING AHEAD

4.1 Naming Binary Compounds

- A. Naming Compounds That Contain a Metal and a Nonmetal
 - Type I Binary Ionic Compounds
 - Type II Binary Ionic Compounds
- B. Naming Binary Compounds That Contain Only Nonmetals (Type III)
- C. Naming Binary Compounds: A Review

4.2 Naming and Writing Formulas for More Complex Compounds

- A. Naming Compounds That Contain Polyatomic Ions
- B. Naming Acids
- C. Writing Formulas from Names



This shell on the beach is made of calcium carbonate.

IN YOUR LIFE

As young children one of the most important things we learned was the names of objects. As we developed, it was important to learn the accepted names for things. For example, if a little girl wants *milk*, but says *bananas*, a great deal of frustration occurs both for the caretaker and the child. To communicate, we need to have names for things that are agreed on by everyone. The sentence “I drove downtown on the freeway, to buy some DVDs” would be meaningless to people who did not know the meaning of downtown, freeway, or DVD.

As we learn new things in school, we have to learn new languages. Mathematics has its own language, computers have a characteristic language, and if we play a particular sport, it has a language of its own.

Learning chemistry requires learning a new language. Although we use the word *plastic* in everyday life to describe many objects, chemists require names for these materials that are more precise. For example, plastic milk jugs, garden furniture, and trash cans are made of polyethylene while disposable foam cups and egg cartons are made of polystyrene. In this chapter we will learn the systematic names of chemicals so we can talk “chemistry” to each other.



WHAT DO YOU KNOW?

Prereading Questions

1. What is the chemical name for table salt? What is the chemical formula for table salt?
2. What is meant by the term *compound*? Provide an example, and name if it you can.
3. Provide the name for each of the following symbols:
Na, K, Ca, O, F, Cl.
4. Using a periodic table, predict the charge of the following atoms when they are part of an ionic compound:
Na, K, Ca, O, F, Cl.
5. Have you ever heard of an *acid*? Can you provide a formula and a name for an acid?

SECTION 4.1

Naming Binary Compounds

Key Terms

- Binary compound
- Binary ionic compound

Objectives

- To learn to name binary compounds of a metal and nonmetal
- To learn to name binary compounds containing only nonmetals
- To summarize the naming of all types of binary compounds

When chemistry was an infant science, there was no system for naming compounds. Names such as sugar of lead, blue vitriol, quicklime, Epsom salts, milk of magnesia, gypsum, and laughing gas were coined by early chemists. Such names are called *common names*. As our knowledge of chemistry grew, it became clear that using common names for compounds was not practical. More than four million chemical compounds are currently known. Memorizing common names for all these compounds would be impossible.

The solution, of course, is a *system* for naming compounds in which the name tells something about the composition of the compound. After learning the system, you should be able to name a compound when you are given its formula. Conversely, you should be able to construct a compound's formula, given its name. In this chapter we will specify the most important rules for naming compounds other than organic compounds (those based on chains of carbon atoms).

A. Naming Compounds That Contain a Metal and a Nonmetal

As we saw earlier, when a metal such as sodium combines with a nonmetal such as chlorine, the resulting compound contains ions. The metal loses one or more electrons to become a cation, and the nonmetal gains one or more electrons to form an anion. The resulting substance is called a **binary ionic compound**. Binary ionic compounds contain a positive ion (cation), which is always written first in the formula, and a negative ion (anion). *To name these compounds we simply name the ions.*

In this section we will consider binary ionic compounds of two types based on the cations they contain. Certain metal atoms form only one cation. For example,

- The Na atom always forms Na^+ , *never* Na^{2+} or Na^{3+} ,
- Cs always forms Cs^+ , Ca always forms Ca^{2+} , and
- Al always forms Al^{3+} .

We will call compounds that contain this type of metal atom Type I binary compounds and the cations they contain Type I cations. Examples of Type I cations are Na^+ , Ca^{2+} , Cs^+ , and Al^{3+} .

Other metal atoms can form two or more cations. For example,

- Cr can form Cr^{2+} and Cr^{3+} and
- Cu can form Cu^+ and Cu^{2+} .

We will call such ions Type II cations and their compounds Type II binary compounds.

Type I compounds: The metal present forms only one type of cation.

Type II compounds: The metal present can form two (or more) cations that have different charges.



An artist using plaster of Paris, a gypsum plaster

Binary ionic compound

A two-element compound consisting of a cation and an anion

Type I	Type II
Na \rightarrow Na^+	Cr \rightarrow Cr^{2+}
Cs \rightarrow Cs^+	\rightarrow Cr^{3+}
Ca \rightarrow Ca^{2+}	Cu \rightarrow Cu^+
Al \rightarrow Al^{3+}	\rightarrow Cu^{2+}

Some common cations and anions and their names are listed in **Table 4.1**. You should memorize these. They are an essential part of your chemical vocabulary.

TABLE 4.1**Common Simple Cations and Anions**

Cation	Name	Anion	Name*
H ⁺	hydrogen	H ⁻	hydride
Li ⁺	lithium	F ⁻	fluoride
Na ⁺	sodium	Cl ⁻	chloride
K ⁺	potassium	Br ⁻	bromide
Cs ⁺	cesium	I ⁻	iodide
Be ²⁺	beryllium	O ²⁻	oxide
Mg ²⁺	magnesium	S ²⁻	sulfide
Ca ²⁺	calcium		
Ba ²⁺	barium		
Al ³⁺	aluminum		
Ag ⁺	silver		

*The root is given in color.

Information

A simple cation has the same name as its parent element.

**Active Reading Question**

Name three metal atoms which form only one cation. Include the charge for each.

Type I Binary Ionic Compounds

The following rules apply for Type I ionic compounds:

Rules for Naming Type I Ionic Compounds

1. The cation is always named first and the anion second.
2. A simple cation (obtained from a single atom) takes its name from the name of the element. For example, Na⁺ is called sodium in the names of compounds containing this ion.
3. A simple anion (obtained from a single atom) is named by taking the first part of the element name (the root) and adding *-ide*. Thus the Cl⁻ ion is called chloride.
4. Write the name for the compound by combining the names of the ions.

We will illustrate these rules by naming a few compounds. For example, the compound NaI is called sodium iodide. It contains Na⁺ (the sodium cation, named for the parent metal) and I⁻ (iodide: the root of iodine plus *-ide*). Similarly, the compound CaO is called calcium oxide because it contains Ca²⁺ (the calcium cation) and O²⁻ (the oxide anion).



The rules for naming binary compounds are also illustrated by the following table.

Examples of Binary Ionic Compounds		
Compound	Ions Present	Name
NaCl	Na^+ , Cl^-	sodium chloride
KI	K^+ , I^-	potassium iodide
CaS	Ca^{2+} , S^{2-}	calcium sulfide
CsBr	Cs^+ , Br^-	cesium bromide
MgO	Mg^{2+} , O^{2-}	magnesium oxide

It is important to note that in the *formulas* of ionic compounds, simple ions are represented by the element symbol: Cl means Cl^- , Na means Na^+ , and so on. However, when *individual ions* are shown, the charge is always included. Thus the formula of potassium bromide is written KBr, but when the potassium and bromide ions are shown individually, they are written K^+ and Br^- .

CHEMISTRY IN YOUR WORLD

Connection to History

Sugar of Lead

In ancient Roman society it was common to boil fermented grape juice in a lead-lined vessel, driving off much of the water to produce a very sweet, thick syrup called *sapa*. This syrup was often used as a sweetener for many types of food and drink.

We now realize that a major component of this syrup was lead acetate, $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$. This compound has a very sweet taste—hence its original name, sugar of lead.

Many historians believe that the fall of the Roman Empire was due at least in part to lead poisoning, which causes lethargy and mental malfunctions. One major source of this lead was the *sapa* syrup. In addition, the Romans' highly advanced plumbing system employed lead water pipes, which allowed lead to be leached into their drinking water.

Sadly, this story is more relevant to today's society than you might think. Lead-based solder was widely used for many years to connect the copper pipes in

water systems in homes and commercial buildings. There is evidence that dangerous amounts of lead can be leached from these soldered joints into drinking water. In fact, large quantities of lead have been found in the water that some drinking fountains and water coolers dispense. In response to these problems, the U.S. Congress has passed a law banning lead from the solder used in plumbing systems for drinking water.



A carving showing how lead-lined vessels were used

EXAMPLE 4.1**Naming Type I Binary Compounds**

Name each binary compound.

- a. CsF b. AlCl₃ c. MgI₂

Solution

We will name these compounds by systematically following the rules given on page 95.

- a. CsF

Step 1 Identify the cation and anion. Cs is in Group 1, so we know it will form the 1+ ion Cs⁺. Because F is in Group 7, it forms the 1− ion F[−].

Step 2 Name the cation. Cs⁺ is simply called cesium, the same as the element name.

Step 3 Name the anion. F[−] is called fluoride: we use the root name of the element plus *-ide*.

Step 4 Name the compound by combining the names of the individual ions. The name for CsF is cesium fluoride. (Remember that the name of the cation is always given first.)

b. <i>Compound</i>	<i>Ions Present</i>	<i>Ion Names</i>	<i>Comments</i>
$\begin{array}{l} \xrightarrow{\text{Cation}} \\ \text{AlCl}_3 \\ \xleftarrow{\text{Anion}} \end{array}$	Al ³⁺	aluminum	Al (Group 3) always forms Al ³⁺ .
	Cl [−]	chloride	Cl (Group 7) always forms Cl [−] .

The name of AlCl₃ is aluminum chloride.

c. <i>Compound</i>	<i>Ions Present</i>	<i>Ion Names</i>	<i>Comments</i>
$\begin{array}{l} \xrightarrow{\text{Cation}} \\ \text{MgI}_2 \\ \xleftarrow{\text{Anion}} \end{array}$	Mg ²⁺	magnesium	Mg (Group 2) always forms Mg ²⁺ .
	I [−]	iodide	I (Group 7) gains one electron to form I [−] .

The name of MgI₂ is magnesium iodide.

Practice Problem • Exercise 4.1

Name the following compounds.

- a. Rb₂O b. SrI₂ c. K₂S

Let's Review

- Compounds formed from metals and nonmetals are ionic.
- In an ionic compound the cation is always named first.
- The *net* charge on an ionic compound is always zero. In CsF, one of each type of ion (Cs⁺ and F[−]) is required: $(1+) + (1-) = 0$ charge. In AlCl₃, however, three Cl[−] ions are needed to balance the charge of Al³⁺: $(3+) + 3(1-) = 0$ charge. In MgI₂, two I[−] ions are needed for each Mg²⁺ ion: $(2+) + 2(1-) = 0$ charge.

How Can Water Be Hard or Soft?

Water, like most liquids we use every day, is not 100% pure. Floating about in our tap water are all kinds of ions and molecules collected from the places where our water has been. Are these ions and molecules harmful to us? It depends on what they are. The most common impurities in tap water are calcium and magnesium ions. They dissolve into the water as it passes through soil rich in limestone. Calcium and magnesium ions are important in our diets, but when dissolved in water they can have some irritating properties as well. It is “hard” to get soaps to lather—and so we call water with lots of these ions “hard water.”

The calcium and magnesium compounds in hard water also leave a white scale or residue inside tea kettles, coffee pots, irons, and other appliances that use water. This scale can build up and keep the appliance from working properly. When soap is used

with hard water, a new compound is formed—calcium or magnesium stearate—that clouds the water and makes a ring in the bathtub and makes clothes appear dingy after washing.

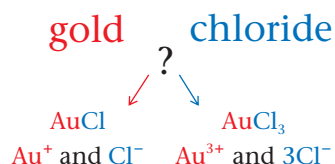
How can we make water soft? Some homeowners install a water softener, which is really an ion exchanger. It is filled with sodium chloride (salt), which exchanges sodium ions for the calcium and magnesium ions in the hard water, making the water soft. How does this exchange alter the behavior of the water? The sodium ions do not react with carbonates to form the white scaly deposits or react with soaps to form soap scum. They produce no negative effects in drinking water, except in people who must limit their sodium intake because of high blood pressure or heart disease. A bonus—your shampoo and soap make much more lather, so less is needed to do the job!

Information Type II Binary Ionic Compounds

Type II binary ionic compounds contain a metal that can form more than one type of cation.

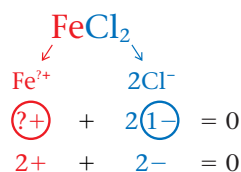
So far we have considered binary ionic compounds (Type I) containing metals that always give the same cation. For example, sodium always forms the Na^+ ion, calcium always forms the Ca^{2+} ion, and aluminum always forms the Al^{3+} ion. As we said previously, we can predict with certainty that each Group 1 metal will give a 1+ cation and each Group 2 metal will give a 2+ cation. And aluminum always forms Al^{3+} .

However, many metals can form more than one type of cation. For example, lead (Pb) can form Pb^{2+} or Pb^{4+} in ionic compounds. Also, iron (Fe) can form Fe^{2+} or Fe^{3+} , chromium (Cr) can form Cr^{2+} or Cr^{3+} , gold (Au) can form Au^+ or Au^{3+} , and so on. This means that if we saw the name gold chloride, we wouldn't know whether it referred to the compound AuCl (containing Au^+ and Cl^-) or the compound AuCl_3 (containing Au^{3+} and three Cl^- ions). Therefore, we need a way of specifying which cation is present in compounds containing metals that can form more than one type of cation.



Chemists have decided to deal with this situation by using a Roman numeral to specify the charge on the cation. To see how this works, consider the compound FeCl_2 . Iron can form Fe^{2+} or Fe^{3+} , so we must first decide which of these cations is present. We can determine the

charge on the iron cation, because we know it must balance the charge on the two 1⁻ anions (the chloride ions).

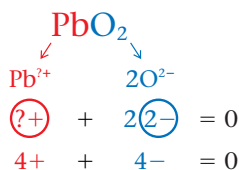


So the charge on Fe = 2+.

The compound FeCl₂, then, contains one Fe²⁺ ion and two Cl⁻ ions. We call this compound iron(II) chloride, where the II tells the charge of the iron cation. That is, Fe²⁺ is called iron(II). Likewise, Fe³⁺ is called iron(III). And FeCl₃, which contains one Fe³⁺ ion and three Cl⁻ ions, is called iron(III) chloride.

The Roman numeral tells the charge on the ion, not the number of ions present in the compound.

Note that in the above examples the Roman numeral for the cation turned out to be the same as the subscript needed for the anion (to balance the charge). This is often not the case. For example, consider the compound PbO₂.



So the charge on Pb = 4+.

The charge on the lead ion must be 4+ to balance the 4⁻ charge of the two oxide ions. The name of PbO₂ is therefore lead(IV) oxide, where the IV indicates the presence of the Pb⁴⁺ cation.

There is an older system for naming ionic compounds containing metals that form two cations. The ion with the higher charge has a name ending in *-ic*, and the one with the lower charge has a name ending in *-ous*. In this system, for example, Fe³⁺ is called the ferric ion, and Fe²⁺ is called the ferrous ion. The names for FeCl₃ and FeCl₂, in this system, are ferric chloride and ferrous chloride, respectively. **Table 4.2** gives both names for many Type II cations. We will use the system of Roman numerals exclusively in this text; the other system is falling into disuse.

Information

FeCl₃ must contain Fe³⁺ to balance the charge of three Cl⁻ ions.



Copper(II) sulfate crystals

TABLE 4.2

Common Type II Cations

Ion	Systematic Name	Older Name	Ion	Systematic Name	Older Name
Fe ³⁺	iron(III)	ferric	Sn ⁴⁺	tin(IV)	stannic
Fe ²⁺	iron(II)	ferrous	Sn ²⁺	tin(II)	stannous
Cu ²⁺	copper(II)	cupric	Pb ⁴⁺	lead(IV)	plumbic
Cu ⁺	copper(I)	cuprous	Pb ²⁺	lead(II)	plumbous
Co ³⁺	cobalt(III)	cobaltic	Hg ²⁺	mercury(II)	mercuric
Co ²⁺	cobalt(II)	cobaltous	Hg ₂ ^{2+*}	mercury(I)	mercurous

*Mercury(I) ions always occur bound together in pairs to form Hg₂²⁺.

To help distinguish between Type I and Type II cations:

- Group 1 and 2 metals are always Type I.
- Transition metals are almost always Type II.

EXAMPLE 4.2

Naming Type II Binary Compounds

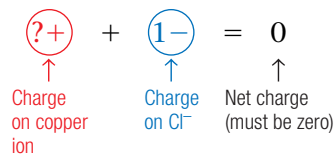
Give the systematic name of each of the following compounds.

- a. CuCl c. Fe₂O₃ e. PbCl₄
 b. HgO d. MnO₂

Solution

All these compounds include a metal that can form more than one type of cation; thus we must first determine the charge on each cation. We do this by recognizing that a compound must be electrically neutral; that is, the positive and negative charges must balance exactly. We will use the known charge on the anion to determine the charge of the cation.

- a. In CuCl we recognize the anion as Cl⁻. To determine the charge on the copper cation, we use the principle of charge balance.



In this case, ?+ must be 1+ because (1+) + (1-) = 0. Thus the copper cation must be Cu⁺. Now we can name the compound by using the regular steps.

CRITICAL THINKING ?

We can use the periodic table to predict the stable ions formed by many atoms. For example, the atoms in column 1 always form 1+ ions. The transition metals, however, can form more than one type of stable ion.

What if each transition metal ion had only one possible charge? How would the naming of compounds be different?

Compound	Ions Present	Ion Names	Comments
CuCl	Cation Cu ⁺	copper(I)	Copper forms other cations (it is a transition metal), so we must include the I to specify its charge.
	Anion Cl ⁻	chloride	

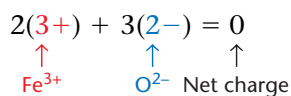
The name of CuCl is copper(I) chloride.

- b. In HgO we recognize the O²⁻ anion. To yield zero net charge, the cation must be Hg²⁺.

Compound	Ions Present	Ion Names	Comments
HgO	Cation Hg ²⁺	mercury(II)	The II is necessary to specify the charge.
	Anion O ²⁻	oxide	

The name of HgO is mercury(II) oxide.

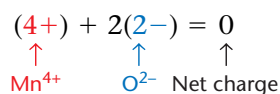
- c. Because Fe_2O_3 contains three O^{2-} anions, the charge on the iron cation must be 3+.



Compound	Ions Present	Ion Names	Comments
Fe_2O_3	Cation → Fe^{3+}	iron(III)	Iron is a transition metal and requires a III to specify the charge on the cation.
	Anion → O^{2-}	oxide	

The name of Fe_2O_3 is iron(III) oxide.

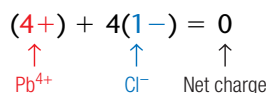
- d. MnO_2 contains two O^{2-} anions, so the charge on the manganese cation is 4+.



Compound	Ions Present	Ion Names	Comments
MnO_2	Cation → Mn^{4+}	manganese (IV)	Manganese is a transition metal and requires a IV to specify the charge on the cation.
	Anion → O^{2-}	oxide	

The name of MnO_2 is manganese(IV) oxide.

- e. Because PbCl_4 contains four Cl^- anions, the charge on the lead cation is 4+.



Compound	Ions Present	Ion Names	Comments
PbCl_4	Cation → Pb^{4+}	lead(IV)	Lead forms both Pb^{2+} and Pb^{4+} , so a Roman numeral is required.
	Anion → Cl^-	chloride	

The name for PbCl_4 is lead(IV) chloride.

The use of a Roman numeral in a systematic name for a compound is required only when more than one ionic compound forms between a given pair of elements. This occurs most often for compounds that contain transition metals, which frequently form more than one cation.

Metals that form only one cation do not need to be identified by a Roman numeral.

Common metals that do not require Roman numerals are the Group 1 elements, which form only 1+ ions; the Group 2 elements, which form only 2+ ions; and such Group 3 metals as aluminum and gallium, which form only 3+ ions.

Active Reading Question

Name three metal atoms which form more than one cation. Include all charges for each.

As shown in Example 4.2, when a metal ion that forms more than one type of cation is present, the charge on the metal ion must be determined by balancing the positive and negative charges of the compound. To do this, you must be able to recognize the common anions and you must know their charges (see Table 4.1).

EXAMPLE 4.3

Naming Binary Ionic Compounds: Summary

Give the systematic name of each of the following compounds.

- a. CoBr_2 c. Al_2O_3
 b. CaCl_2 d. CrCl_3

Solution

Compound	Ions and Names	Compound Name	Comments
a. CoBr_2	Co^{2+} cobalt(II) Br^- bromide	cobalt(II) bromide	Cobalt is a transition metal; the name of the compound must have a Roman numeral. The two Br^- ions must be balanced by a Co^{2+} cation.
b. CaCl_2	Ca^{2+} calcium Cl^- chloride	calcium chloride	Calcium, an alkaline earth metal, forms only the Ca^{2+} ion. A Roman numeral is not necessary.
c. Al_2O_3	Al^{3+} aluminum O^{2-} oxide	aluminum oxide	Aluminum forms only Al^{3+} . A Roman numeral is not necessary.
d. CrCl_3	Cr^{3+} chromium(III) Cl^- chloride	chromium(III) chloride	Chromium is a transition metal. The name of the compound must have a Roman numeral. CrCl_3 contains Cr^{3+} .

Practice Problem • Exercise 4.3

Give the names of the following compounds.

- a. PbBr_2 and PbBr_4 c. AlBr_3 e. CoCl_3
 b. FeS and Fe_2S_3 d. Na_2S

Name Game I: Ionic Compounds**Materials**

- index cards

Procedure

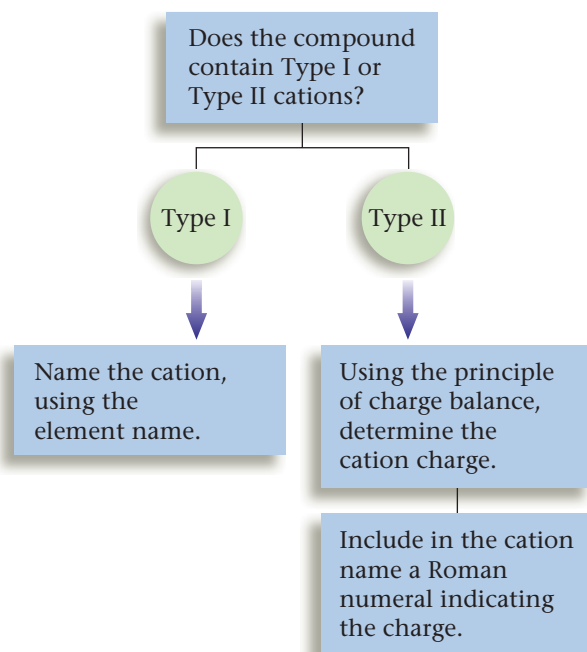
- Make two sets of flashcards. Write the name of an ion on one side of a card and its symbol and charge on the other side.
 - Set A: all of the alkali and alkaline earth metals, and iron(II) and iron(III).
 - Set B: all of the nonmetals from Groups 5A, 6A, and 7A.

- Randomly pick one card from Set A and one card from Set B. Write the proper formula for the ionic compound made from these two ions, and name the compound.
- Repeat step 2 until you answer ten in a row correctly.

Results/Analysis

- Make sure that you can write formulas and names when given either the names or the symbols and charges of the ions.

The following flow chart is useful when you are naming binary ionic compounds:

**Information**

Sometimes transition metals form only one ion, such as silver, which forms Ag^+ ; zinc, which forms Zn^{2+} ; and cadmium, which forms Cd^{2+} . In these cases, chemists do not use a Roman numeral, although it is not “wrong” to do so.

Active Reading Question

What suffix usually tells you the named compound is a binary compound?

B. Naming Binary Compounds That Contain Only Nonmetals (Type III)

Binary compounds that contain only nonmetals are named in accordance with a system similar in some ways to the rules for naming binary ionic compounds, but there are important differences.

Type III binary compounds contain only nonmetals.

The following rules cover the naming of these compounds.

Rules for Naming Type III Binary Compounds

1. The first element in the formula is named first, and the full element name is used.
2. The second element is named as though it were an anion.
3. Prefixes are used to denote the numbers of atoms present. These prefixes are given in **Table 4.3**.
4. The prefix *mono-* is never used for naming the first element. For example, CO is called carbon monoxide, *not* monocarbon monoxide.

EXAMPLE 4.4

Naming Type III Binary Compounds

Name the following binary compounds, which contain two nonmetals (Type III).

- a. BF_3 b. NO c. N_2O_5

Solution

- a. BF_3

Rule 1 Name the first element, using the full element name: boron.

Rule 2 Name the second element as though it were an anion: fluoride.

Rule 3 and 4 Use prefixes to denote numbers of atoms. One boron atom: do not use *mono-* in first position. Three fluorine atoms: use the prefix *tri-*.

The name of BF_3 is boron trifluoride.

- | b. <i>Compound</i> | <i>Individual Names</i> | <i>Prefixes</i> | <i>Comments</i> |
|--------------------|-------------------------|----------------------|---|
| NO | nitrogen
oxide | none
<i>mono-</i> | <i>Mono-</i> is not used for the first element. |

The name for NO is nitrogen monoxide. Note that the second *o* in *mono-* has been dropped for easier pronunciation. The *common* name for NO , which is often used by chemists, is nitric oxide.

- | c. <i>Compound</i> | <i>Individual Names</i> | <i>Prefixes</i> | <i>Comments</i> |
|------------------------|-------------------------|-----------------------------|-----------------------------|
| N_2O_5 | nitrogen
oxide | <i>di-</i>
<i>penta-</i> | two N atoms
five O atoms |

The name for N_2O_5 is dinitrogen pentoxide. The *a* in *penta-* has been dropped for easier pronunciation.

Practice Problem • Exercise 4.4

Name the following compounds.

- a. CCl_4 b. NO_2 c. IF_5

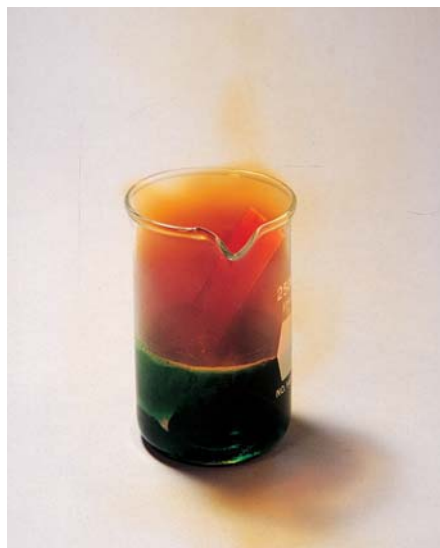
TABLE 4.3

Prefixes Used to Indicate Numbers in Chemical Names

Prefix	Number Indicated
<i>mono-</i>	1
<i>di-</i>	2
<i>tri-</i>	3
<i>tetra-</i>	4
<i>penta-</i>	5
<i>hexa-</i>	6
<i>hepta-</i>	7
<i>octa-</i>	8



A piece of copper metal about to be placed in nitric acid



Copper reacts with nitric acid to produce colorless NO, which immediately reacts with the oxygen in the air to form reddish-brown NO₂ gas.

The previous examples illustrate that, to avoid awkward pronunciation, we often drop the final *o* or *a* of the prefix when the second element is oxygen. For example, N₂O₄ is called dinitrogen tetroxide, *not* dinitrogen tetraoxide, and CO is called carbon monoxide, *not* carbon monoxide.

Some compounds are always referred to by their common names. The two best examples are water and ammonia. The systematic names for H₂O and NH₃ are never used.

EXAMPLE 4.5

Naming Type III Binary Compounds: Summary

Name each of the following compounds.

- a. PCl₅ b. P₄O₆ c. SF₆ d. SO₃ e. SO₂ f. N₂O₃

Solution

Compound	Name
a. PCl ₅	phosphorus pentachloride
b. P ₄ O ₆	tetraphosphorus hexoxide
c. SF ₆	sulfur hexafluoride
d. SO ₃	sulfur trioxide
e. SO ₂	sulfur dioxide
f. N ₂ O ₃	dinitrogen trioxide

Practice Problem • Exercise 4.5

Name the following compounds.

- a. SiO₂ b. O₂F₂ c. XeF₆

C. Naming Binary Compounds: A Review

Because different rules apply for naming various types of binary compounds, we will now consider an overall strategy to use for these compounds. We have considered three types of binary compounds, and naming each type requires a different procedure.

CRITICAL THINKING ?

What if all compounds had only common names? What problems might arise?

Let's Review

Type I: Ionic compounds with metals that always form a cation with the same charge

Type II: Ionic compounds with metals (usually transition metals) that form cations with various charges

Type III: Compounds that contain only nonmetals

Active Reading Question

Why are prefixes such as "mono-" and "di-" unnecessary for a compound made of a metal and a nonmetal?

In trying to determine which type of compound you are naming, use the periodic table to help you identify metals and nonmetals and to determine which elements are transition metals.

The flow chart given in **Figure 4.1** should help you as you name binary compounds of the various types.

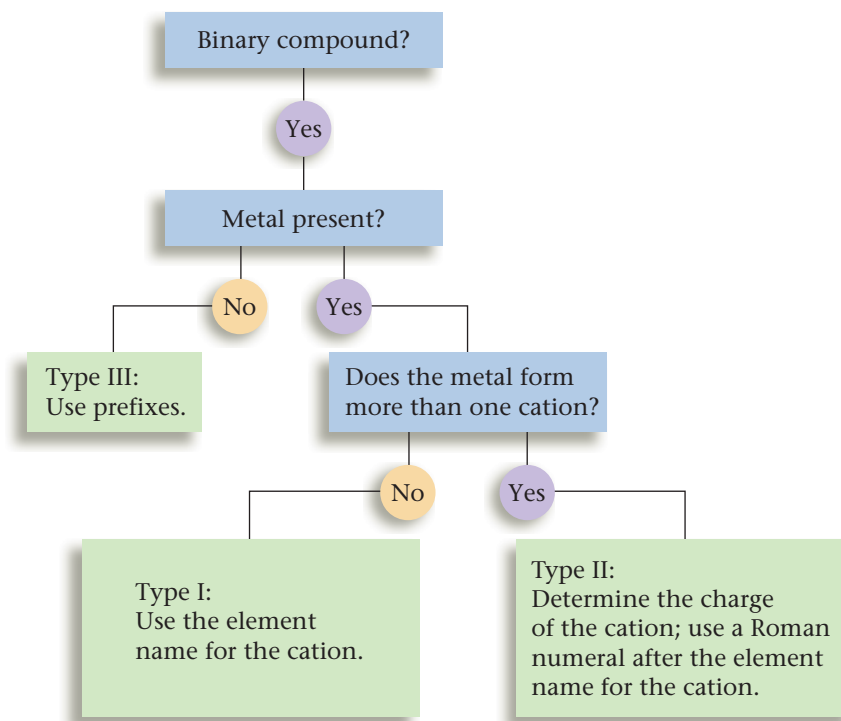


Figure 4.1

A flow chart for naming binary compounds

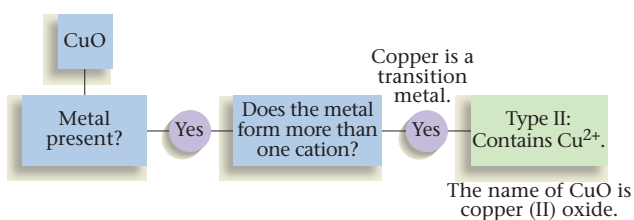
EXAMPLE 4.6**Naming Binary Compounds: Summary**

Name the following binary compounds.

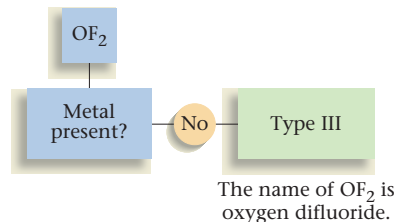
- a. CuO b. SrO c. B
- ₂
- O
- ₃
- d. TiCl
- ₄
- e. K
- ₂
- S f. OF
- ₂
- g. NH
- ₃

Solution

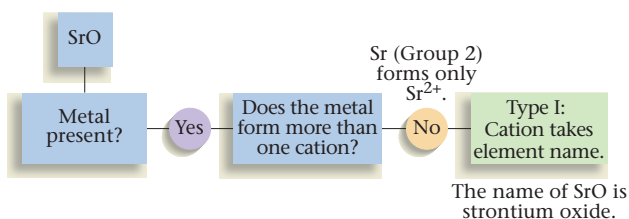
a.



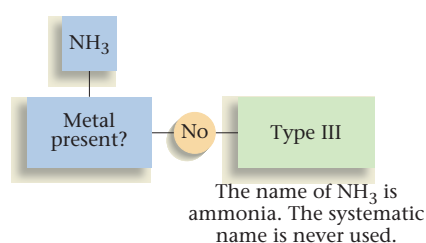
f.



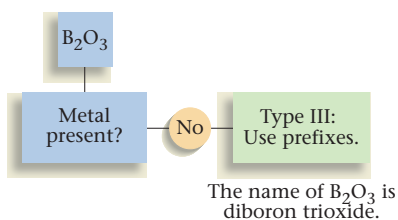
b.



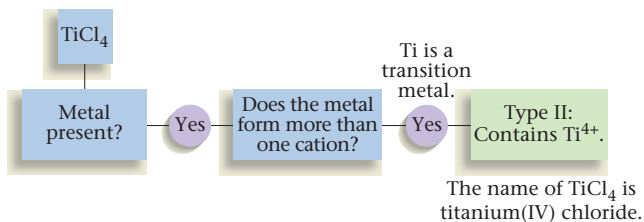
g.



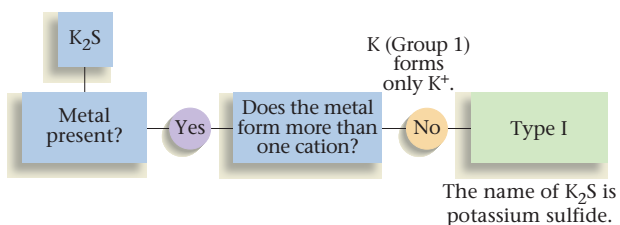
c.



d.



e.

**Practice Problem • Exercise 4.6**

Name the following binary compounds.

- a. ClF
- ₃
- b. VF
- ₅
- c. CuCl d. MnO
- ₂
- e. MgO f. H
- ₂
- O

CELEBRITY CHEMICAL

Nitric Oxide (NO)

Until recently nitric oxide (NO, more correctly called nitrogen monoxide) was primarily viewed as an air pollutant. In the last few years, however, NO has been found to be a potent biological regulator. It turns out that this compound is produced in the body and regulates blood pressure by dilating blood vessels. Studies in the 1990s indicated that NO administered to patients with sickle cell anemia may relieve serious symptoms of that disease. For example, it seems to benefit patients with acute chest syndrome—characterized by chest pain, fever, and high blood pressure in the lungs due to clogged blood vessels—which is the most life-threatening complication of sickle cell disease, especially for children.

Sickle cell anemia is a genetic disease that most often strikes people of West African descent. It causes the red blood cells to be misshapen (“sickle-shaped”), which in turn causes them to stick together and block blood flow. NO is thought to bind to a special site on the hemoglobin molecule (the molecule that carries O_2 from the lungs to the tissues). It is released when blood flow is impaired, causing dilation (expansion) of the blood vessels, thus helping to improve blood flow.

Further tests are now under way to try to discover exactly how NO works to relieve sickle cell symptoms.

A mixture of normal and sickle-shaped red blood cells



SECTION 4.1 REVIEW QUESTIONS

- Why is a system for naming compounds necessary?
- How can you tell whether a substance is a binary compound?
- How are the anions in all types of binary compounds similar?
- How are the cations different in each type of binary compound?
- Use **Figure 4.1** to name these compounds:
 - KBr
 - SnF₂
 - CO
 - CuBr
 - MgI₂
 - PCl₃

- List the charges of the ions formed from atoms in each of the groups.

The diagram shows a periodic table with the following regions:

- Metals:** Grey shaded area covering the left and middle sections.
- Nonmetals:** Yellow shaded area covering the top-right section.
- Transition Region:** Blue shaded area between the metals and nonmetals.

 Labels and arrows:

- a:** Points to the first group (Group 1).
- b:** Points to the last group (Group 18).
- c:** Points to the second group (Group 2).
- d:** Points to the first group of nonmetals (Group 17).

- _____
- _____
- _____
- _____



Naming and Writing Formulas for More Complex Compounds

Objectives

- To learn the names of common polyatomic ions
- To learn to name compounds containing polyatomic ions
- To learn how the anion composition determines an acid's name
- To learn names for common acids
- To learn to write the formula for a compound, given its name

A. Naming Compounds That Contain Polyatomic Ions

A type of ionic compound that we have not yet considered is exemplified by ammonium nitrate, NH_4NO_3 , which contains the **polyatomic ions** NH_4^+ and NO_3^- . As their name suggests, polyatomic ions are charged entities composed of several atoms bound together. Polyatomic ions are assigned special names that you *must memorize* to name the compounds containing them. The most important polyatomic ions and their names are listed in **Table 4.4**.

Note in Table 4.4 that several series of polyatomic anions exist that contain an atom of a given element and different numbers of oxygen atoms. These anions are called **oxyanions**. When there are two members in such a series, the name of the one with the smaller number of oxygen atoms ends in *-ite*, and the name of the one with the larger number ends in *-ate*.

Key Terms

- Polyatomic ions
- Oxyanions
- Acids

Polyatomic ions

An ion consisting of two or more atoms bound together

Oxyanion

A polyatomic ion containing at least one oxygen atom and one or more atoms of at least one other element

TABLE 4.4

Names of Common Polyatomic Ions

Ion	Name	Ion	Name
NH_4^+	ammonium	CO_3^{2-}	carbonate
NO_2^-	nitrite	HCO_3^-	hydrogen carbonate (bicarbonate is a widely used common name)
NO_3^-	nitrate	ClO^-	hypochlorite
SO_3^{2-}	sulfite	ClO_2^-	chlorite
SO_4^{2-}	sulfate	ClO_3^-	chlorate
HSO_4^-	hydrogen sulfate (bisulfate is a widely used common name)	ClO_4^-	perchlorate
OH^-	hydroxide	$\text{C}_2\text{H}_3\text{O}_2^-$	acetate
CN^-	cyanide	MnO_4^-	permanganate
PO_4^{3-}	phosphate	$\text{Cr}_2\text{O}_7^{2-}$	dichromate
HPO_4^{2-}	hydrogen phosphate	CrO_4^{2-}	chromate
H_2PO_4^-	dihydrogen phosphate	O_2^{2-}	peroxide

DID YOU KNOW

Ionic compounds containing polyatomic ions are not binary compounds, because they contain more than two elements.

Information

Note that the SO_3^{2-} anion has very different properties from SO_3 (sulfur trioxide), a pungent, toxic gas.

Information

Except for hydroxide, peroxide, and cyanide, the names of polyatomic ions do not have an *-ide* ending.

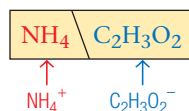
For example, SO_3^{2-} is sulfite and SO_4^{2-} is sulfate. When more than two oxyanions make up a series, *hypo-* (less than) and *per-* (more than) are used as prefixes to name the members of the series with the fewest and the most oxygen atoms, respectively. The best example involves the oxyanions containing chlorine:

ClO^-	<i>hypochlorite</i>
ClO_2^-	<i>chlorite</i>
ClO_3^-	<i>chlorate</i>
ClO_4^-	<i>perchlorate</i>

Naming ionic compounds that contain polyatomic ions is very similar to naming binary ionic compounds. For example, the compound NaOH is called sodium hydroxide, because it contains the Na^+ (sodium) cation and the OH^- (hydroxide) anion. To name these compounds, you must

- learn to recognize the common polyatomic ions.
- learn the *composition* and *charge* of each of the ions in Table 4.4.

Then when you see the formula $\text{NH}_4\text{C}_2\text{H}_3\text{O}_2$, you should immediately recognize its two “parts”:



The correct name is ammonium acetate.

When a metal is present that forms more than one cation, a Roman numeral is required to specify the cation charge, just as in naming Type II binary ionic compounds. For example, the compound FeSO_4 is called iron(II) sulfate, because it contains Fe^{2+} (to balance the $2-$ charge on SO_4^{2-}). Note that to determine the charge on the iron cation, you must know that sulfate has a $2-$ charge.

HANDS-ON CHEMISTRY

• MINI-LAB •

Name Game II: Polyatomic Ions

Materials

- Index cards

Procedure

1. Make flashcards for the common polyatomic ions listed in **Table 4.4**. Write the name on one side and the formula (with charge) on the other side.
2. Group all of the oxyanions together, and arrange these according to suffix (*-ate* or *-ite*). Notice the patterns.

Results/Analysis

1. Learn the names, formulas, and charges for all of the polyatomic ions in **Table 4.4**.

EXAMPLE 4.7**Naming Compounds That Contain Polyatomic Ions**

Give the systematic name of each of the following compounds.

- Na_2SO_4
- KH_2PO_4
- $\text{Fe}(\text{NO}_3)_3$
- $\text{Mn}(\text{OH})_2$
- Na_2SO_3

Solution

<i>Compound</i>	<i>Ions Present</i>	<i>Ion Names</i>	<i>Compound Name</i>
a. Na_2SO_4	two Na^+ SO_4^{2-}	sodium sulfate	sodium sulfate
b. KH_2PO_4	K^+ H_2PO_4^-	potassium dihydrogen phosphate	potassium dihydrogen phosphate
c. $\text{Fe}(\text{NO}_3)_3$	Fe^{3+} three NO_3^-	iron(III) nitrate	iron(III) nitrate
d. $\text{Mn}(\text{OH})_2$	Mn^{2+} two OH^-	manganese(II) hydroxide	manganese(II) hydroxide
e. Na_2SO_3	two Na^+ SO_3^{2-}	sodium sulfite	sodium sulfite

Information

Certain transition metals form only one ion. Common examples are zinc (forms only Zn^{2+}) and silver (forms only Ag^+). For these cases the Roman numeral is omitted from the name.

Practice Problem • Exercise 4.7

Name each of the following compounds.

- $\text{Ca}(\text{OH})_2$
- Na_3PO_4
- KMnO_4
- $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$
- $\text{Co}(\text{ClO}_4)_2$
- KClO_3
- $\text{Cu}(\text{NO}_2)_2$

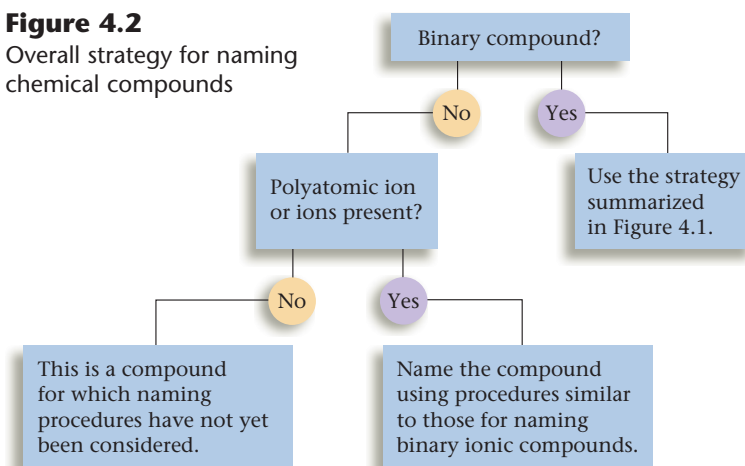
Example 4.7 illustrates that when more than one polyatomic ion appears in a chemical formula, parentheses are used to enclose the ion and a subscript is written after the closing parenthesis. Other examples are $(\text{NH}_4)_2\text{SO}_4$ and $\text{Fe}_3(\text{PO}_4)_2$.

In naming chemical compounds, use the strategy summarized in **Figure 4.2**.

Active Reading Question

How is naming compounds that contain polyatomic ions similar to naming binary ionic compounds? How is it different?

Figure 4.2
Overall strategy for naming
chemical compounds



Let's Review

- If the compound being considered is binary, use the procedure summarized in Figure 4.1.
- If the compound has more than two elements, determine whether it has any polyatomic ions.
- Use Table 4.4 to help you recognize these ions until you have committed them to memory.
- If a polyatomic ion is present, name the compound using procedures very similar to those for naming binary ionic compounds.

EXAMPLE 4.8

Summary of Naming Binary Compounds and Compounds That Contain Polyatomic Ions

Name the following compounds.

- a. Na_2CO_3 c. PCl_3 e. CuSO_4
 b. FeBr_3 d. CsClO_4

Solution

Compound	Name	Comments
a. Na_2CO_3	sodium carbonate	Contains 2Na^+ and CO_3^{2-} .
b. FeBr_3	iron(III) bromide	Contains Fe^{3+} and 3Br^- .
c. CsClO_4	cesium perchlorate	Contains Cs^+ and ClO_4^- .
d. PCl_3	phosphorus trichloride	Type III binary compound (both P and Cl are nonmetals).
e. CuSO_4	copper(II) sulfate	Contains Cu^{2+} and SO_4^{2-} .

Practice Problem • Exercise 4.8

Name the following compounds.

- a. NaHCO_3 c. CsClO_4 e. NaBr g. $\text{Zn}_3(\text{PO}_4)_2$
 b. BaSO_4 d. BrF_5 f. KOCl

B. Naming Acids

When dissolved in water, certain molecules produce H^+ ions (protons). These substances, which are called **acids**, were first recognized by the sour taste of their solutions. For example, citric acid is responsible for the tartness of lemons and limes. Acids will be discussed in detail later. Here we simply present the rules for naming acids.

An acid can be viewed as a molecule with one or more H^+ ions attached to an anion. The rules for naming acids depend on whether the anion contains oxygen.

Acid

A substance that produces hydrogen ions (H^+) in solution; a proton donor

Rules for Naming Acids

1. If the *anion does not contain oxygen*, the acid is named with the prefix *hydro-* and the suffix *-ic* attached to the root name for the element. For example, when gaseous HCl , HCN , and H_2S are dissolved in water, they form the following acids:

<i>Acid</i>	<i>Anion</i>	<i>Name</i>
HCl	Cl^-	hydrochloric acid
HCN	CN^-	hydrocyanic acid
H_2S	S^{2-}	hydrosulfuric acid

2. When the *anion contains oxygen*, the acid name is formed from the root name of the central element of the anion or the anion name, with a suffix of *-ic* or *-ous*. When the anion name ends in *-ate*, the suffix *-ic* is used. For example,

<i>Acid</i>	<i>Anion</i>	<i>Name</i>
H_2SO_4	SO_4^{2-} (sulfate)	sulfuric acid
H_3PO_4	PO_4^{3-} (phosphate)	phosphoric acid
$\text{HC}_2\text{H}_3\text{O}_2$	$\text{C}_2\text{H}_3\text{O}_2^-$ (acetate)	acetic acid

When the anion name ends in *-ite*, the suffix *-ous* is used in the acid name. For example,

<i>Acid</i>	<i>Anion</i>	<i>Name</i>
H_2SO_3	SO_3^{2-} (sulfite)	sulfurous acid
HNO_2	NO_2^- (nitrite)	nitrous acid

The application of rule 2 can be seen in the names of the acids of the oxyanions of chlorine, as shown below.

<i>Acid</i>	<i>Anion</i>	<i>Name</i>
HClO_4	perchlorate	perchloric acid
HClO_3	chlorate	chloric acid
HClO_2	chlorite	chlorous acid
HClO	hypochlorite	hypochlorous acid

The rules for naming acids are given in schematic form in **Figure 4.3**. The names of the most important acids are given in **Table 4.5** and **Table 4.6**. These names should be memorized.

Table 4.5	
Names of Acids That Do Not Contain Oxygen	
Acid	Name
HF	hydrofluoric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydroiodic acid
HCN	hydrocyanic acid
H ₂ S	hydrosulfuric acid

Table 4.6	
Names of Some Oxygen-Containing Acids	
Acid	Name
HNO ₃	nitric acid
HNO ₂	nitrous acid
H ₂ SO ₄	sulfuric acid
H ₂ SO ₃	sulfurous acid
H ₃ PO ₄	phosphoric acid
HC ₂ H ₃ O ₂	acetic acid

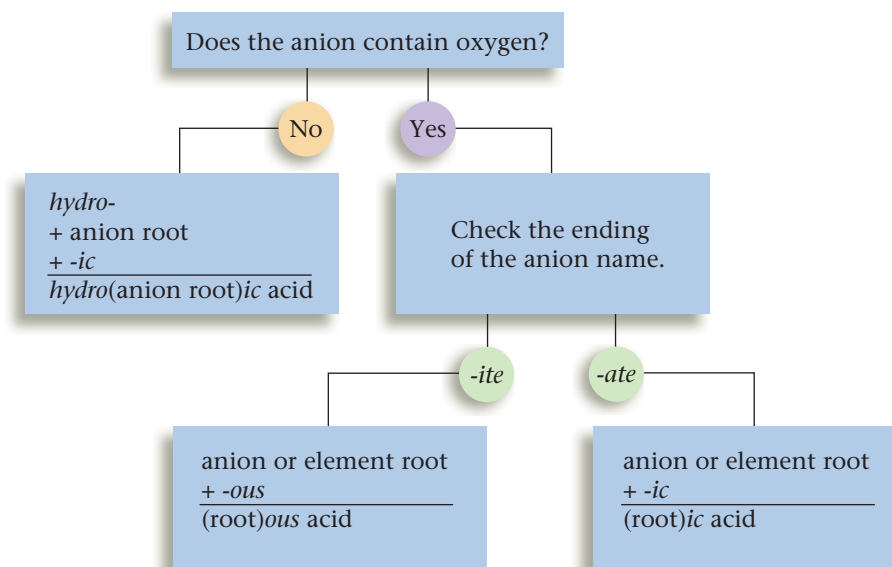


Figure 4.3
A flow chart for naming acids. The acid is pictured as one or more H⁺ ions attached to an anion.

C. Writing Formulas from Names

So far we have started with the chemical formula of a compound and decided on its systematic name. Being able to reverse the process is also important. Often a laboratory procedure describes a compound by name, but the label on the bottle in the lab shows only the formula of the chemical it contains. It is essential that you are able to determine the formula of a compound from its name. In fact, you already know enough about compounds to do this.

For example, given the name calcium hydroxide, you can write the formula as $\text{Ca}(\text{OH})_2$ because you know that calcium forms only Ca^{2+} ions and that, since hydroxide is OH^- , two of these anions are required to give a neutral compound. Similarly, the name iron(II) oxide implies the formula FeO , because the Roman numeral II indicates the presence of the cation Fe^{2+} and the oxide ion is O^{2-} .

We emphasize at this point that it is essential to learn the name, composition, and charge of each of the common polyatomic anions (and the NH_4^+ cation). If you do not recognize these ions by formula and by name, you will not be able to write the compound's name when given its formula or the compound's formula when given its name. You must also learn the names of the common acids.

EXAMPLE 4.9

Writing Formulas from Names

Give the formula for each of the following compounds.

- | | |
|------------------------|-------------------------|
| a. potassium hydroxide | e. calcium chloride |
| b. sodium carbonate | f. lead(IV) oxide |
| c. nitric acid | g. dinitrogen pentoxide |
| d. cobalt(III) nitrate | h. ammonium perchlorate |

Solution

<i>Name</i>	<i>Formula</i>	<i>Comments</i>
a. potassium hydroxide	KOH	Contains K^+ and OH^- .
b. sodium carbonate	Na_2CO_3	We need two Na^+ to balance CO_3^{2-} .
c. nitric acid	HNO_3	Common strong acid; memorize.
d. cobalt(III) nitrate	$\text{Co}(\text{NO}_3)_3$	Cobalt(III) means Co^{3+} ; we need three NO_3^- to balance Co^{3+} .
e. calcium chloride	CaCl_2	We need two Cl^- to balance Ca^{2+} ; Ca (Group 2) always forms Ca^{2+} .
f. lead(IV) oxide	PbO_2	Lead(IV) means Pb^{4+} ; we need two O^{2-} to balance Pb^{4+} .
g. dinitrogen pentoxide	N_2O_5	<i>di-</i> means two; <i>pent(a)-</i> means five.
h. ammonium perchlorate	NH_4ClO_4	Contains NH_4^+ and ClO_4^- .

Practice Problem • Exercise 4.9

Write the formula for each of the following compounds.

- | | |
|-------------------------|----------------------|
| a. ammonium sulfate | d. rubidium peroxide |
| b. vanadium(V) fluoride | e. aluminum oxide |
| c. disulfur dichloride | |

Talking Tadpoles

It is well known that animals communicate by releasing chemicals that others of the same species receive and “understand.” For example, ants use chemicals to signal news about food supplies and danger from predators, and honeybees “recognize” other bees from the same hive by their chemical signals. Now scientists at Yale University have shown that tadpoles send chemical signals to one another.

The experiment involved two groups of tadpoles of red-legged frogs. They were placed in an aquarium partitioned by a screen that allowed water to flow but blocked communications by sight and sound. When a wooden heron (a bird that preys on tadpoles) threatened the tadpoles in the one compartment, those on the other side moved away from the partition and ducked under a shelter. The

researchers concluded that the frightened tadpoles signaled their fear to the tadpoles in the other compartment by releasing a chemical signal that flowed to the other side in the water.

Other water-dwelling animals such as crayfish, hermit crabs, and a fish called the Iowa darter also have been observed to send chemical danger signals when they are threatened. Scientists think that the chemical used to send the signal is the ammonium ion, NH_4^+ . Researchers are now trying to find out whether all aquatic species have a common chemical “language” to signal one another, just as ants and bees do.

A California red-legged frog tadpole



SECTION 4.2 REVIEW QUESTIONS

- What patterns can you see in Table 4.4 that reduce the number of polyatomic ions you need to memorize? (Hint: Try grouping the ions by common elements or suffixes.)
- Once you have identified that a compound contains a polyatomic ion, how is naming it different from naming a binary compound?
- When writing a formula from a chemical name, how can you tell how many of each element or polyatomic ion to put in the formula?
- How can you tell whether a compound should be named an acid?
- Use the flow chart in Figure 4.4 to name the following:
 - HClO_4
 - HNO_2
 - HBr
 - H_2SO_4

- Copy and complete the following table.

Name	Formula
	NaNO_3
potassium sulfite	
	NaHCO_3
aluminum sulfate	
	KClO_4
copper(II) carbonate	
	$\text{NaC}_2\text{H}_3\text{O}_2$
acetic acid	
	CuSO_4
calcium hydroxide	
	NiCO_3
ammonium nitrate	

Chapter 4 Review

Key Terms

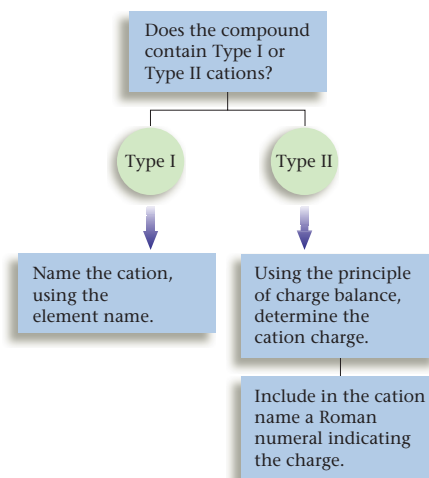
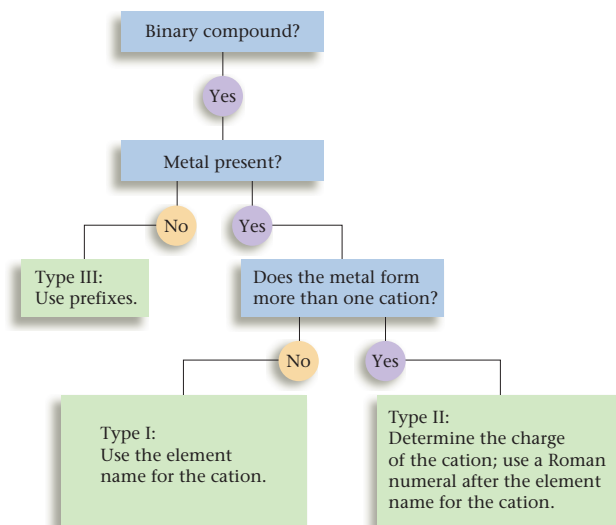
4.1 Binary compounds

Binary ionic compound

Key Ideas

4.1 Naming Binary Compounds

- Binary compounds are named by following a set of rules.
 - For compounds containing both a metal and a nonmetal, the metal is always named first. The nonmetal is named from the root element name.
 - If the metal ion can have more than one charge (Type II), a Roman numeral is used to specify the charge.
 - For binary compounds containing only nonmetals (Type III), prefixes are used to specify the numbers of atoms present.



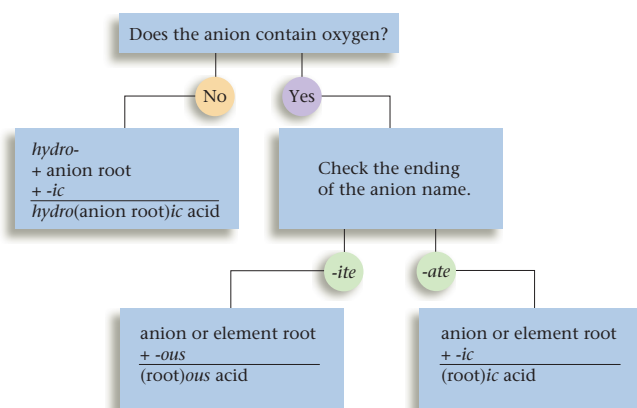
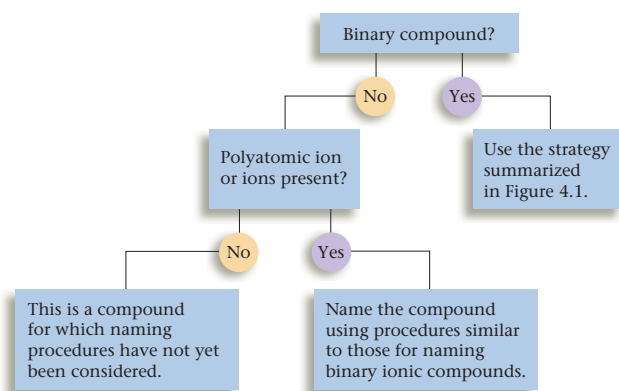
4.2 Polyatomic ions

Oxyanions

Acids

4.2 Naming and Writing Formulas for More Complex Compounds

- Polyatomic ions are charged entities composed of several atoms bound together. They have special names and must be memorized.
- Naming ionic compounds containing polyatomic ions follows rules similar to those for naming binary compounds.
- The names of acids (molecules with one or more H^+ ions attached to an anion) depend on whether the acid contains oxygen.



All exercises with **blue numbers** have answers in the back of this book.

4.1 Naming Binary Compounds

A. Naming Compounds That Contain a Metal and a Nonmetal

- What is a *binary* compound? Give three examples of binary compounds.
- What are the two major types of binary compounds?
- In an ionic compound, the positive ion is called the _____ and the negative ion is called the _____.
- In naming ionic compounds, we always name the _____ first.
- Although we write the formula of sodium chloride as NaCl, we realize that NaCl is an ionic compound and contains no molecules. Explain.
- Give the name of each of the following simple binary ionic compounds.
 - NaI
 - CaF₂
 - Al₂S₃
 - CaBr₂
 - SrO
 - AgCl
 - CsI
 - Li₂O
- Identify each case in which the name is incorrect. Give the correct name.
 - BaH₂, barium hydroxide
 - Na₂O, disodium oxide
 - SnCl₄, tin(IV) chloride
 - SiO₂, silver dioxide
 - FeBr₃, iron(III) bromide
- Write the name of each of the following ionic substances, using the system that includes a Roman numeral to specify the charge of the cation.
 - FeI₃
 - MnCl₂
 - HgO
 - Cu₂O
 - CuO
 - SnBr₄
- Write the name of each of the following ionic substances, using *-ous* or *-ic* endings to indicate the charge of the cation.
 - CoCl₂
 - CrBr₃
 - PbO
 - SnO₂
 - Fe₂O₃
 - FeCl₃

B. Naming Binary Compounds That Contain Only Nonmetals (Type III)

- Name each of the following binary compounds of nonmetallic elements.
 - IF₅
 - AsCl₃
 - SeO
 - XeF₄
 - Nl₃
 - B₂O₃
- Write the name for each of the following binary compounds of nonmetallic elements.
 - GeH₄
 - N₂Br₄
 - P₂S₅
 - SeO₂
 - NH₃
 - SiO₂

C. Naming Binary Compounds: A Review

- Name each of the following binary compounds, using the periodic table to determine whether the compound is likely to be ionic (containing a metal and a nonmetal) or nonionic (containing only nonmetals).
 - B₂H₆
 - Ca₃N₂
 - CBr₄
 - Ag₂S
 - CuCl₂
 - ClF
- Name each of the following binary compounds, using the periodic table to determine whether the compound is likely to be ionic (containing a metal and a nonmetal) or nonionic (containing only nonmetals).
 - RaCl₂
 - SeCl₂
 - PCl₃
 - Na₃P
 - MnF₂
 - ZnO

4.2 Naming and Writing Formulas for More Complex Compounds

A. Naming Compounds That Contain Polyatomic Ions

- What is an *oxyanion*? How is a series of oxyanions named? (Give an example.)
- Which oxyanion of chlorine contains the most oxygen atoms: hypochlorite, chlorite, or perchlorate?
- Copy and complete the table by filling in the missing names or formulas of the indicated oxyanions.

BrO ⁻	
	iodate
IO ₄ ⁻	
	hypoiodite

17. Write the formula for each of the following nitrogen-containing polyatomic ions, including the overall charge of the ion.

- a. nitrate
- b. nitrite
- c. ammonium
- d. cyanide

18. Write the formula for each of the following chlorine-containing ions, including the overall charge of the ion.

- a. chloride
- b. hypochlorite
- c. chlorate
- d. perchlorate

19. Write the formulas of the compounds below (refer to your answers to problem 18).

- a. magnesium chloride
- b. calcium hypochlorite
- c. potassium chlorate
- d. barium perchlorate

20. Give the name of each of the following polyatomic anions.

- a. MnO_4^-
- b. O_2^{2-}
- c. CrO_4^{2-}
- d. $\text{Cr}_2\text{O}_7^{2-}$
- e. NO_3^-
- f. SO_3^{2-}

21. Name each of the following compounds, which contain polyatomic ions.

- a. $\text{Fe}(\text{NO}_3)_3$
- b. $\text{Co}_3(\text{PO}_4)_2$
- c. $\text{Cr}(\text{CN})_3$
- d. $\text{Al}_2(\text{SO}_4)_3$
- e. $\text{Cr}(\text{C}_2\text{H}_3\text{O}_2)_2$
- f. $(\text{NH}_4)_2\text{SO}_3$

B. Naming Acids

22. Give a simple definition of an *acid*.

23. Many acids contain the element _____ in addition to hydrogen.

24. Name each of the following acids.

- a. HCl
- b. H_2SO_4
- c. HNO_3
- d. HI
- e. HNO_2
- f. HClO_3
- g. HBr
- h. HF
- i. $\text{HC}_2\text{H}_3\text{O}_2$

C. Writing Formulas from Names

25. Write formulas for each of the following simple binary ionic compounds.

- a. lithium oxide
- b. aluminum iodide
- c. silver oxide
- d. potassium nitride
- e. calcium phosphide
- f. magnesium fluoride
- g. sodium sulfide
- h. barium hydride

26. Write the formula for each of the following binary compounds of nonmetallic elements.

- a. carbon dioxide
- b. sulfur dioxide
- c. dinitrogen tetrachloride
- d. carbon tetraiodide
- e. phosphorus pentafluoride
- f. diphosphorus pentoxide

27. Write the formula for each of the following compounds that contain polyatomic ions. Be sure to enclose the polyatomic ion in parentheses if more than one such ion is needed.

- a. calcium phosphate
- b. ammonium nitrate
- c. aluminum hydrogen sulfate
- d. barium sulfate
- e. iron(III) nitrate
- f. copper(I) hydroxide

28. Write the formula for each of the following acids.

- a. hydrocyanic acid
- b. nitric acid
- c. sulfuric acid
- d. phosphoric acid
- e. hypochlorous acid
- f. hydrobromic acid
- g. bromous acid
- h. hydrofluoric acid

29. Write the formula for each of the following substances.

- a. lithium chloride
- b. cuprous carbonate
- c. hydrobromic acid
- d. calcium nitrate
- e. sodium perchlorate
- f. aluminum hydroxide
- g. barium hydrogen carbonate
- h. iron(II) sulfate
- i. diboron hexachloride
- j. phosphorus pentabromide
- k. potassium sulfite
- l. barium acetate

Critical Thinking

30. Before an electrocardiogram (ECG) is recorded for a cardiac patient, the ECG leads are usually coated with a moist paste containing sodium chloride. What property of an ionic substance such as NaCl is being made use of?

31. What is a *binary* compound? What is a *polyatomic* anion? What is an *oxyanion*?

32. Name the following compounds.

- AuBr₃
- Co(CN)₃
- MgHPO₄
- B₂H₆
- NH₃
- Ag₂SO₄
- Be(OH)₂

33. Name the following compounds.

- (NH₄)₂CO₃
- NH₄HCO₃
- Ca₃(PO₄)₂
- H₂SO₃
- MnO₂
- HIO₃
- KH

34. Most metallic elements form *oxides*, and often the oxide is the most common compound of the element that is found in the earth's crust. Write the formulas for the oxides of the following metallic elements.

- potassium
- magnesium
- iron(II)
- iron(III)
- zinc
- lead(II)
- aluminum

35. Consider the hypothetical metallic element M, which is capable of forming stable simple cations that have charges of 1+, 2+, and 3+, respectively. Consider also the nonmetallic elements D, E, and F, which form anions that have charges of 1-, 2-, and 3-, respectively. Write the formulas of all possible compounds between metal M and nonmetals D, E, and F.

36. Copy and complete Table 4.A (below) by writing the names and formulas for the ionic compounds formed when the cations listed across the top combine with the anions shown in the left column.

TABLE 4.A

Ions	Fe ²⁺	Al ³⁺	Na ⁺	Ca ²⁺	NH ₄ ⁺	Fe ³⁺	Ni ²⁺	Hg ₂ ²⁺	Hg ²⁺
CO ₃ ²⁻									
BrO ₃ ⁻									
C ₂ H ₃ O ₂ ⁻									
OH ⁻									
HCO ₃ ⁻									
PO ₄ ³⁻									
SO ₃ ²⁻									
ClO ₄ ⁻									
SO ₄ ²⁻									
O ²⁻									
Cl ⁻									

37. Copy and complete Table 4.B (below) by writing the formulas for the ionic compounds formed when the anions listed across the top combine with the cations shown in the left column.

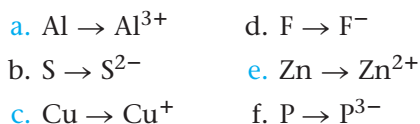
TABLE 4.B

ions	nitrate	sulfate	hydrogen sulfate	dihydrogen phosphate	oxide	chloride
calcium						
strontium						
ammonium						
aluminum						
iron(III)						
nickel(II)						
silver(I)						
gold(III)						
potassium						
mercury(II)						
barium						

38. Copy and complete the following table. For each of the negative ions listed below, use the periodic table to find the total number of electrons that the ion contains.

Negative Ion	Total number of electrons
Se ²⁻	
S ²⁻	
P ³⁻	
O ²⁻	
N ³⁻	
I ⁻	
F ⁻	
Cl ⁻	
Br ⁻	
At ⁻	

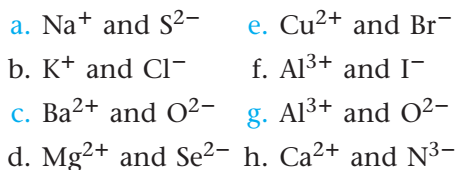
39. For each of the following processes that show the formation of ions, complete the process by indicating the number of electrons that must be gained or lost to form the ion. Indicate the total number of electrons in the ion, and in the atom from which it was made.



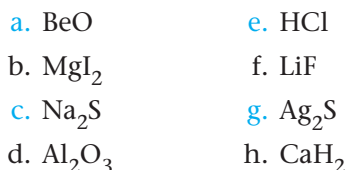
40. For each of the following atomic numbers, use the periodic table to write the formula (including the charge) for the simple *ion* that the element is most likely to form.



41. For the following pairs of ions, use the principle of electrical neutrality to predict the formula of the binary compound that the ions are most likely to form.



42. Give the name of each of the following simple binary ionic compounds.



- 43.** In which of the following pairs is the name incorrect?
- SiI_4 , silver iodide
 - CoCl_2 , copper(II) chloride
 - CaH_2 , hydrocalcic acid
 - $\text{Zn}(\text{C}_2\text{H}_3\text{O}_2)_2$, zinc acetate
 - PH_3 , phosphoric trihydride
- 44.** Write the name of each of the following ionic substances, using the system that includes a Roman numeral to specify the charge of the cation.
- FeBr_2
 - CoS
 - Co_2S_3
 - SnO_2
 - Hg_2Cl_2
 - HgCl_2
- 45.** Name each of the following binary compounds.
- XeF_6
 - OF_2
 - AsI_3
 - N_2O_4
 - Cl_2O
 - SF_6
- 46.** Name each of the following compounds.
- $\text{Fe}(\text{C}_2\text{H}_3\text{O}_2)_3$
 - BrF
 - K_2O_2
 - SiBr_4
 - $\text{Cu}(\text{MnO}_4)_2$
 - CaCrO_4
- 47.** Name each of the following compounds, which contain polyatomic ions.
- LiH_2PO_4
 - $\text{Cu}(\text{CN})_2$
 - $\text{Pb}(\text{NO}_3)_2$
 - Na_2HPO_4
 - NaClO_2
 - $\text{Co}_2(\text{SO}_4)_3$
- 48.** Write the formula for each of the following simple binary ionic compounds.
- calcium chloride
 - silver(I) oxide (usually called silver oxide)
 - aluminum sulfide
 - beryllium bromide
 - hydrosulfuric acid
 - potassium hydride
 - magnesium iodide
 - cesium fluoride
- 49.** Write the formula for each of the following binary compounds of nonmetallic elements.
- sulfur dioxide
 - dinitrogen monoxide
 - xenon tetrafluoride
 - tetraphosphorus decoxide
 - phosphorus pentachloride
 - sulfur hexafluoride
 - nitrogen dioxide
- 50.** Write the formula for each of the following ionic substances.
- sodium dihydrogen phosphate
 - lithium perchlorate
 - copper(II) hydrogen carbonate
 - potassium acetate
 - barium peroxide
 - cesium sulfite
- 51.** Write the formula for each of the following compounds, which contain polyatomic ions. Be sure to enclose the polyatomic ion in parentheses if more than one such ion is needed to balance the oppositely charged ion(s).
- silver(I) perchlorate (usually called silver perchlorate)
 - cobalt(III) hydroxide
 - sodium hypochlorite
 - potassium dichromate
 - ammonium nitrite
 - iron(III) hydroxide
 - ammonium hydrogen carbonate
 - potassium perbromate

- 1** How many of the following compounds are named correctly?
- Na_2SO_4 sodium sulfate
 - CaCO_3 calcium carbonate
 - PCl_5 phosphorus pentachloride
 - CoO cobalt(II) oxide
- 4
 - 3
 - 2
 - 1
- 2** Which of the following chemical formulas is incorrect?
- FeS iron(II) sulfide
 - SF_6 sulfur hexafluoride
 - PbO_2 lead(IV) oxide
 - NH_4SO_4 ammonium sulfate
- 3** Which of the following compounds is named incorrectly?
- PCl_3 phosphorus trichloride
 - MgCl_2 magnesium dichloride
 - CuO copper(II) oxide
 - At least two of the above (A–C) are named incorrectly.
- 4** Which of the following names is incorrect?
- cobalt(II) chloride
 - magnesium oxide
 - aluminum(III) oxide
 - diphosphorus pentoxide
- 5** In which case is the polyatomic ion named incorrectly?
- NH_4^+ ammonium
 - SO_3^{2-} sulfate
 - CO_3^{2-} carbonate
 - NO_2^- nitrite
- 6** What is the correct name for the compound Fe_2O_3 ?
- iron oxide
 - iron(III) oxide
 - iron(II) oxide
 - diiron trioxide
- 7** Which of the following is the correct formula for chlorous acid?
- HClO
 - HClO_2
 - HClO_3
 - HClO_4
- 8** When dihydrogen sulfide is dissolved in water, it is called,
- sulfuric acid.
 - sulfurous acid.
 - hydrosulfuric acid.
 - dihydrosulfuric acid.
- 9** Which of the following acids is named incorrectly?
- HNO_3 nitrous acid
 - HCl hydrochloric acid
 - $\text{HC}_2\text{H}_3\text{O}_2$ acetic acid
 - H_3PO_4 phosphoric acid
- 10** Fill in the missing name or formula
- | Name | Formula |
|-------------------|-----------------------|
| barium phosphate | |
| | N_2O |
| sulfuric acid | |
| | Cu_2O |
| boron trifluoride | |
- 11** In naming binary compounds, you sometimes simply name the ions. In other cases, you need to use Roman numerals, or prefixes.
- Explain why Roman numerals are sometimes needed in naming binary compounds, and provide an example name and formula.
 - Explain why prefixes are sometimes needed in naming binary compounds, and provide an example name and formula.

Measurements and Calculations

LOOKING AHEAD

5.1 Scientific Notation and Units

- A. Scientific Notation
- B. Units
- C. Measurements of Length, Volume, and Mass

5.2 Uncertainty in Measurement and Significant Figures

- A. Uncertainty in Measurement
- B. Significant Figures
 - Rounding Off Numbers
 - Determining Significant Figures in Calculations

5.3 Problem Solving and Unit Conversions

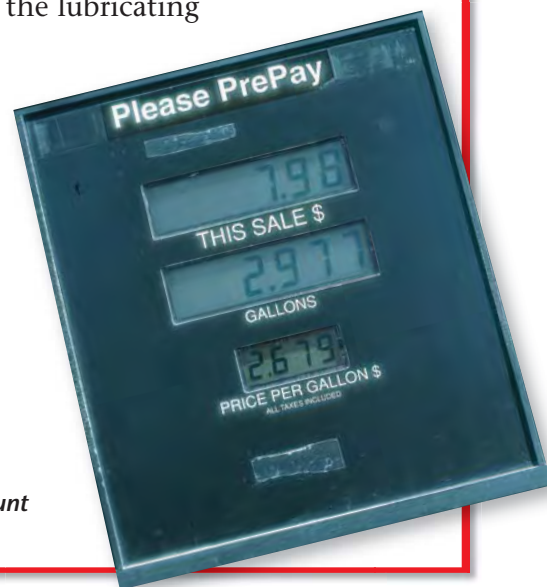
- A. Tools for Problem Solving
 - Problem-Solving Strategy
- B. Temperature Conversions
 - Converting Between the Kelvin and Celsius Scales
 - Converting Between the Fahrenheit and Celsius Scales
- C. Density

There are a variety of measuring devices that you can use to make a quantitative observation.

IN YOUR LIFE

As we pointed out in Chapter 1, making observations is a key part of the scientific process. Sometimes observations are *qualitative* (“the substance is a yellow solid”) and sometimes they are *quantitative* (“the substance weighs 4.3 grams”). A quantitative observation is called a **measurement**. Measurements are very important in our daily lives. For example, we pay for gasoline by the gallon, so the gas pump must accurately measure the gas delivered to our fuel tank. The efficiency of the modern automobile engine depends on various measurements, including the amount of oxygen in the exhaust gases, the temperature of the coolant, and the pressure of the lubricating oil. In addition, cars with traction control systems have devices to measure and compare the rates of rotation of all four wheels.

As we will see in the “Chemistry in Your World” discussion in this chapter, measuring devices have become very sophisticated in dealing with our fast-moving and complicated society.



A gas pump measures the amount of gasoline delivered.

WHAT DO YOU KNOW?



Prereading Questions

1. What are some units that are used in the metric system?
2. A donut shop sells donuts only by the dozen. If you have to buy donuts for 63 people and want each person to get at least one donut, how many dozen do you need to buy?
3. How do you think the word *quantity* relates to the word *quantitative*?

SECTION 5.1

Scientific Notation and Units

Key Terms

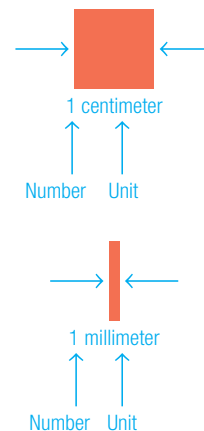
- Measurement
- Scientific notation
- Units
- English system
- Metric system
- International System (SI)
- Volume
- Liter
- Milliliter
- Mass
- Gram

Objectives

- To show how very large or very small numbers can be expressed in scientific notation
- To learn the English, metric, and SI systems of measurement
- To use the metric system to measure length, volume, and mass

A measurement always consists of two parts: a number and a unit. Both parts are necessary to make the measurement meaningful. For example, suppose that a friend tells you that she saw a bug 5 long. This statement is meaningless as it stands. Five what? If it's 5 millimeters, the bug is quite small. If it's 5 centimeters, the bug is quite large. If it's 5 meters, run for cover!

The point is that for a measurement to be meaningful, it must consist of both a number and a unit that tells us the scale being used.



Active Reading Question

What is the difference between a qualitative observation and a quantitative observation?

A. Scientific Notation

The numbers associated with scientific measurements are often very large or very small. For example, the distance from Earth to the sun is approximately 93,000,000 (93 million) miles. Written out, this number is rather bulky. Scientific notation is a method for making very large or very small numbers more compact and easier to write.

To see how this is done, consider the number 125, which can be written as the product:

$$125 = 1.25 \times 100$$

Because $100 = 10 \times 10 = 10^2$, we can write:

$$125 = 1.25 \times 100 = 1.25 \times 10^2$$

Similarly, the number 1700 can be written:

$$1700 = 1.7 \times 1000$$

Because $1000 = 10 \times 10 \times 10 = 10^3$, we can write:

$$1700 = 1.7 \times 1000 = 1.7 \times 10^3$$

Scientific notation expresses a number as a product of a number between 1 and 10 and the appropriate power of 10.

Representing Large Numbers The distance from Earth to the sun is about 93 million miles. The number 93,000,000 can be expressed as:

$$\begin{aligned} 93,000,000 &= 9.3 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 \times 10 \\ &= 9.3 \times 10^7 \end{aligned}$$

Number between 1 and 10 Appropriate power of 10
($10,000,000 = 10^7$)

MATH



When writing a number in scientific notation, keep one digit to the left of the decimal point.

The easiest way to determine the appropriate power of 10 for scientific notation is to start with the number being represented and count the number of places the decimal point must be moved to obtain a number between 1 and 10. For example, for the number

$$\begin{array}{cccccccc} 9 & 3 & 0 & 0 & 0 & 0 & 0 & 0 \\ \underbrace{\hspace{1.5cm}} & & & & & & & \\ 7 & 6 & 5 & 4 & 3 & 2 & 1 & \end{array}$$

we must move the decimal point seven places to the left to get 9.3 (a number between 1 and 10). To compensate for every move of the decimal point to the left, we must multiply by 10. That is, each time we move the decimal point to the left, we make the number smaller by one power of 10. So for each move of the decimal point to the left, we must multiply by 10 to restore the number to its original magnitude. Moving the decimal point seven places to the left means that we must multiply 9.3 by 10 seven times, which equals 10^7 :

$$93,000,000 = 9.3 \times 10^7 \quad \text{We moved the decimal point seven places to the left, so we need to multiply by } 10^7.$$

Remember: when the decimal point is moved to the *left*, the exponent of 10 is *positive*.

Representing Small Numbers We can represent numbers smaller than 1 by using the same convention, but in this case the power of 10 is negative. For example, for the number 0.010 we must move the decimal point two places to the right to obtain a number between 1 and 10:

$$\begin{array}{cccc} 0 & . & 0 & 1 & 0 \\ \underbrace{\hspace{1.5cm}} & & & & \\ 1 & 2 & & & \end{array}$$

This requires an exponent of -2 , so $0.010 = 1.0 \times 10^{-2}$. Remember: when the decimal point is moved to the *right*, the exponent of 10 is *negative*.

Next consider the number 0.000167. In this case we must move the decimal point four places to the right to obtain 1.67 (a number between 1 and 10):

$$\begin{array}{ccccccc} 0 & . & 0 & 0 & 0 & 1 & 6 & 7 \\ \underbrace{\hspace{1.5cm}} & & & & & & & \\ 1 & 2 & 3 & 4 & & & & \end{array}$$

Moving the decimal point four places to the right requires an exponent of -4 . Therefore,

$$0.000167 = 1.67 \times 10^{-4} \quad \text{We moved the decimal point four places to the right.}$$

Let's Review

Using Scientific Notation

- Any number can be represented as the product of a number between 1 and 10 and a power of 10 (either positive or negative).
- The power of 10 depends on the number of places the decimal point is moved and in which direction. The *number of places* the decimal point is moved determines the *power of 10*. The *direction* of the move determines whether the power of 10 is *positive* or *negative*. If the decimal point is moved to the left, the power of 10 is positive; if the decimal point is moved to the right, the power of 10 is negative.

MATH



Moving the decimal point to the *left* requires a *positive* exponent.

MATH



Moving the decimal point to the *right* requires a *negative* exponent.

Information

Turn to the Appendix if you need a further discussion of exponents and scientific notation.

Active Reading Question

What does it mean when a number written in scientific notation has a positive exponent? What does it mean when a number written in scientific notation has a negative exponent?

EXAMPLE 5.1

Scientific Notation: Powers of 10 (Positive)

Represent the following numbers in scientific notation.

- a. 238,000 b. 1,500,000

MATH



1	1	10^0
10	10	10^1
100	10×10	10^2
1000	$10 \times 10 \times 10$	10^3
10000	$10 \times 10 \times 10 \times 10$	10^4

Solution

- a. First we move the decimal point until we have a number between 1 and 10, in this case 2.38.

2 3 8 0 0 0

5 4 3 2 1

The decimal point was moved five places to the left.

Because we moved the decimal point five places to the left, the power of 10 is positive 5. So, $238,000 = 2.38 \times 10^5$.

- b. 1 5 0 0 0 0 0

6 5 4 3 2 1

The decimal point was moved six places to the left, so the power of 10 is 6.

So $1,500,000 = 1.5 \times 10^6$.

EXAMPLE 5.2

Scientific Notation: Powers of 10 (Negative)

Represent the following numbers in scientific notation.

- a. 0.00043 b. 0.089

MATH



1	1	10^0
0.1	$\frac{1}{10}$	10^{-1}
0.01	$\frac{1}{10 \times 10}$	10^{-2}
0.001	$\frac{1}{10 \times 10 \times 10}$	10^{-3}
0.0001	$\frac{1}{10 \times 10 \times 10 \times 10}$	10^{-4}

Solution

- a. First we move the decimal point until we have a number between 1 and 10, in this case 4.3.

0.00043

1 2 3 4

The decimal point was moved four places to the right.

Because we moved the decimal point four places to the right, the power of 10 is negative 4. Therefore, $0.00043 = 4.3 \times 10^{-4}$.

- b. 0.089 The power of 10 is negative 2 because the decimal point was moved two places to the right.

1 2

So $0.089 = 8.9 \times 10^{-2}$.

Practice Problem • Exercise 5.2

Write the numbers 357 and 0.0055 in scientific notation. If you are having difficulty with scientific notation, read the Appendix.

B. Units

The **units** part of a measurement tells us what *scale* or *standard* is being used to represent the results of the measurement. From the earliest days of civilization, trade has required common units. For example, if a farmer from one region wanted to trade grain for the gold of a miner who lived in another region, the two people had to have common standards (units) for measuring the amount of the grain and the weight of the gold.

The need for common units also applies to scientists, who measure quantities such as mass, length, time, and temperature. If every scientist had her or his own personal set of units, complete chaos would result. Unfortunately, although standard systems of units did arise, different systems were adopted in different parts of the world. The two most widely used systems are the **English system** used in the United States and the **metric system** used in most of the rest of the industrialized world.

Active Reading Question

Why is a unit a necessary part of a measurement?

The metric system has long been preferred for most scientific work. In 1960 an international agreement set up a comprehensive system of units called the **International System** (*le Système Internationale* in French), or **SI**. The SI units are based on the metric system and units derived from the metric system. The most important fundamental SI units are listed in **Table 5.1**. Later in this chapter we will discuss how to manipulate some of these units.

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Critical Units!

How important are conversions from one unit to another? If you ask the National Aeronautic and Space Administration (NASA), very important! In 1999, NASA lost a \$125 million Mars Climate Orbiter because of a failure to convert from English to metric units.

The problem arose because two teams working on the Mars mission were using different sets of units. NASA's scientists at the Jet Propulsion Laboratory in Pasadena, California, assumed that the thrust data for the rockets on the Orbiter that they received from Lockheed Martin Astronautics in Denver, which built the spacecraft, were in metric units. In reality, the units were English. As a result, the Orbiter dipped 100 kilometers lower into the Mars atmosphere than planned and the friction from the atmosphere caused the craft to burn up.

NASA's mistake refueled the controversy over whether Congress should require the United States to

switch to the metric system. About 95% of the world now uses the metric system, and the United States is slowly switching from English to metric. For example, the automobile industry has adopted metric fasteners and we buy our soda in two liter bottles.

Units can be very important. In fact, they can mean the difference between life and death on some occasions. In 1983, for example, a Canadian jetliner almost ran out of fuel when someone pumped 22,300 pounds of fuel into the aircraft instead of 22,300 kilograms. Remember to watch your units!

Artist's conception of the lost Mars Climate Orbiter

Unit

The part of the measurement telling us the scale being used

CRITICAL THINKING ?

What if you were not allowed to use units for one day? How would it affect your life for that day?

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TABLE 5.1

Common Fundamental SI Units

Physical Quality	Name of Unit	Abbreviation
mass	kilogram	kg
length	meter	m
time	second	s
temperature	kelvin	K

Because the fundamental units are not always a convenient size, the SI system uses prefixes to change the size of the unit. The most commonly used prefixes are listed in **Table 5.2**. Although the fundamental unit for length is the meter (m), we can also use the decimeter (dm), which represents one-tenth (0.1) of a meter; the centimeter (cm), which represents one one-hundredth (0.01) of a meter; the millimeter (mm), which represents one one-thousandth (0.001) of a meter; and so on. For example, it's much more convenient to specify the diameter of a certain contact lens as 1.0 cm than as 1.0×10^{-2} m.

TABLE 5.2

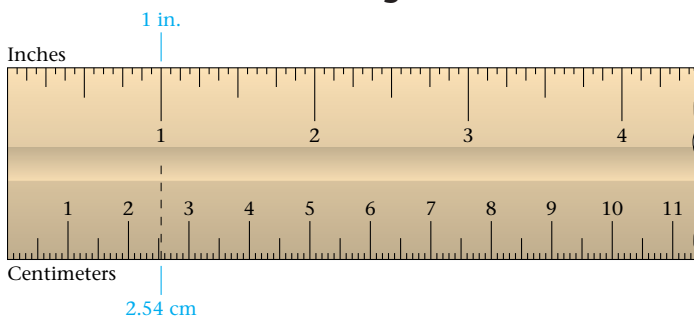
Commonly Used Prefixes in the Metric System

Prefix	Symbol	Meaning	Power of 10 for Scientific Notation
mega	M	1,000,000	10^6
kilo	k	1000	10^3
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}

C. Measurements of Length, Volume, and Mass

The fundamental SI unit of length is the meter, which is a little longer than a yard (1 meter = 39.37 inches). In the metric system fractions of a meter or multiples of a meter can be expressed by powers of 10, as summarized in **Table 5.3**.

The English and metric systems are compared on the ruler shown in **Figure 5.1**. Note that:



$$1 \text{ inch} = 2.54 \text{ centimeters}$$

Other English-metric equivalences are given in Section 5.3.

Figure 5.1

Comparison of English and metric units for length on a ruler

TABLE 5.3

Metric System for Measuring Length

Unit	Symbol	Meter Equivalent
kilometer	km	1000 m or 10^3 m
meter	m	1 m or 1 m
decimeter	dm	0.1 m or 10^{-1} m
centimeter	cm	0.01 m or 10^{-2} m
millimeter	mm	0.001 m or 10^{-3} m
micrometer	μm	0.000001 m or 10^{-6} m
nanometer	nm	0.000000001 m or 10^{-9} m

 Active Reading Question

How does a kilometer compare with a meter? How does a milliliter compare with a liter?

Volume is the amount of three-dimensional space occupied by a substance. The fundamental unit of volume in the SI system is based on the volume of a cube that measures 1 meter in each of the three directions. That is, each edge of the cube is 1 meter in length. The volume of this cube is:

$$1 \text{ m} \times 1 \text{ m} \times 1 \text{ m} = (1 \text{ m})^3 = 1 \text{ m}^3$$

or, in words, one cubic meter. In **Figure 5.2** a cube is divided into 1000 smaller cubes. Each of these small cubes represents a volume of 1 dm^3 , which is commonly called the **liter** (rhymes with “meter” and is slightly larger than a quart). Liter is abbreviated L.

The cube with a volume of 1 dm^3 (1 liter) can in turn be broken into 1000 smaller cubes, each representing a volume of 1 cm^3 . Each liter contains 1000 cm^3 . One cubic centimeter is called a **milliliter** (abbreviated mL), a unit of volume used very commonly in chemistry. This relationship is summarized in **Table 5.4**.

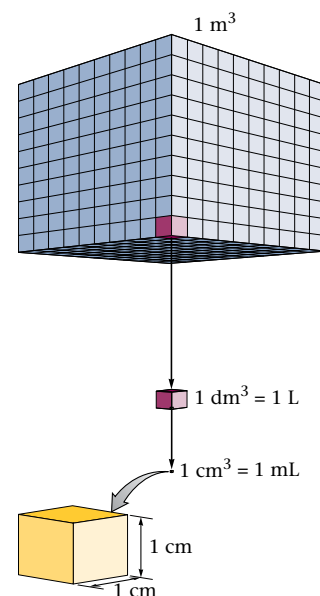


Figure 5.2

The largest drawing represents a cube that has sides 1 m in length and a volume of 1 m^3 . The middle-size cube has sides 1 dm in length and a volume of 1 dm^3 , or 1 L. The smallest cube has sides 1 cm in length and a volume of 1 cm^3 , or 1 mL.

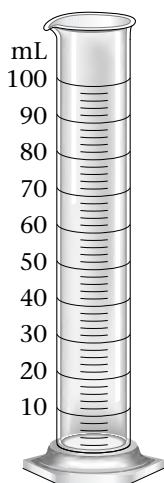


Figure 5.3
A 100-mL graduated cylinder

TABLE 5.4

Relationship Between the Liter and Milliliter

Unit	Symbol	Equivalence
liter	L	1 L = 1000 mL
milliliter	mL	$\frac{1}{1000}$ L = 10^{-3} L = 1 mL

The *graduated cylinder* (see **Figure 5.3**), which is commonly used in chemical laboratories for measuring the volumes of liquids, is marked off in convenient units of volume (usually milliliters). The graduated cylinder is filled to the desired volume with the liquid, which then can be poured out.

Mass

The quantity of matter in an object

Another important measurable quantity is **mass**, which can be defined as the quantity of matter present in an object. The fundamental SI unit of mass is the kilogram. Because the metric system, which existed before the SI system, used the gram as the fundamental unit, the prefixes for the various mass units are based on the **gram**, as shown in **Table 5.5**.

TABLE 5.5

Most Commonly Used Metric Units for Mass

Unit	Symbol	Gram Equivalent
kilogram	kg	$1000 = 10^3\text{g} = 1 \text{ kg}$
gram	g	1 g
milligram	mg	$0.001 \text{ g} = 10^{-3}\text{g} = 1 \text{ mg}$

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Figure 5.4

An electronic analytical balance used in chemistry labs

In the laboratory we determine the mass of an object by using a balance. A balance compares the mass of the object to a set of standard masses (weights). For example, the mass of an object can be determined by using a single-pan balance (see **Figure 5.4**).

To help you get a feeling for the common units of length, volume, and mass, some familiar objects are described in **Table 5.6**.

TABLE 5.6

Examples of Commonly Used Units

length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
mass	A nickel has a mass of about 5 g. A 120-lb woman has a mass of about 55 kg.
volume	A 12-oz can of soda has a volume of about 360 mL. A half gallon of milk is equal to about 2 L of milk.



SECTION 5.1

REVIEW QUESTIONS

- When writing numbers using scientific notation, numbers greater than one have _____ exponents and numbers less than one have _____ exponents.
- Change the following to scientific notation:
 - 8,475,000
 - 0.0000754
 - 1,000
 - 0.35724

- Copy and complete the table below.

	English unit	Metric unit
Length		meter
Volume	quart	
Mass	pound	

- How many millimeters are in 50 cm?
- Which is a larger, a liter or a quart?

Uncertainty in Measurement and Significant Figures

Objectives

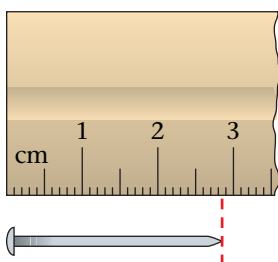
- To learn how uncertainty in a measurement arises
- To learn to indicate a measurement's uncertainty by using significant figures
- To learn to determine the number of significant figures in a calculated result

Key Terms

- Significant figures
- Rounding off

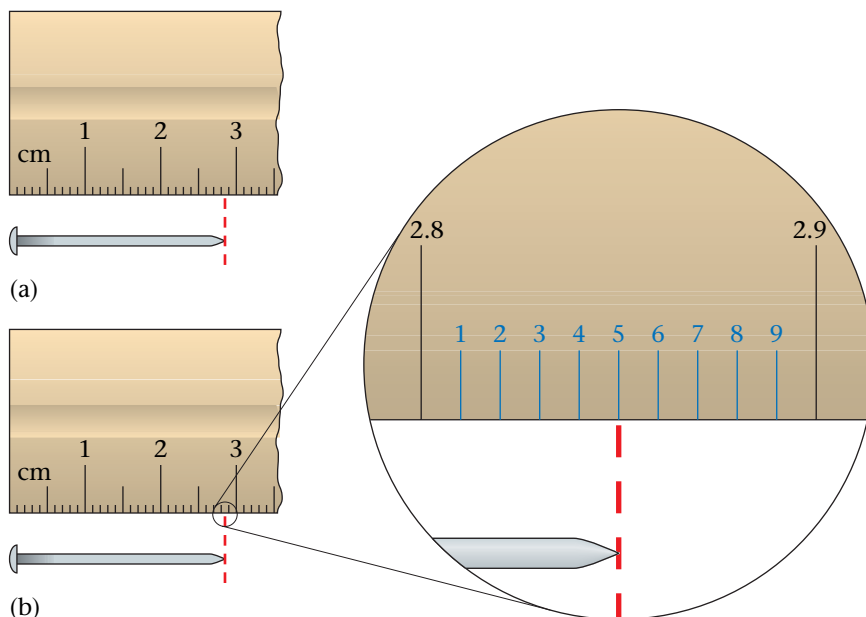
A. Uncertainty in Measurement

Whenever a measurement is made with a device such as a ruler or a graduated cylinder, an estimate is required. We can illustrate this by measuring the pin shown.



We can see from the ruler that the pin is a little longer than 2.8 cm and a little shorter than 2.9 cm. Because there are no graduations on the ruler between 2.8 and 2.9, we must estimate the pin's length between 2.8 and 2.9 cm. We do this by *imagining* that the distance between 2.8 and 2.9 is broken into 10 equal divisions **shown in blue** and estimating to which division the end of the pin reaches.

The end of the pin appears to come about halfway between 2.8 and 2.9, which corresponds to 5 of our 10 imaginary divisions. So we estimate the pin's length as 2.85 cm. The result of our measurement is that the pin is approximately 2.85 cm in length, but we had to rely on a visual estimate, so it might actually be 2.84 or 2.86 cm.



Because the last number is based on a visual estimate, it may be different when another person makes the same measurement. For example, if five different people measured the pin, the results might be

Person	Result of Measurement
1	2.85 cm
2	2.84 cm
3	2.86 cm
4	2.85 cm
5	2.86 cm

Note that the first two digits in each measurement are the same regardless of who made the measurement; these are called the *certain* numbers of the measurement. However, the third digit is estimated and can vary; it is called an *uncertain* number. When one is making a measurement, the custom is to record all of the certain numbers plus the *first* uncertain number. It would not make any sense to try to measure the pin to the third decimal place (thousandths of a centimeter), because this ruler requires an estimate of even the second decimal place (hundredths of a centimeter).

A measurement always has some degree of uncertainty.

The uncertainty of a measurement depends on the measuring device. For example, if our ruler had marks indicating hundredths of a centimeter, the uncertainty in the measurement of the pin would occur in the thousandths place rather than the hundredths place—but some uncertainty would still exist.

Active Reading Question

What do we mean when we say that a measurement always has a degree of uncertainty?



Students working on a lab project

CHEMISTRY EXPLORERS

Cecile Hoover Edwards • 1926–2005

Cecile Hoover Edwards was born in East St. Louis, Illinois, in 1926. She enrolled in college at the age of 15 and received a master's degree in chemistry from what is now Tuskegee University before she turned 21.

After receiving a doctorate in nutrition from Iowa State University in 1950, she returned to Tuskegee as a faculty member and researcher and worked there for the next 20 years.

In 1971 she joined the faculty of Howard University, served as a professor and dean, and established a doctoral program in nutrition. Much of her research focused on improving the nutritional value of the high-protein low-cost foods used by African-Americans eating a traditional southeastern American diet.

In 1984, Illinois declared April 5 “Dr. Cecile Hoover Edwards Day.”



The numbers recorded in a measurement (all the certain numbers plus the first uncertain number) are called **significant figures**. The number of significant figures for a given measurement is determined by the inherent uncertainty of the measuring device. For example, the ruler used to measure the pin can give results only to hundredths of a centimeter. So, when we record the significant figures for a measurement, we automatically give information about the uncertainty in a measurement. The uncertainty in the last number (the estimated number) is usually assumed to be ± 1 unless otherwise indicated.

For example, the measurement 1.86 kilograms can be interpreted as 1.86 ± 0.01 kilograms, where the symbol \pm means plus or minus. That is, it could be $1.86 - 0.01 \text{ kg} = 1.85 \text{ kg}$ or $1.86 + 0.01 \text{ kg} = 1.87 \text{ kg}$.

Significant figures

The certain digits and the first uncertain digit of a measurement

B. Significant Figures

We have seen that any measurement involves an estimate and thus is uncertain to some extent. We signify the degree of certainty for a particular measurement by the number of significant figures we record.

Because doing chemistry requires many types of calculations, we must consider what happens when we use numbers that contain uncertainties. It is important that we know the degree of uncertainty in the final result. Although we will not discuss the process here, mathematicians have studied how uncertainty accumulates and have designed a set of rules to determine how many significant figures the result of a calculation should have. You should follow these rules whenever you carry out a calculation. The first thing we need to do is learn how to count the significant figures in a given number. To do this we use the following rules:

Rules for Counting Significant Figures

- 1. Nonzero integers.** Nonzero integers *always* count as significant figures. For example, the number 1457 has four nonzero integers, all of which count as significant figures.
- 2. Zeros.** There are three classes of zeros:
 - a. Leading zeros** are zeros that *precede* all of the nonzero digits. They *never* count as significant figures. For example, in the number 0.0025, the three zeros simply indicate the position of the decimal point. The number has only two significant figures, the 2 and the 5.
 - b. Captive zeros** are zeros that fall *between* nonzero digits. They *always* count as significant figures. For example, the number 1.008 has four significant figures.
 - c. Trailing zeros** are zeros at the *right end* of the number. They are significant only if the number is written with a decimal point. The number one hundred written as 100 has only one significant figure, but written as 100., it has three significant figures.
- 3. Exact numbers.** Often calculations involve numbers that were not obtained using measuring devices but were determined by counting: 10 experiments, 3 apples, 8 molecules. Such numbers are called *exact numbers*. They can be assumed to have an unlimited number of significant figures. Exact numbers can also arise from definitions. For example, 1 inch is defined as *exactly* 2.54 centimeters. Thus, in the statement $1 \text{ in.} = 2.54 \text{ cm}$, neither 2.54 nor 1 limits the number of significant figures when it is used in a calculation.

information

- Leading zeros are never significant figures.
- Captive zeros are always significant figures.
- Trailing zeros are sometimes significant figures.
- Exact numbers never limit the number of significant figures in a calculation.

Rules for counting significant figures also apply to numbers written in scientific notation. For example, the number 100. can be written as 1.00×10^2 , and both versions have three significant figures. Scientific notation offers two major advantages: the number of significant figures can be indicated easily, and fewer zeros are needed to write a very large or a very small number. For example, the number 0.000060 is much more conveniently represented as 6.0×10^{-5} , and the number has two significant figures, written in either form.

Active Reading Question

The numbers 10.0 and 10 are the same, but the measurements 10.0 g and 10 g are different. Why?

EXAMPLE 5.3**Counting Significant Figures**

Give the number of significant figures for each of the following measurements.

- A sample of orange juice contains 0.0108 g of vitamin C.
- A forensic chemist in a crime lab weighs a single hair and records its mass as 0.0050060 g.
- The distance between two points was found to be 5.030×10^3 ft.
- In yesterday's bicycle race, 110 riders started but only 60 finished.

Solution

- The number contains three significant figures. The zeros to the left of the 1 are leading zeros and are not significant, but the remaining zero (a captive zero) is significant.
- The number contains five significant figures. The leading zeros (to the left of the 5) are not significant. The captive zeros between the 5 and the 6 are significant, and the trailing zero to the right of the 6 is significant because the number contains a decimal point.
- This number has four significant figures. Both zeros in 5.030 are significant.
- Both numbers are exact (they were obtained by counting the riders). Thus these numbers have an unlimited number of significant figures.

Practice Problem • Exercise 5.3

Give the number of significant figures for each of the following measurements.

- 0.00100 m
- 2.0800×10^2 L
- 480 cars

Rounding Off Numbers

When you perform a calculation on your calculator, the number of digits displayed is usually greater than the number of significant figures that the result should possess. So you must “round off” the number (reduce it to fewer digits). The rules for **rounding off** follow.

Rules for Rounding Off

- If the digit to be removed
 - is less than 5, the preceding digit stays the same. For example, 1.33 rounds to 1.3.
 - is equal to or greater than 5, the preceding digit is increased by 1. For example, 1.36 rounds to 1.4, and 3.15 rounds to 3.2.
- In a series of calculations, carry the extra digits through to the final result and *then* round off.* This means that you should carry all of the digits that show on your calculator until you arrive at the final number (the answer) and then round off, using the procedures in rule 1.

Information

The rules for rounding off reflect the way calculators round off.

**This practice will not be followed in the worked-out examples in this text, because we want to show the correct number of significant figures in each step of the example.*

We need to make one more point about rounding off to the correct number of significant figures. Suppose that the number 4.348 needs to be rounded to two significant figures. In doing this, we look at *only* the *first number* to the right of the 3:

4.348



Look at this number to round off to two significant figures.

Information

Do not round off sequentially. The number 6.8347 rounded to three significant figures is 6.83, not 6.84.

The number is rounded to 4.3 because 4 is less than 5. It is incorrect to round sequentially. For example, do *not* round the 4 to 5 to give 4.35 and then round the 3 to 4 to give 4.4.

When rounding off, use *only the first number to the right of the last significant figure*.

Determining Significant Figures in Calculations

Next we will learn how to determine the correct number of significant figures in the result of a calculation. To do this we will use the following rules.

For multiplication and division, significant figures are counted.
For addition and subtraction, the decimal places are counted.

Rules for Significant Figures in Multiplication and Division Calculations

For *multiplication* or *division*, the number of significant figures in the result is the same as that in the measurement with the *smallest number* of significant figures. We say this measurement is *limiting*, because it limits the number of significant figures in the result. For example, consider this calculation:

$$\begin{array}{ccccccc}
 4.56 & \times & 1.4 & = & 6.384 & \xrightarrow{\text{Round off}} & 6.4 \\
 \text{Three} & & \text{Limiting (two)} & & & & \text{Two} \\
 \text{significant} & & \text{significant} & & & & \text{significant} \\
 \text{figures} & & \text{figures)} & & & & \text{figures}
 \end{array}$$

Because 1.4 has only two significant figures, it limits the result to two significant figures. Thus the product is correctly written as 6.4, which has two significant figures.

Consider another example. In the division $\frac{8.315}{298}$, how many significant figures should appear in the answer?

Because 8.315 has four significant figures, the number 298 (with three significant figures) limits the result. The calculation is correctly represented as:

$$\begin{array}{ccccccc}
 \frac{8.315}{298} & = & 0.0279027 & \xrightarrow{\text{Round off}} & 2.79 \times 10^{-2} \\
 \text{Limiting (three} & & \text{Result} & & \text{Three} \\
 \text{significant} & & \text{shown on} & & \text{significant} \\
 \text{figures)} & & \text{calculator} & & \text{figures}
 \end{array}$$

Rules for Significant Figures in Addition and Subtraction Calculations

For *addition* or *subtraction*, the limiting term is the one with the smallest number of decimal places. For example, consider the following sum:

$$\begin{array}{r} 12.11 \\ 18.0 \\ \underline{1.013} \\ 31.123 \end{array}$$

Limiting term (has one decimal place)

Round off → 31.1

↑

One decimal place

The correct result is 31.1 (it is limited to one decimal place because 18.0 has only one decimal place). Consider another example:

$$\begin{array}{r} 0.6875 \\ -0.1 \\ \hline 0.5875 \end{array}$$

Limiting term (one decimal place)

Round off → 0.6

EXAMPLE 5.4

Counting Significant Figures in Calculations

Without performing the calculations, tell how many significant figures each answer should contain.

- $$\begin{array}{r} 5.19 \\ 1.9 \\ + 0.842 \end{array}$$
- $1081 - 7.25$
- 2.3×3.14
- the total cost of 3 boxes of candy at \$2.50 per box

Solution

- The answer will have one digit after the decimal place. The limiting number is 1.9, which has one decimal place, so the answer has two significant figures.
- The answer will have no digits after the decimal point. The number 1081 has no digits to the right of the decimal point and limits the result, so the answer has four significant figures.
- The answer will have two significant figures because the number 2.3 has only two significant figures (3.14 has three).
- The answer will have three significant figures. The limiting factor is 2.50 because 3 (boxes of candy) is an exact number.

CHEMISTRY IN YOUR WORLD

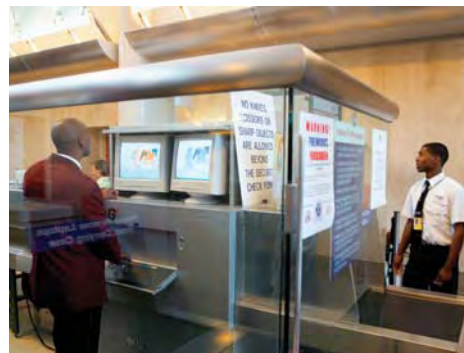
Consumer Connection

Measuring Changes

Measurement lies at the heart of doing science. We obtain the data for formulating laws and testing theories by doing measurements. Measurements also have very practical importance; they tell us whether our drinking water is safe, whether we are anemic, and the exact amount of gasoline we put in our cars at the filling station.



Engines in modern automobiles have oxygen sensors that analyze the oxygen content in the exhaust gases. This information is sent to the computer that controls the engine functions so that instantaneous adjustments can be made in spark timing and air-fuel mixtures to provide efficient power with minimum air pollution.



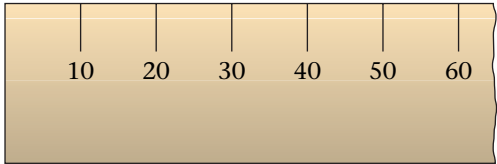
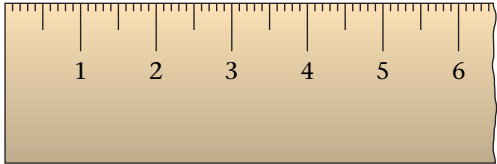
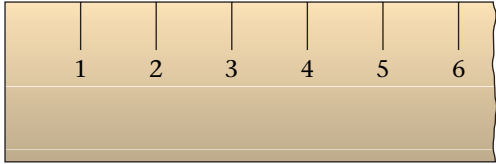
Scientists are now developing a screening procedure that bombards airline luggage with high-energy particles that cause any substance present to emit radiation characteristic of that substance. This radiation is monitored to identify luggage with unusually large quantities of nitrogen because most chemical explosives are based on compounds containing nitrogen.



Scientists are examining the natural world to find supersensitive detectors because many organisms are sensitive to tiny amounts of chemicals in their environments—for example, the sensitive noses of bloodhounds. One of these natural measuring devices uses the sensory hairs from Hawaiian red swimming crabs, which are connected to electrical analyzers and used to detect hormones at levels as low as 10^{-8} g/L.



SECTION 5.2 REVIEW QUESTIONS

- 1 Why are all measurements uncertain?
- 2 What are the three types of zeros in a number? Which type is always significant? Which type is never significant? Which type is sometimes significant?
- 3 Copy the following numbers, and mark each significant zero with an arrow.
 - a. 0.003042
 - b. 1.4030
 - c. 1000
 - d. 0.060
 - e. 50.0
 - f. 10.47020
 - g. 250.
- 4 How many significant figures are in each number below?
 - a. 0.003042
 - b. 1.4030
 - c. 1000
 - d. 0.060
 - e. 50.0
 - f. 10.47020
 - g. 250.
- 5 Without doing the calculation, tell how many significant figures should be in each result.
 - a. $3.2 + 4.17 + 1.243$
 - b. $1.3478 - 0.02$
 - c. 4.6×3.435
 - d. $(4.2 \times 10^{-5}) \times 3.74 \div 6.783$
 - e. $50 - 0.00473$
- 6 Perform the calculations, and determine the results from question 5.
- 7 State the number of significant figures in a measurement made with each ruler:
 - a. 
 - b. 
 - c. 



RESEARCH LINKS
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Problem Solving and Unit Conversions

Objectives

- To learn how dimensional analysis can be used to solve problems
- To learn the three temperature scales
- To learn to convert from one temperature scale to another
- To practice using problem-solving techniques
- To define density and its units

Key Terms

- Conversion factor
- Equivalence statement
- Dimensional analysis
- Fahrenheit scale
- Celsius scale
- Kelvin scale (absolute)
- Density
- Specific gravity

A. Tools for Problem Solving

Suppose that the boss at the store where you work on weekends asks you to pick up two dozen doughnuts on the way to work. However, you find that the doughnut shop sells by the doughnut. How many doughnuts do you need?

Problem-Solving Strategy

One of the most important things we do in life is solve problems. We confront problems constantly in school, at our jobs and in our personal lives. The more creative we are at solving these problems, the more successful we will be. It turns out that most of the problems we encounter can be attacked systematically. It is helpful to use the following steps when solving a problem:

Where do we want to go?

- To solve a problem, we need to be able to state the problem clearly—what exactly are we trying to do? What is the goal? Often drawing a picture can help.

What do we know?

- This represents the starting point. What facts do we know that relate to the problem?

How do we get there?

- How do we use the facts we know and the understanding we have of the relevant concepts to proceed from the starting point to the eventual goal—the solution to the problem. In many types of problems we will discuss the relevant fundamental ideas and a method (often in the form of steps) for solving the problem before you are asked to try a similar one on your own. You might think of these steps as the tools for solving the problem.

Does it make sense?

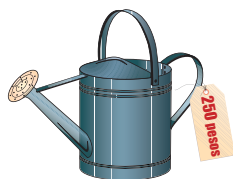
- After we have reached the solution, we need to evaluate it. Is the answer we have obtained reasonable? For example, if we are trying to calculate the volume of water in a certain glass and our answer is 65 L, that would be unreasonable. This number cannot be correct. We must have made a mistake somewhere. When we arrive at an answer to a problem, it is important to test the reasonableness of the answer.

Cooking



3 cups = ? pints

Traveling



250 pesos = ? dollars

Sports



3215 km = ? miles

In doing problems in this text, we will use this general strategy where it fits. It usually works best for numerical problems such as the examples in the remainder of this chapter.

Converting Units of Measurement The “doughnut problem” mentioned previously is an example of something you encounter all the time: converting from one unit of measurement to another.

How do we convert from one unit of measurement to another? Let’s explore this process by using the doughnut problem.

2 dozen doughnuts = ? individual doughnuts

The essential information you must have is the definition of a dozen:

$$1 \text{ dozen} = 12$$

You can use this information to make the needed conversion as follows:

$$2 \text{ dozen doughnuts} \times \frac{12}{1 \text{ dozen}} = 24 \text{ doughnuts}$$

You need to buy 24 doughnuts.

Note two important things about this process.

1. The factor $\frac{12}{1 \text{ dozen}}$ is a conversion factor based on the definition of the term *dozen*. This conversion factor is a ratio of the two parts of the definition of a dozen given above.
2. The unit *dozen* itself cancels.

To change from one unit to another, we will use a conversion factor.

$$\text{Unit}_1 \times \text{conversion factor} = \text{Unit}_2$$

The **conversion factor** is a ratio of the two parts of the statement that relates the two units.

Choosing Conversion Factors Earlier in this chapter we considered a pin that measured 2.85 cm in length. What is the length of the pin in inches? We can represent this problem as

$$2.85 \text{ cm} \rightarrow ? \text{ in.}$$

The question mark stands for the number we want to find. To solve this problem, we must know the relationship between inches and centimeters. In **Table 5.7**, which gives several equivalents between the English and metric systems, we find the relationship:

$$2.54 \text{ cm} = 1 \text{ in.}$$

This is called an **equivalence statement**. In other words, 2.54 cm and 1 in. stand for *exactly the same distance* (see **Figure 5.1** on page 130). The respective numbers are different because they refer to different *scales (units)* of distance.

The equivalence statement $2.54 \text{ cm} = 1 \text{ in.}$ can lead to either of two conversion factors:

$$\frac{2.54 \text{ cm}}{1 \text{ in.}} \text{ or } \frac{1 \text{ in.}}{2.54 \text{ cm}}$$

Conversion factors are ratios of the two parts of the equivalence statement that relates the two units.

Conversion factor

A ratio used to convert from one unit to another

TABLE 5.7

English-Metric and English-English Equivalents

Length	1 m = 1.094 yd 2.54 cm = 1 in. 1 mi = 5280. ft 1 mi = 1760. yd
Mass	1 kg = 2.205 lb 453.6 g = 1 lb
Volume	1 L = 1.06 qt 1 ft ³ = 28.32 L

Which of the two possible conversion factors do we need? Recall our problem:

$$2.85 \text{ cm} = ? \text{ in.}$$

That is, we want to convert from units of centimeters to inches:

$$2.85 \text{ cm} \times \text{conversion factor} = ? \text{ in.}$$

We choose a conversion factor that cancels the units we want to eliminate and leaves the units we want in the result.

Thus we do the conversion as follows:

$$2.85 \text{ cm} \times \frac{1 \text{ in.}}{2.54 \text{ cm}} = \frac{2.85 \text{ in.}}{2.54} = 1.12 \text{ in.}$$

Note two important facts about this conversion.

1. The centimeter units cancel to give inches for the result. This is exactly what we wanted to accomplish. Using the other conversion factor:

$$\left(2.85 \text{ cm} \times \frac{2.54 \text{ cm}}{1 \text{ in.}}\right)$$

would not work because the units would not cancel to give inches in the result.

2. As the units changed from centimeters to inches, the number changed from 2.85 to 1.12. Thus 2.85 cm has exactly the same value (is the same length) as 1.12 in. Notice that in this conversion, the number decreased from 2.85 to 1.12. This makes sense because the inch is a larger unit of length than the centimeter is. That is, it takes fewer inches to make the same length in centimeters.

The result in the previous conversion has three significant figures as required. Caution: Noting that the term 1 appears in the conversion, you might think that because this number appears to have only one significant figure, the result should have only one significant figure. That is, the answer should be given as 1 in. rather than 1.12 in. However, in the equivalence statement $1 \text{ in.} = 2.54 \text{ cm}$, the 1 is an exact number (by definition). In other words, exactly 1 in. equals 2.54 cm. Therefore, the 1 does not limit the number of significant digits in the result.

We have seen how to convert from centimeters to inches. What about the reverse conversion? For example, if a pencil is 7.00 in. long, what is its length in centimeters? In this case, the conversion we want to make is:

$$7.00 \text{ in.} \rightarrow ? \text{ cm}$$

What conversion factor do we need to make this conversion?

Remember that two conversion factors can be derived from each equivalence statement. In this case, the equivalence statement $2.54 \text{ cm} = 1 \text{ in.}$ gives

$$\frac{2.54 \text{ cm}}{1 \text{ in.}} \text{ or } \frac{1 \text{ in.}}{2.54 \text{ cm}}$$

Again, we choose which to use by looking at the *direction* of the required change. For us to change from inches to centimeters, the inches must

cancel. Thus the factor $\frac{2.54 \text{ cm}}{1 \text{ in.}}$ is used, and the conversion is done as follows:

$$7.00 \text{ in.} \times \frac{2.54 \text{ cm}}{1 \text{ in.}} = (7.00)(2.54) \text{ cm} = 17.8 \text{ cm}$$

Here the inch units cancel, leaving centimeters as required.

MATH



Units can cancel, just as numbers can.

Information

When you finish a calculation, always check to make sure that the answer makes sense.

MATH



When exact numbers are used in a calculation, they never limit the number of significant digits.

Note that in this conversion, the number increased (from 7.00 to 17.8). This increase makes sense because the centimeter is a smaller unit of length than the inch. That is, it takes more centimeters to make the same length in inches.

Always take a moment to think about whether your answer makes sense. It will help you avoid errors.

Dimensional Analysis

The process of using conversion factors to change from one unit to another

Changing from one unit to another via conversion factors (based on the equivalence statements between the units) is often called **dimensional analysis**. We will use this method throughout our study of chemistry.

We can now state some general steps for doing conversions by dimensional analysis. These steps are the tools for “How do we get there” in our problem-solving strategy.

Tools for Converting from One Unit to Another

Step 1 To convert from one unit to another, use the equivalence statement that relates the two units. The conversion factor needed is a ratio of the two parts of the equivalence statement.

Step 2 Choose the appropriate conversion factor by looking at the direction of the required change (make sure that the unwanted units cancel).

Step 3 Multiply the quantity to be converted by the conversion factor to give the quantity with the desired units.

Step 4 Make sure that you have the correct number of significant figures.

Information

Consider the direction of the required change to select the correct conversion factor.

Active Reading Questions

1. Why must a conversion factor come from an equivalence statement?
2. How can you decide which conversion factor to choose in a problem?

HANDS-ON CHEMISTRY

• MINI-LAB •

Cooking in a Metric World

Materials

- recipe cards
- conversion table
- calculator

Procedure

1. Obtain a recipe card from your teacher.
2. Convert all of the measurements to metric units. Be careful in determining the number of significant figures in your answers.

EXAMPLE 5.6

An Italian bicycle has its frame size given as 62 cm. What is the frame size in inches?

Solution**Where do we want to go?**

The problem can be represented as:

$$62 \text{ cm} = ? \text{ in.}$$

What do we know?

- The frame size is 62 cm.
- 1 in. = 2.54 cm

**How do we get there?**

In this problem we want to convert from centimeters to inches.

$$62 \text{ cm} \times \text{conversion factor} = ? \text{ in.}$$

Step 1 To convert from centimeters to inches, we need the equivalence statement 1 in. = 2.54 cm. This leads to two conversion factors:

$$\frac{1 \text{ in.}}{2.54 \text{ cm}} \text{ and } \frac{2.54 \text{ cm}}{1 \text{ in.}}$$

Step 2 In this case, the direction we want is centimeters \rightarrow inches

so we need the conversion factor $\frac{1 \text{ in.}}{2.54 \text{ cm}}$. We know that this is the one we want because using it will make the units of centimeters cancel, leaving units of inches.

Step 3 The conversion is carried out as follows:

$$62 \text{ cm} \times \frac{1 \text{ in.}}{2.54 \text{ cm}} = 24 \text{ in.}$$

Step 4 The result is limited to two significant figures by the number 62. The centimeters cancel, leaving inches as required.

Does it make sense?

Note that the number decreased in this conversion. This makes sense because the inch is a larger unit of length than the centimeter.

Practice Problem • Exercise 5.6

Water is often bottled in 0.750-L containers. Using the appropriate equivalence statement from Table 5.7, calculate the volume of such a water bottle in quarts.

We can also use dimensional analysis and this problem-solving plan to make multiple conversions.

EXAMPLE 5.7

Conversion Factors: Multiple-Step Problems

The length of the marathon race is approximately 26.2 mi. What is this distance in kilometers?

Solution

Where do we want to go?

The problem can be represented as: 26.2 mi = ? km

What do we know?

- The race is 26.2 mi.
- 1 mi = 1760 yd
- 1 m = 1.094 yd

How do we get there?

Step 1 Because **Table 5.7** gives the equivalence statements 1 mi = 1760 yd and 1 m = 1.094 yd, we will proceed as follows:

Miles → yards → meters → kilometers

Step 2 and Step 3 for each conversion

Miles → Yards:

Result shown
on calculator

$$26.2 \cancel{\text{mi}} \times \frac{1760 \text{ yd}}{1 \cancel{\text{mi}}} = 46,112 \text{ yd}$$

$$46,112 \text{ yd} \xrightarrow{\text{Round off}} 46,100 \text{ yd} = 4.61 \times 10^4 \text{ yd}$$

Yards → Meters:

Result shown
on calculator

$$4.61 \times 10^4 \cancel{\text{yd}} \times \frac{1 \text{ m}}{1.094 \cancel{\text{yd}}} = 4.213894 \times 10^4 \text{ m}$$

$$4.213894 \times 10^4 \text{ m} \xrightarrow{\text{Round off}} 4.21 \times 10^4 \text{ m}$$

Meters → Kilometers:

$$4.21 \times 10^4 \cancel{\text{m}} \times \frac{1 \text{ km}}{10^3 \cancel{\text{m}}} = 4.21 \times 10^1 \text{ km} \\ = 42.1 \text{ km}$$

Thus the marathon (26.2 mi) is 42.1 km.

When you feel comfortable with the conversion process, you can combine the steps. The combined expression is miles → yards → meters → kilometers.

Step 4 Note that the answer has 3 significant figures. This matches the number of significant figures in the original measurement (26.2 mi).

Does it make sense?

The problem asked for the distance in kilometers. Our answer is given in kilometers and has 3 significant figures as required.

MATH



Remember that we are rounding off at the end of each step to show the correct number of significant figures. However, in doing a multistep calculation, you should retain the extra numbers that show on your calculator and round off only at the end of the calculation.

information

Units provide a very valuable check on the validity of your solution. Always use them.

Practice Problem • Exercise 5.7

Racing cars at the Indianapolis Motor Speedway routinely travel around the track at an average speed of 225 mi/h. What is this speed in kilometers per hour?

Let's Review

Whenever you work problems, remember the following points:

- Always include the units (a measurement always has two parts: a number *and* a unit).
- Cancel units as you carry out the calculations.
- Check that your final answer has the correct units. If it doesn't, you have done something wrong.
- Check that your final answer has the correct number of significant figures.
- Think about whether your answer makes sense.

B. Temperature Conversions

When the doctor tells you that your temperature is 102 degrees and the meteorologist on TV says it will be 75 degrees tomorrow, they are using the **Fahrenheit scale**. Water boils at 212 °F and freezes at 32 °F, and normal body temperature is 98.6 °F (where °F signifies “Fahrenheit degrees”). This temperature scale is widely used in the United States and Great Britain, and it is the scale employed in most of the engineering sciences.

Another temperature scale, used in Canada and Europe and in the physical and life sciences in most countries, is the **Celsius scale**. In keeping with the metric system, which is based on powers of 10, the freezing and boiling points of water on the Celsius scale are assigned as 0 °C and 100. °C, respectively. On both the Fahrenheit and the Celsius scales, the unit of temperature is called a degree, and the symbol for it is followed by the capital letter representing the scale on which the units are measured: °C or °F.

Still another temperature scale used in the sciences is the **absolute** or **Kelvin scale**. On this scale water freezes at 273 K and boils at 373 K. On the Kelvin scale, the unit of temperature is called a kelvin and is symbolized by K. Thus, on the three scales, the boiling point of water is stated as 212 Fahrenheit degrees (212 °F), 100. Celsius degrees (100. °C), and 373 kelvins (373 K). The three temperature scales are compared in **Figure 5.5**.

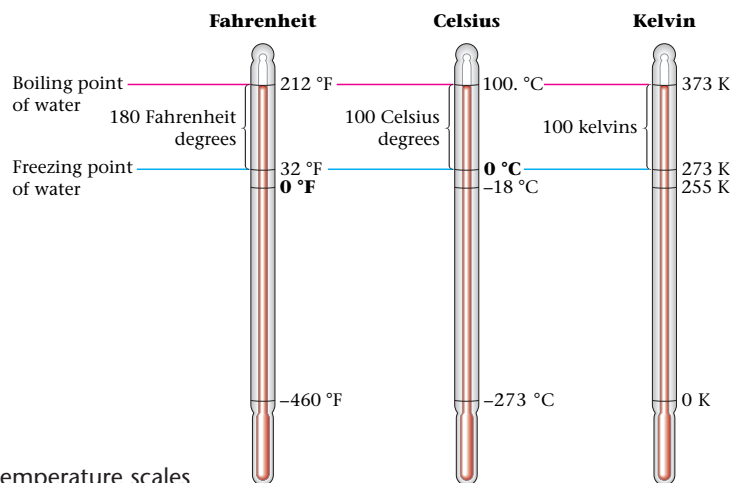


Figure 5.5

The three major temperature scales

MATH

Although 373 K is often stated as 373 degrees kelvin, it is more correct to say 373 kelvins.

Let's Review

- The size of each temperature unit (each degree) is the same for the Celsius and Kelvin scales. This follows from the fact that the *difference* between the boiling and freezing points of water is 100 units on both of these scales.
- The Fahrenheit degree is smaller than the Celsius and Kelvin unit. Note that on the Fahrenheit scale there are 180 Fahrenheit degrees between the boiling and freezing points of water, as compared with 100 units on the other two scales.
- The zero points are different on all three scales.

In your study of chemistry, you will sometimes need to convert from one temperature scale to another. We will consider in some detail how this conversion is done. In addition to learning how to change temperature scales, you should also use this section as an opportunity to further develop your skills in problem solving.

Converting Between the Kelvin and Celsius Scales

It is simple to convert between the Celsius and Kelvin scales because

- the temperature unit is the same size.
- the zero points are different. Look at **Figure 5.5**. To go from each Celsius temperature marked on the middle thermometer to the Kelvin temperature on the right thermometer, what do we need to do?

$$100. \text{ } ^\circ\text{C} + 273 = 373 \text{ K}$$

$$0 \text{ } ^\circ\text{C} + 273 = 273 \text{ K}$$

$$-18 \text{ } ^\circ\text{C} + 273 = 255 \text{ K}$$

$$-273 \text{ } ^\circ\text{C} + 273 = 0 \text{ K}$$

In each case we need to add 273 to the Celsius temperature. Because $0 \text{ } ^\circ\text{C}$ corresponds to 273 K, this addition of 273 corrects for the difference in zero points.

$$T_{\text{C}} + 273 = T_{\text{K}}$$

We will illustrate this procedure in Example 5.8.

Active Reading Question

Which change is the largest, a change of $1 \text{ } ^\circ\text{C}$, $1 \text{ } ^\circ\text{F}$, or 1 K ?

EXAMPLE 5.8**Temperature Conversion: Celsius to Kelvin**

The boiling point of water at the top of Mt. Everest is 70. °C. Convert this temperature to the Kelvin scale.

Solution**Where do we want to go?**

- 70. °C = ? K

What do we know?

- The boiling point is 70. °C.
- The relationship between the Celsius and Kelvin scales is:

$$T_{\text{°C}} + 273 = T_{\text{K}}$$

How do we get there?

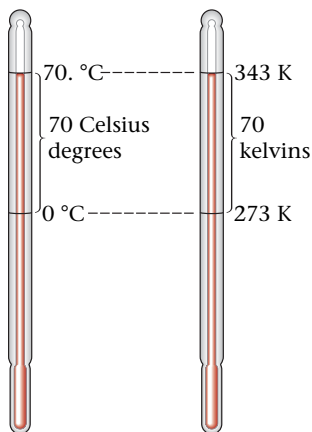
Celsius-to-Kelvin conversions require us to adjust for different zero points by using the formula

$$T_{\text{°C}} + 273 = T_{\text{K}}$$

Using this formula to solve this problem gives

$$70. + 273 = 343$$

343 K is the correct answer. This conversion is summarized in **Figure 5.6**.

**Figure 5.6**

Converting 70. °C to units measured on the Kelvin scale

Does it make sense?

The problem asked for the boiling point in kelvins. Our answer is given in kelvins. Note that the answer is to the units place as required by the original measurement (70. °C).

MATH

The decimal point after the temperature reading indicates that the trailing zero is significant.

Let's Review

To convert from the Celsius to the Kelvin scale, we use the formula:

$$T_{\text{°C}} + 273 = T_{\text{K}}$$

Temperature in Celsius degrees Temperature in kelvins

EXAMPLE 5.9

Temperature Conversion: Kelvin to Celsius

Liquid nitrogen boils at 77 K. What is the boiling point of nitrogen on the Celsius scale?

Solution

Where do we want to go?

- 77 K = ? °C

What do we know?

- Liquid nitrogen boils at 77 K.
- The relationship between the Celsius and Kelvin scales is $T_{\text{C}} + 273 = T_{\text{K}}$

How do we get there?

We solve this problem by using the formula $T_{\text{C}} + 273 = T_{\text{K}}$

However, in this case we want to solve for the Celsius temperature T_{C} . That is, we want to isolate T_{C} on one side of the equal sign. To do this, we use an important mathematical principle: doing *the same thing on both sides of the equal sign* preserves the equality. In other words, it's always okay to perform the same operation on both sides of the equal sign.

To isolate T_{C} , we need to subtract 273 from both sides.

$$T_{\text{C}} + 273 - 273 = T_{\text{K}} - 273$$

 ↑ ↑
Sum is zero

to give $T_{\text{C}} = T_{\text{K}} - 273$

Using this equation to solve the problem, we have

$$T_{\text{C}} = T_{\text{K}} - 273 = 77 - 273 = -196$$

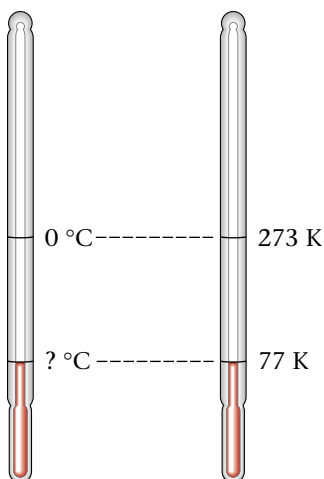
So we have shown that 77 K = -196 °C.

Does it make sense?

The problem asked for the boiling point in degrees Celsius. Our answer is given in degrees Celsius. Our answer (-196 °C) has the correct units. Note that the answer is to the units place as required by the original measurement (77 K).

Practice Problem • Exercise 5.9

Which temperature is colder, 172 K or -75 °C?



In summary, because the Kelvin and Celsius scales have the same size unit, to switch from one scale to the other we must simply account for the different zero points. We must add 273 to the Celsius temperature to obtain the temperature on the Kelvin scale:

$$T_{\text{K}} = T_{\text{C}} + 273$$

To convert from the Kelvin scale to the Celsius scale, we must subtract 273 from the Kelvin temperature:

$$T_{\text{C}} = T_{\text{K}} - 273$$

Converting Between the Fahrenheit and Celsius Scales

The conversion between the Fahrenheit and Celsius temperature scales requires two adjustments:

- For the different size units
- For the different zero points

To see how to adjust for the different unit sizes, consider the diagram in

Figure 5.7. Note that because $212\text{ }^{\circ}\text{F} = 100\text{ }^{\circ}\text{C}$ and $32\text{ }^{\circ}\text{F} = 0\text{ }^{\circ}\text{C}$,

$$212 - 32 = 180. \text{ Fahrenheit degrees} = 100. - 0 = 100. \text{ Celsius degrees}$$

Thus

$$180. \text{ Fahrenheit degrees} = 100. \text{ Celsius degrees}$$

Dividing both sides of this equation by 100. gives

$$\frac{180.}{100.} \text{ Fahrenheit degrees} = \frac{100.}{100.} \text{ Celsius degrees}$$

or

$$1.80 \text{ Fahrenheit degrees} = 1.00 \text{ Celsius degree}$$

The factor 1.80 is used to convert from one degree size to the other.

Next we have to account for the fact that $0\text{ }^{\circ}\text{C}$ is *not* the same as $0\text{ }^{\circ}\text{F}$.

In fact, $32\text{ }^{\circ}\text{F} = 0\text{ }^{\circ}\text{C}$.

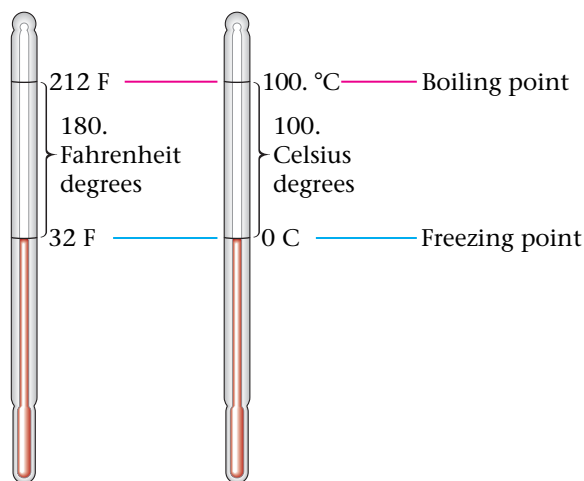


Figure 5.7

Comparison of the Celsius and Fahrenheit scales

Celsius to Fahrenheit Although we will not show how to derive it, the equation to convert a temperature in Celsius degrees to the Fahrenheit scale is:

$$T_{\text{F}} = 1.80(T_{\text{C}}) + 32$$

↑
↑
 Temperature in $^{\circ}\text{F}$ Temperature in $^{\circ}\text{C}$

In this equation the term $1.80(T_{\text{C}})$ adjusts for the difference in degree size between the two scales. The 32 in the equation accounts for the different zero points. We will now show how to use this equation.

To convert from Celsius to Fahrenheit, we have used the equation

$$T_{\text{F}} = 1.80(T_{\text{C}}) + 32)$$

MATH



Note that because 180 and 100 here are both exact numbers obtained by counting the number of degrees, the ratio $180/100$ has an unlimited number of significant digits. This ratio is typically shown as 1.80.

EXAMPLE 5.10

Temperature Conversion: Celsius to Fahrenheit

On a summer day the temperature in the laboratory, as measured on a lab thermometer, is 28 °C. Express this temperature on the Fahrenheit scale.

Solution

Where do we want to go?

- 28 °C = ? °F

What do we know?

- The temperature is 28 °C.
- The relationship between the Celsius and Kelvin scales is $T_{\text{°F}} + 1.80(T_{\text{°C}}) + 32$

How do we get there?

$$T_{\text{°F}} = 1.80(T_{\text{°C}}) + 32$$

In this case,

$$\begin{aligned} T_{\text{°F}} &= ? \text{ °F} = 1.80(T_{\text{°C}}) + 32 = 50.4 + 32 \\ &= 50. + 32 = 82 \end{aligned}$$

$T_{\text{°C}}$
↓
Rounds off to 50.

Thus 28 °C = 82 °F.

Does it make sense?

The problem asked for the temperature in degrees Fahrenheit. Our answer is given in degrees Fahrenheit. Our answer (82 °F) has the correct units. Note that the answer is to the units place as required by the original measurement (28 °C).

MATH



Note that 28 °C is approximately equal to 82 °F. Because the numbers are reversed, this is an easy reference point to remember for the two scales.

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Ersatz Fog

Did you know that a knowledge of chemistry might win you an Academy Award? In fact, James F. Foley, Charles E. Converse, and F. Edward Gardner of Praxair Corporation were given Technical Achievement Awards by the Academy of Motion Picture Arts and Sciences for their fake fog concoction used in the movie "Batman and Robin." The Praxair employees created a special low-lying fog for scenes involving the villain Mr. Freeze, played by Arnold Schwarzenegger. The fog resulted from a special mixture of steam, liquid nitrogen, and liquid oxygen. About 27.5 million cubic feet of the

mixture was used in making the film. The liquid oxygen and nitrogen were hauled to the studio daily from a nearby Praxair air liquefaction plant in a cryogenic tank truck. A patent is now pending for the fog recipe, which looks like it has a "bright" future in Hollywood.

Note the ersatz fog behind Mr. Freeze.



EXAMPLE 5.11**Temperature Conversion: Celsius to Fahrenheit**

Express the temperature $-40.^\circ\text{C}$ on the Fahrenheit scale.

Solution**Where do we want to go?**

- $-40.^\circ\text{C} = ?^\circ\text{F}$

What do we know?

- The temperature is $-40.^\circ\text{C}$.
- The relationship between the Fahrenheit and Celsius scales is
 $T_{\circ\text{F}} = 1.80(T_{\circ\text{C}}) + 32$

How do we get there?

$$T_{\circ\text{F}} = 1.80(T_{\circ\text{C}}) + 32$$

In this case,

$$\begin{aligned} T_{\circ\text{F}} = ?^\circ\text{F} &= 1.80(\overset{T_{\circ\text{C}}}{\downarrow} -40.) + 32 \\ &= -72 + 32 = -40. \end{aligned}$$

So $-40.^\circ\text{C} = -40.^\circ\text{F}$. This is a very interesting result and is another useful reference point.

Does it make sense?

The problem asked for the temperature in degrees Fahrenheit. Our answer is given in degrees Fahrenheit. Our answer ($-40.^\circ\text{F}$) has the correct units. Note that the answer is to the units place as required by the original measurement ($-40.^\circ\text{C}$).

Practice Problem • Exercise 5.11

Hot tubs are often maintained at 41°C . What is this temperature in Fahrenheit degrees?

Fahrenheit to Celsius To convert a Fahrenheit temperature to Celsius, we need to rearrange this equation to isolate Celsius degrees ($T_{\circ\text{C}}$). Remember that we can always do the same operation to both sides of the equation. First subtract 32 from each side:

$$T_{\circ\text{F}} - 32 = 1.80(T_{\circ\text{C}}) + 32 - 32$$

$\uparrow \quad \uparrow$
 Sum is zero

to give

$$T_{\circ\text{F}} - 32 = 1.80(T_{\circ\text{C}})$$

Next divide both sides by 1.80

$$\frac{T_{\circ\text{F}} - 32}{1.80} = \frac{1.80(T_{\circ\text{C}})}{1.80} = T_{\circ\text{C}}$$

$$\text{Temperature in } ^\circ\text{C} = T_{\circ\text{C}} = \frac{T_{\circ\text{F}} - 32}{1.80}$$

EXAMPLE 5.12

Temperature Conversion: Fahrenheit to Celsius

One of the body's responses to an infection or injury is to elevate its temperature. A certain flu victim has a body temperature of 101 °F. What is this temperature on the Celsius scale?

Solution

Where do we want to go?

- 101 °F = ? °C

What do we know?

- The temperature is 101 °F.
- The relationship between the Celsius and Fahrenheit scales is:

$$T_{\text{°C}} = \frac{T_{\text{°F}} - 32}{1.80}$$

How do we get there?

$$T_{\text{°C}} = \frac{T_{\text{°F}} - 32}{1.80}$$

yields $T_{\text{°F}}$

$$T_{\text{°C}} = ? \text{ °C} = \frac{101 - 32}{1.80} = \frac{69}{1.80} = 38$$

That is, 101 °F = 38 °C.

Does it make sense?

The problem asked for the temperature in degrees Celsius. Our answer is given in degrees Celsius. Our answer (38 °C) has the correct units. Note that the answer is to the units place as required by the original measurement (101 °F).

Practice Problem • Exercise 5.12

An antifreeze solution in a car's radiator boils at 239 °F. What is this temperature on the Celsius scale?

Let's Review

Temperature Conversion Formulas

Celsius to Kelvin $T_K = T_{\text{°C}} + 273$

Kelvin to Celsius $T_{\text{°C}} = T_K - 273$

Celsius to Fahrenheit $T_{\text{°F}} = 1.80(T_{\text{°C}}) + 32$

Fahrenheit to Celsius $T_{\text{°C}} = \frac{T_{\text{°F}} - 32}{1.80}$

C. Density

When you were in elementary school, you may have been embarrassed by your answer to the question, “Which is heavier, a pound of lead or a pound of feathers?” If you said lead, you were undoubtedly thinking about density, not mass. **Density** can be defined as the amount of matter present in a given volume of substance. That is, density is mass per unit volume, or the ratio of the mass of an object to its volume:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

It takes a much bigger volume to make a pound of feathers than to make a pound of lead. This is because lead has a much greater mass per unit volume—a greater density.

Density

The mass of a substance per a given volume of that substance

Active Reading Question

Which is heavier, a pound of lead or a pound of feathers? What is the “trick” about this question?

The density of a liquid can be determined easily by weighing a known volume of the substance as illustrated in Example 5.13.

EXAMPLE 5.13

Calculating Density

Suppose that a student finds that 23.50 mL of a certain liquid weighs 35.062 g. What is the density of this liquid?

Solution

Where do we want to go?

- Density of liquid = ? g/mL

What do we know?

- Volume = 23.50 mL
- Mass = 35.062 g
- Density = $\frac{\text{mass}}{\text{volume}}$

How do we get there?

We can calculate the density of this liquid simply by applying the definition.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{35.062 \text{ g}}{23.50 \text{ mL}} = 1.492 \text{ g/mL}$$

This result could also be expressed as 1.492 g/cm³ because 1 mL = 1 cm³.

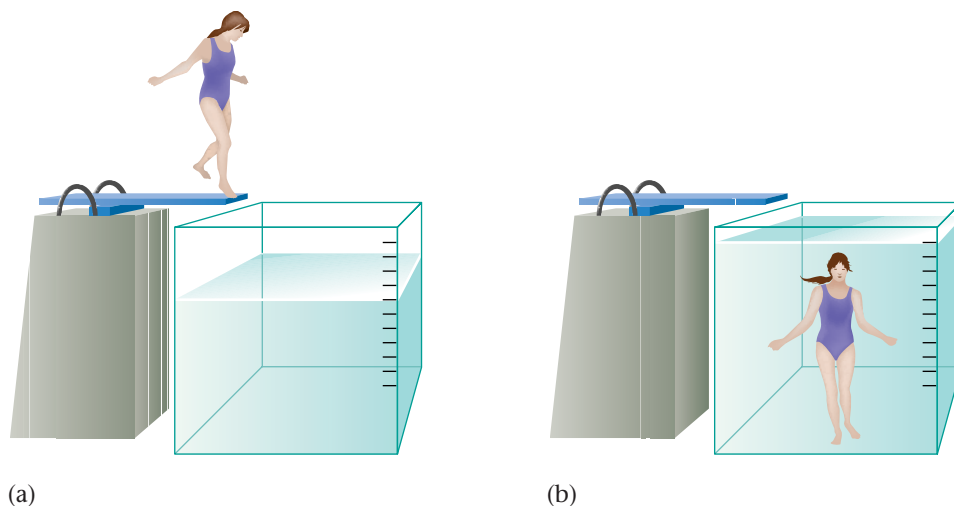
Does it make sense?

Density is defined as $\frac{\text{mass}}{\text{volume}}$. Our units ($\frac{\text{g}}{\text{mL}}$) express

mass (g) over volume (mL). Our answer has 4 significant figures as required by the 4 significant figures in the volume measurement (23.50 mL).

Figure 5.8

- (a) Tank of water
(b) Person submerged in the tank, raising the level of the water



The volume of a solid object is often determined indirectly by submerging it in water and measuring the volume of water displaced. In fact, this is the most accurate method for measuring a person's percentage of body fat. The person is submerged momentarily in a tank of water, and the increase in volume is measured (see **Figure 5.8**). It is possible to calculate the body density by using the person's weight (mass) and the volume of the person's body determined by submersion. Fat, muscle, and bone have different densities (fat is less dense than muscle tissue, for example), so the fraction of the person's body that is fat can be calculated. The more muscle and the less fat a person has, the higher his or her body density. For example, a muscular person weighing 150 lb has a smaller body volume (and thus a higher density) than a non-muscular person weighing 150 lb.

The densities of various common substances are given in **Table 5.8**.

TABLE 5.8**Densities of Various Common Substances at 20 °C**

Substance	Physical State	Density (g/cm ³)	Substance	Physical State	Density (g/cm ³)
oxygen	gas	0.00133*	aluminum	solid	2.70
hydrogen	gas	0.0000084*	iron	solid	7.87
ethanol	liquid	0.78	copper	solid	8.96
benzene	liquid	0.880	silver	solid	10.5
water	liquid	1.000	lead	solid	11.34
magnesium	solid	1.74	mercury	liquid	13.6
salt (sodium chloride)	solid	2.16	gold	solid	19.32

*At 1 atmosphere pressure

EXAMPLE 5.14**Determining Density**

At a local pawn shop, a student finds a medallion that the shop owner insists is pure platinum. However, the student suspects that the medallion may actually be silver and thus much less valuable. The student buys the medallion only after the shop owner agrees to refund the price if the medallion is returned within two days. The student, a chemistry student, then takes the medallion to her lab and measures the density as follows:

- She weighs the medallion and finds that its mass is 55.64 g.
- She places some water in a graduated cylinder and reads the volume as 75.2 mL.
- She drops the medallion into the water in the cylinder and reads the new volume as 77.8 mL.

Is the medallion platinum (density = 21.4 g/cm³) or silver (density = 10.5 g/cm³)?

Solution**Where do we want to go?**

- We want to determine the density of the medallion to see whether it is platinum or silver.

What do we know?

- Mass (medallion) = 55.64 g
- Initial volume of water = 75.2 mL
- Final volume (water + medallion) = 77.8 mL
- Density = $\frac{\text{mass}}{\text{volume}}$
- Density (Pt) = 21.4 g/cm³
- Density (Ag) = 10.5 g/cm³
- 1 g/mL = 1 g/cm³

How do we get there?**Calculating the density of the medallion**

- Density = $\frac{\text{mass}}{\text{volume}} = \frac{55.64 \text{ g}}{\text{volume}}$
- The volume of medallion is obtained by finding the difference between the volume reading of the water in the graduated cylinder.

$$\text{Volume (medallion)} = 77.8 \text{ mL} - 75.2 \text{ mL} = 2.6 \text{ mL}$$

- Density = mass/volume = $\frac{55.64 \text{ g}}{2.6 \text{ mL}} = 21 \frac{\text{g}}{\text{mL}} = 21 \frac{\text{g}}{\text{cm}^3}$

Comparing densities

- Densities
 - Density (medallion) = 21 g/cm³
 - Density (Pt) = 21.4 g/cm³
 - Density (Ag) = 10.5 g/cm³
- The medallion must be platinum because its density is the same as the platinum.

Does it make sense?

The density of the medallion is the same as platinum to two significant figures (as required by the volume measurement).

Information

The most common units for density are g/mL = g/cm³.

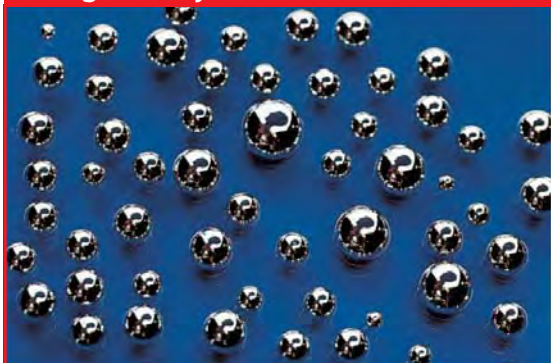
Practice Problem • Exercise 5.14

A student wants to identify the main component in a commercial liquid cleaner. He finds that 35.8 mL of the cleaner weighs 28.1 g. Of the following possibilities, which is the main component of the cleaner?

Substance	Density, g/cm ³
chloroform	1.483
diethyl ether	0.714
isopropyl alcohol	0.785
toluene	0.867

EXAMPLE 5.15

Using Density in Calculations



Spherical droplets of mercury, a very dense liquid

Mercury has a density of 13.6 g/mL. What volume of mercury must be taken to obtain 225 g of the metal?

Solution

Where do we want to go?

- volume of mercury → 225 g

What do we know?

- Mass (Hg) = 225 g
- Density (Hg) = 13.6 g/mL
- Density = $\frac{\text{mass}}{\text{volume}}$

How do we get there?

To solve this problem, start with the definition of density:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

Then rearrange this equation to isolate the required quantity. In this case we want to find the volume. Remember that we maintain an equality when we do the same thing to both sides. For example, if we multiply *both sides* of the density definition by volume,

$$\text{Volume} \times \text{density} = \frac{\text{mass}}{\text{volume}} \times \text{volume}$$

volume cancels on the right, leaving

$$\text{Volume} \times \text{density} = \text{mass}$$

We want the volume, so we now divide both sides by density:

$$\frac{\text{Volume} \times \text{density}}{\text{density}} = \frac{\text{mass}}{\text{density}}$$

to give

$$\text{Volume} = \frac{\text{mass}}{\text{density}}$$

Now we can solve the problem by substituting the given numbers:

$$\begin{aligned} \text{Volume} &= \frac{225 \text{ g}}{13.6 \text{ g/mL}} \\ &= 16.5 \text{ mL} \end{aligned}$$

We must take 16.5 mL of mercury to obtain an amount that has a mass of 225 g.

Does it make sense?

The correct unit for volume is milliliters, which we have in our answer (16.5 mL). Our answer has 3 significant figures as required.

Besides being a tool for the identification of substances, density has many other uses. For example, the liquid in your car's lead storage battery (a solution of sulfuric acid) changes density because the sulfuric acid is consumed as the battery discharges. In a fully charged battery, the density of the solution is about 1.30 g/cm^3 . When the density falls below 1.20 g/cm^3 , the battery has to be recharged. Density measurement is also used to determine the amount of antifreeze, and thus the level of protection against freezing, in the cooling system of a car. Water and antifreeze have different densities, so the measured density of the mixture tells us how much of each is present. The device used to test the density of the solution—a hydrometer—is shown in **Figure 5.9**.

In certain situations, the term *specific gravity* is used to describe the density of a liquid. **Specific gravity** is defined as the ratio of the density of a given liquid to the density of water at 4°C . Because it is a ratio of densities, specific gravity has no units.

Specific gravity

The ratio of the density of given liquid to the density of water at 4°C

Image not available for electronic use. Please refer to the image in the textbook.

Figure 5.9

A hydrometer being used to determine the density of the antifreeze solution in a car's radiator

Top Ten Elements with Highest Densities		
Element	Year Discovered	Density (g/cm^3)
Osmium	1804	22.59
Iridium	1804	22.56
Platinum	1748	21.45
Rhenium	1925	21.01
Neptunium	1940	20.26
Plutonium	1940	20.26
Gold	prehistoric	19.32
Tungsten	1783	19.26
Uranium	1789	19.05
Tantalum	1802	16.67



HANDS-ON CHEMISTRY

• MINI-LAB •

The Density of Clay

Materials

- clay
- student-determined materials

Procedure

1. Obtain a piece of clay from your teacher.
2. Develop a method to determine the density of clay. You will be able to use a balance, a graduated cylinder, and water. Discuss your method with your teacher.
3. Determine the density of your clay.

Results/Analysis

1. Does the amount of clay you use in your determination affect the density you calculate? Why or why not?
2. Does the shape of the clay used in your determination affect the density you calculate? Why or why not?
3. If you make a hollow ball out of clay, will it have the same density as you calculated? Explain.

CELEBRITY CHEMICAL

Carvone (C₁₀H₁₄O)

Carvone is the main component of spearmint oil. It occurs naturally in caraway seeds, dill seeds, gingergrass, and spearmint. A pleasant-smelling liquid at room temperature, this chemical is often used as a flavoring agent in liqueurs and chewing gum and is added to soaps and perfumes to improve their aromas. Carvone is one of the “essential oils” that have been used in spices, perfumes, and medicines for thousands of years.

Carvone is a member of a class of compounds called *terpenes* that are produced by plants, often

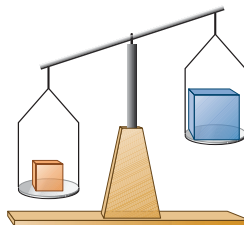
to attract beneficial insects. For example, certain plants give off these compounds when they are attacked by caterpillars. The released terpenes attract the attention of parasitic wasps that kill the caterpillars and inject their eggs. The caterpillar bodies then serve as food for the developing wasp larvae.

A gulf fritillary caterpillar

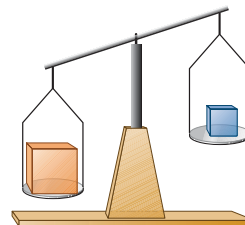


SECTION 5.3 REVIEW QUESTIONS

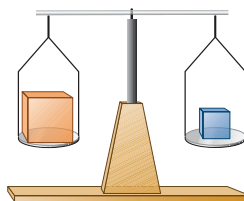
- 1 What is an equivalence statement? How many conversion factors can be created from one equivalence statement? Choose an equivalence statement from Table 5.7, and write all the possible conversion factors from it.
- 2 Write a conversion factor or factors for these problems. Do not do the calculation.
 - a. 37 cm \times _____ = ? in.
 - b. 4.2 qt \times _____ = ? L
 - c. 2.2 kg \times _____ = ? pounds
 - d. 3.5 mi \times _____ = ? m
- 3 Complete the conversions for each of the problems in question 2.
- 4 You run a 100-yd dash in 9.8 seconds. What is your speed in miles per hour?
- 5 Convert the following temperatures.
 - a. 30. °C to °F
 - b. 72 °F to K
 - c. 14 K to °C
 - d. 342 K to °F
- 6 A certain liquid has a density of 0.79 g/mL. What volume of this liquid has a mass of 15.0 g?
- 7 For each of the following, decide which block is more dense: the orange block or the blue block, or state that it cannot be determined. Explain your answers.



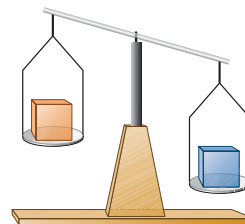
a.



c.



b.



d.

Chapter 5 Review

Key Terms

5.1 Measurement
Scientific notation
Units
English system
Metric system
International System (SI)
Volume
Liter
Milliliter
Mass
Gram

5.2 Significant figures
Rounding off

5.3 Conversion factor
Equivalence statement
Dimensional analysis
Fahrenheit scale
Celsius scale
Kelvin scale (absolute)
Density
Specific gravity

Key Ideas

5.1 Scientific Notation and Units

- A quantitative observation is called a measurement and consists of a number and a unit.
- Very large or very small numbers are conveniently expressed by using scientific notation.
 - The number is expressed as a number between 1 and 10 multiplied by 10 and raised to a power.
- Units provide a scale on which to represent the results of a measurement. There are three commonly used unit systems.
 - English
 - Metric (uses prefixes to change the size of the unit)
 - SI (uses prefixes to change the size of the unit)

5.2 Uncertainty in Measurement and Significant Figures

- All measurements have some uncertainty, which is reflected by the number of significant figures used to express the number.
- Rules exist for rounding off to the correct number of significant figures in a calculated result.

5.3 Problem Solving and Unit Conversions

- We can convert from one system of units to another by a method called dimensional analysis using conversion factors.
- Conversion factors are built from an equivalence statement which shows the relationship between the units in different systems.

English-Metric and English-English Equivalents	
Length	1 m = 1.094 yd 2.54 cm = 1 in. 1 mi = 5280. ft 1 mi = 1760. yd
Mass	1 kg = 2.205 lb 453.6 g = 1 lb
Volume	1 L = 1.06 qt 1 ft ³ = 28.32 L

- There are three commonly used temperature scales: Fahrenheit, Celsius, and Kelvin.
- We can convert among the temperature scales by adjusting the zero point and the size of the unit. Useful equations for conversions are:
 - $^{\circ}T_C + 273 = T_K$
 - $^{\circ}T_F = 1.80(^{\circ}T_C) + 32$
 - Density represents the amount of matter present in a given volume:
density = $\frac{\text{mass}}{\text{volume}}$

All exercises with **blue numbers** have answers in the back of this book.

5.1 Scientific Notation and Units

A. Scientific Notation

- When the number 98,145 is written in standard scientific notation, the exponent indicating the power of 10 will be _____.
- When the number 4.512×10^3 is written in ordinary decimal notation, it is expressed as _____.
- Write each of the following as an “ordinary” decimal number.

a. 6.235×10^{-2}	c. 5.001×10^{-6}
b. 7.229×10^3	d. 8.621×10^4
- For each of the following numbers, if the number is rewritten in standard scientific notation, what will be the value of the exponent (for the power of 10)?

a. 0.000067	c. 1/10,000
b. 9,331,442	d. 163.1×10^2
- Express each of the following numbers in *standard* scientific notation.

a. 9,367,421	e. 6.24×10^2
b. 7241	f. 6319×10^{-2}
c. 0.0005519	g. 0.000000007215
d. 5.408	h. 0.721
- Express each of the following as an “ordinary” decimal number.

a. 4.83×10^2	g. 9.999×10^3
b. 7.221×10^{-4}	h. 1.016×10^{-5}
c. 6.1×10^0	i. 1.016×10^5
d. 9.11×10^{-8}	j. 4.11×10^{-1}
e. 4.221×10^6	k. 9.71×10^4
f. 1.22×10^{-3}	l. 9.71×10^{-4}
- Write each of the following numbers in *standard* scientific notation.

a. 142.3×10^3	e. 0.0251×10^4
b. 0.0007741×10^{-9}	f. $97,522 \times 10^{-3}$
c. 22.7×10^3	g. 0.0000097752×10^6
d. 6.272×10^{-5}	h. $44,252 \times 10^4$

- Write each of the following numbers in standard scientific notation.

a. $\frac{1}{0.00032}$	e. $\frac{(10^5)(10^4)(10^{-4})}{(10^{-2})}$
b. $\frac{10^3}{10^{-3}}$	f. $\frac{43.2}{(4.32 \times 10^{-5})}$
c. $\frac{10^3}{10^3}$	g. $\frac{(4.32 \times 10^{-5})}{432}$
d. $\frac{1}{55,000}$	h. $\frac{1}{(10^5)(10^{-6})}$

B. Units

- Indicate the meaning (as a power of 10) for each of the following metric prefixes.

a. kilo	d. deci
b. centi	e. nano
c. milli	f. micro
- Give the metric prefix that corresponds to each of the following:

a. 1,000,000	d. 10^6
b. 10^{-3}	e. 10^{-2}
c. 10^{-9}	f. 0.000001

C. Measurements of Length, Volume, and Mass

- Which distance is greater, 100 mi or 100 km?
- One liter of volume in the metric system is approximately equivalent to one _____ in the English system.
- The length 52.2 mm can also be expressed as _____ cm.
- Who is taller, a man who is 1.62 m tall or a woman who is 5 ft 6 in. tall?
- A 1-kg package of hamburger has a mass closest to which of the following?

a. 8 oz	c. 2 lb
b. 1 lb	d. 10 lb
- A 2-L bottle of soda contains a volume closest to which of the following?

a. 5 gal	c. 2 pt
b. 1 qt	d. 2 qt
- A recipe written in metric units calls for 250 mL of milk. Which of the following best approximates this amount?

a. 1 qt	c. 1 cup
b. 1 gal	d. 1 pt

18. Which metric system unit is most appropriate for measuring the distance between two cities?

- a. meters c. centimeters
b. millimeters d. kilometers

For exercises 19 and 20, some examples of simple approximate metric to English equivalents are given in the table.

Examples of Commonly Used Units	
length	A dime is 1 mm thick. A quarter is 2.5 cm in diameter. The average height of an adult man is 1.8 m.
mass	A nickel has a mass of about 5 g. A 120-lb woman has a mass of about 55 kg.
volume	A 12-oz can of soda has a volume of about 360 mL. A half gallon of milk is equal to about 2 L of milk.

19. What is the value in dollars of a stack of dimes that is 10 cm high?

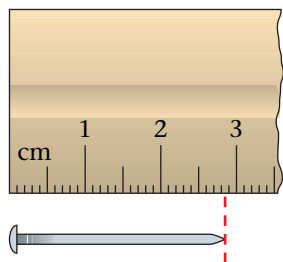
20. How many quarters must be lined up in a row to reach a length of 1 meter?

5.2 Uncertainty in Measurement and Significant Figures

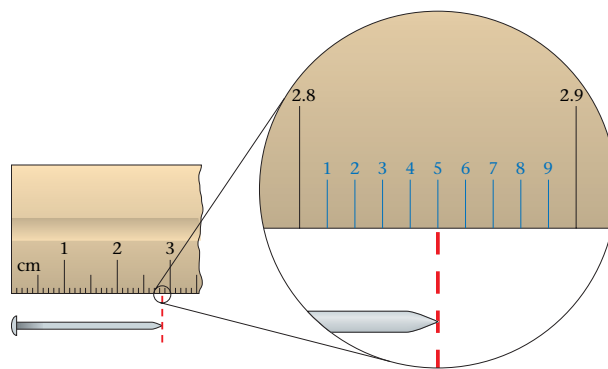
A. Uncertainty in Measurement

21. When a measuring scale is used properly to the limit of precision, the last significant digit recorded for the measurement is said to be *uncertain*. Explain.

22. For the pin shown below, why is the third digit determined for the length of the pin uncertain? Considering that the third digit is uncertain, explain why the length of the pin is indicated as 2.85 cm rather than, for example, 2.83 or 2.87 cm.



23. Why can the length of the pin shown below not be recorded as 2.850 cm?



B. Significant Figures

24. Indicate the number of significant figures in each of the following:

- a. 1422 e. 200.
b. 65,321 f. 2.00×10^2
c. 1.004×10^5 g. 435.662
d. 200 h. 56.341

25. Indicate the number of significant figures implied in each of the following statements:

- a. The population of the United States is 250 million.
b. One hour is equivalent to 60 minutes.
c. There are 5280 feet in 1 mile.
d. Jet airliners fly at 500 mi/h.
e. The Daytona 500 is a 500-mile race.

26. Round off each of the following numbers to three significant digits.

- a. 1,566,311 d. 0.0011672
b. 2.7651×10^{-3} e. 0.07759
c. 84,592

27. Round off each of the following numbers to the indicated number of significant digits, and write the answer in standard scientific notation.

- a. 0.00034159 to three digits
b. 103.351×10^2 to four digits
c. 17.9915 to five digits
d. 3.365×10^5 to three digits

28. When the calculation $(0.0043)(0.0821)(298)$ is performed, the answer should be reported to _____ significant figures.

- 29.** The quotient $\frac{(2.3733 \times 10^2)}{(343)}$ should be written with _____ significant figures.
- 30.** How many digits after the decimal point should be reported when the calculation $(199.0354 + 43.09 + 121.2)$ is performed?
- 31.** How many digits after the decimal point should be reported when the calculation $(10,434 - 9.3344)$ is performed?
- 32.** Evaluate each of the following, and write the answer to the appropriate number of significant figures.
- $212.2 + 26.7 + 402.09$
 - $1.0028 + 0.221 + 0.10337$
 - $52.331 + 26.01 - 0.9981$
 - $(2.01 \times 10^2) + (3.014 \times 10^3)$
- 33.** Evaluate each of the following, and write the answer to the appropriate number of significant figures.
- $(4.031)(0.08206)(373.1)/(0.995)$
 - $(12.011)/(6.022 \times 10^{23})$
 - $(0.500)/(44.02)$
 - $(0.15)/(280.62)$
- 34.** Evaluate each of the following, and write the answer to the appropriate number of significant figures.
- $(2.0944 + 0.0003233 + 12.22)/(7.001)$
 - $[(1.42 \times 10^2) + (1.021 \times 10^3)]/(3.1 \times 10^{-1})$
 - $(9.762 \times 10^{-3})/[(1.43 \times 10^2) + (4.51 \times 10^1)]$
 - $(6.1982 \times 10^{-4})^2$

5.3 Problem Solving and Unit Conversions

A. Tools for Problem Solving

- 35.** How many significant figures are understood for the numbers in the following definition:
1 mi = 5280 ft?
- 36.** Given that 1 mi = 1760 yd, determine what conversion factor is appropriate to convert 1849 yd to miles and to convert 2.781 mi to yards.
- For exercises 37 and 38, apples cost \$0.79 per pound.
- 37.** What conversion factor is appropriate to express the cost of 5.3 lb of apples?

- 38.** What conversion factor could be used to determine how many pounds of apples could be bought for \$2.00?
- 39.** Perform each of the following conversions, being sure to set up clearly the appropriate conversion factor in each case.
- 2.23 m to yards
 - 46.2 yd to meters
 - 292 cm to inches
 - 881.2 in. to centimeters
 - 1043 km to miles
 - 445.5 mi to kilometers
 - 36.2 m to kilometers
 - 0.501 km to centimeters
- 40.** Perform each of the following conversions, being sure to set up clearly the appropriate conversion factor in each case.
- 254.3 g to kilograms
 - 2.75 kg to grams
 - 2.75 kg to pounds
 - 2.75 kg to ounces
 - 534.1 g to pounds
 - 1.75 lb to grams
 - 8.7 oz to grams
 - 45.9 g to ounces
- 41.** If \$1.00 is equivalent to 1.2 euros, what is \$20.00 worth in euros? What is the value in dollars of 100 euros?
- 42.** Boston and New York City are 190 miles apart. What is this distance in kilometers? In meters? In feet?
- 43.** The United States will soon have high-speed trains running between Boston and New York capable of speeds up to 160 mi/h. Will these trains be faster or slower than the fastest trains in the United Kingdom, which reach speeds of 225 km/h?

B. Temperature Conversions

- 44.** The normal boiling point of water is _____ °F, or _____ °C.
- 45.** On both the Celsius and Kelvin temperature scales, there are _____ degrees between the normal freezing and boiling points of water.

- 46.** On which temperature scale ($^{\circ}\text{F}$, $^{\circ}\text{C}$, or K) does 1 degree represent the smallest change in temperature?
- 47.** Convert the following temperatures to kelvins.
- $-155\text{ }^{\circ}\text{C}$
 - $200\text{ }^{\circ}\text{C}$
 - $-52\text{ }^{\circ}\text{C}$
 - $101\text{ }^{\circ}\text{F}$
 - $-52\text{ }^{\circ}\text{F}$
 - $-196\text{ }^{\circ}\text{F}$
- 48.** Convert the following Kelvin temperatures to Celsius degrees.
- 275 K
 - 445 K
 - 0 K
 - 77 K
 - $10,000\text{ K}$
 - 2 K
- 49.** Carry out the indicated temperature conversions.
- $-40\text{ }^{\circ}\text{C}$ to Fahrenheit degrees
 - $-40\text{ }^{\circ}\text{F}$ to Celsius degrees
 - 232 K to Celsius degrees
 - 232 K to Fahrenheit degrees
- 50.** Carry out the indicated temperature conversions.
- $-201\text{ }^{\circ}\text{F}$ to kelvins
 - $-201\text{ }^{\circ}\text{C}$ to kelvins
 - $351\text{ }^{\circ}\text{C}$ to Fahrenheit degrees
 - $-150\text{ }^{\circ}\text{F}$ to Celsius degrees

C. Density

- 51.** The ratio of an object's mass to its _____ is called the *density* of the object.
- 52.** A kilogram of lead occupies a much smaller volume than a kilogram of water because _____ has a much higher density.
- 53.** Typically, gases have very (high/low) densities compared with those of solids and liquids (see Table 5.8).
- 54.** What property of density makes it useful as an aid in identifying substances?
- 55.** Referring to Table 5.8, determine whether air, water, ethanol, or aluminum is the most dense.
- 56.** For the masses and volumes indicated, calculate the density in grams per cubic centimeter.
- mass = 4.53 kg ; volume = 225 cm^3
 - mass = 26.3 g ; volume = 25.0 mL
 - mass = 1.00 lb ; volume = $500.\text{ cm}^3$
 - mass = 352 mg ; volume = 0.271 cm^3

- 57.** If 89.2 mL of a liquid has a mass of 75.2 g , calculate the liquid's density.
- 58.** A cube of metal weighs 1.45 kg and displaces 542 mL of water when immersed. Calculate the density of the metal.
- 59.** A material will float on the surface of a liquid if the material has a density less than that of the liquid. Given that the density of water is approximately 1.0 g/mL , will a block of material having a volume of $1.2 \times 10^4\text{ in.}^3$ and weighing 3.5 lb float or sink when placed in a reservoir of water?
- 60.** The density of pure silver is 10.5 g/cm^3 at $20\text{ }^{\circ}\text{C}$. If 5.25 g of pure silver pellets is added to a graduated cylinder containing 11.2 mL of water, to what volume level will the water in the cylinder rise?

Use the information in the table for questions 61–62.

Densities of Various Common Substances at $20\text{ }^{\circ}\text{C}$		
Substance	Physical State	Density (g/cm^3)
benzene	liquid	0.880
salt (sodium chloride)	solid	2.16
iron	solid	7.87
silver	solid	10.5
mercury	liquid	13.6
gold	solid	19.32

- 61.** Calculate the volume of 50.0 g of each of the following substances.
- sodium chloride
 - mercury
 - benzene
 - silver
- 62.** Calculate the mass of 50.0 cm^3 of each of the following substances.
- gold
 - iron
 - lead
 - aluminum

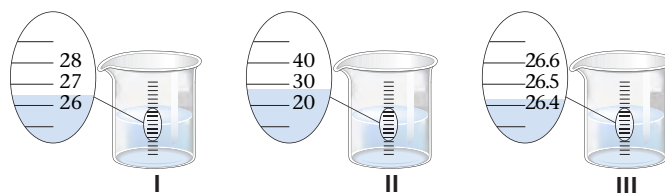
Critical Thinking

- 63.** Which unit of length in the metric system would be most appropriate in size for measuring each of the following items?
- the dimensions of this page
 - the size of the room in which you are sitting
 - the distance from New York to London
 - the diameter of a baseball
 - the diameter of a common pin
- 64.** Suppose that your car is rated at 45 mi/gal for highway use and 38 mi/gal for city driving. If you wanted to write to your friend in Spain about your car's mileage, what ratings in kilometers per liter would you report?
- 65.** You are in Paris, and you want to buy some peaches for lunch. The sign in the fruit stand indicates that peaches are 2.76 Euros per kilogram. Given that there are approximately 11.2 Euros to the dollar, calculate what a pound of peaches will cost in dollars.
- 66.** Suppose a sample of antibiotic pills has a mass of 15.6 g. If the average mass per pill is 0.65 g, determine how many pills are in the sample.
- 67.** For a material to float on the surface of water, the material must have a density less than that of water (1.0 g/mL) and must not react with the water or dissolve in it. A spherical ball has a radius of 0.50 cm and weighs 2.0 g. Will this ball float or sink when placed in water?
(Note: Volume of a sphere = $\frac{4}{3}\pi r^3$)
- 68.** Ethanol and benzene dissolve in each other. When 100. mL of ethanol is dissolved in 1.00 L of benzene, what is the mass of the mixture? (See Table 5.8.)
- 69.** For each of the following numbers, by how many places must the decimal point be moved to express the number in standard scientific notation? In each case, will the exponent be positive, negative, or zero?
- 55,651
 - 0.000008991
 - 2.04
 - 883,541
 - 0.09814
- 70.** Which weighs more, 4.25 grams of gold or 425 milligrams of gold?
- 71.** Indicate the number of significant figures in each of the following:
- This book contains over 500 pages.
 - A mile is just over 5000 ft.
 - A liter is equivalent to 1.059 qt.
 - The population of the United States is approaching 250 million.
 - A kilogram is 1000 g.
 - The Boeing 747 cruises at around 600 mi/h.
- 72.** An organic solvent has density 1.31 g/mL. What volume is occupied by 50.0 g of the liquid?
- 73.** A sample containing 33.42 g of metal pellets is poured into a graduated cylinder initially containing 12.7 mL of water, causing the water level in the cylinder to rise to 21.6 mL. Calculate the density of the metal.
- 74.** You place 10.00 mL of water in a graduated cylinder and add a 15.0 g sample of iron to the water. The new water level is 11.91 mL. What is the density of the iron?
- 75.** You are a passenger in a car traveling at 55 mph (miles per hour). How many feet do you travel in 1 second? There are 5280 feet in a mile.

- 1 Express 1,870,000 in scientific notation.
- A 5.49×10^{-8}
 B 1.87×10^{-6}
 C 1.87×10^6
 D 187×10^6
- 2 The number 0.00003007 expressed in exponential notation is
- A 3.007×10^{-5}
 B 3.7×10^{-5}
 C 3.007×10^{-5}
 D 3.007×10^{-4}
- 3 The SI prefix that corresponds to a factor of 10^{-3} is
- A kilo
 B deci
 C centi
 D milli
- 4 The number of milliliters in 0.0367 liter is
- A 3.67×10^{-5} mL
 B 36.7 mL
 C 3.67 mL
 D 367 mL
- 5 Consider the numbers 23.68 and 4.12. The sum of these numbers has ___ significant figures, and the product of these numbers has ___ significant figures.
- A 3, 3
 B 4, 4
 C 3, 4
 D 4, 3
- 6 An equivalence statement is written so that on each side of the equal sign
- A there are the same numbers and same units.
 B there are the same numbers but different units.
 C there are generally different numbers and different units.
 D there are generally different numbers but the same units.

- 7 Which of the following statements is true?
- A A change of 1°C is larger than a change of 1°F , but smaller than a change of 1 K.
 B A change of 1°F is larger than a change of 1°C , but smaller than a change of 1 K.
 C A change of 1°C is larger than a change of 1°F , and equal to a change of 1 K.
 D A change of 1°F is larger than a change of 1°C , and equal to a change of 1 K.
- 8 A 15.0 mL sample of a liquid has a mass of 12.3 g. What is the density of the liquid?
- A 0.820 g/mL
 B 1.22 g/mL
 C 185 g/mL
 D 12.3 g/mL

- 9 The beakers shown below have different precisions as shown.



- a. Report the volume of water in each beaker with the proper number of significant figures.
- b. Label the digits in each measurement from part A as “certain” or “uncertain.”
- c. You pour the water from these three beakers into one container. What is the volume of liquid in the container reported to the correct number of significant figures?

LOOKING AHEAD

6.1 Atoms and Moles

A. Counting by Weighing

- Averaging the Mass of Similar Objects
- Averaging the Mass of Different Objects
- A Closer Look

B. Atomic Masses: Counting Atoms by Weighing

- Calculating Mass

C. The Mole

- Masses with the Same Number of Atoms
- Defining the Mole
- Using the Mole
- A Closer Look

6.2 Molar Mass and Percent Composition

A. Molar Mass

B. Percent Composition of Compounds

6.3 Formulas of Compounds

A. Empirical Formulas

- A Closer Look
- Finding the Formula

B. Calculation of Empirical Formulas

C. Calculation of Molecular Formulas

Pearls are made of layers of calcium carbonate. They can be measured by counting or by weighing.



IN YOUR LIFE

One very important chemical activity is the synthesis of new substances. Nylon, the artificial sweetener Splenda, Kevlar used in bulletproof vests and the body parts of exotic cars, polyvinyl chloride (PVC) for plastic water pipes, Teflon, and so many other materials that make our lives easier all originated in some chemist's laboratory. When a chemist makes a new substance, the first order of business is to identify it. What is its composition? What is its chemical formula?

The body of this sports car is made from new lightweight compounds to make it go faster on less fuel.

**WHAT DO YOU KNOW?** **Prereading Questions**

1. Explain how to find the average of the following data set: 2, 8, 6, 4, 2, 4, 5.
2. What units are used for mass measurements?
3. Consider a sample of pennies with a mass of 857.2 g. The average mass of a penny is 2.73 g. How many pennies are in the sample?
4. Consider a 12.01 g sample of carbon. The average mass of a carbon atom is 1.994×10^{23} g. How many carbon atoms are in the sample?
5. In a class of 30 chemistry students, 17 of the students are female. What percentage of students is female?

SECTION 6.1

Atoms and Moles

Key Terms

- Atomic mass unit
- Average atomic mass
- Mole
- Avogadro's number

Objectives

- To understand the concept of average mass
- To learn how counting can be done by weighing
- To understand atomic mass and learn how it is determined
- To understand the mole concept and Avogadro's number
- To learn to convert among moles, mass, and number of atoms

A. Counting by Weighing

In this chapter we will learn to determine a compound's formula. Before we can do that, however, we need to think about counting atoms. How do we determine the number of each type of atom in a substance so that we can write its formula? Of course, atoms are too small to count individually. As we will see in this chapter, we typically count atoms by weighing them. So let us first consider the general principle of counting by weighing.

Suppose you work in a candy store that sells gourmet jelly beans by the bean. People come in and ask for 50 beans, 100 beans, 1000 beans, and so on, and you have to count them out—a tedious process at best. As a good problem solver, you try to come up with a better system. It occurs to you that it might be far more efficient to buy a scale and count the jelly beans by weighing them. How can you count jelly beans by weighing them? What information about the individual beans do you need to know?

Averaging the Mass of Similar Objects

Assume that all of the jelly beans are identical and that each has a mass of 5 g. If a customer asks for 1000 jelly beans, what mass of jelly beans would be required? Each bean has a mass of 5 g, so you would need $1000 \text{ beans} \times 5 \text{ g/bean}$, or 5000 g (5 kg). It takes just a few seconds to weigh out 5 kg of jelly beans. It would take much longer to count out 1000 of them.

In reality, jelly beans are not identical. For example, let's assume that you weigh 10 beans individually and get the following results:

Bean	1	2	3	4	5	6	7	8	9	10
Mass (g)	5.1	5.2	5.0	4.8	4.9	5.0	5.0	5.1	4.9	5.0

Can we count these nonidentical beans by weighing? Yes. The key piece of information we need is the *average* mass of the jelly beans. Let's compute the average mass for our 10-bean sample.

$$\begin{aligned}\text{Average mass} &= \frac{\text{total mass of beans}}{\text{number of beans}} \\ &= \frac{5.1 \text{ g} + 5.2 \text{ g} + 5.0 \text{ g} + 4.8 \text{ g} + 4.9 \text{ g} + 5.0 \text{ g} + 5.0 \text{ g} + 5.1 \text{ g} + 4.9 \text{ g} + 5.0 \text{ g}}{10} \\ &= \frac{50.0}{10} \text{ g} = 5.0 \text{ g}\end{aligned}$$

The average mass of a jelly bean is 5.0 g.



Jellybeans can be counted by weighing.

MATH



To find an average, add up all the individual measurements and divide by the total number of measurements.



So, to count out 1000 beans, we need to weigh out 5000 g of beans. This sample of beans, in which the beans have an average mass of 5.0 g, can be treated exactly like a sample where all of the beans are identical.

Objects do not need to have identical masses to be counted by weighing. We simply need to know the average mass of the objects.

For purposes of counting, the objects *behave as though they were all identical*, as though they each actually had the average mass.

Averaging the Mass of Different Objects

Suppose a customer comes into the store and says, “I want to buy a bag of candy for each of my kids. One of them likes jelly beans and the other one likes mints. Please put a scoopful of jelly beans in a bag and a scoopful of mints in another bag.” Then the customer recognizes a problem. “Wait! My kids will fight unless I bring home exactly the same number of candies for each one. Both bags must have the same number of pieces because they’ll definitely count them and compare. But I’m really in a hurry, so we don’t have time to count them here. Is there a simple way you can be sure the bags will contain the same number of candies?”

You need to solve this problem quickly. Suppose you know the average masses of the two kinds of candy:

Jelly beans: average mass = 5 g

Mints: average mass = 15 g

You fill the scoop with jelly beans and dump them onto the scale, which reads 500 g. Now the key question:

What mass of mints do you need to give the same number of mints as there are jelly beans in 500 g of jelly beans?

Comparing the average masses of the mints (15 g) and jelly beans (5 g), you realize that each mint has three times the mass of each jelly bean:

$$\frac{\text{average mass of mints}}{\text{average mass of jelly beans}} = \frac{15 \text{ g}}{5 \text{ g}} = 3$$

This means that you must weigh out an amount of mints that is three times the mass of the jelly beans:

$$3 \times 500 \text{ g} = 1500 \text{ g}$$

You weigh out 1500 g of mints and put them in a bag. The customer leaves with your assurance that both the bag containing 500 g of jelly beans and the bag containing 1500 g of mints contain the same number of candies.

In solving this problem, you have discovered a principle that is very important in chemistry:

Two samples containing different types of components, A and B, both contain the same number of components if the ratio of the sample masses is the same as the ratio of the masses of the individual components of A and B.









Active Reading Question

What information is needed in order to count by weighing?

A Closer Look

Let's look at the relationship between mass and number of components more closely by using our candy example. The individual candies have the masses 15 g (mints) and 5 g (jelly beans). Consider several cases.

1 Candy	10 Candies	100 Candies
15 g =  $(1 \text{ mint}) \times \left(\frac{15 \text{ g}}{\text{mint}}\right) = 15 \text{ g}$	150 g =  10 mints $(10 \text{ mints}) \times \left(\frac{15 \text{ g}}{\text{mint}}\right) = 150 \text{ g}$	1500 g =  100 mints $(100 \text{ mints}) \times \left(\frac{15 \text{ g}}{\text{mint}}\right) = 1500 \text{ g}$
5 g =  $(1 \text{ jelly bean}) \times \left(\frac{5 \text{ g}}{\text{jelly bean}}\right) = 5 \text{ g}$	50 g =  10 jelly beans $(10 \text{ jelly beans}) \times \left(\frac{5 \text{ g}}{\text{jelly bean}}\right) = 50 \text{ g}$	500 g =  100 jelly beans $(100 \text{ jelly beans}) \times \left(\frac{5 \text{ g}}{\text{jelly bean}}\right) = 500 \text{ g}$

Note in each case that the ratio of the masses is always 3 to 1,

$$\frac{15}{5} = \frac{150}{50} = \frac{1500}{500} = \frac{3}{1}$$

which is the ratio of the masses of the individual components:

$$\text{Mass of mint/mass of jelly bean} = \frac{15}{5} = \frac{3}{1}$$

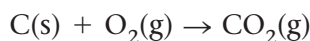
Any two samples, one of mints and one of jelly beans, that have a mass ratio of $15/5 = 3/1$ will contain the same number of components. And these same ideas apply also to atoms, as we will see below.

READING TIP

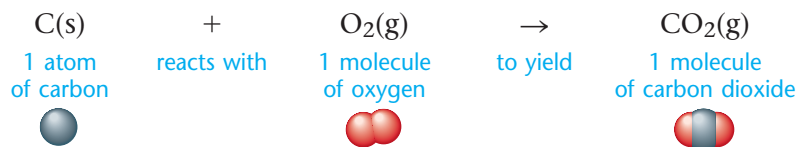
Reading Math The division sign is read aloud as "per." For example, $15 \frac{\text{g}}{\text{mint}}$ is read aloud as "15 grams per mint."

B. Atomic Masses: Counting Atoms by Weighing

The balanced chemical equation for the reaction of solid carbon and gaseous oxygen to form gaseous carbon dioxide is as follows:



Now suppose you have a small pile of solid carbon and want to know how many oxygen molecules are required to convert all of this carbon into carbon dioxide. The balanced equation tells us that one oxygen molecule is required for each carbon atom.



To determine the number of oxygen molecules required, we must know how many carbon atoms are present in the pile of carbon. But individual atoms are far too small to see. We must learn to count atoms by weighing samples containing large numbers of them.

Earlier we saw that we can easily count things like jelly beans and mints by weighing. Exactly the same principles can be applied to counting atoms.

Because atoms are so tiny, the normal units of mass—the gram and the kilogram—are much too large to be convenient. For example, the mass of a single carbon atom is 1.99×10^{-23} g. To avoid using terms like 10^{-23} when describing the mass of an atom, scientists have defined a much smaller unit of mass called the **atomic mass unit**, which is abbreviated **amu**. In terms of grams,

$$1 \text{ amu} = 1.66 \times 10^{-24} \text{ g}$$

Counting Carbon Atoms To count carbon atoms by weighing, we need to know the mass of individual atoms, just as we needed to know the mass of the individual jelly beans. We learned from Chapter 3 that the atoms of a given element exist as isotopes.

The isotopes of carbon are $^{12}_6\text{C}$, $^{13}_6\text{C}$, and $^{14}_6\text{C}$. Any sample of carbon contains a mixture of these isotopes, always in the same proportions. Each of these isotopes has a slightly different mass. Therefore, just as with the non-identical jelly beans, we need to use an average mass for the carbon atoms.

The **average atomic mass** for carbon atoms is 12.01 amu. This means that any sample of carbon from nature can be treated as though it were composed of identical carbon atoms, each with a mass of 12.01 amu.

Calculating Mass

Now that we know the average mass of the carbon atom (12.01 amu), we can count carbon atoms by weighing samples of natural carbon. For example, what mass of natural carbon must we take to have 1000 carbon atoms present?

$$\begin{aligned} \text{Mass of 1000 natural carbon atoms} &= (1000 \text{ atoms}) \frac{12.01 \text{ amu}}{\text{atom}} \\ &= 12,010 \text{ amu} \\ &= 1.201 \times 10^4 \text{ amu} \end{aligned}$$

Using Mass to Count Atoms Now let's assume that when we weigh the pile of natural carbon mentioned earlier, the result is 3.00×10^{20} amu. How many carbon atoms are present in this sample?

We know that an average carbon atom has the mass 12.01 amu, so we can compute the number of carbon atoms by using the equivalence statement

$$1 \text{ carbon atom} = 12.01 \text{ amu}$$

to construct the appropriate conversion factor,

$$\frac{1 \text{ carbon atom}}{12.01 \text{ amu}}$$

Atomic mass unit (amu)

A small unit of mass equal to 1.66×10^{-24} grams

Average atomic mass

The weighted average of the masses of all the isotopes of an element

Information

Remember that 1000 is an exact number here.

The calculation is carried out as follows:

$$\begin{array}{ccccccc} 3.00 \times 10^{20} & \text{amu} & \times & \frac{1 \text{ carbon atom}}{12.01 \text{ amu}} & = & 2.50 \times 10^{19} & \text{carbon atoms} \\ \text{mass of sample} & & \times & \text{conversion} & = & \text{Number of} & \\ & & & \text{factor} & & \text{atoms in sample} & \end{array}$$

The principles we have just discussed for carbon apply to all the other elements as well. All the elements as found in nature typically consist of a mixture of various isotopes. Values for the average masses of the atoms of the elements are listed in Appendix H of this book.

To count the atoms in a sample of a given element by weighing, we must know the mass of the sample and the average mass for that element.

EXAMPLE 6.1

Calculating Mass Using Atomic Mass Units (amu)

Calculate the mass, in amu, of a sample of aluminum that contains 75 atoms.

Solution

Where do we want to go?

75 atoms Al \rightarrow ? amu

What do we know?

- Mass (Al) = 26.98 amu
- 75 atoms Al

How do we get there?

We set up the equivalence statement

$$1 \text{ Al atom} = 26.98 \text{ amu}$$

which gives the conversion factor we need:

$$75 \text{ Al atoms} \times \frac{26.98 \text{ amu}}{1 \text{ Al atom}} = 2024 \text{ amu}$$

Does it make sense?

The correct unit for mass in this case is amu which we have in our answer (2024 amu). Our answer has 4 significant figures as required since the mass is 26.98 and 75 is an exact number.

Practice Problem • Exercise 6.1

Calculate the mass of a sample that contains 23 nitrogen atoms.

The opposite calculation can also be carried out. That is, if we know the mass of a sample, we can determine the number of atoms present. This procedure is illustrated in Example 6.2.

EXAMPLE 6.2**Calculating the Number of Atoms from the Mass**

Calculate the number of sodium atoms present in a sample that has a mass of 1172.49 amu.

Solution**Where do we want to go?**

1172.49 amu \rightarrow atoms Na

What do we know?

- Mass (Na) = 1172.49 amu
- Average atomic mass (Na) = 22.99 amu

How do we get there?

The appropriate equivalence statement is

1 Na atom = 22.99 amu

which gives the conversion factor we need:

$$1172.49 \text{ amu} \times \left(\frac{1 \text{ Na atom}}{22.99 \text{ amu}} \right) = 51.00 \text{ Na atoms}$$

Does it make sense?

The correct unit for our answer in this case is atoms. Our answer (51.00 atoms) has 4 significant figures as required by the average mass of 22.99.

Practice Problem • Exercise 6.2

Calculate the number of oxygen atoms in a sample that has a mass of 288 amu.

**HANDS-ON CHEMISTRY****• MINI-LAB •****Counting Pennies Without Counting****Materials**

- sealed container with pennies
- empty container similar to the sealed container
- ten pennies
- balance

Procedure

1. Obtain a sample of pennies in a sealed container. Do not count the pennies.
2. Obtain an empty container that is similar to the one containing the pennies. Obtain one additional penny. Devise a method using a balance to determine the number of pennies in the sealed container without opening it. Determine the number of pennies in the sealed container.

3. Repeat step 2 using ten additional pennies instead of one penny. Use your method to determine the number of pennies in the sealed container.
4. Open the sealed container and count the pennies.

Results/Analysis

1. Which method (using one penny or ten pennies) allowed you to more accurately determine the number of pennies in the sealed container? Why? How does this finding relate to counting atoms by weighing?

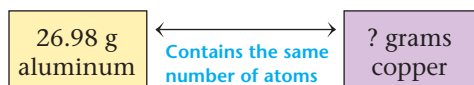
C. The Mole

Earlier we used atomic mass units for mass, because atoms are so tiny. In the laboratory, a much larger unit, the gram, is the convenient unit for mass. Now we will learn to count atoms in samples with masses given in grams.

Masses with the Same Number of Atoms

What mass of copper contains the same number of atoms as 26.98 g of aluminum?

Where do we want to go?



What do we know?

- Our Al sample has a mass of 26.98 g.
- Atomic masses for the elements in this problem:

Al 26.98 amu

Cu 63.55 amu

How do we get there?

- Which atom has the greater atomic mass, aluminum or copper?

Copper

- If we have 26.98 g of Al, do we need more or less than 26.98 g Cu to have the same number of copper atoms as aluminum atoms?

Each copper atom has a greater mass than each aluminum atom. We will need more mass of copper than aluminum for the same number of atoms.

- How much copper do we need?

The average atomic masses for the elements are:

Al 26.98 g

Cu 63.55 g

These masses contain the same number of atoms.

We need 63.55 g of copper.

Does it make sense?

The mass of copper found is greater than the mass of aluminum. This makes sense because a copper atom has a greater mass than an aluminum atom.

Earlier we learned that samples in which the ratio of the masses is the same as the ratio of the masses of the individual atoms always contain the same number of atoms. In the case just considered, the ratios are

$$\frac{26.98\text{g}}{63.55\text{g}} = \frac{26.98\text{ amu}}{63.55\text{ amu}}$$

Ratio of sample masses = Ratio of atomic masses

Therefore, 26.98 g of aluminum contains the same number of aluminum atoms as 63.55 g of copper contains copper atoms.

Active Reading Question

What mass of copper has the same number of atoms as 40.47 g of aluminum?

Defining the Mole

Now compare silver (average atomic mass, 107.9 amu) and lead (average atomic mass, 207.2 amu). A sample of 107.9 g of silver contains the same number of atoms as 207.2 g of lead.

	Silver (Ag)	Lead (Pb)
		
Average atomic mass of this element	107.9 amu	207.2 amu
Mass of this sample	107.9 g	207.2 g
Number of atoms in this sample	6.022×10^{23} atoms	6.022×10^{23} atoms

In fact, if we weigh out samples of all the elements such that each sample has a mass equal to that element's average atomic mass in grams, these samples all contain the same number of atoms. This number (the number of atoms present in all of these samples) assumes special importance in chemistry. It is called the mole, the unit all chemists use in describing numbers of atoms.

The **mole** (abbreviated **mol**) can be defined as *the number equal to the number of carbon atoms in 12.01 grams of carbon*.

Techniques for counting atoms very precisely have been used to determine this number to be 6.022×10^{23} . This number is called **Avogadro's number**.

One mole of something consists of 6.022×10^{23} units of that substance.

Just as a dozen eggs is 12 eggs, a mole of eggs is 6.022×10^{23} eggs. And a mole of water contains 6.022×10^{23} H₂O molecules.

The magnitude of the number 6.022×10^{23} is very difficult to imagine. To give you some idea, 1 mol of seconds represents a span of time 4 million times as long as the earth has already existed! One mole of marbles is enough to cover the entire earth to a depth of 50 miles! However, because atoms are so tiny, a mole of atoms or molecules is a perfectly manageable quantity to use in a reaction.

Mole

The number equal to the number of carbon atoms in exactly 12 grams of pure ¹²C: Avogadro's number. One mole represents 6.022×10^{23} units.

Avogadro's number

The number of atoms in exactly 12 grams of pure ¹²C, equal to 6.022×10^{23}

CRITICAL THINKING ?

What if you were offered \$1 million to count from 1 to 6×10^{23} at a rate of one number each second? Determine your hourly wage. Would you do it?

Using the Mole

How do we use the mole in chemical calculations? Recall that Avogadro's number is defined such that a sample of *any* element that weighs a number of grams equal to the average atomic mass of that element contains 6.022×10^{23} atoms (1 mol) of that element. For example:



Table 6.1 shows the masses of several elements that contain 1 mol of atoms.

A sample of an element with a mass equal to that element's average atomic mass expressed in grams contains 1 mol of atoms.

Information

Avogadro's number (to four significant figures) is 6.022×10^{23} . One mole of anything is 6.022×10^{23} units of that substance.

Table 6.1

Comparison of 1-Mol Samples of Various Elements

Element	Number of Atoms Present	Mass of Sample (g)
Aluminum	6.022×10^{23}	26.98
Gold	6.022×10^{23}	196.97
Iron	6.022×10^{23}	55.85
Sulfur	6.022×10^{23}	32.07
Boron	6.022×10^{23}	10.81
Xenon	6.022×10^{23}	131.3

A Closer Look

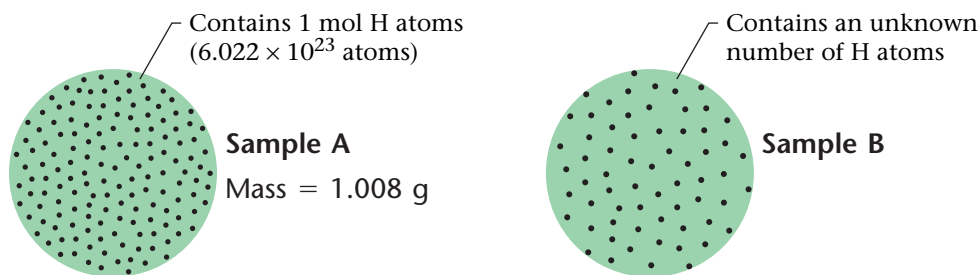
To do chemical calculations, you must

- understand what the mole means
- determine the number of moles in a given mass of a substance

However, before we do any calculations, let's be sure that the process of counting by weighing is clear.

Consider the "bag" of H atoms (symbolized by dots in the green circle), which contains 1 mol (6.022×10^{23}) of H atoms and has a mass of 1.008 g. Assume the bag itself has no mass (**Sample A**).

Now consider another “bag” of hydrogen atoms in which the number of hydrogen atoms is unknown (**Sample B**).



We want to find out how many H atoms are present in Sample (“bag”) B. How can we do that? We can do it by weighing the sample. We find the mass of Sample B to be 0.500 g.

How does this measured mass help us determine the number of atoms in Sample B? We know that 1 mol of H atoms has a mass of 1.008 g. Sample B has a mass of 0.500 g, which is approximately half the mass of a mole of H atoms.

Sample A
Mass = 1.008 g

Sample B
Mass = 0.500 g

Contains 1 mol
of H atoms

Must contain about
1/2 mol of H atoms

We carry out the calculation by using the equivalence statement

$$1 \text{ mol H atoms} = 1.008 \text{ g H}$$

to construct the conversion factor we need:

$$0.500 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.496 \text{ mol H in Sample B}$$

mass of Sample B
×
conversion factor
=
number of moles in Sample B

Let’s summarize. We know the mass of 1 mol of H atoms, so we can determine the number of moles of H atoms in any other sample of pure hydrogen by weighing the sample and *comparing* its mass to 1.008 g (the mass of 1 mol of H atoms). We can follow this same process for any element, because we know the mass of 1 mol for each of the elements.

Also, because we know that 1 mol is 6.022×10^{23} units, once we know the *moles* of atoms present, we can easily determine the *number* of atoms present. In the case considered above, we have approximately 0.5 mol of H atoms in Sample B. This means that about $\frac{1}{2}$ of 6×10^{23} , or 3×10^{23} , H atoms are present. We carry out the actual calculation by using the equivalence statement

$$1 \text{ mol} = 6.022 \times 10^{23}$$

to determine the conversion factor we need:

$$0.496 \text{ mol H atoms} \times \frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H atoms}} = 2.99 \times 10^{23} \text{ H atoms in Sample B}$$

These procedures are illustrated in Example 6.3.

CRITICAL THINKING ?

What if you discovered Avogadro’s number was not 6.02×10^{23} but 3.01×10^{23} ? Would your discovery affect the relative masses given on the periodic table? If so, how? If not, why not?

EXAMPLE 6.3

Calculating Moles and Number of Atoms

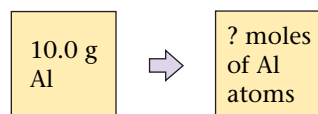


A bicycle with an aluminum frame

Aluminum (Al), a metal with a high strength-to-weight ratio and a high resistance to corrosion, is often used for structures such as high-quality bicycle frames. Compute both the number of moles of atoms and the number of atoms in a 10.0-g sample of aluminum.

Solution

Where do we want to go?



What do we know?

- The mass of 1 mol (6.022×10^{23} atoms) of aluminum is 26.98 g.
- The sample we are considering has a mass of 10.0 g.

How do we get there?

We calculate the number of moles of aluminum atoms in 10.0 g by using the equivalence statement

$$1 \text{ mol Al} = 26.98 \text{ g Al}$$

to construct the appropriate conversion factor,

$$10.0 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 0.371 \text{ mol Al}$$

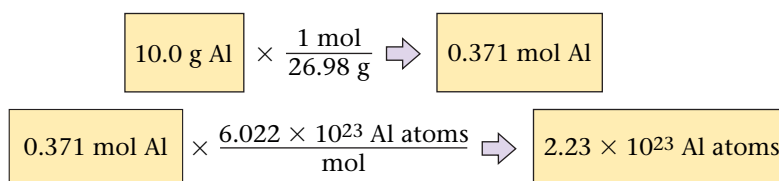
Next we convert from moles of atoms to the number of atoms, using the equivalence statement

$$6.022 \times 10^{23} \text{ Al atoms} = 1 \text{ mol Al atoms}$$

We have

$$0.371 \text{ mol Al} \times \frac{6.022 \times 10^{23} \text{ Al atoms}}{1 \text{ mol Al}} = 2.23 \times 10^{23} \text{ Al atoms}$$

We can summarize this calculation as follows:



Does it make sense?

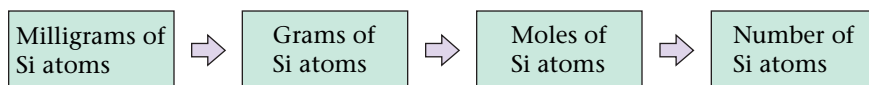
Since 10 g is smaller than 26.98 g (the mass of 1 mol), we should have fewer than 6.022×10^{23} Al atoms. Our answer is 2.23×10^{23} Al atoms which makes sense.

EXAMPLE 6.4**Calculating the Number of Atoms**

A silicon chip used in an integrated circuit of a computer has a mass of 5.68 mg. How many silicon (Si) atoms are present in this chip? The average atomic mass for silicon is 28.09 amu.

Solution**Where do we want to go?**

Our strategy for doing this problem is to convert from milligrams of silicon to grams of silicon, then to moles of silicon, and finally to atoms of silicon:

**What do we know?**

- Our sample has a mass of 5.68 mg.
- The atomic mass for silicon is 28.09 amu.

How do we get there?

Each arrow in the schematic represents a conversion factor. Using $1 \text{ g} = 1000 \text{ mg}$, we can convert the mass to grams.

$$5.68 \text{ mg Si} \times \frac{1 \text{ g Si}}{1000 \text{ mg Si}} = 5.68 \times 10^{-3} \text{ g Si}$$

Next, because the average mass of silicon is 28.09 amu, we know that 1 mol of Si atoms weighs 28.09 g. This leads to the equivalence statement

$$1 \text{ mol Si} = 28.09 \text{ g Si}$$

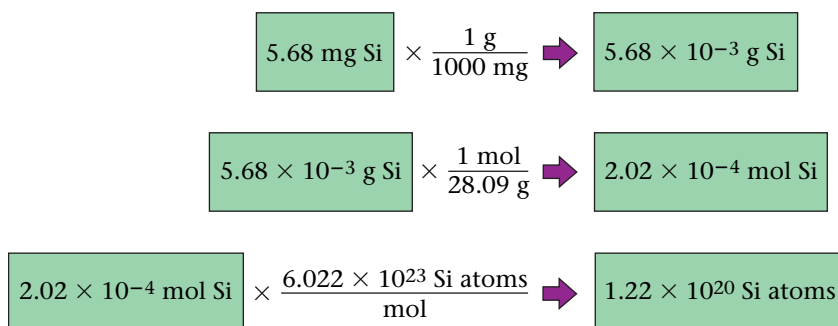
Thus,

$$5.68 \times 10^{-3} \text{ g Si} \times \frac{1 \text{ mol Si}}{28.09 \text{ g Si}} = 2.02 \times 10^{-4} \text{ mol Si}$$

Using the definition of a mole ($1 \text{ mol} = 6.022 \times 10^{23}$), we have

$$2.02 \times 10^{-4} \text{ mol Si} \times \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol Si}} = 1.22 \times 10^{20} \text{ Si atoms}$$

We can summarize this calculation as follows:

**Does it make sense?**

Because 5.68 mg is much smaller than 28.09 g (the mass of 1 mol Si), we expect the number of atoms to be less than 6.022×10^{23} . So our answer 1.22×10^{20} Si atoms makes sense.



A silicon chip of the type used in electronic equipment

Practice Problem • Exercise 6.4

Chromium (Cr) is a metal that is added to steel to improve its resistance to corrosion (for example, to make stainless steel). Calculate both the number of moles in a sample of chromium containing 5.00×10^{20} atoms and the mass of the sample.

information

The values for the average masses of the atoms of the elements are listed in the inside front cover of this book.

Using Units in Problem Solving When you finish a problem, always think about the “reasonableness” of your answer.

In Example 6.4, 5.68 mg of silicon is clearly much less than 1 mol of silicon (which has a mass of 28.09 g), so the final answer of 1.22×10^{20} atoms (compared to 6.022×10^{23} atoms in a mole) at least lies in the right direction. That is, 1.22×10^{20} atoms is a smaller number than 6.022×10^{23} .

Always include the units as you do calculations and make sure the correct units are obtained at the end. Paying careful attention to units and making this type of general check can help you find errors such as an inverted conversion factor or a number that was incorrectly entered into your calculator.

SECTION 6.1 REVIEW QUESTIONS

1 Suppose you work in a hardware store. The manager asks you to fill an order for an important customer who is waiting impatiently. You need 1200 matched sets of nuts and bolts. Unfortunately the nuts and bolts are not boxed—they are loose in big buckets. How can you earn a bonus from your boss and make the customer happy by filling the order in less than 2 minutes?

2 Why is the average atomic mass of any element not a whole number (for example, C, 12 amu; H, 1 amu; N, 14 amu)?

3 A mole of any substance contains Avogadro’s number of units.

- Write an equivalence statement for this definition.
- Write the possible conversion factors from this relationship.
- To determine the following, tell which conversion factor you would need to use:
 - Moles Al from atoms Al
 - Atoms Au from mol Au

4 Copy and complete the following table.

Mass of Sample	Moles of Sample	Atoms in Sample
	0.250 mol Al	
25.4 g Fe		
		2.13×10^{24} atoms Au
	1.28 mol Ca	
4.28 mg Na		
		3.14×10^{23} atoms C

5 Which sample contains more atoms: 3.89 g of nickel or 6.61 g of silver? Defend your answer.



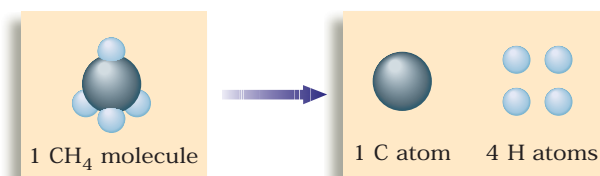
Molar Mass and Percent Composition

Objectives

- To understand the definition of molar mass
- To learn to convert between moles and mass
- To learn to mass percent of an element in a compound

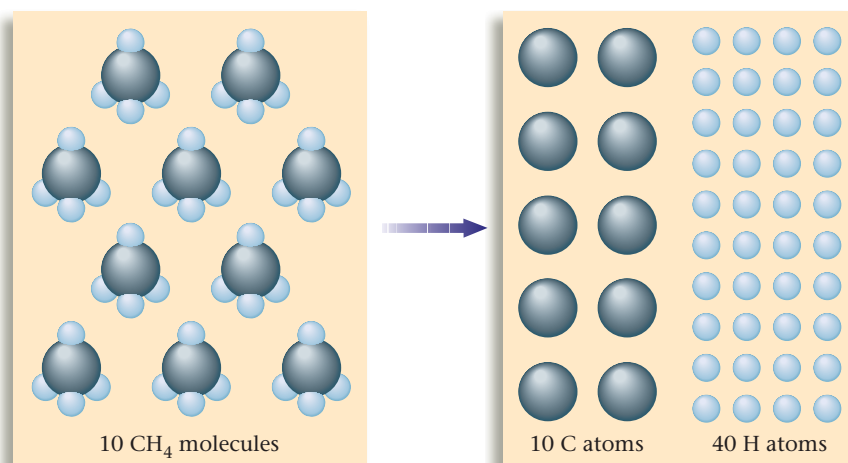
A. Molar Mass

A chemical compound is, fundamentally, a collection of atoms. For example, methane (the major component of natural gas) consists of molecules each containing one carbon atom and four hydrogen atoms (CH_4).



How can we calculate the mass of 1 mol of methane? That is, what is the mass of 6.022×10^{23} CH_4 molecules? Because each CH_4 molecule contains one carbon atom and four hydrogen atoms, 1 mol of CH_4 molecules consists of 1 mol of carbon atoms and 4 mol of hydrogen atoms (see **Figure 6.1**). The mass of 1 mol of methane can be found by summing the masses of carbon and hydrogen present:

$$\begin{aligned} \text{Mass of 1 mol of C} &= 1 \times 12.01 \text{ g} = 12.01 \text{ g} \\ \text{Mass of 4 mol of H} &= 4 \times 1.008 \text{ g} = 4.032 \text{ g} \\ \text{Mass of 1 mol of CH}_4 &= 16.04 \text{ g} \end{aligned}$$



1 mol CH_4 molecules
(6.022×10^{23} CH_4 molecules)

1 mol C atoms
(6.022×10^{23} C atoms)

4 mol H atoms
4 (6.022×10^{23} H atoms)

Figure 6.1
Various numbers of methane molecules showing their constituent atoms

Key Terms

- Molar mass
- Mass percent

Information

Note that when we say 1 mol of methane, we mean 1 mol of methane molecules.

MATH

Remember that the least number of decimal places limits the number of significant figures in addition.

Molar mass

The mass in grams of 1 mole of the substance

The quantity 16.04 g is called the molar mass for methane: the mass of 1 mol of CH_4 molecules. The **molar mass** of any substance is the *mass (in grams) of 1 mol of the substance*.

The molar mass is obtained by summing the masses of the component atoms.

Active Reading Question

Explain how to calculate the molar mass of a compound and give an example.

EXAMPLE 6.5

Calculating Molar Mass

Information

A substance's molar mass (in grams) is the mass of 1 mol of that substance.

Calculate the molar mass of sulfur dioxide, a gas produced when sulfur-containing fuels are burned. Unless "scrubbed" from the exhaust, sulfur dioxide can react with moisture in the atmosphere to produce acid rain.

Solution

Where do we want to go?

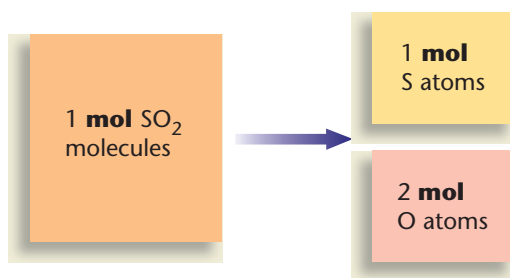
Molar mass (SO_2) \rightarrow ? g/mol

What do we know?

- The formula for sulfur dioxide is SO_2
- Atomic masses
S = 32.07 g/mol
O = 16.00 g/mol

How do we get there?

We need to find the mass of 1 mol of SO_2 molecules—the molar mass for sulfur dioxide.



$$\text{Mass of 1 mol of S} = 1 \times 32.07 = 32.07 \text{ g}$$

$$\text{Mass of 2 mol of O} = 2 \times 16.00 = \underline{32.00 \text{ g}}$$

$$\text{Mass of 1 mol of SO}_2 = 64.07 \text{ g} = \text{molar mass} = 64.07 \text{ g/mol}$$

The molar mass of SO_2 is 64.07 g. It represents the mass of 1 mol of SO_2 molecules.

Does it make sense?

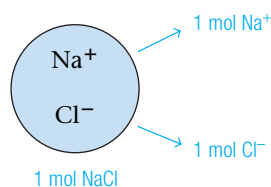
Note the units obtained here (g/mol) are correct for molar mass. The answer is given to two decimal places as required by the atomic masses.

Practice Problem • Exercise 6.5

Polyvinyl chloride, called PVC, which is widely used for floor coverings (“vinyl”) as well as for plastic pipes in plumbing systems, is made from a molecule with the formula C_2H_3Cl . Calculate the molar mass of this substance.

Some substances exist as a collection of ions rather than as separate molecules. For example, ordinary table salt, sodium chloride ($NaCl$), is composed of an array of Na^+ and Cl^- ions. There are no $NaCl$ molecules present. In some books the term *formula weight* is used instead of *molar mass* for ionic compounds. However, in this book we will apply the term *molar mass* to both ionic and molecular substances.

To calculate the molar mass for sodium chloride, we must realize that 1 mol of $NaCl$ contains 1 mol of Na^+ ions and 1 mol of Cl^- ions.



Therefore, the molar mass (in grams) for sodium chloride represents the sum of the mass of 1 mol of sodium ions and the mass of 1 mol of chloride ions.

$$\begin{aligned}\text{Mass of 1 mol of } Na^+ &= 22.99 \text{ g} \\ \text{Mass of 1 mol of } Cl^- &= 35.45 \text{ g} \\ \text{Mass of 1 mol of } NaCl &= 58.44 \text{ g} = \text{molar mass}\end{aligned}$$

The molar mass of $NaCl$ is 58.44 g. It represents the mass of 1 mol of sodium chloride.

EXAMPLE 6.6

Calculating Mass from Moles

Calcium carbonate, $CaCO_3$ (also called calcite), is the principal mineral found in limestone, marble, chalk, pearls, and the shells of marine animals such as clams.

- Calculate the molar mass of calcium carbonate.
- A certain sample of calcium carbonate contains 4.86 mol. What is the mass in grams of this sample?

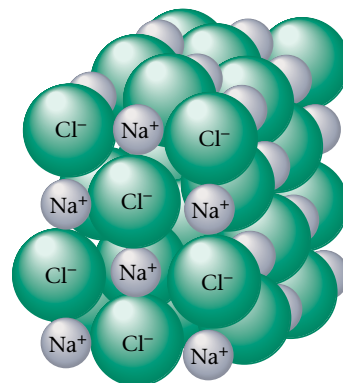
Solution

- a. **Where do we want to go?**

$$\text{Molar mass } (CaCO_3) = ? \text{ g/mol}$$

What do we know?

- The formula for calcium carbonate is $CaCO_3$.
- Atomic masses
 - $Ca = 40.08 \text{ g/mol}$
 - $C = 12.01 \text{ g/mol}$
 - $O = 16.00 \text{ g/mol}$



Information

The mass of the electron is so small that Na^+ and Na have the same mass for our purposes, even though Na^+ has one electron less than Na . Also, the mass of Cl^- virtually equals the mass of Cl even though it has one more electron than Cl .



How do we get there?

Calcium carbonate is an ionic compound composed of Ca^{2+} and CO_3^{2-} ions. One mole of calcium carbonate contains 1 mol of Ca^{2+} and 1 mol of CO_3^{2-} ions. We calculate the molar mass by summing the masses of the components.

$$\text{Mass of 1 mol of Ca}^{2+} = 1 \times 40.08 \text{ g} = 40.08 \text{ g}$$

Mass of 1 mol of CO_3^{2-} (contains 1 mol of C and 3 mol of O):

$$1 \text{ mol of C} = 1 \times 12.01 \text{ g} = 12.01 \text{ g}$$

$$3 \text{ mol of O} = 3 \times 16.00 \text{ g} = \underline{48.00 \text{ g}}$$

$$\text{Mass of 1 mol of CaCO}_3 = 100.09 \text{ g} = \text{molar mass}$$

Does it make sense?

Note the units obtained here (g/mol) are correct for molar mass. The answer is given to two decimal places as required by the atomic masses.

b. Where do we want to go?

$$\text{Mass of 4.86 mol CaCO}_3 = ? \text{ g}$$

What do we know?

- Molar mass $\text{CaCO}_3 = 100.09 \text{ g/mol}$
- Moles $\text{CaCO}_3 = 4.86 \text{ mol}$

How do we get there?

We determine the mass of 4.86 mol of CaCO_3 by using the molar mass,

$$4.86 \text{ mol CaCO}_3 \times \frac{100.09 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} = 486 \text{ g CaCO}_3$$

which can be diagrammed as follows:

$$\boxed{4.86 \text{ mol CaCO}_3} \times \frac{100.09 \text{ g}}{\text{mol}} \Rightarrow \boxed{486 \text{ g CaCO}_3}$$

Does it make sense?

Note that this sample contains nearly 5 mol and thus should have a mass of nearly 500 g (since the molar mass is about 100 g) so our answer makes sense. The number of significant figures in our answer (486 g) is three as required by the initial moles (4.86).

Practice Problem • Exercise 6.6

Calculate the molar mass for sodium sulfate, Na_2SO_4 . A sample of sodium sulfate with a mass of 300.0 g represents what number of moles of sodium sulfate?

information

For average atomic masses, look inside the front cover of this book.

Let's Review

- The molar mass of a substance can be obtained by summing the masses of the component atoms.
- The molar mass (in grams) represents the mass of 1 mol of the substance.
- Once we know the molar mass of a compound, we can compute the number of moles present in a sample of known mass.

EXAMPLE 6.7**Calculating Moles from Mass**

Juglone, a dye known for centuries, is produced from the husks of black walnuts. It is also a natural herbicide (weed killer) that kills off competitive plants around the black walnut tree but does not affect grass and other noncompetitive plants. The formula for juglone is $C_{10}H_6O_3$.

- Calculate the molar mass of juglone.
- A sample of 1.56 g of pure juglone was extracted from black walnut husks. How many moles of juglone does this sample represent?



Juglone, a natural herbicide, is produced from the husks of black walnuts.

Solution**a. Where do we want to go?**

Molar mass (juglone) = ? g/mol

What do we know?

- The formula for juglone is $C_{10}H_6O_3$.
- Atomic masses

C = 12.01 g/mol H = 1.008 g/mol O = 16.00 g/mol

How do we get there?

The molar mass is obtained by summing the masses of the component atoms.

1 mol	→	10 mol C
juglone		6 mol H
$C_{10}H_6O_3$		3 mol O

Mass of 10 mol C = $10 \times 12.01 \text{ g} = 120.1 \text{ g}$

Mass of 6 mol H = $6 \times 1.008 \text{ g} = 6.048 \text{ g}$

Mass of 3 mol O = $3 \times 16.00 \text{ g} = \underline{48.00 \text{ g}}$

Mass of 1 mol $C_{10}H_6O_3$ = 174.1 g = molar mass = 174.1 g/mol

Does it make sense?

Note the units obtained here (g/mol) are correct for molar mass. The answer is given to one decimal place as required by the mass of 10 mol of C (120.1 g).

b. Where do we want to go?

1.56 g juglone = ? mol juglone

What do we know?

- Molar mass (juglone) = 174.1 g/mol
- Mass of juglone sample = 1.56 g

How do we get there?

We can determine the exact fraction of a mole by using the equivalence statement

1 mol = 174.1 g juglone

$$1.56 \text{ g juglone} \times \frac{1 \text{ mol juglone}}{174.1 \text{ g juglone}} = 0.00896 \text{ mol juglone}$$

$$= 8.96 \times 10^{-3} \text{ mol juglone}$$

1.56 g juglone	$\times \frac{1 \text{ mol}}{174.1 \text{ g}}$	\Rightarrow	$8.96 \times 10^{-3} \text{ mol juglone}$
----------------	--	---------------	---

Does it make sense?

The mass of 1 mol of this compound is 174.1 g, so 1.56 g is much less than a mole. Our answer (8.96×10^{-3} mol) has units of mol as expected and has three significant figures as required by the original sample size (1.56 g).

EXAMPLE 6.8

Calculating Number of Molecules



Isopentyl acetate, $C_7H_{14}O_2$, the compound responsible for the scent of bananas, can be produced commercially. Interestingly, bees release about 1 mg (1×10^{-6} g) of this compound when they sting. This attracts other bees, which then join the attack. How many moles and how many molecules of isopentyl acetate are released in a typical bee sting?

Solution

Where do we want to go?

- 1×10^{-6} g $C_7H_{14}O_2 \rightarrow ?$ mol $C_7H_{14}O_2$

What do we know?

- The formula for isopentyl acetate is $C_7H_{14}O_2$.

- Atomic masses

$$C = 12.01 \text{ g/mol} \quad H = 1.008 \text{ g/mol} \quad O = 16.00 \text{ g/mol}$$

- Sample size = 1×10^{-6} g

How do we get there?

- We are given a mass of isopentyl acetate and want the number of molecules, so we must first compute the molar mass.

$$7 \text{ mol C} \times 12.01 \frac{\text{g}}{\text{mol}} = 84.07 \text{ g C}$$

$$14 \text{ mol H} \times 1.008 \frac{\text{g}}{\text{mol}} = 14.11 \text{ g H}$$

$$2 \text{ mol O} \times 16.00 \frac{\text{g}}{\text{mol}} = \underline{32.00 \text{ g O}}$$

$$\text{Mass of 1 mol } C_7H_{14}O_2 = 130.18 \text{ g} = \text{molar mass}$$

This means that 1 mol of isopentyl acetate (6.022×10^{23} molecules) has a mass of 130.18 g. The molar mass is 130.18 g/mol.

- Next we determine the number of moles of isopentyl acetate in $1 \mu\text{g}$, which is 1×10^{-6} g. To do this we use the equivalence statement

$$1 \text{ mol isopentyl acetate} = 130.18 \text{ g isopentyl acetate}$$

which yields the conversion factor we need:

$$1 \times 10^{-6} \text{ g } C_7H_{14}O_2 \times \frac{1 \text{ mol } C_7H_{14}O_2}{130.18 \text{ g } C_7H_{14}O_2} = 8 \times 10^{-9} \text{ mol } C_7H_{14}O_2$$



Using the equivalence statement $1 \text{ mol} = 6.022 \times 10^{23}$ units, we can determine the number of molecules:

$$8 \times 10^{-9} \text{ mol } C_7H_{14}O_2 \times \frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{ mol } C_7H_{14}O_2} = 5 \times 10^{15} \text{ molecules}$$

This very large number of molecules is released in each bee sting.

Does it make sense?

The mass of $C_7H_{14}O_2$ released in each sting (1×10^{-6} g) is much less than the mass of one mole of $C_7H_{14}O_2$ so the number of moles (8×10^{-9}) is much smaller than one and the number of molecules (5×10^{15}) is much smaller than 6×10^{23} . The number of significant figures in our answer is 1 as determined by the mass of $C_7H_{14}O_2$ given (1×10^{-6} g).



Practice Problem • Exercise 6.8

The substance Teflon, the nonstick coating on many frying pans, is made from the C_2F_4 molecule. Calculate the number of C_2F_4 units present in 135 g of Teflon.



HANDS-ON CHEMISTRY

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Relative Masses

Materials

- cotton balls
- paper clips
- rubber stoppers
- balance

Procedure

1. Obtain cotton balls, paper clips, and rubber stoppers from your teacher.
2. Devise a method to find the average mass of each of these objects. Discuss this method with your teacher.
3. Determine the average mass of each object.

Results/Analysis

1. Copy the table onto your paper.
2. Record the average masses of each of the objects in your table.

	Mass of one object	Relative mass of one object	Mass of n objects
Cotton balls	_____ g		_____ lb
Paper clips	_____ g		_____ lb
Rubber stoppers	_____ g		_____ lb

3. Give the lightest object a relative mass of 1.0 in your table. Determine the relative masses of the remaining objects and record these in your table.
4. How many of the lightest object would you need to have a pound of that object? Call this number n .
5. If you had n of each of the other objects, how much would each sample weigh? Fill these numbers into your table.
6. Determine which columns correspond to the chemical terms "atomic mass" and "molar mass."
7. Which number represents "Avogadro's number"?

MATH



$$\text{Percent} = \frac{\text{Part}}{\text{Whole}} \times 100\%$$

B. Percent Composition of Compounds

So far we have discussed the composition of compounds in terms of the numbers of each type of atom. It is often useful to know a compound's composition in terms of the *masses* of its elements. We can obtain this information from the formula of the compound by comparing the mass of each element present in 1 mol of the compound to the total mass of 1 mol of the compound. The mass fraction for each element is calculated as follows:

$$\text{Mass fraction for a given element} = \frac{\text{mass of the element present in 1 mol of compound}}{\text{mass of 1 mol of compound}}$$

The mass fraction is converted to *mass percent* by multiplying by 100%.

We will illustrate this concept using the compound ethanol, an alcohol obtained by fermenting the sugar in grapes, corn, and other fruits and grains. Ethanol is often added to gasoline as an octane enhancer to form a fuel called gasohol. The added ethanol has the effect of increasing the octane of the gasoline and also lowering the carbon monoxide in automobile exhaust.

Note from its formula ($\text{C}_2\text{H}_5\text{OH}$) that each molecule of ethanol contains two carbon atoms, six hydrogen atoms, and one oxygen atom.



We calculate the mass of each element present and the molar mass for ethanol as follows:

$$\text{Mass of C} = 2 \text{ mol} \times 12.01 \frac{\text{g}}{\text{mol}} = 24.02 \text{ g}$$

$$\text{Mass of H} = 6 \text{ mol} \times 1.008 \frac{\text{g}}{\text{mol}} = 6.048 \text{ g}$$

$$\text{Mass of O} = 1 \text{ mol} \times 16.00 \frac{\text{g}}{\text{mol}} = 16.00 \text{ g}$$

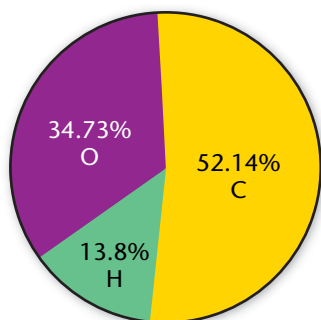
$$\text{Mass of 1 mol of } \text{C}_2\text{H}_5\text{OH} = 46.07 \text{ g} = \text{molar mass}$$

The **mass percent** (sometimes called the weight percent) of carbon in ethanol can be computed by comparing the mass of carbon in 1 mol of ethanol with the total mass of 1 mol of ethanol and multiplying the result by 100%.

$$\begin{aligned} \text{Mass percent of C} &= \frac{\text{mass of C in 1 mol } \text{C}_2\text{H}_5\text{OH}}{\text{mass of 1 mol } \text{C}_2\text{H}_5\text{OH}} \times 100\% \\ &= \frac{24.02 \text{ g}}{46.07 \text{ g}} \times 100\% = 52.14\% \end{aligned}$$

That is, ethanol contains 52.14% by mass of carbon. The mass percents of hydrogen and oxygen in ethanol are obtained in a similar manner.

$$\begin{aligned} \text{Mass percent of H} &= \frac{\text{mass of H in 1 mol } \text{C}_2\text{H}_5\text{OH}}{\text{mass of 1 mol } \text{C}_2\text{H}_5\text{OH}} \times 100\% \\ &= \frac{6.048 \text{ g}}{46.07 \text{ g}} \times 100\% = 13.13\% \end{aligned}$$



Information

The formula for ethanol is written $\text{C}_2\text{H}_5\text{OH}$, although you might expect it to be written simply as $\text{C}_2\text{H}_6\text{O}$.

$$\begin{aligned}\text{Mass percent of O} &= \frac{\text{mass of O in 1 mol C}_2\text{H}_5\text{OH}}{\text{mass of 1 mol C}_2\text{H}_5\text{OH}} \times 100\% \\ &= \frac{16.00 \text{ g}}{46.07 \text{ g}} \times 100\% = 34.73\%\end{aligned}$$

The mass percentages of all the elements in a compound add up to 100%, although rounding-off effects may produce a small deviation. Adding up the percentages is a good way to check the calculations. In this case, the sum of the mass percents is $52.14\% + 13.13\% + 34.73\% = 100.00\%$.

Active Reading Question

In the compound $\text{C}_2\text{H}_5\text{OH}$, the number of hydrogen atoms in a molecule is greater than the number of carbon atoms and oxygen atoms in the molecule. However, the mass percent of hydrogen is lower than the mass percent of carbon and the mass percent of oxygen. Why is this?

MATH

Sometimes, because of rounding-off effects, the sum of the mass percents in a compound is not exactly 100%.

EXAMPLE 6.9

Calculating Mass Percent

Carvone is a substance that occurs in two forms, both of which have the same molecular formula ($\text{C}_{10}\text{H}_{14}\text{O}$) and molar mass. One type of carvone gives caraway seeds their characteristic smell; the other is responsible for the smell of spearmint oil. Compute the mass percent of each element in carvone.

Solution

Where do we want to go?

For carvone

- %C = ? %
- %H = ? %
- %O = ? %

What do we know?

- The formula for carvone is $\text{C}_{10}\text{H}_{14}\text{O}$.
- Atomic masses
C = 12.01 g/mol H = 1.008 g/mol O = 16.00 g/mol

How do we get there?

Because the formula for carvone is $\text{C}_{10}\text{H}_{14}\text{O}$, the masses of the various elements in 1 mol of carvone are

$$\text{Mass of C in 1 mol} = 10 \cancel{\text{ mol}} \times 12.01 \frac{\text{g}}{\cancel{\text{mol}}} = 120.1 \text{ g}$$

$$\text{Mass of H in 1 mol} = 14 \cancel{\text{ mol}} \times 1.008 \frac{\text{g}}{\cancel{\text{mol}}} = 14.11 \text{ g}$$

$$\text{Mass of O in 1 mol} = 1 \cancel{\text{ mol}} \times 16.00 \frac{\text{g}}{\cancel{\text{mol}}} = 16.00 \text{ g}$$

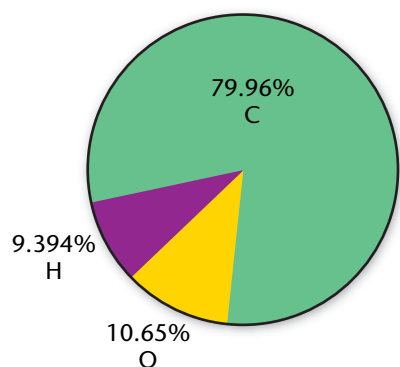
$$\text{Mass of 1 mol of C}_{10}\text{H}_{14}\text{O} = 150.21 \text{ g}$$

$$\text{Molar mass} = 150.2 \text{ g/mol}$$

(rounding to the correct number of significant figures)

MATH

The 120.1 limits the sum to one decimal place. Remember, in addition the number with the fewest decimal places determines the number of significant figures in the answer.



Next we find the fraction of the total mass contributed by each element and convert it to a percentage.

$$\text{Mass percent of C} = \frac{120.1 \text{ g C}}{150.2 \text{ g C}_{10}\text{H}_{14}\text{O}} \times 100\% = 79.96\%$$

$$\text{Mass percent of H} = \frac{14.11 \text{ g H}}{150.2 \text{ g C}_{10}\text{H}_{14}\text{O}} \times 100\% = 9.394\%$$

$$\text{Mass percent of O} = \frac{16.00 \text{ g O}}{150.2 \text{ g C}_{10}\text{H}_{14}\text{O}} \times 100\% = 10.65\%$$

Does it make sense?

Add the individual mass percent values—they should total 100% within a small range due to rounding off. In this case the percentages add up to 100.0%.

Practice Problem • Exercise 6.9

Penicillin, an important antibiotic (antibacterial agent), was discovered accidentally by the Scottish bacteriologist Alexander Fleming in 1928, although he was never able to isolate it as a pure compound. This and similar antibiotics have saved millions of lives that would otherwise have been lost to infections. Penicillin, like many of the molecules produced by living systems, is a large molecule containing many atoms. One type of penicillin, penicillin F, has the formula $\text{C}_{14}\text{H}_{20}\text{N}_2\text{SO}_4$. Compute the mass percent of each element in this compound.



HANDS-ON CHEMISTRY

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And the Winner Is . . .

Materials

- three different types of sandwich cookies
- balance
- aluminum foil
- spoon

Procedure

1. Obtain three different types of sandwich cookies from your teacher.
2. Determine the percent composition of each cookie. Assume the cookie consists of two parts, the crunchy cookie and the cream filling. You will need a balance, foil, and a spoon. Place the foil on the balance before you place any part of the cookie or cream filling on the balance.

Results/Analysis

1. How does the percent composition of each cookie compare?
2. Which would you say was the best cookie? Why do you say this?
3. Your teacher will tell you the price per cookie for each of the three types. Use this information to decide which cookie is the best buy. Defend your answer.



SECTION 6.2

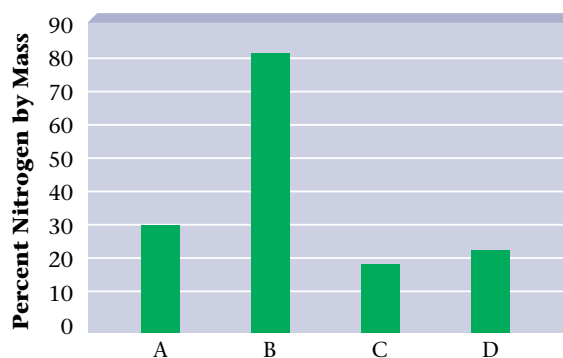
REVIEW QUESTIONS

- Tell how many moles of each element are in the following amounts of each compound:
 - 3 moles NH_3
 - 0.25 moles H_2O
 - 5.0 moles H_2SO_4
 - 0.75 moles $\text{Ca}(\text{NO}_3)_2$
- Determine the molar mass for each of the following:
 - NH_3
 - H_2O
 - H_2SO_4
 - $\text{Ca}(\text{NO}_3)_2$
- Determine the mass of each sample in question 1.
- Explain how you would find the mass percent of each element in water.

- Copy and complete the following table.

Moles	Mass	Number of molecules
4.7 moles NH_3		
	1.80 g H_2O	
		5.21×10^{22} molecules CH_4

- Use the graph to match compounds A, B, C, and D with the following choices: NaCN , HNO_3 , NH_3 , $\text{Ca}(\text{NO}_3)_2$.



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SECTION 6.3

Formulas of Compounds

Key Terms

- Empirical formula
- Molecular formula

Objectives

- To understand the meaning of empirical formula
- To learn to calculate empirical formulas
- To learn to calculate the molecular formula of a compound

Assume that you have mixed two solutions, and a solid product forms. How can you find out what the solid is? What is its formula? Although an experienced chemist can often predict the product expected in a chemical reaction, the only sure way to identify the product is to perform experiments. Usually we compare the physical properties of the product to the properties of known compounds.

A. Empirical Formulas

Sometimes a chemical reaction gives a product that has never been obtained before. In such a case, a chemist determines what compound has been formed by determining which elements are present and how much of each is there. These data can be used to obtain the formula of the compound. Previously we used the formula of the compound to determine the mass of each element present in a mole of the compound. To obtain the formula of an unknown compound, we do the opposite. That is, we use the measured masses of the elements present to determine the formula.

A Closer Look

Recall that the formula of a compound represents the relative numbers of the various types of atoms present. For example, the molecular formula CO_2 tells us that for each carbon atom there are two oxygen atoms in each molecule of carbon dioxide. So to determine the formula of a substance we need to count the atoms. As we have seen in this chapter, we can do this by weighing.

Suppose we know that a compound contains only the elements carbon, hydrogen, and oxygen, and we weigh out a 0.2015-g sample. Using methods we will not discuss here, we find that this 0.2015-g sample of compound contains 0.0806 g of carbon, 0.01353 g of hydrogen, and 0.1074 g of oxygen. We have just learned how to convert these masses to numbers of atoms by using the atomic mass of each element.

Carbon

$$(0.0806 \text{ g C}) \times \frac{1 \text{ mol C atoms}}{12.01 \text{ g C}} = 0.00671 \text{ mol C atoms}$$

Hydrogen

$$(0.01353 \text{ g H}) \times \frac{1 \text{ mol H atoms}}{1.008 \text{ g H}} = 0.01342 \text{ mol H atoms}$$

Oxygen

$$(0.1074 \text{ g O}) \times \frac{1 \text{ mol O atoms}}{16.00 \text{ g O}} = 0.006713 \text{ mol O atoms}$$

Let's review what we have established. We now know that 0.2015 g of the compound contains 0.00671 mol of C atoms, 0.01342 mol of H atoms, and 0.006713 mol of O atoms.

Carbon

$$(0.00671 \text{ mol C atoms}) \times \frac{6.022 \times 10^{23} \text{ C atoms}}{1 \text{ mol C atoms}} = 4.04 \times 10^{21} \text{ C atoms}$$

Hydrogen

$$(0.01342 \text{ mol H atoms}) \times \frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H atoms}} = 8.08 \times 10^{21} \text{ H atoms}$$

Oxygen

$$(0.006713 \text{ mol O atoms}) \times \frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O atoms}} = 4.043 \times 10^{21} \text{ O atoms}$$

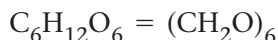
These are the numbers of the various atoms *in 0.2015 g of compound*. What do these numbers tell us about the formula of the compound? Note the following:

- The compound contains the same number of C and O atoms.
- There are twice as many H atoms as C atoms or O atoms.

Finding the Formula

We can represent this information by the formula CH_2O , which expresses the *relative* numbers of C, H, and O atoms present. Is this the true formula for the compound? In other words, is the compound made up of CH_2O molecules? It may be. However, it might also be made up of $\text{C}_2\text{H}_4\text{O}_2$ molecules, $\text{C}_3\text{H}_6\text{O}_3$ molecules, $\text{C}_4\text{H}_8\text{O}_4$ molecules, $\text{C}_5\text{H}_{10}\text{O}_5$ molecules, $\text{C}_6\text{H}_{12}\text{O}_6$ molecules, and so on. Note that each of these molecules has the required 1:2:1 ratio of carbon to hydrogen to oxygen atoms (the ratio shown by experiment to be present in the compound).

When we break a compound down into its separate elements and “count” the atoms present, we learn only the ratio of atoms—we get only the *relative* numbers of atoms. The formula of a compound that expresses the smallest whole-number ratio of the atoms present is called the **empirical formula** or *simplest formula*. A compound that contains the molecules $\text{C}_4\text{H}_8\text{O}_4$ has the same empirical formula as a compound that contains $\text{C}_6\text{H}_{12}\text{O}_6$ molecules. The empirical formula for both is CH_2O . The actual formula of a compound—the one that gives the composition of the molecules that are present—is called the **molecular formula**. The sugar called glucose is made of molecules with the molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$ (see **Figure 6.2**). Note from the molecular formula for glucose that the empirical formula is CH_2O . We can represent the molecular formula as a multiple (by 6) of the empirical formula:



Empirical formula

The formula of a compound expressing the smallest whole-number ratio of atoms in a compound

Molecular formula

The actual formula of a compound giving the types of atoms and numbers of each type of atom

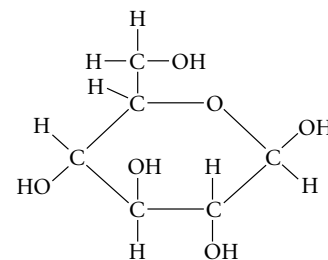


Figure 6.2

The glucose molecule. The molecular formula is $\text{C}_6\text{H}_{12}\text{O}_6$, as can be verified by counting the atoms. The empirical formula for glucose is CH_2O .

Active Reading Question

Provide two examples when empirical and molecular formulas are different. Provide two examples when empirical and molecular formulas are the same.

EXAMPLE 6.10

Determining Empirical Formulas

Information

When a subscript is outside of the parentheses, it applies to everything in the parentheses. Each subscript inside is multiplied by the subscript outside the parentheses.

In each case below, the molecular formula for a compound is given. Determine the empirical formula for each compound.

Compound	Use	Molecular Formula	Empirical Formula
Benzene	Starting material for organic chemicals	C_6H_6	?
Dioxin	Poison	$C_{12}H_4Cl_4O_2$?

Solution

$C_6H_6 = (CH)_6$; CH is the empirical formula. Each subscript in the empirical formula is multiplied by 6 to obtain the molecular formula.

$C_{12}H_4Cl_4O_2 = (C_6H_2Cl_2O)_2$; $C_6H_2Cl_2O$ is the empirical formula. Each subscript in the empirical formula is multiplied by 2 to obtain the molecular formula.

B. Calculation of Empirical Formulas

One of the most important things we can learn about a new compound is its chemical formula. To calculate the empirical formula of a compound, we first determine the relative masses of the various elements that are present.

One way to do this is to measure the masses of elements that react to form the compound. For example, suppose we weigh out 0.2636 g of pure nickel metal into a crucible and heat this metal in the air so that the nickel can react with oxygen to form a nickel oxide compound. After the sample has cooled, we weigh it again and find its mass to be 0.3354 g. The gain in mass is due to the oxygen that reacts with the nickel to form the oxide. Therefore, the mass of oxygen present in the compound is the total mass of the product minus the mass of the nickel:

Total mass of nickel oxide	–	Mass of nickel originally present	=	Mass of oxygen that reacted with the nickel
0.3354 g		0.2636 g		0.0718 g

Note that the mass of nickel present in the compound is the nickel metal originally weighed out. So we know that the nickel oxide contains 0.2636 g of nickel and 0.0718 g of oxygen. What is the empirical formula of this compound?

To answer this question we must convert the masses to numbers of atoms, using atomic masses:

$$0.2636 \text{ g Ni} \times \frac{1 \text{ mol Ni atoms}}{58.69 \text{ g Ni}} = 0.004491 \text{ mol Ni atoms}$$

(Four significant figures allowed.)

$$0.0718 \text{ g O} \times \frac{1 \text{ mol O atoms}}{16.00 \text{ g O}} = 0.00449 \text{ mol O atoms}$$

(Three significant figures allowed.)

These mole quantities represent numbers of atoms (remember that a mole of atoms is 6.022×10^{23} atoms). It is clear from the moles of atoms that the compound contains an equal number of Ni and O atoms, so the formula is NiO. This is the *empirical formula*; it expresses the smallest whole-number ratio of atoms:

$$\frac{0.004491 \text{ mol Ni atoms}}{0.00449 \text{ mol O atoms}} = \frac{1 \text{ Ni}}{1 \text{ O}}$$

That is, this compound contains equal numbers of nickel atoms and oxygen atoms. We say the ratio of nickel atoms to oxygen atoms is 1:1 (1 to 1).

EXAMPLE 6.11

Calculating Empirical Formulas

An oxide of aluminum is formed by the reaction of 4.151 g of aluminum with 3.692 g of oxygen. Calculate the empirical formula for this compound.

Solution

Where do we want to go?

Empirical formula = Al_xO_y

- $x = ?$
- $y = ?$

What do we know?

- The compound contains:

4.151 g Al	3.692 g O
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- Atomic masses

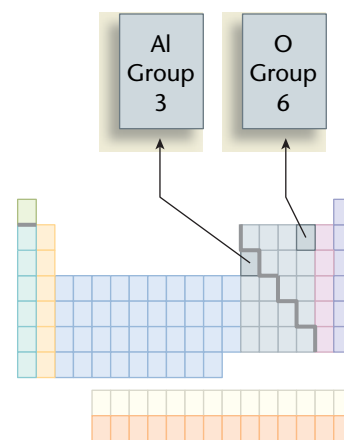
Al = 26.98 g/mol	O = 16.00 g/mol
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How do we get there?

We need to know the relative numbers of each type of atom to write the formula, so we must convert the masses of the elements to moles of atoms to get the empirical formula. We carry out the conversion by using the atomic masses of the elements.

$$4.151 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} = 0.1539 \text{ mol Al atoms}$$

$$3.692 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.2308 \text{ mol O atoms}$$



Because chemical formulas use only whole numbers, we next find the whole-number ratio of the atoms. To do this we start by dividing both numbers by the smallest of the two. This converts the smallest number to 1.

$$\frac{0.1539 \text{ mol Al}}{0.1539} = 1.000 \text{ mol Al atoms}$$

$$\frac{0.2308 \text{ mol O}}{0.1539} = 1.500 \text{ mol O atoms}$$

Note that dividing both numbers of moles of atoms by the *same* number does not change the *relative* numbers of oxygen and aluminum atoms. That is,

$$\frac{0.2308 \text{ mol O}}{0.153 \text{ mol Al}} = \frac{1.500 \text{ mol O}}{1.000 \text{ mol Al}}$$

Thus we know that the compound contains 1.500 mol of O atoms for every 1.000 mol of Al atoms, or, in terms of individual atoms, we could say that the compound contains 1.500 O atoms for every 1.000 Al atom. However, because only *whole* atoms combine to form compounds, we must find a set of *whole numbers* to express the empirical formula. When we multiply both 1.000 and 1.500 by 2, we get the integers we need.

$$1.500 \text{ O} \times 2 = 3.000 = 3 \text{ O atoms}; \quad x = 3$$

$$1.000 \text{ Al} \times 2 = 2.000 = 2 \text{ Al atoms}; \quad y = 2$$

Therefore, this compound contains two Al atoms for every three O atoms, and the empirical formula is Al_2O_3 . Note that the *ratio* of atoms in this compound is given by each of the following fractions:

$$\frac{0.2308 \text{ O}}{0.1539 \text{ Al}} = \frac{1.500 \text{ O}}{1.000 \text{ Al}} = \frac{\left(\frac{3}{2} \text{ O}\right)}{(1 \text{ Al})} \times \frac{2}{2} = \frac{3 \text{ O}}{2 \text{ Al}}$$

The smallest whole-number ratio corresponds to the subscripts of the empirical formula, Al_2O_3 .

MATH



Multiplying both the numerator and the denominator of a fraction (ratio) by the same number is just multiplying by 1:

$$\frac{2}{2} = 1, \quad \frac{3}{3} = 1, \quad \text{and so on.}$$

It does not change the value of the fraction but it can help us get whole numbers in the ratio.

Sometimes the relative numbers of moles you get when you calculate an empirical formula will not turn out to be whole numbers, as is the case in Example 6.11. When this happens, you must convert to the appropriate whole numbers. This is done by multiplying all the numbers by the same smallest whole number, which can be found by trial and error. The multiplier needed is almost always between 1 and 6. We will now summarize what we have learned about calculating empirical formulas.

Tools for Determining the Empirical Formula of a Compound

- Step 1** Obtain the mass of each element present (in grams).
- Step 2** Determine the number of moles of each type of atom present.
- Step 3** Divide the number of moles of each element by the smallest number of moles to convert the smallest number to 1. If all of the numbers so obtained are whole numbers, they are the subscripts in the empirical formula. If one or more of these numbers are not whole numbers, go on to step 4.
- Step 4** Multiply the numbers from step 3 by the smallest whole number that will convert all of them to whole numbers. This set of whole numbers represents the subscripts in the empirical formula.

CHEMISTRY EXPLORER

William Perkin • 1838–1907

Sometimes a little chemical knowledge can lead to a new and profitable career—as happened to a young chemistry student named William Perkin. In one of his chemistry lectures, Perkin's professor said it would be wonderful to make quinine artificially (in 1856, the only source for quinine—a drug effective against malaria—was the bark of the cinchona tree). Perkin, who was only 18 at the time, decided to make quinine in his home laboratory as a project over Easter vacation.

In his attempts to make quinine Perkin started with materials obtained from coal tar. However, instead of getting quinine he got red-brown and black sludges. Fortunately Perkins noticed that one of these messes turned a beautiful purple color when dissolved in alcohol or water. Perkin liked the color so much that he was distracted from his original quest to make quinine. In testing the purple solutions, he found he could use them to dye silk and cotton.

Perkin decided to quit school, against the advice of his professors, and convinced his wealthy father to invest in him. He patented his dye and built a factory to produce it on a large scale. His dye became known as *mauve*.

Until Perkin's discovery, all permanent purple or lavender dyes were outrageously expensive. The only source for the natural dye was a mollusk in the Mediterranean Sea. It took 9000 mollusks for 1 g of dye! No wonder purple became the color for royalty. Perkin changed everything by producing a purple dye from coal tar that just about everyone could afford.

Perkin went on to develop another popular dye called alizarin in 1869. This red dye was being produced by the ton in his factory by 1871. Perkin sold his factory in 1874 (at the age of 36) and had enough money to spend the rest of his life doing pure research. He bought a new house and continued working in his old laboratory. Eventually he made coumarin, the first perfume from coal tar and cinnamic acid. Perkin's work started an explosion in the manufacture of organic chemicals.



EXAMPLE 6.12

Calculating Empirical Formulas for Binary Compounds

When a 0.3546-g sample of vanadium metal is heated in air, it reacts with oxygen to reach a final mass of 0.6330 g. Calculate the empirical formula of this vanadium oxide.

Solution

Where do we want to go?

Empirical formula = V_xO_y

- $x = ?$
- $y = ?$

What do we know?

- The sample of compound contains 0.3546 g V.
- Mass of the sample is 0.6330 g
- Atomic masses
V = 50.94 g/mol O = 16.00 g/mol

How do we get there?

Step 1 All the vanadium that was originally present will be found in the final compound, so we can calculate the mass of oxygen that reacted by taking the following difference:

Total mass of compound	−	Mass of vanadium in compound	=	Mass of oxygen in compound
------------------------	---	------------------------------	---	----------------------------

$$0.6330 \text{ g} - 0.3546 \text{ g} = 0.2784 \text{ g}$$

Step 2 Using the atomic masses (50.94 for V and 16.00 for O), we obtain

$$0.3546 \text{ g V} \times \frac{1 \text{ mol V atoms}}{50.94 \text{ g V}} = 0.006961 \text{ mol V atoms}$$

$$0.2784 \text{ g O} \times \frac{1 \text{ mol O atoms}}{16.00 \text{ g O}} = 0.01740 \text{ mol O atoms}$$

Step 3 Then we divide both numbers of moles by the smaller, 0.006961.

$$\frac{0.006961 \text{ mol V atoms}}{0.006961} = 1.000 \text{ mol V atoms}$$

$$\frac{0.01740 \text{ mol O atoms}}{0.006961} = 2.500 \text{ mol O atoms}$$

Because one of these numbers (2.500) is not a whole number, we go on to step 4.

Step 4 We note that $2 \times 2.500 = 5.000$ and $2 \times 1.000 = 2.000$, so we multiply both numbers by 2 to get whole numbers.

$$2 \times 1.000 \text{ V} = 2.000 \text{ V} = 2 \text{ V} \quad x = 2$$

$$2 \times 2.500 \text{ O} = 5.000 \text{ O} = 5 \text{ O} \quad y = 5$$

This compound contains 2 V atoms for every 5 O atoms, and the empirical formula is V_2O_5 .

MATH



$V_{1.000}O_{2.500}$ becomes V_2O_5 when we multiply both subscripts by 2.

Practice Problem • Exercise 6.12

In a lab experiment it was observed that 0.6884 g of lead combines with 0.2356 g of chlorine to form a binary compound. Calculate the empirical formula of this compound.

The same procedures we have used for binary compounds also apply to compounds containing three or more elements, as Example 6.13 illustrates.

EXAMPLE 6.13

Calculating Empirical Formulas for Compounds Containing Three or More Elements

A sample of lead arsenate, an insecticide used against the potato beetle, contains 1.3813 g of lead, 0.00672 g of hydrogen, 0.4995 g of arsenic, and 0.4267 g of oxygen. Calculate the empirical formula for lead arsenate.

Solution

Where do we want to go?

Empirical formula = $\text{Pb}_a\text{H}_b\text{As}_c\text{O}_d$

- $a = ?$
- $b = ?$
- $c = ?$
- $d = ?$

What do we know?

- The sample of compound contains:

1.3813 g Pb	0.00672 g H
0.4995 g As	0.4267 g O

- Atomic masses

Pb = 207.2 g/mol	H = 1.008 g/mol
As = 74.92 g/mol	O = 16.00 g/mol

How do we get there?

Step 1 The compound contains 1.3813 g Pb, 0.00672 g H, 0.4995 g As, and 0.4267 g O.

Step 2 We use the atomic masses of the elements present to calculate the moles of each.

$$1.3813 \text{ g Pb} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} = 0.006667 \text{ mol Pb}$$

$$0.00672 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.00667 \text{ mol H}$$

$$0.4995 \text{ g As} \times \frac{1 \text{ mol As}}{74.92 \text{ g As}} = 0.006667 \text{ mol As}$$

$$0.4267 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.02667 \text{ mol O}$$

information

Only three significant figures are allowed because the mass of H has only three significant figures.

Step 3 Now we divide by the smallest number of moles.

$$\frac{0.006667 \text{ mol Pb}}{0.006667} = 1.000 \text{ mol Pb} \quad a = 1$$

$$\frac{0.006667 \text{ mol H}}{0.006667} = 1.00 \text{ mol H} \quad b = 1$$

$$\frac{0.006667 \text{ mol As}}{0.006667} = 1.000 \text{ mol As} \quad c = 1$$

$$\frac{0.02667 \text{ mol O}}{0.006667} = 4.000 \text{ mol O} \quad d = 4$$

The empirical formula is PbHAsO_4 .

Practice Problem • Exercise 6.13

Sevin, the commercial name for an insecticide used to protect crops such as cotton, vegetables, and fruit, is made from carbamic acid. A chemist analyzing a sample of carbamic acid finds 0.8007 g of carbon, 0.9333 g of nitrogen, 0.2016 g of hydrogen, and 2.133 g of oxygen. Determine the empirical formula for carbamic acid.

Information

Percent by mass for a given element means the grams of that element in 100 g of the compound.

When a compound is analyzed to determine the relative amounts of the elements present, the results are usually given in terms of percentages by masses of the various elements. Previously we learned to calculate the percent composition of a compound from its formula. Now we will do the opposite. Given the percent composition, we will calculate the empirical formula.

To understand this procedure, you must understand the meaning of *percent*. Remember that percent means parts of a given component per 100 parts of the total mixture. For example, if a given compound is 15% carbon (by mass), the compound contains 15 g of carbon per 100 g of compound.

Calculation of the empirical formula of a compound when one is given its percent composition is illustrated in Example 6.14.

EXAMPLE 6.14

Calculating Empirical Formulas from Percent Composition

Cisplatin, the common name for a platinum compound that is used to treat cancerous tumors, has the composition (mass percent) 65.02% platinum, 9.34% nitrogen, 2.02% hydrogen, and 23.63% chlorine. Calculate the empirical formula for cisplatin.

Solution

Where do we want to go?

Empirical formula = $\text{Pt}_a\text{N}_b\text{H}_c\text{Cl}_d$

- $a = ?$
- $b = ?$
- $c = ?$
- $d = ?$

What do we know?

- The sample of compound contains:

65.02% Pt 9.34% N 2.02% H 23.63% Cl

- Atomic masses

$$\text{Pt} = 195.1 \text{ g/mol}$$

$$\text{N} = 14.01 \text{ g/mol}$$

$$\text{H} = 1.008 \text{ g/mol}$$

$$\text{Cl} = 35.45 \text{ g/mol}$$

How do we get there?

Step 1 Determine how many grams of each element are present in 100 g of the compound. For 100 g of cisplatin:

- 65.02% Pt → 65.02 g Pt
- 2.02% H → 2.02 g H
- 9.34% N → 9.34 g N
- 23.63% Cl → 23.63 g Cl

Step 2 Determine the number of moles of each type of atom. We use the atomic masses to calculate moles.

$$65.02 \text{ g Pt} \times \frac{1 \text{ mol Pt}}{195.1 \text{ g Pt}} = 0.3333 \text{ mol Pt}$$

$$2.02 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 2.00 \text{ mol H}$$

$$9.34 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.667 \text{ mol N}$$

$$23.63 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.6666 \text{ mol Cl}$$

Step 3 Divide through by the smallest number of moles.

$$\frac{0.3333 \text{ mol Pt}}{0.3333} = 1.000 \text{ mol Pt} \quad a = 1$$

$$\frac{2.00 \text{ mol H}}{0.3333} = 6.01 \text{ mol H} \quad c = 6$$

$$\frac{0.667 \text{ mol N}}{0.3333} = 2.00 \text{ mol N} \quad b = 2$$

$$\frac{0.6666 \text{ mol Cl}}{0.3333} = 2.000 \text{ mol Cl} \quad d = 2$$

The empirical formula for cisplatin is $\text{PtN}_2\text{H}_6\text{Cl}_2$. Note that the number for hydrogen is slightly greater than 6 because of rounding.

Practice Problem • Exercise 6.14

The most common form of nylon (Nylon-6) is 63.68% carbon, 12.38% nitrogen, 9.80% hydrogen, and 14.4% oxygen. Calculate the empirical formula for Nylon-6.

C. Calculation of Molecular Formulas

If we know the composition of a compound in terms of the masses (or mass percentages) of the elements present, we can calculate the empirical formula but not the molecular formula. For reasons that will become clear as we consider Example 6.15, to obtain the molecular formula we must know the molar mass. Next we will consider compounds where both the percent composition and the molar mass are known.

EXAMPLE 6.15

Calculating Molecular Formulas

A white powder is analyzed and found to have an empirical formula of P_2O_5 . The compound has a molar mass of 283.88 g/mol. What is the compound's molecular formula?

Solution

Where do we want to go?

Molecular formula = P_xO_y

- $x = ?$
- $y = ?$

What do we know?

- Empirical formula = P_2O_5
- Molar mass = 283.88 g/mol
- Atomic masses

$$P = 30.97 \text{ g/mol}$$

$$O = 16.00 \text{ g/mol}$$

How do we get there?

To obtain the molecular formula, we must compare the empirical formula mass to the molar mass. The empirical formula mass for P_2O_5 is the mass of 1 mol of P_2O_5 units.

$$\begin{array}{r} \text{Atomic} \\ \text{mass of P} \\ \hline 2 \text{ mol P: } 2 \times 30.97 \text{ g} = 61.94 \text{ g} \\ \hline 5 \text{ mol O: } 5 \times 16.00 \text{ g} = 80.00 \text{ g} \\ \hline \text{Atomic} \\ \text{mass of O} \\ \hline 141.94 \text{ g} \quad \text{Mass of 1 mol of } P_2O_5 \text{ units} \end{array}$$

Recall that the molecular formula contains a whole number of empirical formula units. That is,

$$\text{Molecular formula} = (\text{empirical formula})_n$$

where n is a small whole number. Now, because

$$\text{Molecular formula} = n \times \text{empirical formula}$$

then

$$\text{Molar mass} = n \times \text{empirical formula mass}$$

Solving for n gives

$$n = \frac{\text{molar mass}}{\text{empirical formula mass}}$$

Thus, to determine the molecular formula, we first divide the molar mass by the empirical formula mass. This tells us how many empirical formula masses are present in one molar mass.

$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{283.88 \text{ g}}{141.94 \text{ g}} = 2$$

This result means that $n = 2$ for this compound, so the molecular formula consists of two empirical formula units, and the molecular formula is $(\text{P}_2\text{O}_5)_2$, or P_4O_{10} . Thus $x = 4$, $y = 10$. The structure of this interesting compound is shown in **Figure 6.3**.

Practice Problem • Exercise 6.15

A compound used as an additive for gasoline to help prevent engine knock shows the following percentage composition:

71.65% Cl 24.27% C 4.07% H

The molar mass is known to be 98.96 g. Determine the empirical formula and the molecular formula for this compound.

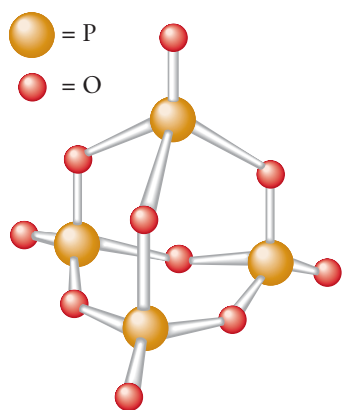
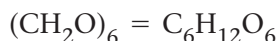


Figure 6.3

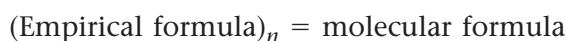
The structure of P_4O_{10} as a “ball-and-stick” model. This compound has a great affinity for water and is often used as a desiccant, or drying agent.

The molecular formula is always a whole number multiple of the empirical formula.

For example, the sugar glucose (see Figure 6.2) has the empirical formula CH_2O and the molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$. In this case there are six empirical formula units in each glucose molecule:



In general, we can represent the molecular formula in terms of the empirical formula as follows:



where n is a whole number. If $n = 1$, the molecular formula is the same as the empirical formula. For example, for carbon dioxide the empirical formula (CO_2) and the molecular formula (CO_2) are the same, so $n = 1$. On the other hand, for tetraphosphorus decoxide the empirical formula is P_2O_5 and the molecular formula is $\text{P}_4\text{O}_{10} = (\text{P}_2\text{O}_5)_2$. In this case $n = 2$.

information

Molecular formula =
(empirical formula) $_n$,
where n is a whole
number.

Active Reading Question

What data are needed to determine the empirical formula of a compound? What additional information is needed to determine the molecular formula of this compound?

Nutrition Facts

The food you eat contains all kinds of nutrients. But how much of what things? And is it healthful or “junk food”? To help you make informed decisions about the food you choose to eat, the U.S. Food and Drug Administration requires manufacturers to place nutrition facts on the packages of processed foods. Look at the label shown here from a package of macaroni. Notice that the values of several substances are given as % Daily Value. Because it is important to know the percent composition for food, the labels

must list the percentages of fat, trans fat, protein, carbohydrate, and other vital chemicals. You can use these guides to help you consume a healthy diet.

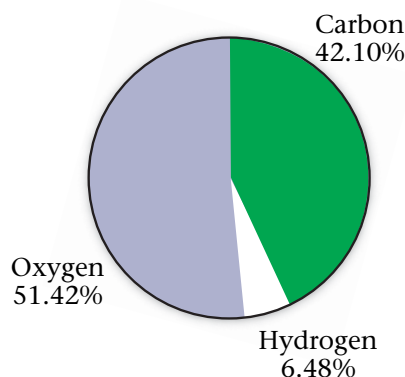
The law requires that the content of food (per serving) be printed on the label.

Amount Per Serving	
Calories	320
Calories from Fat	80
% Daily Value*	
Total Fat	9g 14%
Saturated Fat	4g 20%
Trans Fat	0g
Cholesterol	20mg 6%
Sodium	880mg 37%
Total Carbohydrate	45g 15%
Dietary Fiber	4g 16%
Sugars	5g
Protein	14g
Vitamin A	6%
Vitamin C	0%
Calcium	20%
Iron	10%

SECTION 6.3 REVIEW QUESTIONS

- Is the empirical formula for a compound ever the same as its molecular formula? Explain.
- Why are the subscripts in a chemical formula whole numbers (for example, N_2O_5 instead of $NO_{2.5}$)?
- What critical piece of information must be known (or given in the problem) in order to determine the molecular formula for a compound from its empirical compound?
- When 1.50 g of copper is heated in air, it reacts with oxygen to achieve a final mass of 1.88 g. Calculate the empirical formula of this copper oxide.
- A certain hydrocarbon (a compound consisting of carbon and hydrogen) is 82.66% carbon by mass and has a molar mass of 58.12 g. Determine the molecular formula of this hydrocarbon.
- The following graph shows the mass percent of the elements in sucrose (table sugar). Determine the empirical formula for sucrose.

Percent Composition of Sucrose



Key Terms

6.1 Atomic mass unit
Average atomic mass
Mole
Avogadro's number

6.2 Molar mass
Mass percent

6.3 Empirical formula
Molecular formula

Key Ideas

6.1 Atoms and Moles

- Objects do not need to have identical masses to be counted by weighing. All we need to know is the average mass of the objects.
- To count the atoms in a sample of a given element by weighing we must know the mass of the sample and the average mass for that element.
- Samples in which the ratio of the masses is the same as the ratio of the masses of the individual atoms always contain the same number of atoms.
- One mole of anything contains 6.022×10^{23} units of that substance.
- A sample of an element with a mass equal to that element's average atomic mass (expressed in g) contains one mole of atoms.

6.2 Molar Mass and Percent Composition

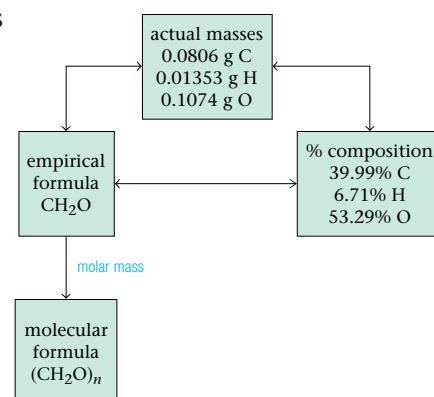
- The molar mass of any compound is the mass in grams of one mole of the compound.
- The molar mass of a compound is the sum of the masses of the component atoms.
- Moles of a compound = $\frac{\text{mass of the sample (g)}}{\text{molar mass of the compound (g/mol)}}$
- Mass of a sample (g) = (moles of sample) \times (molar mass of compound)
- Percent composition consists of the mass percent of each element in a compound:

Mass percent =

$$\frac{\text{mass of a given element in 1 mol of a compound}}{\text{mass of 1 mol of compound}} \times 100\%$$

6.3 Formulas of Compounds

- The empirical formula of a compound is the simplest whole number ratio of the atoms present in the compound.
- The empirical formula can be found from the percent composition of the compound.
- The molecular formula is the exact formula of the molecules present in a substance.
- The molecular formula is always a whole number multiple of the empirical formula.
- The following diagram shows these different ways of expressing the same information.



All exercises with **blue numbers** have answers in the back of this book.

6.1 Atoms and Moles

A. Counting by Weighing

1. Merchants usually sell small nuts, washers, and bolts by weight (like jelly beans!) rather than by individually counting the items. Suppose a particular type of washer weighs 0.110 g on the average. What would 100 such washers weigh? How many washers would there be in 100 g of washers?
2. A particular small laboratory cork weighs 1.63 g, whereas a rubber lab stopper of the same size weighs 4.31 g. How many corks would there be in 500 g of such corks? How many rubber stoppers would there be in 500 g of similar stoppers? How many grams of rubber stoppers would be needed to contain the same number of stoppers as there are corks in 1.00 kg of corks?

B. Atomic Masses: Counting Atoms by Weighing

3. Define the *amu*. What is one amu equivalent to in grams?
4. Why do we use the *average* atomic mass of the elements when performing calculations?
5. Using average atomic masses for each of the following elements (see the table in the inside front cover of this book), calculate the mass, in amu, of each of the following samples.
 - a. 635 atoms of hydrogen
 - b. 1.261×10^4 atoms of tungsten
 - c. 42 atoms of potassium
 - d. 7.213×10^{23} atoms of nitrogen
 - e. 891 atoms of iron
6. Using average atomic masses for each of the following elements (see the table in the inside front cover of this book), calculate the number of atoms present in each of the following samples.
 - a. 10.81 amu of boron
 - b. 320.7 amu of sulfur
 - c. 19,697 amu of gold
 - d. 19,695 amu of xenon
 - e. 3588.3 amu of aluminum

7. If an average atom of sulfur weighs 32.07 amu, how many sulfur atoms are contained in a sample with mass 8274 amu? What is the mass of 5.213×10^{24} sulfur atoms?

C. The Mole

8. In 24.02 g of carbon, there are _____ carbon atoms.
9. A sample equal to the atomic mass of an element in grams contains _____ atoms.
10. What mass of calcium metal contains the same number of atoms as 12.16 g of magnesium? What mass of calcium metal contains the same number of atoms as 24.31 g of magnesium?
11. What mass of cobalt contains the same number of atoms as 57.0 g of fluorine?
12. Calculate the average mass in grams of 1 atom of oxygen.
13. Which weighs more, 0.50 mol of oxygen atoms or 4 mol of hydrogen atoms?
14. Using the average atomic masses given in the inside front cover of this book, calculate the number of moles of each element in samples with the following masses.
 - a. 26.2 g of gold
 - b. 41.5 g of calcium
 - c. 335 mg of barium
 - d. 1.42×10^{23} g of palladium
 - e. 3.05×10^{25} mg of nickel
 - f. 1.00 lb of iron
 - g. 12.01 g of carbon
15. Using the average atomic masses given in the inside front cover of this book, calculate the *mass in grams* of each of the following samples.
 - a. 2.00 mol of iron
 - b. 0.521 mol of nickel
 - c. 1.23×10^{-3} mol of platinum
 - d. 72.5 mol of lead
 - e. 0.00102 mol of magnesium
 - f. 4.87×10^3 mol of aluminum
 - g. 211.5 mol of lithium
 - h. 1.72×10^{-6} mol of sodium

- 16.** Using the average atomic masses given in the inside front cover of this book, calculate the indicated quantities.
- the number of cobalt atoms in 0.00103 g of cobalt
 - the number of cobalt atoms in 0.00103 mol of cobalt
 - the number of moles of cobalt in 2.75 g of cobalt
 - the number of moles of cobalt represented by 5.99×10^{21} cobalt atoms
 - the mass of 4.23 mol of cobalt
 - the number of cobalt atoms in 4.23 mol of cobalt
 - the number of cobalt atoms in 4.23 g of cobalt

6.2 Molar Mass and Percent Composition

A. Molar Mass

- 17.** The _____ of a substance is the mass (in grams) of 1 mol of the substance.
- 18.** The molar mass of a substance can be obtained by _____ the atomic masses of the component atoms.
- 19.** Calculate the molar mass for each of the following substances.
- sodium nitride, Na_3N
 - carbon disulfide, CS_2
 - ammonium bromide, NH_4Br
 - ethyl alcohol, $\text{C}_2\text{H}_5\text{OH}$
 - sulfurous acid, H_2SO_3
 - sulfuric acid, H_2SO_4
- 20.** Calculate the molar mass for each of the following substances.
- barium perchlorate
 - magnesium sulfate
 - lead(II) chloride
 - copper(II) nitrate
 - tin(IV) chloride
 - phenol, $\text{C}_6\text{H}_6\text{O}$
- 21.** Calculate the number of moles of the indicated substance in each of the following samples.
- 49.2 mg of sulfur trioxide
 - 7.44×10^4 kg of lead(IV) oxide
 - 59.1 g of chloroform, CHCl_3
 - 3.27 mg of trichloroethane, $\text{C}_2\text{H}_3\text{Cl}_3$
 - 4.01 g of lithium hydroxide
- 22.** Calculate the number of moles of the indicated substance in each of the following samples.
- 4.26×10^{-3} g of sodium dihydrogen phosphate
 - 521 g of copper(I) chloride
 - 151 kg of iron
 - 8.76 g of strontium fluoride
 - 1.26×10^4 g of aluminum
- 23.** Calculate the *mass in grams* of each of the following samples.
- 1.50 mol of aluminum iodide
 - 1.91×10^{-3} mol of benzene, C_6H_6
 - 4.00 mol of glucose, $\text{C}_6\text{H}_{12}\text{O}_6$
 - 4.56×10^5 mol of ethanol, $\text{C}_2\text{H}_5\text{OH}$
 - 2.27 mol of calcium nitrate
- 24.** Calculate the *mass in grams* of each of the following samples.
- 1.27×10^{-3} mol of carbon dioxide
 - 4.12×10^3 mol of nitrogen trichloride
 - 0.00451 mol of ammonium nitrate
 - 18.0 mol of water
 - 62.7 mol of copper(II) sulfate
- 25.** Calculate the number of *molecules* present in each of the following samples.
- 6.37 mol of carbon monoxide
 - 6.37 g of carbon monoxide
 - 2.62×10^{-6} g of water
 - 2.62×10^{-6} mol of water
 - 5.23 g of benzene, C_6H_6
- 26.** Calculate the number of *moles* of sulfur atoms present in each of the following samples.
- 2.01 g of sodium sulfate
 - 2.01 g of sodium sulfite
 - 2.01 g of sodium sulfide
 - 2.01 g of sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$

6.3 Formulas of Compounds

A. Percent Composition of Compounds

27. Calculate the percent by mass of each element in the following compounds.
- sodium sulfate
 - sodium sulfite
 - sodium sulfide
 - sodium thiosulfate, $\text{Na}_2\text{S}_2\text{O}_3$
 - potassium phosphate
 - potassium hydrogen phosphate
 - potassium dihydrogen phosphate
 - potassium phosphide
28. Calculate the percent by mass of the element listed *first* in the formulas for each of the following compounds.
- copper(II) bromide, CuBr_2
 - copper(I) bromide, CuBr
 - iron(II) chloride, FeCl_2
 - iron(III) chloride, FeCl_3
 - cobalt(II) iodide, CoI_2
 - cobalt(III) iodide, CoI_3
 - tin(II) oxide, SnO
 - tin(IV) oxide, SnO_2
29. Calculate the percent by mass of the element listed first in the formulas for each of the following compounds.
- adipic acid, $\text{C}_6\text{H}_{10}\text{O}_4$
 - ammonium nitrate, NH_4NO_3
 - caffeine, $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$
 - chlorine dioxide, ClO_2
 - cyclohexanol, $\text{C}_6\text{H}_{11}\text{OH}$
 - dextrose, $\text{C}_6\text{H}_{12}\text{O}_6$
 - eicosane, $\text{C}_{20}\text{H}_{42}$
 - ethanol, $\text{C}_2\text{H}_5\text{OH}$
30. For each of the following ionic substances, calculate the percentage of the overall molar mass of the compound that is represented by the positive ions the compound contains.
- ammonium chloride
 - copper(II) sulfate
 - gold(III) chloride
 - silver nitrate

B. Formulas of Compounds

31. What experimental evidence about a new compound must be known before its formula can be determined?
32. What does the *empirical* formula of a compound represent? How does the *molecular* formula differ from the empirical formula?
33. Give the empirical formula that corresponds to each of the following molecular formulas.
- sodium peroxide, Na_2O_2
 - terephthalic acid, $\text{C}_8\text{H}_6\text{O}_4$
 - phenobarbital, $\text{C}_{12}\text{H}_{12}\text{N}_2\text{O}_3$
 - 1,4-dichloro-2-butene, $\text{C}_4\text{H}_6\text{Cl}_2$
34. Which of the following pairs of compounds have the same *empirical* formula?
- acetylene, C_2H_2 , and benzene, C_6H_6
 - ethane, C_2H_6 , and butane, C_4H_{10}
 - nitrogen dioxide, NO_2 , and dinitrogen tetroxide, N_2O_4
 - diphenyl ether, $\text{C}_{12}\text{H}_{10}\text{O}$, and phenol, $\text{C}_6\text{H}_5\text{OH}$

C. Calculation of Empirical Formulas

35. A new compound has been prepared. A 0.4791-g sample was analyzed and was found to contain the following masses of elements:
carbon, 0.1929 g
hydrogen, 0.01079 g
oxygen, 0.08566 g
chlorine, 0.1898 g
Determine the empirical formula of the new compound.
36. In an experiment, a 2.514-g sample of calcium was heated in a stream of pure oxygen, and was found to increase in mass by 1.004 g. Calculate the empirical formula of calcium oxide.
37. A compound has the following percentages by mass: barium, 58.84%; sulfur, 13.74%; oxygen, 27.43%. Determine the empirical formula of the compound.
38. If a 1.271-g sample of aluminum metal is heated in a chlorine gas atmosphere, the mass of aluminum chloride produced is 6.280 g. Calculate the empirical formula of aluminum chloride.
39. If cobalt metal is mixed with excess sulfur and heated strongly, a sulfide is produced that contains 55.06% cobalt by mass. Calculate the empirical formula of the sulfide.

- 40.** If 2.461 g of metallic calcium is heated in a stream of chlorine gas, 4.353 g of Cl_2 is absorbed in forming the metal chloride. Calculate the empirical formula of calcium chloride.
- 41.** If 10.00 g of copper metal is heated strongly in the air, the sample gains 2.52 g of oxygen in forming an oxide. Determine the empirical formula of this oxide.
- 42.** A compound has the following percentage composition by mass: copper, 33.88%; nitrogen, 14.94%; oxygen, 51.18%. Determine the empirical formula of the compound.
- 43.** Sodium and nitrogen form two binary compounds. The percentages of the elements in these compounds are:

Compound	Na	N
1	83.12%	16.88%
2	35.36%	64.64%

Calculate the empirical formula of each of the compounds.

D. Calculation of Molecular Formulas

- 44.** How does the molecular formula of a compound differ from the empirical formula? Can a compound's empirical and molecular formulas be the same? Explain.
- 45.** What information do we need to determine the molecular formula of a compound if we know only the empirical formula?
- 46.** A compound with the empirical formula CH_2O was found to have a molar mass between 89 and 91 g. What is the molecular formula of the compound?
- 47.** A compound with the empirical formula CH_2 was found to have a molar mass of approximately 84 g. What is the molecular formula of the compound?
- 48.** A compound with the empirical formula CH_4O was found in a subsequent experiment to have a molar mass of approximately 192 g. What is the molecular formula of the compound?
- 49.** A compound having an approximate molar mass of 165–170 g has the following percentage composition by mass: carbon, 42.87%; hydrogen, 3.598%; oxygen, 28.55%; nitrogen, 25.00%. Determine the empirical and molecular formulas of the compound.
- 50.** A compound consists of 65.45% C, 5.492% H, and 29.06% O on a mass basis and has a molar mass of approximately 110. Determine the molecular formula of the compound.

Critical Thinking

- 51.** Use the periodic table inside the back cover of this text to determine the atomic mass (per mole) or molar mass of each of the substances in column 1, and find that mass in column 2.

Column 1

- (1) molybdenum
- (2) lanthanum
- (3) carbon tetrabromide
- (4) mercury(II) oxide
- (5) titanium(IV) oxide
- (6) manganese(II) chloride
- (7) phosphine, PH_3
- (8) tin(II) fluoride
- (9) lead(II) sulfide
- (10) copper(I) oxide

Column 2

- (a) 33.99 g
- (b) 79.9 g
- (c) 95.94 g
- (d) 125.84 g
- (e) 138.9 g
- (f) 143.1 g
- (g) 156.7 g
- (h) 216.6 g
- (i) 239.3 g
- (j) 331.6 g

- 52.** Copy and complete the following table.

Mass of Sample	Moles of Sample	Atoms in Sample
5.00 g Al		
	0.00250 mol Fe	
		2.6×10^{24} atoms Cu
0.00250 g Mg		
	2.7×10^{23} mol Na	
		1.00×10^4 atoms U

53. Copy and complete the following table.

Mass of Sample	Moles of Sample	Molecules in Sample	Atoms in Sample
4.24 g C_6H_6			
	0.224 mol H_2O		
		2.71×10^{22} molecules CO_2	
	1.26 mol HCl		
		4.21×10^{24} molecules H_2O	
0.297 g CH_3OH			












54. A binary compound of magnesium and nitrogen is analyzed, and 1.2791 g of the compound is found to contain 0.9240 g of magnesium. When a second sample of this compound is treated with water and heated, the nitrogen is driven off as ammonia, leaving a compound that contains 60.31% magnesium and 39.69% oxygen by mass. Calculate the empirical formulas of the two magnesium compounds.
55. When a 2.118-g sample of copper is heated in an atmosphere in which the amount of oxygen present is restricted, the sample gains 0.2666 g of oxygen in forming a reddish-brown oxide. However, when 2.118 g of copper is heated in a stream of pure oxygen, the sample gains 0.5332 g of oxygen. Calculate the empirical formulas of the two oxides of copper.
56. A molecule of an organic compound has twice as many hydrogen atoms as carbon atoms, the same number of oxygen atoms as carbon atoms, and one-eighth as many sulfur atoms as hydrogen atoms. The molar mass of the compound is 152 g/mol. What are the empirical and molecular formulas for this compound?
57. Calculate the number of grams of iron that contain the same number of atoms as 2.24 g of cobalt.
58. Calculate the number of grams of cobalt that contain the same number of atoms as 2.24 g of iron.
59. A strikingly beautiful copper compound with the common name "blue vitriol" has the following elemental composition: 25.45% Cu, 12.84% S, 4.036% H, 57.67% O. Determine the empirical formula of the compound.
60. A 0.7221-g sample of a new compound has been analyzed and found to contain the following masses of elements: carbon, 0.2990 g; hydrogen, 0.05849 g; nitrogen, 0.2318 g; oxygen, 0.1328 g. Calculate the empirical formula of the compound.
61. When 4.01 g of mercury is strongly heated in air, the resulting oxide weighs 4.33 g. Calculate the empirical formula of the oxide.
62. When barium metal is heated in chlorine gas, a binary compound forms that consists of 65.95% Ba and 34.05% Cl by mass. Calculate the empirical formula of the compound.
63. A particular compound in the chemistry laboratory is found to contain 7.2×10^{24} atoms of oxygen, 56.0 g of nitrogen, and 4.0 mol of hydrogen. What is its empirical formula?
64. The compound A_2O is 63.7% A (a mystery element) and 36.3% oxygen. What is the identity of element A?
65. Find the item in list B that best explains or completes the statement or question in list A.

List A

- (1) 1 amu
- (2) 1008 amu
- (3) mass of the "average" atom of an element
- (4) number of carbon atoms in 12.01 g of carbon
- (5) 6.022×10^{23} molecules
- (6) total mass of all atoms in 1 mol of a compound
- (7) smallest whole-number ratio of atoms present in a molecule
- (8) formula showing the actual number of atoms present in a molecule
- (9) product formed when any carbon-containing compound is burned in O_2
- (10) have the same empirical formulas, but different molecular formulas

List B

- (a) 6.022×10^{23}
- (b) atomic mass
- (c) mass of 1000 hydrogen atoms
- (d) benzene, C_6H_6 , and acetylene, C_2H_2
- (e) carbon dioxide
- (f) empirical formula
- (g) 1.66×10^{-24} g
- (h) molecular formula
- (i) molar mass
- (j) 1 mol


- 1 You have been saving dimes for a long time and take all of them to the bank to exchange for dollar bills. The teller weighs the dimes and records a total mass of 3358 g. If a dime has an average mass of 2.30 g, how much money are you owed?
 A \$7723.40 C \$2.30
 B \$1460 D \$146
- 2 How many atoms of neon (Ne) are in a sample of neon with a mass of 1452.96 amu?
 A 29,321 C 65
 B 72 D 6.02×10^{23}
- 3 Consider a 20.0 g sample of silver (Ag) metal and a 20.0 g sample of another “mystery” metal. You are told that there are exactly four times as many atoms of the “mystery” metal as of silver metal. Identify the “mystery” metal.
 A Iron (Fe) C Platinum (Pt)
 B Lithium (Li) D Aluminum (Al)
- 4 Which of the following is the most accurate description of a mole?
 A The number of atoms in 55.85 grams of iron (Fe).
 B The mass of carbon in a measured sample of carbon.
 C The number of atoms in any given mass of a molecule.
 D 6.022×10^{23} grams of a sample of any element.
- 5 The molar mass of calcium nitrate, $\text{Ca}(\text{NO}_3)_2$, is:
 A 102.09 g/mol
 B 70.09 g/mol
 C 164.1 g/mol
 D 100.1 g/mol
- 6 A single molecule of a particular compound has a mass of 4.65×10^{-23} g. Which of the following could be this compound?
 A CO
 B H_2O
 C NO_2
 D NH_3
- 7 What is the percent by mass of carbon in table sugar (sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$)?
 A 27.80% C 56.69%
 B 42.10% D 71.38%
- 8 Which of the following compounds have the same empirical formulas?
 A H_2O  and H_2O_2 
 B N_2O_4  and NO_2 
 C CO  and CO_2 
 D CH_4  and C_2H_6 
- 9 The percent by mass of nitrogen is 46.7% for a species containing only nitrogen and oxygen. Which of the following could be this species?
 A N_2O_5 C NO 
 B NO_2  D N_2O 
- 10 A compound consisting of carbon, hydrogen, and oxygen is found to contain 40.00% carbon by mass, and 6.71% hydrogen by mass. A 0.0500 mol sample of the compound has a mass of 9.00 g. Determine the molecular formula for the compound.
 A $\text{C}_8\text{H}_{16}\text{O}_4$
 B C_6HO_8
 C $\text{C}_6\text{H}_{14}\text{O}_6$
 D $\text{C}_6\text{H}_{12}\text{O}_6$
- 11 You are given the percent by mass of the elements in a compound and wish to determine the empirical formula. Label each of the following statements as true or false and explain your answers.
 a. You must assume exactly 100.0 g of the compound.
 b. You must convert percent by mass to relative numbers of atoms.
 c. You must divide all of the percent by mass numbers by the smallest percent by mass.
 d. You cannot solve for the empirical formula without the molar mass.
- 12 An oxide with the formula MO_2 is 40.1% oxygen by mass. Determine the identity of M.

LOOKING AHEAD

7.1 Evidence for a Chemical Reaction

7.2 Chemical Equations

7.3 Balancing Chemical Equations



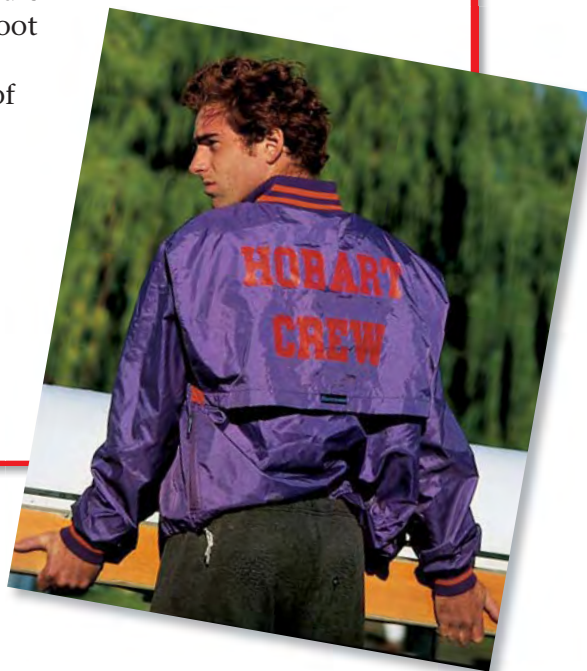
Aluminum reacting with bromine

IN YOUR LIFE

Chemistry is about change. Grass grows. Steel rusts. Hair is bleached, dyed, “permed,” or straightened. Natural gas burns to heat houses. Nylon is produced for jackets, swimsuits, and pantyhose. Water is decomposed to hydrogen and oxygen gas by an electric current. The bombardier beetle concocts a toxic spray to shoot at its enemies.

These are just a few examples of chemical changes that affect each of us. Chemical reactions are the heart and soul of chemistry, and in this chapter we will discuss fundamental ideas about chemical reactions.

Nylon jackets are sturdy and dry quickly. These characteristics make them ideal for athletic wear.



WHAT DO YOU KNOW?

Prereading Questions

1. How can you tell that a chemical reaction has taken place?
2. Hydrogen gas reacts with oxygen gas to form water. Write a chemical equation for this reaction.
3. Write the law of conservation of matter in your own words.

Evidence for a Chemical Reaction

Objective

- To learn the signals that show a chemical reaction has occurred

How do we know when a chemical change (a reaction) has occurred? That is, what are the clues that a chemical change has taken place? A glance back at the processes in the introduction suggests that *chemical reactions often give a visual signal*. Steel changes from a smooth, shiny material to a reddish-brown, flaky substance when it rusts. Hair changes color when it is bleached. Solid nylon is formed when two particular liquid solutions are brought into contact. A blue flame appears when natural gas reacts with oxygen. Chemical reactions, then, often give *visual* clues:

- A color changes.
- A solid forms.
- Bubbles are produced.
- A flame occurs.

However, reactions are not always visible. Sometimes the only signal that a reaction is occurring is a change in temperature as heat is produced or absorbed (see **Figure 7.1**).

**Figure 7.1**

(a) An injured girl uses a cold pack to help prevent swelling. The pack is activated by breaking an ampule; which starts a chemical reaction that absorbs heat rapidly, lowering the temperature of the area to which the pack is applied.

(b) A hot pack is used to warm hands and feet in winter. When the package is opened, oxygen from the air penetrates a bag containing solid chemicals. The resulting reaction produces heat.

Evidence for a Chemical Reaction



color change



solid forms

TABLE 7.1

Some Clues That a Chemical Reaction Has Occurred

1. The color changes.
2. A solid forms.
3. Bubbles form.
4. Heat and/or a flame is produced, or heat is absorbed.



bubbles form



heat or flame occurs

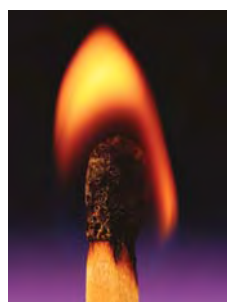
SECTION 7.1 REVIEW QUESTIONS

1. What types of evidence indicate that a chemical reaction has taken place?
2. Give an example of each type of clue from Table 7.1 that a chemical reaction has taken place.
3. Look at the following photographs and identify any clues that tell you a chemical reaction has taken place.

a.



b.



c.



d.



SECTION 7.2

Chemical Equations

Key Terms

- Chemical reaction
- Chemical equation
- Reactants
- Products
- Balancing the chemical equation

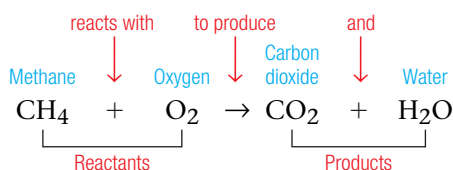
Objectives

- To learn to identify the characteristics of a chemical reaction
- To learn the information given by a chemical equation

Chemists have learned that a chemical change always involves a rearrangement of the ways in which the atoms are grouped. For example, when the methane, CH_4 , in natural gas combines with oxygen, O_2 , in the air and burns, carbon dioxide, CO_2 , and water, H_2O , are formed. Such a chemical change is called a **chemical reaction**. We represent a chemical reaction by writing a **chemical equation** in which the chemicals present before the reaction, the **reactants**, are shown to the left of an arrow and the chemicals formed by the reaction, the **products**, are shown to the right of an arrow. The arrow indicates the direction of the change and is read as “yields” or “produces”:

Reactants \rightarrow Products

For the reaction of methane with oxygen, we have

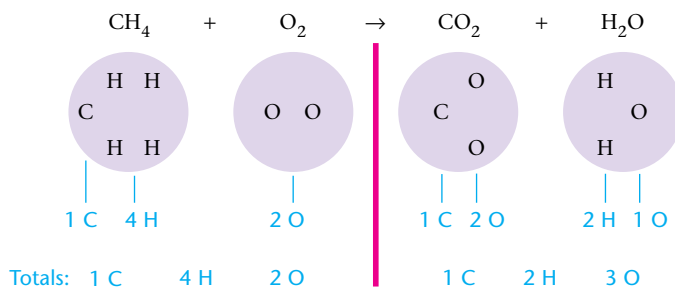


Note from this equation that the products contain the same atoms as the reactants but that the atoms are associated in different ways. That is, *a chemical reaction involves changing the ways the atoms are grouped.*

In a chemical reaction, atoms are neither created nor destroyed. All atoms present in the reactants must be accounted for among the products.

In other words, there must be the same number of each type of atom on the product side as on the reactant side of the arrow. Making sure that the equation for a reaction obeys this rule is called **balancing the chemical equation** for a reaction.

The equation that we have shown for the reaction between CH_4 and O_2 is not balanced. We can see that it is not balanced by taking apart the reactants and products.



The reaction cannot happen this way because, as it stands, this equation states that one oxygen atom is created and that two hydrogen atoms are destroyed. A reaction is only a rearrangement of the way the atoms are grouped; atoms are not created or destroyed. The total number of each type of atom must be

Chemical reaction

One or more substances are changed into one or more new substances by the reorganization of component atoms.

Chemical equation

A representation of a chemical reaction using the formulas of the starting substances that react and the new substances that are formed

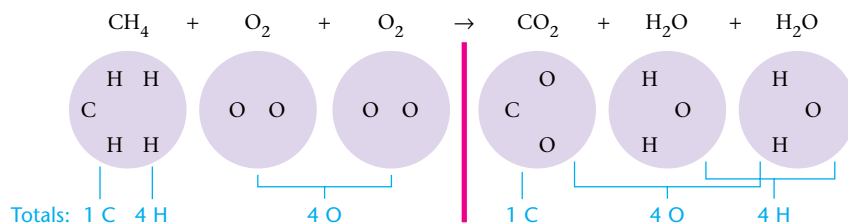
Reactant

The starting substance of a chemical reaction, shown to the left of the reaction arrow

Product

The new substance formed by a chemical reaction, shown to the right of the reaction arrow

the same on both sides of the arrow. We can fix the imbalance in this equation by involving one more O_2 molecule on the left and by showing the production of one more H_2O molecule on the right.

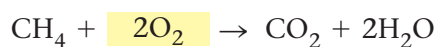


This *balanced chemical equation* shows the actual numbers of molecules involved in this reaction (see **Figure 7.2**).

When we write the balanced equation for a reaction, we group like molecules together:



is written



Active Reading Question

Why must a chemical equation be balanced?

Let's Review

The chemical equation for a reaction provides us with two important types of information:

- The identities of the reactants and products
- The relative numbers of each

Besides specifying the compounds involved in the reaction, we often indicate in the equation the *physical states* of the reactants and products by using the following symbols.

Physical States	
Symbol	State
(s)	solid
(l)	liquid
(g)	gas
(aq)	dissolved in water (in aqueous solution)

For example, when solid potassium reacts with liquid water, the products are hydrogen gas and potassium hydroxide; the latter remains dissolved in the water. From this information about the reactants and products, we can write the equation for the reaction.

- Solid potassium is represented by $K(s)$; liquid water is written as $H_2O(l)$;
- hydrogen gas contains diatomic molecules and is represented as $H_2(g)$;
- potassium hydroxide dissolved in water is written as $KOH(aq)$.

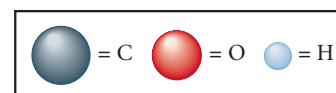
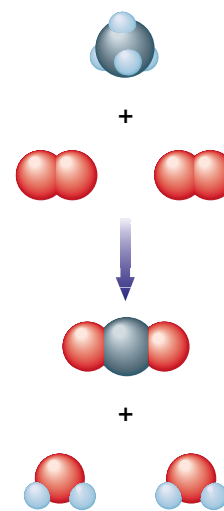
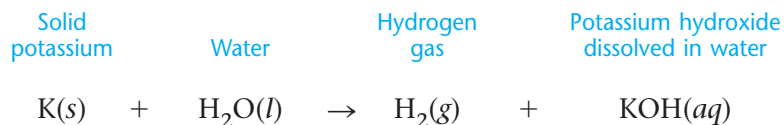


Figure 7.2

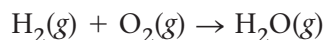
The reaction between methane and oxygen yields water and carbon dioxide. Note that there are four oxygen atoms in the products *and* in the reactants; none has been gained or lost in the reaction. Similarly, there are four hydrogen atoms and one carbon atom in the reactants *and* in the products. The reaction simply changes the way the atoms are grouped.

So the *unbalanced* equation for the reaction is



This reaction is shown in **Figure 7.3**.

The hydrogen gas produced in this reaction then reacts with the oxygen gas in the air, producing gaseous water and a flame. The *unbalanced* equation for this second reaction is



Both of these reactions produce a great deal of heat. In Example 7.1 we will practice writing the unbalanced equations for reactions. Then, in the next section, we will discuss systematic procedures for balancing equations.

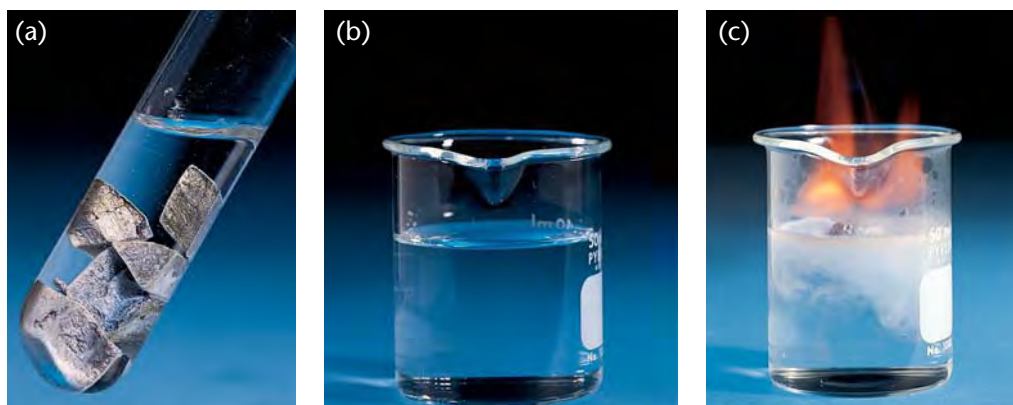


Figure 7.3

The reactants (a) potassium metal (stored in mineral oil to prevent oxidation) and (b) water. (c) The reaction of potassium with water. The flame occurs because the hydrogen gas, $\text{H}_2(g)$, produced by the reaction burns in air, reacts with $\text{O}_2(g)$, at the high temperatures caused by the reaction.

EXAMPLE 7.1

Chemical Equations: Recognizing Reactants and Products

Information

Because Zn forms only the Zn^{2+} ion, a Roman numeral is usually not used. Thus ZnCl_2 is commonly called zinc chloride.

Write the *unbalanced* chemical equation for each of the following reactions:

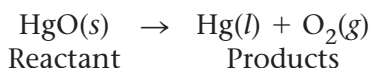
- Solid mercury(II) oxide decomposes to produce liquid mercury metal and gaseous oxygen.
- Solid carbon reacts with gaseous oxygen to form gaseous carbon dioxide.
- Solid zinc is added to an aqueous solution containing dissolved hydrogen chloride to produce gaseous hydrogen that bubbles out of the solution and zinc chloride that remains dissolved in the water.

Solution

- Reactants** In this case we have only one reactant, mercury(II) oxide. The name mercury(II) oxide means that the Hg^{2+} cation is present, so one O^{2-} ion is required for a zero net charge. The formula is HgO , which is written as $\text{HgO}(s)$ in this case because it is given as a solid.

Products The products are liquid mercury, written as $\text{Hg}(l)$, and gaseous oxygen, written as $\text{O}_2(g)$. Remember that under normal conditions, oxygen exists as a diatomic molecule.

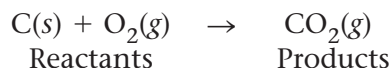
Chemical Equation The unbalanced equation is



- b. **Reactants** In this case solid carbon, written as $\text{C}(s)$, reacts with gaseous oxygen, written as $\text{O}_2(g)$.

Products The product is gaseous carbon dioxide, written as $\text{CO}_2(g)$.

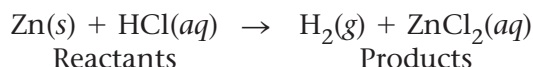
Chemical Equation The equation (which happens to be balanced) is



- c. **Reactants** In this reaction solid zinc, $\text{Zn}(s)$, is added to an aqueous solution of hydrogen chloride, which is written as $\text{HCl}(aq)$ and called hydrochloric acid.

Products The products are gaseous hydrogen, written as $\text{H}_2(g)$, and aqueous zinc chloride. The name *zinc chloride* means that the Zn^{2+} ion is present, so two Cl^- ions are needed to achieve a zero net charge. Thus zinc chloride dissolved in water is written as $\text{ZnCl}_2(aq)$.

Chemical Equation The unbalanced equation is



Practice Problem • Exercise 7.1

Identify the reactants and products and write the *unbalanced* equation (including symbols for states) for each of the following chemical reactions:

- Solid magnesium metal reacts with liquid water to form solid magnesium hydroxide and hydrogen gas.
- Solid ammonium dichromate (review Table 4.4 if this compound is unfamiliar) decomposes to solid chromium(III) oxide, gaseous nitrogen, and gaseous water.
- Gaseous ammonia reacts with gaseous oxygen to form gaseous nitrogen monoxide and gaseous water.



Zinc metal reacts with hydrochloric acid to produce bubbles of hydrogen gas.



SECTION 7.2 REVIEW QUESTIONS

- Identify the reactants and products and write chemical equations for each of the following reactions:
 - Gaseous chlorine reacts with an aqueous solution of potassium bromide to form liquid bromine and an aqueous solution of potassium chloride.
 - Solid aluminum reacts with solid iodine to produce solid aluminum iodide.
 - Solid magnesium reacts with an aqueous solution of hydrochloric acid to form an aqueous solution of magnesium chloride and bubbles of hydrogen gas.
- List several pieces of information you can determine about a chemical reaction just by reading the chemical equation.
- What is the difference between a coefficient and a subscript in a chemical equation? Provide an example of each.
- What is the difference between an aqueous solution and a liquid? Write the symbols for each.



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SECTION 7.3

Balancing Chemical Equations

Key Term

- Coefficients

Objective

- To learn to write a balanced equation for a chemical reaction

As we saw in the previous section, an unbalanced chemical equation is not an accurate representation of the reaction that occurs. Whenever you see an equation for a reaction, you should ask yourself whether it is balanced.

Atoms are conserved in a chemical reaction.

That is, atoms are neither created nor destroyed. They are just grouped differently. The same number of each type of atom is found among the reactants and among the products.

Chemists determine the identity of the reactants and products of a reaction by experimental observation. For example, when methane (natural gas) is burned in the presence of sufficient oxygen gas, the products are always carbon dioxide and water.

The identities (formulas) of the compounds must never be changed in balancing a chemical equation.

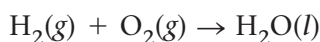
DID YOU KNOW

Trial and error is often useful for solving problems. It's okay to make a few wrong turns before you get to the right answer.

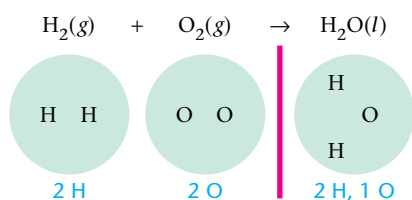
In other words, the subscripts in a formula cannot be changed, nor can atoms be added to or subtracted from a formula.

Most chemical equations can be balanced by trial and error—that is, by inspection. Keep trying until you find the numbers of reactants and products that give the same number of each type of atom on both sides of the arrow.

The Combustion of Hydrogen Consider the reaction of hydrogen gas and oxygen gas to form liquid water. First, we write the unbalanced equation from the description of the reaction.

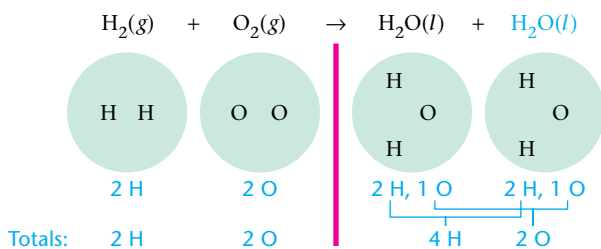


We can see that this equation is unbalanced by counting the atoms on both sides of the arrow.

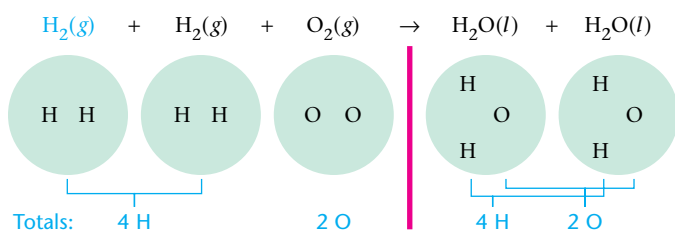


We have one more oxygen atom in the reactants than in the products. Because we cannot create or destroy atoms and because we *cannot change the formulas* of the reactants or products, we must balance the equation by adding more molecules of reactants and/or products. In this case we need one more oxygen atom on the right, so we add another water molecule (which contains one O atom). Then we count all of the atoms again.

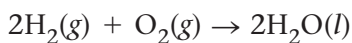
Reactants	Products
2 H	2 H
2 O	1 O



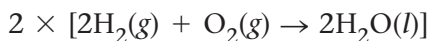
We have balanced the oxygen atoms, but now the hydrogen atoms have become unbalanced. There are more hydrogen atoms on the right than on the left. We can solve this problem by adding another hydrogen molecule (H_2) to the reactant side.



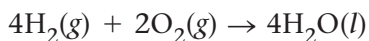
The equation is now balanced. We have the same numbers of hydrogen and oxygen atoms represented on both sides of the arrow. Collecting like molecules, we write the balanced equation as



Consider next what happens if we multiply every part of this balanced equation by 2,

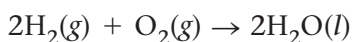


to give



This equation is balanced (count the atoms to verify this). In fact, we can multiply or divide *all parts* of the original balanced equation by any number to give a new balanced equation. Thus each chemical reaction has many possible balanced equations. Is one of the many possibilities preferred over the others? Yes.

The accepted convention is that the “best” balanced equation is the one with the *smallest whole numbers*. These whole numbers are called the **coefficients** for the balanced equation. Therefore, for the reaction of hydrogen and oxygen to form water, the “correct” balanced equation is



The coefficients 2, 1 (never written), and 2, respectively, are the smallest *whole numbers* that give a balanced equation for this reaction.

Active Reading Question

When hydrogen gas reacts with oxygen gas to form water, why must two water molecules be formed for every oxygen molecule reacted?

Reactants	Products
2 H	4 H
2 O	2 O

Reactants	Products
4 H	4 H
2 O	2 O

CRITICAL THINKING ?

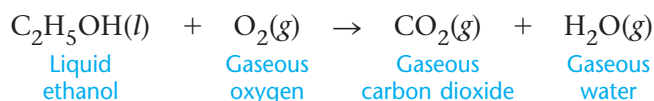
What if mass was not conserved? How would it change how you balance equations?

Coefficient

The number written in front of the chemical formulas in a balanced chemical equation—coefficients indicate the relative numbers of reactants and products in the reaction.

Combustion of Ethanol Next we will balance the equation for the reaction of liquid ethanol, C_2H_5OH , with oxygen gas to form gaseous carbon dioxide and water. This reaction, among many others, occurs in engines that burn a gasoline–ethanol mixture called gasohol.

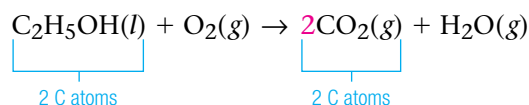
The first step in obtaining the balanced equation for a reaction is always to identify the reactants and products from the description given for the reaction. In this case we are told that liquid ethanol, $C_2H_5OH(l)$, reacts with gaseous oxygen, $O_2(g)$, to produce gaseous carbon dioxide, $CO_2(g)$, and gaseous water, $H_2O(g)$. Therefore, the unbalanced equation is



Information

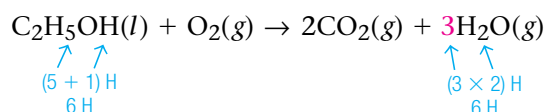
When balancing equations, start by looking at the most complicated molecule.

When one molecule in an equation is more complicated (contains more elements) than the others, it is best to start with that molecule. The most complicated molecule here is C_2H_5OH , so we begin by considering the products that contain the atoms in C_2H_5OH . We start with carbon. The only product that contains carbon is CO_2 . Because C_2H_5OH contains two carbon atoms, we place a 2 before the CO_2 to balance the carbon atoms.

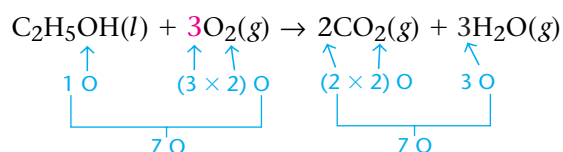


We cannot change the formula of any reactant or product when we balance an equation. We can only place coefficients in front of the formulas.

Next we consider hydrogen. The only product containing hydrogen is H_2O . C_2H_5OH contains six hydrogen atoms, so we need six hydrogen atoms on the right. Because each H_2O contains two hydrogen atoms, we need three H_2O molecules to yield six hydrogen atoms. So we place a 3 before the H_2O .

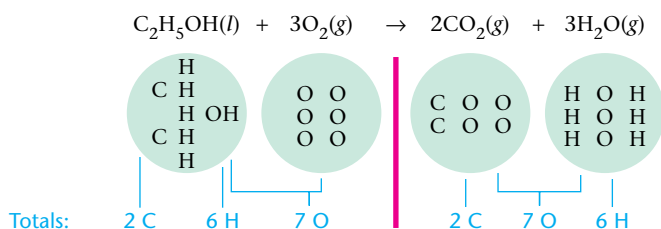


Finally, we count the oxygen atoms. On the left we have three oxygen atoms (one in C_2H_5OH and two in O_2), and on the right we have seven oxygen atoms (four in $2CO_2$ and three in $3H_2O$). We can correct this imbalance if we have three O_2 molecules on the left. That is, we place a coefficient of 3 before the O_2 to produce the balanced equation.



At this point you may have a question: why did we choose O_2 on the left when we balanced the oxygen atoms? Why not use C_2H_5OH , which has an oxygen atom? The answer is that if we had changed the coefficient in front of C_2H_5OH , we would have unbalanced the hydrogen and carbon atoms.

Now we count all of the atoms as a check to make sure the equation is balanced.



The equation is now balanced. We have the same numbers of all types of atoms on both sides of the arrow. Notice that these coefficients are the smallest whole numbers that give a balanced equation.

Reactants	Products
2 C	2 C
6 H	6 H
7 O	7 O

Active Reading Question

Why can we not change the formulas for reactants and products to balance a chemical equation?

How to Write and Balance Equations

Step 1 Read the description of the chemical reaction.

- What are the reactants, the products, and their states?
- Write the appropriate formulas.

Step 2 Write the *unbalanced* equation that summarizes the information from step 1.

Step 3 Balance the equation by inspection, starting with the most complicated molecule.

- Proceed element by element to determine what coefficients are necessary so that the same number of each type of atom appears on both the reactant side and the product side.
- Do not change the identities (formulas) of any of the reactants or products.

Step 4 Check to see that the coefficients used give the same number of each type of atom on both sides of the arrow.

- Note that an “atom” may be present in an element, a compound, or an ion.
- Check to see that the coefficients used are the smallest whole numbers that give the balanced equation. This can be done by determining whether all coefficients can be divided by the same whole numbers to give a set of smaller *whole number* coefficients.

DID YOU KNOW

To check that an equation is balanced, count the number of atoms on both sides of the equation.

Active Reading Question

How do we determine which substance to begin with when balancing a chemical equation?

EXAMPLE 7.2

Balancing Chemical Equations I

For the following reaction:

Solid potassium reacts with liquid water to form gaseous hydrogen and potassium hydroxide that dissolves in the water.

- Write the unbalanced chemical equation.
- Balance the equation.

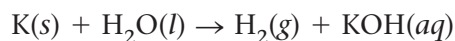
CRITICAL THINKING ?

What if a friend is balancing chemical equations by changing the values of the subscripts instead of using the coefficients? How would you explain to your friend that this tactic is the wrong approach?

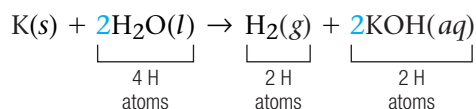
Solution

Step 1 From the description given for the reaction, we know that the reactants are solid potassium, $K(s)$, and liquid water, $H_2O(l)$. The products are gaseous hydrogen, $H_2(g)$, and dissolved potassium hydroxide, $KOH(aq)$.

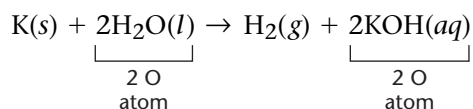
Step 2 The unbalanced equation for the reaction is



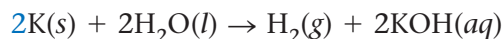
Step 3 Although none of the reactants or products is very complicated, we will start with KOH because it contains the most elements (three). We will arbitrarily consider hydrogen first. Note that on the reactant side of the equation in step 2, there are two hydrogen atoms, but on the product side, there are three. If we place a coefficient of 2 in front of both H_2O and KOH , we now have four H atoms on each side.



Also note that the oxygen atoms balance.

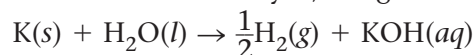


However, the K atoms do not balance; we have one on the left and two on the right. We can fix this easily by placing a coefficient of 2 in front of $K(s)$ to give the balanced equation:



Step 4

Check: There are 2 K, 4 H, and 2 O on both sides of the arrow, and the coefficients are the smallest whole numbers that give a balanced equation. We know this because we cannot divide through by a given whole number to give a set of smaller whole-number coefficients. For example, if we divide all of the coefficients by 2, we get



This is not acceptable because the coefficient for H_2 is not a whole number.

Reactants	Products
2 K	2 K
4 H	4 H
2 O	2 O



HANDS-ON CHEMISTRY

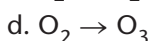
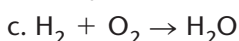
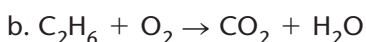
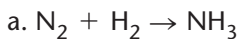
• MINI-LAB •

Modeling Equations**Materials**

- four different colors of modeling clay
- toothpicks

Procedure

1. Obtain four different colors of modeling clay.
2. With the clay and toothpicks, make models of each of the reactants and products in the following equations, and use these models to balance the equations.

**Results/Analysis**

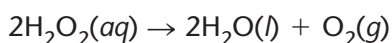
1. How do your models show the difference between a subscript and a coefficient?
2. How do your models show why we cannot change a subscript to balance an equation?

CHEMISTRY IN YOUR WORLD

Connection to Biology

The Bombardier Beetle

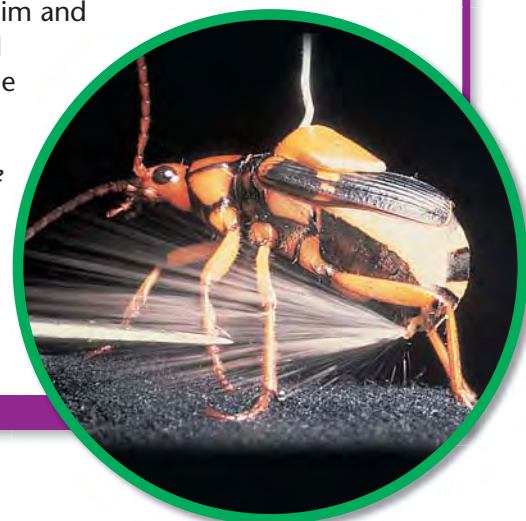
If someone said to you, “Name something that protects itself by spraying its enemies,” your answer would almost certainly be “a skunk.” Of course, you would be correct, but there is another correct answer—the bombardier beetle. When threatened, this beetle shoots a boiling stream of toxic chemicals at its enemy. How does this clever beetle accomplish this? Obviously, the boiling mixture cannot be stored inside the beetle’s body all the time. Instead, when endangered, the beetle mixes chemicals that produce the hot spray. The chemicals involved are stored in two compartments. One compartment contains the chemicals hydrogen peroxide (H_2O_2) and methylhydroquinone ($\text{C}_7\text{H}_8\text{O}_2$). The key reaction is the decomposition of hydrogen peroxide to form oxygen gas and water:



Hydrogen peroxide also reacts with the hydroquinones to produce other compounds that become part of the toxic spray.

However, none of these reactions occurs very fast unless certain enzymes are present. (Enzymes are natural substances that speed up biological reactions by means we will not discuss here.) When the beetle mixes the hydrogen peroxide and hydroquinones with the enzyme, the decomposition of H_2O_2 occurs rapidly, producing a hot mixture pressurized by the formation of oxygen gas. When the gas pressure becomes high enough, the hot spray is ejected in one long stream or in short bursts. The beetle has a highly accurate aim and can shoot several attackers with one batch of spray.

A bombardier beetle defending itself



EXAMPLE 7.3

Balancing Chemical Equations II

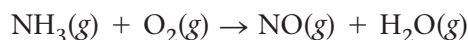
Under appropriate conditions at 1000 °C, ammonia gas reacts with oxygen gas to produce gaseous nitrogen monoxide (common name, nitric oxide) and gaseous water.

- Write the unbalanced equation.
- Balance the equation for this reaction.

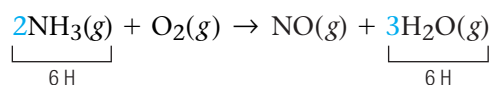
Solution

Step 1 The reactants are gaseous ammonia, $\text{NH}_3(\text{g})$, and gaseous oxygen, $\text{O}_2(\text{g})$. The products are gaseous nitrogen monoxide, $\text{NO}(\text{g})$, and gaseous water, $\text{H}_2\text{O}(\text{g})$.

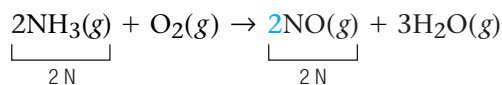
Step 2 The unbalanced equation for the reaction is



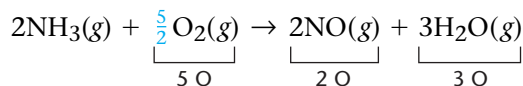
Step 3 In this equation there is no molecule that is obviously the most complicated. Three molecules contain two elements, so we arbitrarily start with NH_3 . We arbitrarily begin by looking at hydrogen. A coefficient of 2 for NH_3 and a coefficient of 3 for H_2O give six atoms of hydrogen on both sides.



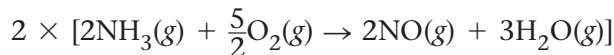
We can balance the nitrogen by giving NO a coefficient of 2.



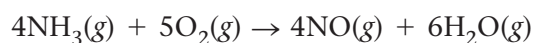
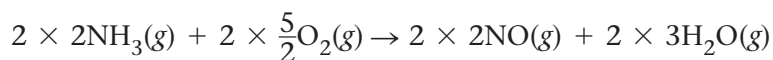
Finally, we note that there are two atoms of oxygen on the left and five on the right. The oxygen can be balanced with a coefficient of $\frac{5}{2}$ for O_2 , because $\frac{5}{2} \times \text{O}_2$ gives five oxygen atoms.



However, the convention is to have whole-number coefficients, so we multiply the entire equation by 2.



or



Step 4

Check: There are 4 N, 12 H, and 10 O atoms on both sides, so the equation is balanced. These coefficients are the smallest whole numbers that give a balanced equation.

Reactants	Products
1 N	1 N
3 H	2 H
2 O	1 O

Reactants	Products
4 N	4 N
12 H	12 H
10 O	10 O

This description of a chemical process contains many words, some of which are crucial to solving the problem and some of which are not. First sort out the important information and use symbols to represent it.

Practice Problem • Exercise 7.3

Propane, C_3H_8 , a liquid at 25 °C under high pressure, is often used for gas grills and as a fuel in rural areas where there is no natural gas pipeline. When liquid propane is released from its storage tank, it changes to propane gas that reacts with oxygen gas (it “burns”) to give gaseous carbon dioxide and gaseous water. Write and balance the equation for this reaction.

EXAMPLE 7.4

Balancing Chemical Equations III

Glass is sometimes decorated by etching patterns on its surface. Etching occurs when hydrofluoric acid, an aqueous solution of HF, reacts with the silicon dioxide in the glass to form gaseous silicon tetrafluoride and liquid water.

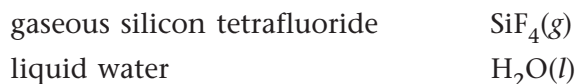
- Write the equation for the reaction.
- Balance the equation.

Solution

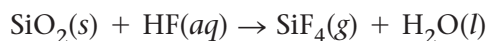
Step 1 From the description of the reaction we can identify the reactants:



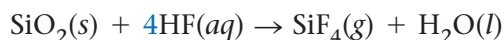
and the products:



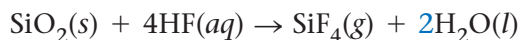
Step 2 The unbalanced equation is



Step 3 There is no clear choice here for the most complicated molecule. We arbitrarily start with the elements in SiF_4 . The silicon is balanced (one atom on each side), but the fluorine is not. To balance the fluorine, we need a coefficient of 4 before the HF.

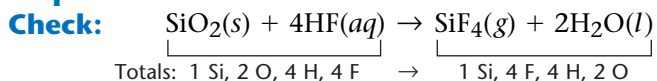


Hydrogen and oxygen are not balanced. Because we have four hydrogen atoms on the left and two on the right, we place a 2 before the H_2O :



This balances the hydrogen *and* the oxygen (two atoms on each side).

Step 4



All atoms check, so the equation is balanced.



Decorations on glass are produced by etching with hydrofluoric acid.

Practice Problem • Exercise 7.4

Give the balanced equation for each of the following reactions.

- When solid ammonium nitrite is heated, it produces nitrogen gas and water vapor.
- Gaseous nitrogen monoxide (common name, nitric oxide) decomposes to produce dinitrogen monoxide gas (common name, nitrous oxide) and nitrogen dioxide gas.
- Liquid nitric acid decomposes to reddish-brown nitrogen dioxide gas, liquid water, and oxygen gas. (This is why bottles of nitric acid become yellow upon standing.)

SECTION 7.3 REVIEW QUESTIONS

- Write and balance each of the following chemical equations.
 - Gaseous chlorine reacts with an aqueous solution of potassium bromide to form liquid bromine and an aqueous solution of potassium chloride.
 - Solid aluminum reacts with solid iodine to produce solid aluminum iodide.



- Solid magnesium reacts with an aqueous solution of hydrochloric acid to form an aqueous solution of magnesium chloride bubbles of hydrogen gas.

- Which can be changed when balancing a chemical equation, the coefficients or the subscripts? Why?
- Why does each chemical reaction have many possible balanced equations? Of all possible balanced equations, which is preferred?
- Explain what each of the following symbols (shown in red) represents:
$$\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(g)$$
- When balancing the equation

$$\text{C}_2\text{H}_5\text{OH}(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)$$
,
which substance should you start with? Why? Balance the chemical equation.



Key Terms

7.1

7.2 Chemical reaction

Chemical equation

Reactants

Products

Balancing the chemical equation

7.3 Coefficients

Key Ideas

7.1 Evidence for a Chemical Reaction

- A chemical reaction produces a signal that it has occurred. These signals include
 - Color change
 - Solid formation
 - Bubble formation
 - Heat
 - Flame

7.2 Chemical Equations

- The physical states of reactants and products in a reaction are indicated by the following symbols.

Physical States	
Symbol	State
(s)	solid
(l)	liquid
(g)	gas
(aq)	dissolved in water (in aqueous solution)

- Chemical reactions involve a rearrangement of the ways atoms are grouped together.
- A chemical equation represents a chemical reaction.
 - Reactants are shown to the left of an arrow.
 - Products are shown to the right of the arrow.
- In a chemical reaction atoms are neither created nor destroyed. A balanced chemical equation must have the same number of each type of atom on the reactant and product sides.
- A balanced chemical equation uses numbers (coefficients) in front of the reactant and product formulas to show the relative numbers of each.

7.3 Balancing Chemical Equations

- A chemical reaction is balanced by using a systematic approach.
 - Write the formulas of the reactants and products to give the unbalanced chemical equation.
 - Balance by trial and error, starting with the most complicated molecule(s).
 - Check to be sure the equation is balanced (same numbers of all types of atoms on the reactant and product sides).

All exercises with **blue numbers** have answers in the back of this book.

7.1 Evidence for a Chemical Reaction

1. Review how we defined *chemical change* in Chapter 3. Explain how the visual clues mentioned in this chapter are indicators of chemical change as defined in Chapter 3.
2. If a piece of blackboard chalk is heated strongly, the mass of the chalk decreases substantially, and the chalk is converted into a fine powder. What evidence is there for a chemical reaction having taken place when the chalk is heated?
3. If you have had a clogged sink drain at your home, your parents may have tried using a commercial drain cleaner to dissolve the clog. What evidence is there that such drain cleaners work by chemical reaction?
4. If a bottle of aspirin is left open to the air, eventually one of the products is acetic acid (vinegar is a mixture of acetic acid and water). Is there evidence that this change represents a chemical reaction?

7.2 Chemical Equations

5. The substances present before a chemical reaction takes place are called the _____, and the substances present after the reaction takes place are called the _____.
6. In an ordinary chemical reaction, _____ are neither created nor destroyed.
7. The notation “(g)” after a substance’s formula indicates it exists in the _____ state.
8. In a chemical equation for a reaction, the notation “(aq)” after a substance’s formula means that the substance is dissolved in _____.

Note: In some of the following problems, you will need to write a chemical formula from the name of the compound. It may help to review Chapter 4.

9. If calcium carbonate is heated strongly, carbon dioxide gas is driven off, leaving a residue of calcium oxide. Write the unbalanced chemical equation for this process.
10. Liquefied propane gas is often used for cooking in rural areas away from natural gas lines. Propane (C_3H_8) burns in oxygen gas, producing carbon dioxide gas, water vapor, and heat. Write the unbalanced chemical equation for this process.

11. If a sample of pure hydrogen gas is ignited very carefully, the hydrogen burns gently, combining with the oxygen gas of the air to form water vapor. Write the unbalanced chemical equation for this reaction.
12. Solid ammonium carbonate, $(NH_4)_2CO_3$, is used as the active ingredient in “smelling salts.” When solid ammonium carbonate is heated, it decomposes into ammonia gas, carbon dioxide gas, and water vapor. Write the unbalanced chemical equation for this process.
13. Silver oxide may be decomposed by strong heating into silver metal and oxygen gas. Write the unbalanced chemical equation for this process.
14. Methanol (CH_3OH , wood alcohol) is an important industrial chemical. Methanol may be synthesized from carbon monoxide gas and hydrogen gas under certain conditions of temperature and pressure. Write the unbalanced chemical equation for this process.
15. Elemental boron is produced in one industrial process by heating diboron trioxide with magnesium metal, also producing magnesium oxide. Write the unbalanced chemical equation for this process.
16. Calcium metal is moderately reactive. If pieces of calcium are added to water, the metal begins to bubble as hydrogen gas is formed. The water begins to turn cloudy as solid calcium hydroxide begins to form. Write the unbalanced chemical equation for the reaction of calcium metal with water.



17. Phosphorus trichloride is used in the manufacture of certain pesticides and may be synthesized directly from its elements. Write the unbalanced chemical equation for this process.
18. Magnesium hydroxide has been used for many years as an antacid (“milk of magnesia”) because it reacts with the hydrochloric acid in the stomach, producing magnesium chloride and water. Write the unbalanced chemical equation for this process.
19. Nitrous oxide gas (systematic name: dinitrogen monoxide) is used by some dentists as an anesthetic. Nitrous oxide (and water vapor) can be produced in small amounts in the laboratory by careful heating of ammonium nitrate. Write the unbalanced chemical equation for this reaction.

- 20.** Hydrogen sulfide gas is responsible for the odor of rotten eggs. Hydrogen sulfide burns in air, producing sulfur dioxide gas and water vapor. Write the unbalanced chemical equation for this process.
- 21.** Acetylene gas (C_2H_2) is often used by plumbers, welders, and glass blowers because it burns in oxygen with an intensely hot flame. The products of the combustion of acetylene are carbon dioxide and water vapor. Write the unbalanced chemical equation for this process.



- 22.** If ferric oxide is heated strongly in a stream of carbon monoxide gas, it produces elemental iron and carbon dioxide gas. Write the unbalanced chemical equation for this process.
- 23.** The Group 2 metals (Ba, Ca, Sr) can be produced in the elemental state by the reaction of their oxides with aluminum metal at high temperatures, also producing solid aluminum oxide. Write the unbalanced chemical equations for the reactions of barium oxide, calcium oxide, and strontium oxide with aluminum.
- 24.** Ozone gas is a form of elemental oxygen containing molecules with *three* oxygen atoms, O_3 . Ozone is produced from atmospheric oxygen gas, O_2 , by the high-energy outbursts found in lightning storms. Write the unbalanced equation for the formation of ozone gas from oxygen gas.
- 25.** Carbon tetrachloride may be prepared by the reaction of natural gas (methane, CH_4) and chlorine gas in the presence of ultraviolet light. Write the unbalanced chemical equation for this process. (HCl is also a product.)
- 26.** Ammonium nitrate is used as a “high-nitrogen” fertilizer, despite the fact that it is quite explosive if not handled carefully. Ammonium nitrate can be synthesized by the reaction of ammonia gas and nitric acid. Write the unbalanced chemical equation for this process.
- 27.** The principal natural ore of lead is galena, which is primarily lead(II) sulfide. Lead(II) sulfide can be converted to lead(II) oxide and sulfur dioxide gas by heating strongly in air. Write the unbalanced chemical equation for the reaction of lead(II) sulfide with oxygen gas.

- 28.** At high temperatures, xenon gas will combine directly with fluorine gas to produce solid xenon tetrafluoride. Write the unbalanced chemical equation for this process.
- 29.** Ammonium nitrate is highly explosive if not handled carefully, breaking down into nitrogen gas, oxygen gas, and water vapor. The expansion of the three gases produced yields the explosive force in this case. Write the unbalanced chemical equation for this process.
- 30.** Silver nitrate is used in “styptic pencils,” which help to cauterize small nicks and cuts that occur during shaving. Silver nitrate can be prepared by dissolving metallic silver in concentrated nitric acid, with hydrogen gas being an additional product of the reaction. Write the unbalanced chemical equation for this process.

7.3 Balancing Chemical Equations

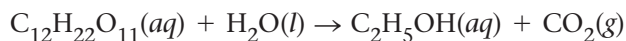
- 31.** When balancing a chemical equation, one must never change the _____ of any reactant or product.
- 32.** After balancing a chemical equation, we ordinarily make sure that the coefficients are the smallest _____ possible.
- 33.** Balance each of the following chemical equations.
- $H_2O_2(aq) \rightarrow H_2O(l) + O_2(g)$
 - $Ag(s) + H_2S(g) \rightarrow Ag_2S(s) + H_2(g)$
 - $FeO(s) + C(s) \rightarrow Fe(l) + CO_2(g)$
 - $Cl_2(g) + KI(aq) \rightarrow KCl(aq) + I_2(s)$
- 34.** Balance each of the following chemical equations.
- $CaF_2(s) + H_2SO_4(l) \rightarrow CaSO_4(s) + HF(g)$
 - $KBr(s) + H_3PO_4(aq) \rightarrow K_3PO_4(aq) + HBr(g)$
 - $TiCl_4(l) + Na(s) \rightarrow NaCl(s) + Ti(s)$
 - $K_2CO_3(s) \rightarrow K_2O(s) + CO_2(g)$
- 35.** Balance each of the following chemical equations.
- $SiI_4(s) + Mg(s) \rightarrow Si(s) + MgI_2(s)$
 - $MnO_2(s) + Mg(s) \rightarrow Mn(s) + MgO(s)$
 - $Ba(s) + S_8(s) \rightarrow BaS(s)$
 - $NH_3(g) + Cl_2(g) \rightarrow NH_4Cl(s) + NCl_3(g)$
- 36.** Balance each of the following chemical equations.
- $Ba(NO_3)_2(aq) + Na_2CrO_4(aq) \rightarrow BaCrO_4(s) + NaNO_3(aq)$
 - $C_2H_5OH(l) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
 - $CaC_2(s) + H_2O(l) \rightarrow Ca(OH)_2(s) + C_2H_2(g)$

Critical Thinking

37. Many ships are built with aluminum superstructures to save weight. Aluminum, however, burns in oxygen if there is a sufficiently hot ignition source, which has led to several tragedies at sea. Write the unbalanced chemical equation for the reaction of aluminum with oxygen, producing aluminum oxide as the product.

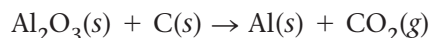
38. Crude gunpowders often contain a mixture of potassium nitrate and charcoal (carbon). When such a mixture is heated until reaction occurs, a solid residue of potassium carbonate is produced. The explosive force of the gunpowder comes from the fact that two gases are also produced (carbon monoxide and nitrogen), which increase in volume with great force and speed. Write the unbalanced chemical equation for the process.

39. The sugar sucrose, which is present in many fruits and vegetables, reacts in the presence of certain yeast enzymes to produce ethyl alcohol (ethanol) and carbon dioxide gas. Balance the following equation for this reaction of sucrose.

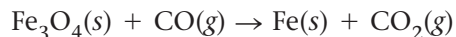
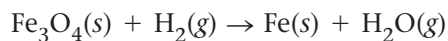


40. Methanol (methyl alcohol), CH_3OH , is a very important industrial chemical. Today, methanol is synthesized from carbon monoxide and elemental hydrogen. Write the balanced chemical equation for this process.

41. The Hall process is an important method by which pure aluminum is prepared from its oxide (alumina, Al_2O_3) by indirect reaction with graphite (carbon). Balance the following equation, which is a simplified representation of this process.



42. Iron oxide ores, commonly a mixture of FeO and Fe_2O_3 , are given the general formula Fe_3O_4 . They yield elemental iron when heated to a very high temperature with either carbon monoxide or elemental hydrogen. Balance the following equations for these processes.

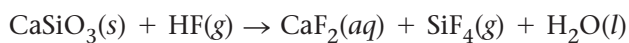


43. The elements of Group 1 all react with sulfur to form the metal sulfides. Write balanced chemical equations for the reactions of the Group 1 elements with sulfur.

44. A common experiment in chemistry classes involves heating a weighed mixture of potassium chlorate, KClO_3 , and potassium chloride. Potassium chlorate decomposes when heated, producing potassium chloride and oxygen gas. By measuring the volume of oxygen gas produced in this experiment, students can calculate the relative percentage of KClO_3 and KCl in the original mixture. Write the balanced chemical equation for this process.

45. A common demonstration in chemistry classes involves adding a tiny speck of manganese(IV) oxide to a concentrated hydrogen peroxide, H_2O_2 , solution. Hydrogen peroxide decomposes quite spectacularly under these conditions to produce oxygen gas and steam (water vapor). Manganese(IV) oxide speeds up the decomposition of hydrogen peroxide but is not consumed in the reaction. Write the balanced equation for the decomposition reaction of hydrogen peroxide.

46. Glass is a mixture of several compounds, but a major constituent of most glass is calcium silicate, CaSiO_3 . Glass can be etched by treatment with hydrogen fluoride: HF attacks the calcium silicate of the glass, producing gaseous and water-soluble products that can be removed by washing the glass. For example, the volumetric glassware in chemistry laboratories is often graduated by using this process. Balance the following equation for the reaction of hydrogen fluoride with calcium silicate.



47. If you had a “sour stomach,” you might try an antacid tablet to relieve the problem. Can you think of evidence that the action of such an antacid is a chemical reaction?

48. If solutions of potassium chromate and barium chloride are mixed, a bright yellow solid (barium chromate) forms and settles out of the mixture, leaving potassium chloride in solution. Write a balanced chemical equation for this process.

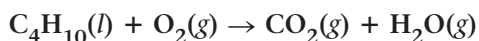
49. When a strip of magnesium metal is heated in oxygen, it bursts into an intensely white flame and produces a finely powdered dust of magnesium oxide. Write the unbalanced chemical equation for this process.

- 1 Which of the following is not a clue that a chemical reaction has taken place?
- A A solid is formed when two clear solutions are mixed.
- B A clear solution is added to a red solution, and the result is a blue solution.
- C A solid is added to water, and bubbles form.
- D A pure solid is heated and turns into a pure liquid.

- 2 What does it mean for a chemical equation to be balanced?
- A The sums of the coefficients on each side of the equation are equal.
- B The same number of each type of atom appears on both sides of the equation.
- C The formulas of the reactants and the products are the same.
- D The sums of the subscripts on each side of the equation are equal.

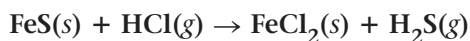
- 3 Which of the following physical states is not used for H_2O in a chemical equation?
- A (s)
- B (l)
- C (aq)
- D (g)

- 4 Which of the following is the coefficient for O_2 when the equation is balanced?



- A 4
- B 8
- C 10
- D 13

- 5 What is the sum of the coefficients of the following equation when it is balanced?

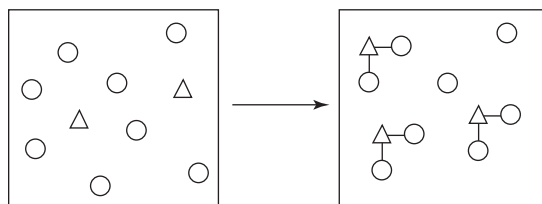


- A 5
- B 4
- C 3
- D 2

- 6 When the equation representing the reaction between iron(II) oxide and carbon to form iron and carbon dioxide is balanced, the coefficient for iron is

- A 1
- B 2
- C 3
- D 4

- 7 The reaction of an element X (Δ) with element Y (\circ) is represented in the following diagram. Which of the equations best describes this reaction?



- A $3\text{X} + 8\text{Y} \rightarrow \text{X}_3\text{Y}_8$
- B $3\text{X} + 6\text{Y} \rightarrow \text{X}_3\text{Y}_6$
- C $\text{X} + 2\text{Y} \rightarrow \text{XY}_2$
- D $3\text{X} + 8\text{Y} \rightarrow 3\text{XY}_2 + 2\text{Y}$

- 8 Explain the law of conservation of mass and discuss how it applies to balancing a chemical equation.

For questions 9–12, write and balance the chemical equation as described. Include the state of each substance.

- 9 When magnesium metal is added to water, hydrogen gas is produced along with a solution of magnesium hydroxide.
- 10 When an aqueous solution of lead(II) nitrate is added to an aqueous solution of sodium iodide, the yellow solid lead(II) iodide is produced along with a solution of sodium nitrate.
- 11 Propane gas (C_3H_8) reacts with oxygen in the air to form carbon dioxide and water vapor.
- 12 Calcium metal reacts with oxygen gas to produce solid calcium oxide.

Reactions in Aqueous Solutions

LOOKING AHEAD

8.1 Understanding Reactions in Aqueous Solutions

- A. Predicting Whether a Reaction Will Occur
- B. Reactions in Which a Solid Forms
 - What Happens When an Ionic Compound Dissolves in Water?
 - How to Decide What Products Form
 - Using Solubility Rules
- C. Describing Reactions in Aqueous Solutions
 - A Closer Look

8.2 Other Reactions in Aqueous Solutions

- A. Reactions that Form Water: Acids and Bases
- B. Reactions of Metals with Nonmetals (Oxidation–Reduction)

8.3 Classifying Reactions

- A. Ways to Classify Reactions
- B. Other Ways to Classify Reactions
 - Combustion Reactions
 - Synthesis (Combination) Reactions
 - Decomposition Reactions

Potassium metal reacts vigorously with water.



IN YOUR LIFE

The chemical reactions that are most important to us occur in water—in aqueous solutions. Virtually all of the chemical reactions that keep each of us alive and well take place in the aqueous medium present in our bodies. For example, the oxygen you breathe dissolves in your blood, where it associates with the hemoglobin in the red blood cells. While attached to the hemoglobin it is transported to your cells, where it reacts with fuel from the food you eat to provide energy for living. However, the reaction between oxygen and fuel is not direct—the cells are not tiny furnaces. Instead, electrons are transferred from the fuel to a series of molecules that pass them along (this is called the respiratory chain) until they eventually reach oxygen. Many other reactions are also crucial to our health and well-being. You will see numerous examples of these as you continue your study of chemistry.



Oxygen is dissolved in the bloodstream of these runners to react with food chemicals to provide energy.

WHAT DO YOU KNOW?



Prereading Questions

1. If you could see an extremely magnified view of a solution of sodium chloride, what would it look like? What about a solution of barium nitrate?
2. Have you ever heard of *acid* or *base*? Give an example of each.
3. Have you heard of the terms *oxidation* or *reduction*? What do they mean?

SECTION 8.1

Understanding Reactions in Aqueous Solutions

Key Terms

- Precipitation
- Precipitate
- Precipitation reaction
- Strong electrolyte
- Soluble solid
- Insoluble solid (slightly soluble solid)
- Molecular equation
- Complete ionic equation
- Spectator ions
- Net ionic equation

Objectives

- To learn about some of the factors that cause reactions to occur
- To learn to identify the solid that forms in a precipitation reaction
- To learn to write molecular, complete ionic, and net ionic equations

A. Predicting Whether a Reaction Will Occur

In this text we have already seen many chemical reactions. Now let's consider an important question: Why does a chemical reaction occur? What causes reactants to “want” to form products? As chemists have studied reactions, they have recognized several “tendencies” in reactants that drive them to form products. That is, there are several “driving forces” that pull reactants toward products—changes that tend to make reactions go in the direction of the arrow. The most common of these driving forces are

- Formation of a solid
- Formation of water
- Transfer of electrons
- Formation of a gas

When two or more chemicals are brought together, if any of these things can occur, a chemical change (a reaction) is likely to take place. Accordingly, when we are confronted with a set of reactants and want to predict whether a reaction will occur and what products might form, we will consider these driving forces. They will help us organize our thoughts as we encounter new reactions.



Figure 8.1

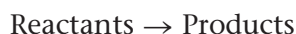
The precipitation reaction that occurs when yellow potassium chromate, $\text{K}_2\text{CrO}_4(aq)$, is mixed with a colorless barium nitrate solution, $\text{Ba}(\text{NO}_3)_2(aq)$

Active Reading Question

Provide an example for each of the driving forces.

B. Reactions in Which a Solid Forms

One driving force for a chemical reaction is the formation of a solid, a process called **precipitation**. The solid that forms is called a **precipitate**, and the reaction is known as a **precipitation reaction**. For example, when an aqueous (water) solution of potassium chromate, $\text{K}_2\text{CrO}_4(aq)$, which is yellow, is added to a colorless aqueous solution containing barium nitrate, $\text{Ba}(\text{NO}_3)_2(aq)$, a yellow solid forms (see **Figure 8.1**). The fact that a solid forms tells us that a reaction—a chemical change—has occurred. That is, we have a situation where



What is the equation that describes this chemical change? To write the equation, we must decipher the identities of the reactants and products. The reactants have already been described: $\text{K}_2\text{CrO}_4(aq)$ and $\text{Ba}(\text{NO}_3)_2(aq)$. Is there some way in which we can predict the identities of the products? What is the yellow solid? The best way to predict the identity of this solid is to first *consider what products are possible*. To do this we need to know what chemical species are present in the solution that results when the reactant solutions are mixed. First, let's think about the nature of each reactant in an aqueous solution.

Precipitation

The formation of a solid in a chemical reaction

Precipitate

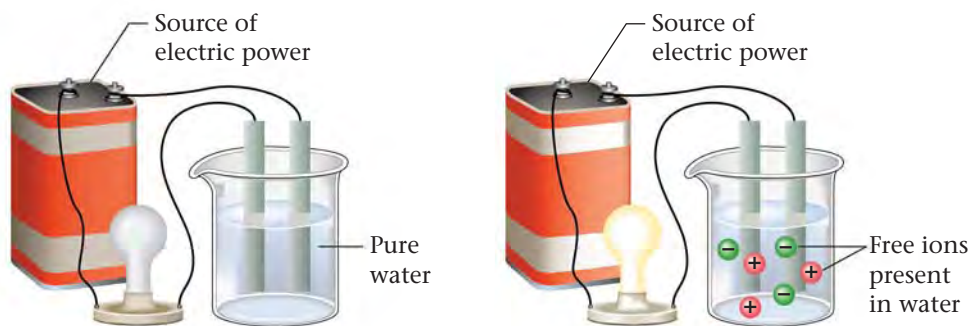
The solid that forms in a precipitation reaction

Precipitation reaction

A reaction in which a solid forms and separates from the solution

What Happens When an Ionic Compound Dissolves in Water?

The designation $\text{Ba}(\text{NO}_3)_2(aq)$ means that barium nitrate (a white solid) has been dissolved in water. Note from its formula that barium nitrate contains the Ba^{2+} and NO_3^- ions. *In virtually every case when a solid containing ions dissolves in water, the ions separate and move around independently.* We say that the ions of the solid *dissociate* when the solid dissolves in water. That is, $\text{Ba}(\text{NO}_3)_2(aq)$ does not contain $\text{Ba}(\text{NO}_3)_2$ units. Rather, it contains separated Ba^{2+} and NO_3^- ions. In the solution there are two NO_3^- ions for every Ba^{2+} ion. Chemists know that separated ions are present in this solution because it is an excellent conductor of electricity (see **Figure 8.2**). Pure water does not conduct an electric current. Ions must be present in water for a current to flow.



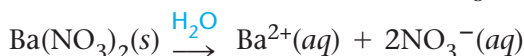
(a) Pure water does not conduct an electric current. The lamp does not light.

(b) When an ionic compound is dissolved in water, current flows and the lamp lights.

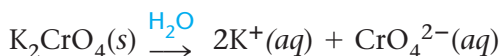
Figure 8.2

Electrical conductivity of aqueous solutions. The result of this experiment is strong evidence that ionic compounds dissolved in water exist in the form of separated ions.

When each unit of a substance that dissolves in water produces separated ions, the substance is called a **strong electrolyte**. Barium nitrate is a strong electrolyte in water, because each $\text{Ba}(\text{NO}_3)_2$ unit produces the separated ions (Ba^{2+} , NO_3^- , NO_3^-).



Similarly, aqueous K_2CrO_4 also behaves as a strong electrolyte. Potassium chromate contains the K^+ and CrO_4^{2-} ions, so an aqueous solution of potassium chromate (which is prepared by dissolving solid K_2CrO_4 in water) contains these separated ions.



That is, $\text{K}_2\text{CrO}_4(aq)$ does not contain K_2CrO_4 units but instead contains K^+ cations and CrO_4^{2-} anions, which move around independently. (There are two K^+ ions for each CrO_4^{2-} ion.) The idea introduced here is very important:

When ionic compounds dissolve, the resulting solution contains the separated ions.

Therefore, we can represent the mixing of $\text{K}_2\text{CrO}_4(aq)$ and $\text{Ba}(\text{NO}_3)_2(aq)$ in two ways. We usually write these reactants as:



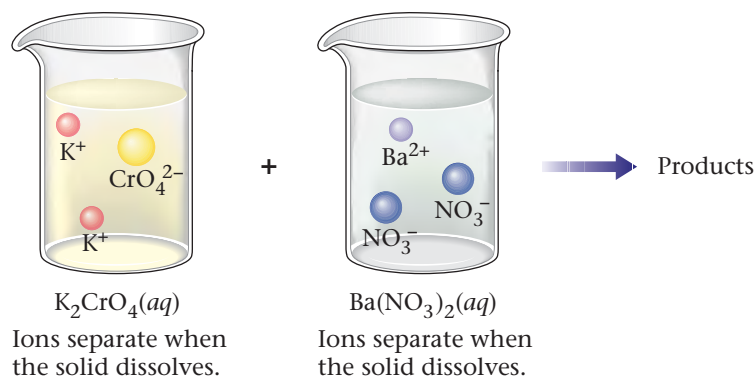
Strong electrolyte

A substance that dissolves in water by dissociating completely into ions

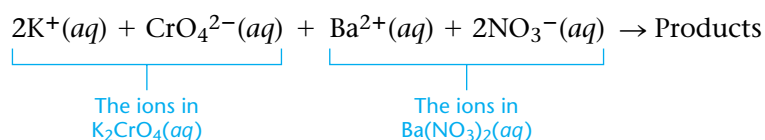
Information

Note that the H_2O over the arrow here indicates that the substance is being dissolved in water.

However, a more accurate representation of the situation is:



We can express this information in equation form as follows:



Thus the *mixed solution* contains four types of ions: K^+ , CrO_4^{2-} , Ba^{2+} , and NO_3^- . Now that we know what the reactants are, we can make some educated guesses about the possible products.

How to Decide What Products Form

Which of these ions combine to form the yellow solid observed when the original solutions are mixed? This is not an easy question to answer. Even an experienced chemist is not sure what will happen in a new reaction. The chemist tries to think of the various possibilities, considers the likelihood of each possibility, and then makes a prediction (an educated guess). Only after identifying each product experimentally can the chemist be sure what reaction actually has taken place. However, an educated guess

CHEMISTRY IN YOUR WORLD

Consumer Connection

Instant Cooking—On Demand

Can you think of foods you've recently prepared that were ready to go—except for adding water? From breakfast through dessert, chemists have found ways to prepare mixes that stay unreacted and ready to use on our shelves. All we need to do is place them in a bowl, add water, and cook! Why don't the ingredients react in the box?

Most of the ingredients used in mixes are solids. We've already seen that solids don't usually react until they are dissolved in water. So as long as moisture is kept out of the mix, it will remain unreacted.

Sometimes manufacturers also dehydrate ingredients such as milk and eggs so they can be stored in unreactive form in the mix as well.



is very useful because it indicates what kinds of products are most likely. It gives us a place to start. So the best way to proceed is first to think of the various possibilities and then to decide which of them is most likely.

Possible Products What are the possible products of the reaction between $\text{K}_2\text{CrO}_4(aq)$ and $\text{Ba}(\text{NO}_3)_2(aq)$ or, more accurately, what reaction can occur among the ions K^+ , CrO_4^{2-} , Ba^{2+} , and NO_3^- ? We already know some things that will help us decide.

A solid compound must have a zero net charge. This means that the product of our reaction must contain both anions and cations (negative and positive ions).

For example, K^+ and Ba^{2+} could not combine to form the solid because such a solid would have a positive charge. Similarly, CrO_4^{2-} and NO_3^- could not combine to form a solid because that solid would have a negative charge.

Something else that will help us is an observation that chemists have made by examining many compounds: *most ionic materials contain only two types of ions*—one type of cation and one type of anion. This idea is illustrated by the following compounds (among many others):

Ionic Materials		
Compound	Cation	Anion
NaCl	Na^+	Cl^-
KOH	K^+	OH^-
Na_2SO_4	Na^+	SO_4^{2-}
NH_4Cl	NH_4^+	Cl^-
Na_2CO_3	Na^+	CO_3^{2-}

All the possible combinations of a cation and an anion to form uncharged compounds from among the ions K^+ , CrO_4^{2-} , Ba^{2+} , and NO_3^- are shown below:

	NO_3^-	CrO_4^{2-}
K^+	KNO_3	K_2CrO_4
Ba^{2+}	$\text{Ba}(\text{NO}_3)_2$	BaCrO_4

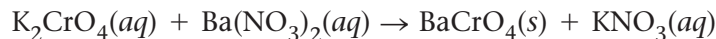
So the compounds in red *might* be the solid.

Which of these possibilities is most likely to represent the yellow solid? We know it's not K_2CrO_4 or $\text{Ba}(\text{NO}_3)_2$; these are the reactants. They were present (dissolved) in the separate solutions that were mixed initially. The only real possibilities are KNO_3 and BaCrO_4 . To decide which of these is more likely to represent the yellow solid, we need more facts. An experienced chemist, for example, knows that KNO_3 is a white solid. On the other hand, the CrO_4^{2-} ion is yellow. Therefore, the yellow solid must be BaCrO_4 .

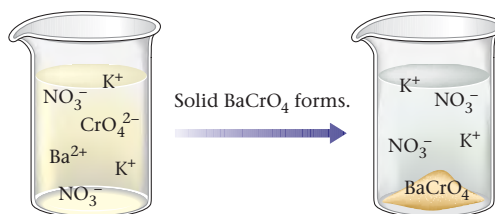
We have determined that one product of the reaction between $\text{K}_2\text{CrO}_4(aq)$ and $\text{Ba}(\text{NO}_3)_2(aq)$ is $\text{BaCrO}_4(s)$, but what happened to the K^+ and NO_3^- ions? The answer is that these ions are left dissolved in the solution. That is, KNO_3 does not form a solid when the K^+ and NO_3^- ions are present in water. In other words, if we took the white solid $\text{KNO}_3(s)$ and put it in water, it would

totally dissolve (the white solid would “disappear,” yielding a colorless solution). So when we mix $\text{K}_2\text{CrO}_4(aq)$ and $\text{Ba}(\text{NO}_3)_2(aq)$, $\text{BaCrO}_4(s)$ forms but KNO_3 is left behind in solution—we write it as $\text{KNO}_3(aq)$. If we poured the mixture through a filter to remove the solid BaCrO_4 and then evaporated all of the water, we would obtain the white solid KNO_3 .

After all this thinking, we can finally write the unbalanced equation for the precipitation reaction.



We can represent this reaction in pictures as follows:



Note that the K^+ and NO_3^- ions are not involved in the chemical change. They are present in the solution before and after the reaction.

Active Reading Question

How do we know that K_2CrO_4 and $\text{Ba}(\text{NO}_3)_2$ are not the solids?

Using Solubility Rules

In our example above we were finally able to identify the products of the reaction by using two types of chemical knowledge:

- 1. Facts:** Knowing the colors of the compounds was very helpful.
- 2. Concepts:** Knowing the concept that solids always have a net charge of zero was also important.

These two kinds of knowledge allowed us to make a good guess about the identity of the solid that formed. As you continue to study chemistry, you will see that a balance of factual and conceptual knowledge is always required. You must both *memorize* important facts and *understand* crucial concepts to succeed.

Information

Solids must contain both anions and cations in the relative numbers necessary to produce zero net charge.

Does a Solid Always Form? In the present case we are dealing with a reaction in which an ionic solid forms—that is, a process in which ions that are dissolved in water combine to give a solid. We know that for a solid to form, both positive and negative ions must be present in relative numbers that give zero net charge. However, oppositely charged ions in water do not always react to form a solid, as we have seen for K^+ and NO_3^- . In addition, Na^+ and Cl^- can coexist in water in very large numbers with no formation of solid NaCl . In other words, when solid NaCl (common salt) is placed in water, it dissolves—the white solid “disappears” as the Na^+ and Cl^- ions are dispersed throughout the water. The following two statements, then, are really saying the same thing:

1. Solid NaCl is very soluble in water.
2. Solid NaCl does not form when one solution containing Na^+ is mixed with another solution containing Cl^- .

Predicting Precipitates To predict whether a given pair of dissolved ions will form a solid when mixed, we must know some facts about the solubilities of various types of ionic compounds. In this text we will use the term **soluble solid** to mean a solid that readily dissolves in water; the solid “disappears” as the ions are dispersed in the water. The terms **insoluble solid** and **slightly soluble solid** are taken to mean the same thing: a solid where such a tiny amount dissolves in water that it is undetectable with the naked eye. The solubility information about common solids that is summarized in **Table 8.1** is based on observations of the behavior of many compounds. This is factual knowledge that you will need to predict what will happen in chemical reactions where a solid might form. This information is summarized in **Figure 8.3**.

Soluble solid

A solid that readily dissolves in water

Insoluble solid (slightly soluble solid)

A solid that dissolves to such a small degree that it is not detectable to the naked eye

Table 8.1

General Rules for Solubility of Ionic Compounds (Salts) in Water at 25 °C

1. Most nitrate (NO_3^-) salts are soluble.
2. Most salts of Na^+ , K^+ , and NH_4^+ are soluble.
3. Most chloride salts are soluble. Notable exceptions are AgCl , PbCl_2 , and Hg_2Cl_2 .
4. Most sulfate salts are soluble. Notable exceptions are BaSO_4 , PbSO_4 , and CaSO_4 .
5. Most hydroxide compounds are only slightly soluble.* The important exceptions are NaOH and KOH . $\text{Ba}(\text{OH})_2$ and $\text{Ca}(\text{OH})_2$ are moderately soluble.
6. Most sulfide (S^{2-}), carbonate (CO_3^{2-}), and phosphate (PO_4^{3-}) salts are only slightly soluble.*

*The terms *insoluble* and *slightly soluble* really mean the same thing: such a tiny amount dissolves that it is not possible to detect it with the naked eye.

(a) Soluble compounds

NO_3^- salts	
Na^+ , K^+ , NH_4^+ salts	
Cl^- , Br^- , I^- salts	Except for those containing Ag^+ , Hg_2^{2+} , Pb^{2+}
SO_4^{2-} salts	Except for those containing Ba^{2+} , Pb^{2+} , Ca^{2+}

(b) Insoluble compounds

S^{2-} , CO_3^{2-} , PO_4^{3-} salts	
OH^- salts	Except for those containing Na^+ , K^+ , Ca^{2+}

Figure 8.3

Solubilities of common compounds

Notice that in Table 8.1 and Figure 8.3 the term *salt* is used to mean *ionic compound*. Many chemists use the terms *salt* and *ionic compound* interchangeably. In Example 8.1, we will illustrate how to use the solubility rules to predict the products of reactions among ions.

Information

In predicting what reaction will occur when ions are mixed in water, focus on the ions that are in solution before any reaction occurs.

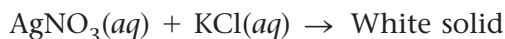
EXAMPLE 8.1

Identifying Precipitates in Reactions Where a Solid Forms

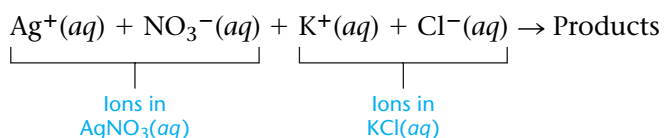
When an aqueous solution of silver nitrate is added to an aqueous solution of potassium chloride, a white solid forms. Identify the white solid and write the balanced equation for the reaction that occurs.

Solution

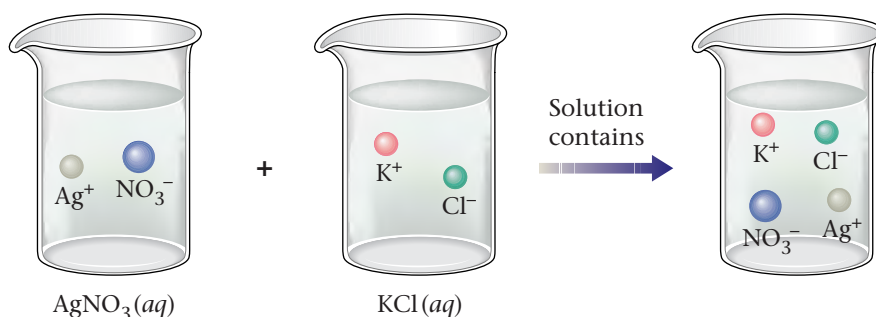
First let's write the equation for the reaction:



To answer the question "What is the white solid?" we must decide what ions are present in the mixed solution. Remember that *when ionic substances dissolve in water, the ions separate*. So we can write the equation



or use pictures to represent the ions present in the mixed solution before any reaction occurs.



Information

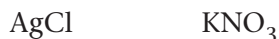
Use a table like this to figure out the possible combinations.

	NO_3^-	Cl^-
Ag^+	AgNO_3	AgCl
K^+	KNO_3	KCl

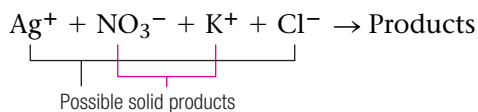
Now we will consider what solid *might* form from this collection of ions. Because the solid must contain both positive and negative ions, the possible compounds that can be assembled from this collection of ions are



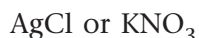
AgNO_3 and KCl are the substances already dissolved in the reactant solutions, so we know that they do not represent the white solid product. We are left with two possibilities:



Another way to obtain these two possibilities is by *ion interchange*. This means that in the reaction of $\text{AgNO}_3(aq)$ and $\text{KCl}(aq)$, we take the cation from one reactant and combine it with the anion of the other reactant.



Ion interchange leads to the following possible solids:



To decide whether AgCl or KNO_3 is the white solid, we need the solubility rules (Table 8.1).

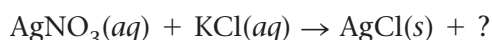
Rule 2 states that most salts containing K^+ are soluble in water.

Rule 1 says that most nitrate salts (those containing NO_3^-) are soluble.

So the salt KNO_3 is water-soluble. That is, when K^+ and NO_3^- are mixed in water, a solid (KNO_3) does *not* form.

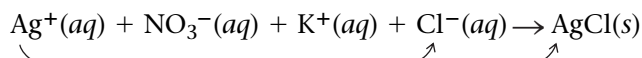
Rule 3 states that although most chloride salts (salts that contain Cl^-) are soluble, AgCl is an exception. That is, $\text{AgCl}(s)$ is insoluble in water.

Thus the white solid must be AgCl . Now we can write



What is the other product?

To form $\text{AgCl}(s)$, we have used the Ag^+ and Cl^- ions:



This leaves the K^+ and NO_3^- ions. What do they do? Nothing. Because KNO_3 is very soluble in water (rules 1 and 2), the K^+ and NO_3^- ions remain separate in the water; the KNO_3 remains dissolved and we represent it as $\text{KNO}_3(aq)$. We can now write the full equation:

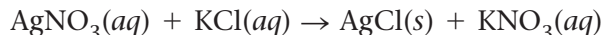


Figure 8.4 shows the precipitation of $\text{AgCl}(s)$ that occurs when this reaction takes place. In graphic form, the reaction is



Figure 8.4

Precipitation of silver chloride occurs when solutions of silver nitrate and potassium chloride are mixed. The K^+ and NO_3^- ions remain in solution.



Let's Review

How to Predict Precipitates When Solutions of Two Ionic Compounds Are Mixed

- Step 1** Write the reactants as they actually exist before any reaction occurs. Remember that when a salt dissolves, its ions separate.
- Step 2** Consider the various solids that could form. To do this, simply *exchange the anions* of the added salts.
- Step 3** Use the solubility rules (Table 8.1) to decide whether a solid forms and, if so, to predict the identity of the solid.

EXAMPLE 8.2

Using Solubility Rules to Predict the Products of Reactions

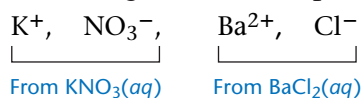
Using the solubility rules in Table 8.1, predict what will happen when the following solutions are mixed. Write the balanced equation for any reaction that occurs.

- a. $\text{KNO}_3(aq)$ and $\text{BaCl}_2(aq)$ c. $\text{KOH}(aq)$ and $\text{Fe}(\text{NO}_3)_3(aq)$
b. $\text{Na}_2\text{SO}_4(aq)$ and $\text{Pb}(\text{NO}_3)_2(aq)$

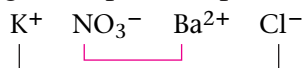
Solution

- a. $\text{KNO}_3(aq)$ and $\text{BaCl}_2(aq)$

Step 1 $\text{KNO}_3(aq)$ represents an aqueous solution obtained by dissolving solid KNO_3 in water to give the ions $\text{K}^+(aq)$ and $\text{NO}_3^-(aq)$. Likewise, $\text{BaCl}_2(aq)$ is a solution formed by dissolving solid BaCl_2 in water to produce $\text{Ba}^{2+}(aq)$ and $\text{Cl}^-(aq)$. When these two solutions are mixed, the following ions will be present:

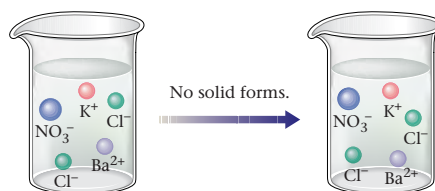


Step 2 To get the possible products, we exchange the anions.



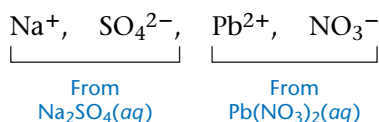
This yields the possibilities KCl and $\text{Ba}(\text{NO}_3)_2$. These are the solids that *might* form. Notice that two NO_3^- ions are needed to balance the 2+ charge on Ba^{2+} .

Step 3 The rules listed in Table 8.1 indicate that both KCl and $\text{Ba}(\text{NO}_3)_2$ are soluble in water. So no precipitate forms when $\text{KNO}_3(aq)$ and $\text{BaCl}_2(aq)$ are mixed. All of the ions remain dissolved in the solution. This means that no reaction takes place. That is, no chemical change occurs.

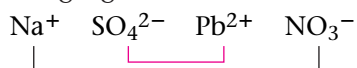


- b. $\text{Na}_2\text{SO}_4(aq)$ and $\text{Pb}(\text{NO}_3)_2(aq)$

Step 1 The following ions are present in the mixed solution before any reaction occurs:

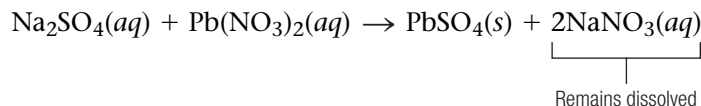


Step 2 Exchanging anions as follows:

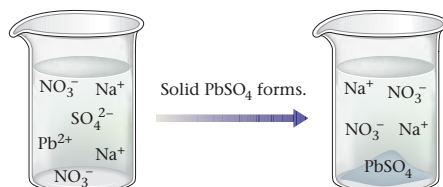


yields the *possible* solid products PbSO_4 and NaNO_3 .

Step 3 Using Table 8.1, we see that NaNO_3 is soluble in water (rules 1 and 2) but that PbSO_4 is only slightly soluble (rule 4). Thus, when these solutions are mixed, solid PbSO_4 forms. The balanced reaction is

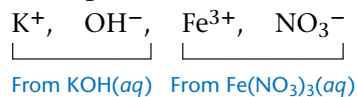


which can be represented as

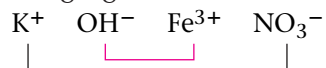


c. $\text{KOH}(aq)$ and $\text{Fe}(\text{NO}_3)_3(aq)$

Step 1 The ions present in the mixed solution before any reaction occurs are

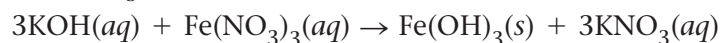


Step 2 Exchanging anions as follows:

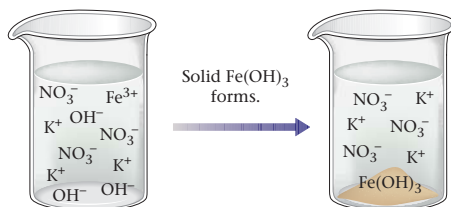


yields the possible solid products KNO_3 and $\text{Fe}(\text{OH})_3$.

Step 3 Rules 1 and 2 (Table 8.1) state that KNO_3 is soluble, whereas $\text{Fe}(\text{OH})_3$ is only slightly soluble (rule 5). Thus, when these solutions are mixed, solid $\text{Fe}(\text{OH})_3$ forms. The balanced equation for the reaction is



which can be represented as



Practice Problem • Exercise 8.2

Predict whether a solid will form when the following pairs of solutions are mixed. If so, identify the solid and write the balanced equation for the reaction.

- $\text{Ba}(\text{NO}_3)_2(aq)$ and $\text{NaCl}(aq)$
- $\text{Na}_2\text{S}(aq)$ and $\text{Cu}(\text{NO}_3)_2(aq)$
- $\text{NH}_4\text{Cl}(aq)$ and $\text{Pb}(\text{NO}_3)_2(aq)$

C. Describing Reactions in Aqueous Solutions

Molecular equation

A chemical equation showing the complete (undissociated) forms of all reactants and products

Complete ionic equation

A chemical equation for a reaction in solution representing all strong electrolytes as ions

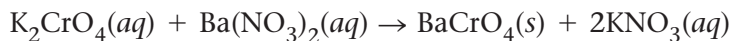
Spectator ion

An ion present in solution that does not participate in a reaction

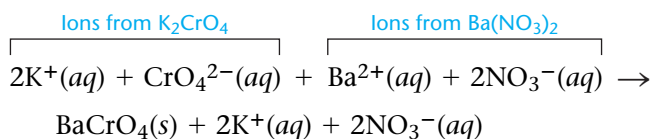
Net ionic equation

A chemical equation for a reaction in solution showing only those components that are directly involved in the reaction. Strong electrolytes are represented as ions.

Much important chemistry, including virtually all of the reactions that make life possible, occurs in aqueous solutions. We will now consider the types of equations used to represent reactions that occur in water. For example, as we saw earlier, when we mix aqueous potassium chromate with aqueous barium nitrate, a reaction occurs to form solid barium chromate and dissolved potassium nitrate. One way to represent this reaction is by the equation

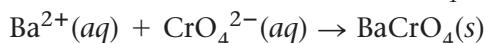


This is called the **molecular equation** for the reaction; it shows the complete formulas of all reactants and products. However, although this equation shows the reactants and products of the reaction, it does not give a very clear picture of what actually occurs in solution. As we have seen, aqueous solutions of potassium chromate, barium nitrate, and potassium nitrate contain the individual ions, not molecules as is implied by the molecular equation. Thus the **complete ionic equation**,



better represents the actual forms of the reactants and products in solution. *In a complete ionic equation, all substances that are strong electrolytes are represented as ions.* Notice that BaCrO_4 is not written as the separate ions, because it is present as a solid; it is not dissolved.

The complete ionic equation reveals that only some of the ions participate in the reaction. Notice that the K^+ and NO_3^- ions are present in solution both before and after the reaction. Ions such as these, which do not participate directly in a reaction in solution, are called **spectator ions**. The ions that participate in this reaction are the Ba^{2+} and CrO_4^{2-} ions, which combine to form solid BaCrO_4 :



This equation, called the **net ionic equation**, includes only those components that are directly involved in the reaction. Chemists usually write the net ionic equation for a reaction in solution, because it gives the actual forms of the reactants and products and includes only the species that undergo a change.

Information

A strong electrolyte is a substance that completely breaks apart into ions when dissolved in water. The resulting solution readily conducts an electric current.

Active Reading Questions

1. A solution of lead(II) nitrate is reacted with a solution of potassium iodide. Write a molecular equation, complete ionic equation, and net ionic equation for this reaction.
2. According to the solubility rules, which ions are generally spectator ions?

Let's Review

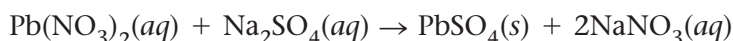
Types of Equations for Reactions in Aqueous Solutions

Three types of equations are used to describe reactions in solutions.

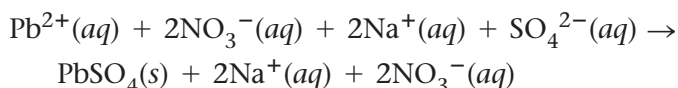
- The *molecular equation* shows the overall reaction but not necessarily the actual forms of the reactants and products in solution.
- The *complete ionic equation* represents all reactants and products that are strong electrolytes as ions. All reactants and products are included.
- The *net ionic equation* includes only those components that undergo a change. Spectator ions are not included.

A Closer Look

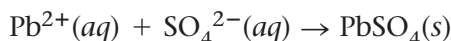
In Example 8.2 we considered the reaction between aqueous solutions of lead nitrate and sodium sulfate. The molecular equation for this reaction is



Because any ionic compound that is dissolved in water is present as the separated ions, we can write the complete ionic equation as follows:



The PbSO_4 is not written as separate ions because it is present as a solid. The ions that take part in the chemical change are the Pb^{2+} and the SO_4^{2-} ions, which combine to form solid PbSO_4 . Thus the net ionic equation is



The Na^{+} and NO_3^{-} ions do not undergo any chemical change; they are spectator ions.

Information

The net ionic equation includes only those components that undergo a change in the reaction.

Active Reading Question

What is the advantage of writing net ionic equations for chemical reactions?

EXAMPLE 8.3

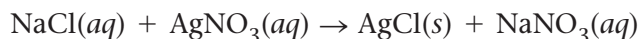
Writing Equations for Reactions

For each of the following reactions, write the molecular equation, the complete ionic equation, and the net ionic equation.

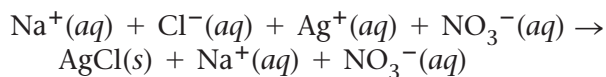
- Aqueous sodium chloride is added to aqueous silver nitrate to form solid silver chloride plus aqueous sodium nitrate.
- Aqueous potassium hydroxide is mixed with aqueous iron(III) nitrate to form solid iron(III) hydroxide and aqueous potassium nitrate.

Solution

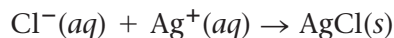
a. *Molecular equation:*



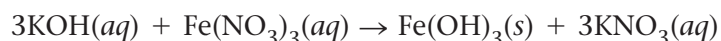
Complete ionic equation:



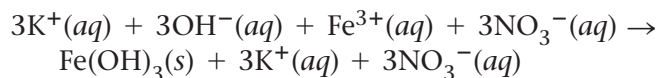
Net ionic equation:



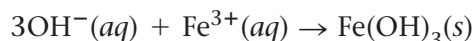
b. *Molecular equation:*



Complete ionic equation:



Net ionic equation:



Practice Problem • Exercise 8.3

For each of the following reactions, write the molecular equation, the complete ionic equation, and the net ionic equation.

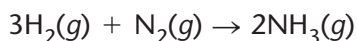
- Aqueous sodium sulfide is mixed with aqueous copper(II) nitrate to produce solid copper(II) sulfide and aqueous sodium nitrate.
- Aqueous ammonium chloride and aqueous lead(II) nitrate react to form solid lead(II) chloride and aqueous ammonium nitrate.

CELEBRITY CHEMICAL

Ammonia (NH_3)

Ammonia is a colorless gas with a strong odor that can be liquified at -34°C . Ammonia (dissolved in water) is found in many household cleaning products.

Ammonia is manufactured by combining the elements nitrogen and hydrogen:



The major use of ammonia is as a fertilizer to furnish nitrogen atoms to growing plants. Approximately 30 billion pounds of ammonia are produced every year for this purpose. For use as a

fertilizer, the ammonia is liquified (at high pressures) and stored in mobile tanks that can be pulled through the fields by a tractor. It is then injected into the ground to serve as an additional source of nitrogen for the crop.

A cross section showing how ammonia is injected into the soil to act as a fertilizer



HANDS-ON CHEMISTRY

• MINI-LAB •

Forecast: Precipitation**Materials**

- index cards (or flashcards from the Chapter 4 activities if available)

Procedure

1. Make two sets of flashcards. Put the name of an ion on one side and its symbol and charge on the other (use the cards from the Chapter 4 activities if you did them).

Set A: all of the alkali and alkaline earth metals, and iron(II), iron(III), lead(II), and silver.

Set B: all of the nonmetals from Groups 5A, 6A, and 7A, and the polyatomic ions listed in Table 4.4

2. Randomly pick one card from Set A and one card from Set B. Write the proper formula for the ionic compound made from these two ions, and name the compound.
3. Repeat step 2. Are both ionic compounds soluble? If so, go to step 4, if not, go back to step 2.
4. Use the two soluble ionic compounds as reactants. Write the names and formulas for the possible products of the reaction, along with the molecular equation and complete ionic equation. Is at least one of the products insoluble? If so, go to step 5. If not, go back to step 2.
5. Write the net ionic equation for the reaction. Go back to step 2 and continue this process until you have 5 net ionic equations.

**SECTION 8.1****REVIEW QUESTIONS**

1. What are the driving forces that indicate a chemical reaction is likely to occur?
2. Use the solubility rules in Table 8.1 or the information in Figure 8.3 to predict which of the following will be soluble in water:
 - a. potassium nitrate
 - b. zinc hydroxide
 - c. calcium carbonate
 - d. ammonium chloride
3. Consider two separate beakers, one containing an aqueous solution of hydrochloric acid, the second an aqueous solution of lead(II) nitrate.
 - a. Draw a picture of each solution showing the ions present.
 - b. Draw a picture after the solutions are mixed showing what is present.
 - c. Predict the products for any reaction that occurs.
 - d. Write a net ionic equation for the reaction.
4. How is a molecular equation different from a complete ionic equation?
5. What is a spectator ion and what happens to it in a net ionic equation?
6. Consider the following reaction: aqueous sodium sulfate is added to aqueous barium bromide to form solid barium sulfate and aqueous sodium bromide.
 - a. Write the molecular equation.
 - b. Write the complete ionic equation.
 - c. List the spectator ions.
 - d. Write the net ionic equation.



RESEARCH LINKS
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SECTION 8.2

Other Reactions in Aqueous Solutions

Key Terms

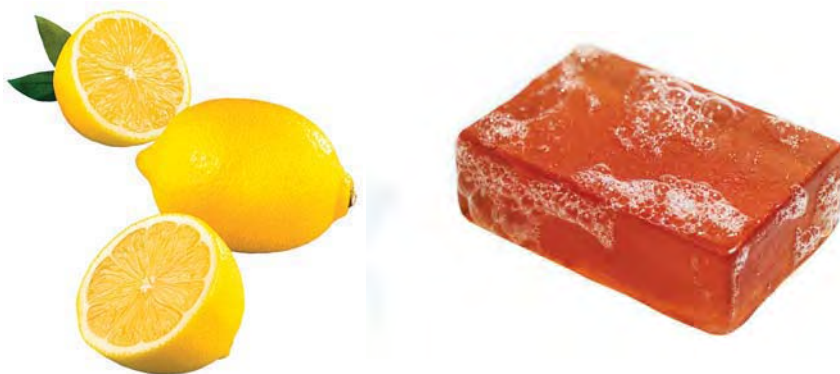
- Acid
- Strong acids
- Base
- Strong bases
- Salt
- Oxidation–reduction reaction

Objectives

- To learn about the reactions between strong acids and strong bases
- To learn about the reaction between a metal and a nonmetal
- To understand how electron transfer produces a chemical reaction

A. Reactions That Form Water: Acids and Bases

In this section we encounter two very important classes of compounds: acids and bases. Acids were first associated with the sour taste of citrus fruits. In fact, the word *acid* comes from the Latin word *acidus*, which means “sour.” Vinegar tastes sour because it is a dilute solution of acetic acid; citric acid is responsible for the sour taste of a lemon. Bases, sometimes called *alkalis*, are characterized by their bitter taste and slippery feel, like wet soap. Most commercial preparations for unclogging drains are highly basic.

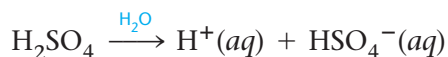
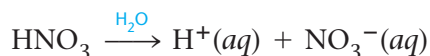
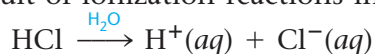


Information

Don't taste chemicals!

Acids have been known for hundreds of years. For example, the *mineral acids* sulfuric acid, H_2SO_4 , and nitric acid, HNO_3 , so named because they were originally obtained by the treatment of minerals, were discovered around 1300. However, it was not until the late 1800s that the essential nature of acids was discovered by Svante Arrhenius, then a Swedish graduate student in physics.

Arrhenius Acids and Bases Arrhenius, who was trying to discover why only certain solutions could conduct an electric current, found that conductivity arose from the presence of ions. In his studies of solutions, Arrhenius observed that when the substances HCl , HNO_3 , and H_2SO_4 were dissolved in water, they behaved as strong electrolytes. He suggested that this was the result of ionization reactions in water.

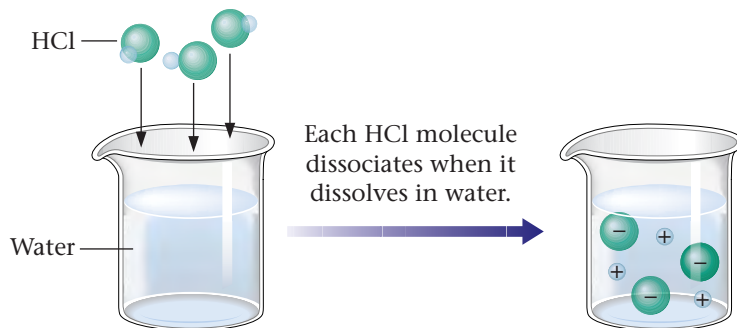


Arrhenius proposed that an **acid** is a substance that produces H^+ ions (protons) when it is dissolved in water.

Information

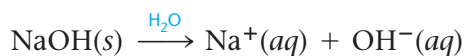
The Arrhenius definition of an acid: a substance that produces H^+ ions in aqueous solution.

Studies show that when HCl, HNO₃, and H₂SO₄ are placed in water, *virtually every molecule* dissociates to give ions. This means that when 100 molecules of HCl are dissolved in water, 100 H⁺ ions and 100 Cl⁻ ions are produced. Virtually no HCl molecules exist in aqueous solution.



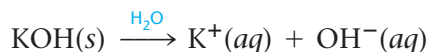
Because these substances are strong electrolytes that produce H⁺ ions, they are called **strong acids**.

Arrhenius also found that *aqueous solutions that exhibit basic behavior* always contain hydroxide ions. He defined a **base** as a *substance that produces hydroxide ions (OH⁻) in water*. The base most commonly used in the chemical laboratory is sodium hydroxide, NaOH, which contains Na⁺ and OH⁻ ions and is very soluble in water. Sodium hydroxide, like all ionic substances, produces separated cations and anions when it is dissolved in water.



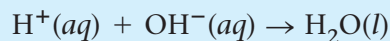
Although dissolved sodium hydroxide is usually represented as NaOH(aq), you should remember that the solution really contains separated Na⁺ and OH⁻ ions. In fact, for every 100 units of NaOH dissolved in water, 100 Na⁺ ions and 100 OH⁻ ions are produced.

Potassium hydroxide (KOH) has properties markedly similar to those of sodium hydroxide. It is very soluble in water and produces separated ions.



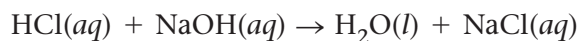
Because these hydroxide compounds are strong electrolytes that contain OH⁻ ions, they are called **strong bases**.

When strong acids and strong bases (hydroxides) are mixed, the fundamental chemical change that always occurs is that H⁺ ions react with OH⁻ ions to form water.



Water is a very stable compound, as evidenced by the abundance of it on the earth's surface. Therefore, when substances that can form water are mixed, there is a strong tendency for the reaction to occur. In particular, the hydroxide ion OH⁻ has a high affinity for the H⁺ ion to produce water.

The tendency to form water is the second of the driving forces for reactions that we mentioned before. Any compound that produces OH⁻ ions in water reacts vigorously to form H₂O with any compound that can furnish H⁺ ions. For example, the reaction between hydrochloric acid and aqueous sodium hydroxide is represented by the following molecular equation:



Strong acid

An acid that completely dissociates to produce H⁺ ions in solution.

Strong base

A base that completely dissociates to produce OH⁻ ions in solution.

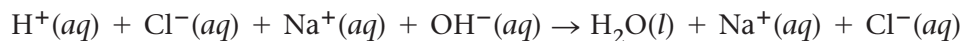
DID YOU KNOW

The Nobel Prize in Chemistry was awarded to Arrhenius in 1903 for his studies of solution conductivity.

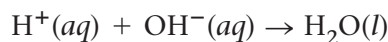
Information

Hydrochloric acid is an aqueous solution that contains dissolved hydrogen chloride. It is a strong electrolyte.

Because HCl, NaOH, and NaCl exist as completely separated ions in water, the complete ionic equation for this reaction is



Notice that the Cl^- and Na^+ are spectator ions (they undergo no changes), so the net ionic equation is



Thus the only chemical change that occurs when these solutions are mixed is that water is formed from H^+ and OH^- ions.

Active Reading Question

What is always a product of an acid–base reaction?

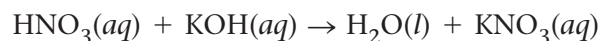
EXAMPLE 8.4

Writing Equations for Acid–Base Reactions

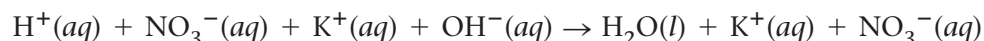
Nitric acid is a strong acid. Write the molecular, complete ionic, and net ionic equations for the reaction of aqueous nitric acid and aqueous potassium hydroxide.

Solution

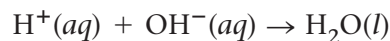
Molecular equation:



Complete ionic equation:



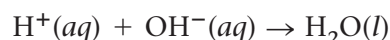
Net ionic equation:



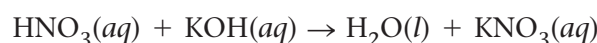
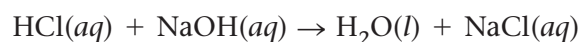
Note that K^+ and NO_3^- are spectator ions and that the formation of water is the driving force for this reaction.

There are two important things to note as we examine the reaction of hydrochloric acid with aqueous sodium hydroxide and the reaction of nitric acid with aqueous potassium hydroxide:

1. The net ionic equation is the same in both cases; water is formed.



2. Besides water, which is *always a product* of the reaction of an acid with OH^- , the second product is an ionic compound, which might precipitate or remain dissolved, depending on its solubility.



Dissolved ionic compounds

Salt

An ionic compound

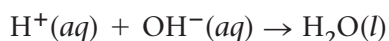
This ionic compound is called a **salt**. In the first case the salt is sodium chloride, and in the second case the salt is potassium nitrate. We can obtain these soluble salts in solid form (both are white solids) by evaporating the water.

Let's Review

Strong Acids and Strong Bases

The following points about strong acids and strong bases are particularly important.

- The common strong acids are aqueous solutions of HCl, HNO₃, and H₂SO₄.
- A strong acid is a substance that completely dissociates (ionizes) in water. (Each molecule breaks up into an H⁺ ion plus an anion.)
- A strong base is a metal hydroxide compound that is very soluble in water. The most common strong bases are NaOH and KOH, which completely break up into separated ions (Na⁺ and OH⁻ or K⁺ and OH⁻) when they are dissolved in water.
- The net ionic equation for the reaction of a strong acid and a strong base is always the same: it shows the production of water.



- In the reaction of a strong acid and a strong base, one product is always water and the other is always an ionic compound called a salt, which remains dissolved in the water. This salt can be obtained as a solid by evaporating the water.
- The reaction of H⁺ and OH⁻ is often called an acid–base reaction, where H⁺ is the acidic ion and OH⁻ is the basic ion.

Information

Both strong acids and strong bases are strong electrolytes.

CELEBRITY CHEMICAL

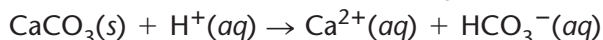
Calcium Carbonate (CaCO₃)

Calcium carbonate, which contains the Ca²⁺ and CO₃²⁻ ions, is very common in nature, occurring in eggshells, limestone, marble, seashells, and coral. The spectacular formations seen in limestone caves are also composed of calcium carbonate.

Limestone caves form when underground limestone deposits come in contact with water made acidic by dissolved carbon dioxide. When rainwater absorbs CO₂ from the atmosphere, the following reaction occurs:

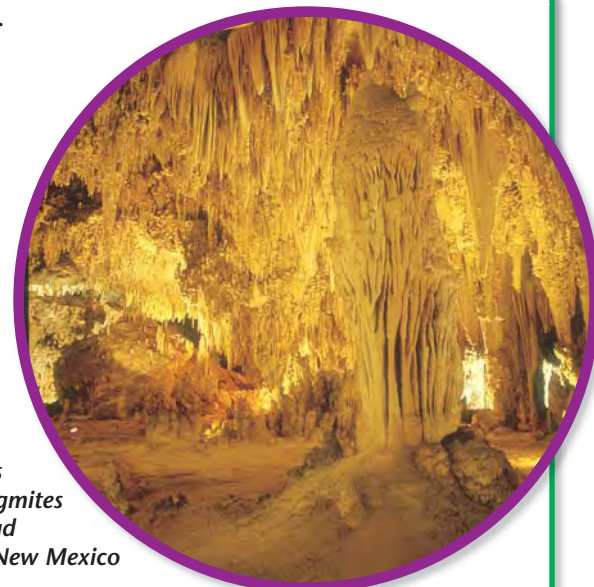


This reaction leads to the presence of H⁺ in the groundwater. The acidic groundwater then causes the limestone (which is made of CaCO₃) to dissolve:



Underground caverns then form. In the process of dissolving the limestone and creating the cave, the water containing the dissolved CaCO₃ drips from the ceiling of the cave. As the water forms drops, it tends to lose some of the dissolved CO₂, which lowers the amount of H⁺ present (by the reversal

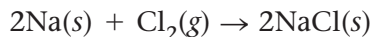
of the first reaction). This in turn leads to the reversal of the second reaction, which then reforms the solid CaCO₃. This process causes stalactites to “grow” from the ceiling of the cave. Water that drips to the floor before losing its dissolved CO₂ forms stalagmites that build up from the floor of the cave.



Stalactites and stalagmites in Carlsbad Caverns, New Mexico

B. Reactions of Metals with Nonmetals (Oxidation–Reduction)

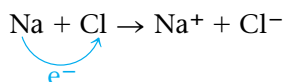
In Chapter 3 we spent considerable time discussing ionic compounds—compounds formed in the reaction of a metal and a nonmetal. A typical example is sodium chloride, formed by the reaction of sodium metal and chlorine gas:



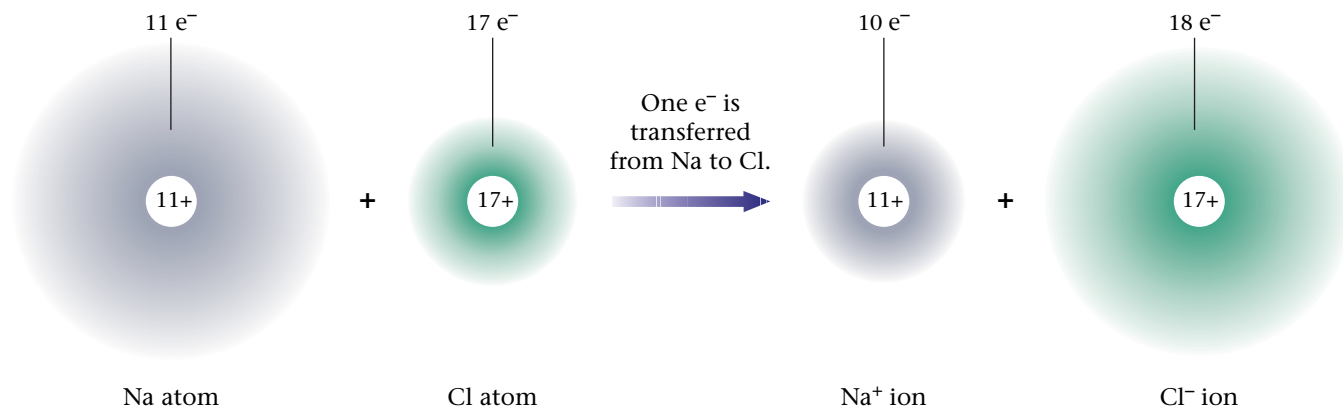
Let's examine what happens in this reaction.

- Sodium metal is composed of sodium atoms, each of which has a net charge of zero. (The positive charges of the eleven protons in its nucleus are exactly balanced by the negative charges on the eleven electrons.)
- Similarly, the chlorine molecule consists of two uncharged chlorine atoms (each has seventeen protons and seventeen electrons).
- However, in the product (sodium chloride), the sodium is present as Na^+ and the chlorine as Cl^- .

By what process do the neutral atoms become ions? The answer is that one electron is transferred from each sodium atom to each chlorine atom.



After the electron transfer, each sodium atom has ten electrons and eleven protons (a net charge of 1+), and each chlorine atom has eighteen electrons and seventeen protons (a net charge of 1-).



Thus the reaction of a metal with a nonmetal to form an ionic compound involves the transfer of one or more electrons from the metal (which forms a cation) to the nonmetal (which forms an anion). This tendency to transfer electrons from metals to nonmetals is the third driving force for reactions that we listed in Section 8.1. A reaction that *involves a transfer of electrons* is called an **oxidation–reduction reaction**.

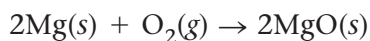
Oxidation–reduction reaction

A chemical reaction involving the transfer of electrons

Active Reading Question

Why can metal–nonmetal reactions always be assumed to be oxidation–reduction reactions?

There are many examples of oxidation–reduction reactions in which a metal reacts with a nonmetal to form an ionic compound. Consider the reaction of magnesium metal with oxygen,

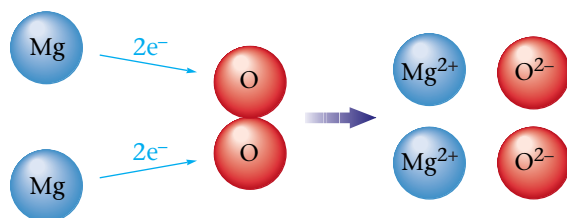


Note that the reactants contain uncharged atoms, but the product contains ions:

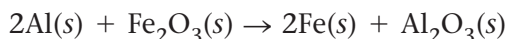


Contains Mg^{2+} , O^{2-}

Therefore, in this reaction, each magnesium atom loses two electrons ($\text{Mg} \rightarrow \text{Mg}^{2+} + 2e^-$) and each oxygen atom gains two electrons ($\text{O} + 2e^- \rightarrow \text{O}^{2-}$). We might represent this reaction as follows:



Another example is



which is a reaction (called the thermite reaction) that produces so much energy (heat) that the iron is initially formed as a liquid (see **Figure 8.5**). In this case the aluminum is originally present as the elemental metal (which contains uncharged Al atoms) and ends up in Al_2O_3 , where it is present as Al^{3+} cations (the 2Al^{3+} ions just balance the charge of the 3O^{2-} ions). Therefore, in the reaction each aluminum atom loses three electrons.

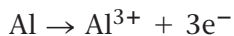


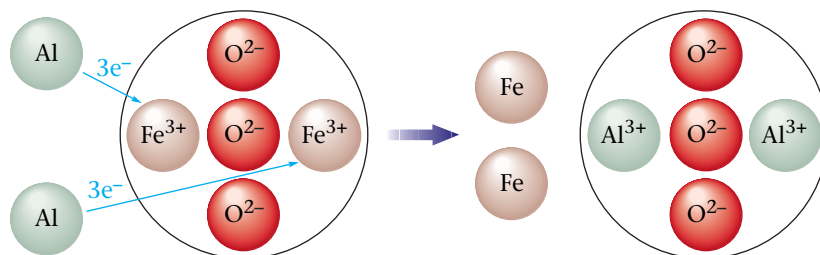
Figure 8.5

The thermite reaction gives off so much heat that the iron formed is molten.

The opposite process occurs with the iron, which is initially present as Fe^{3+} ions in Fe_2O_3 and ends up as uncharged atoms in the elemental iron. Thus each iron cation gains three electrons to form an uncharged atom:



We can represent this reaction in schematic form as follows:



EXAMPLE 8.5

Identifying Electron Transfer in Oxidation–Reduction Reactions



Figure 8.6

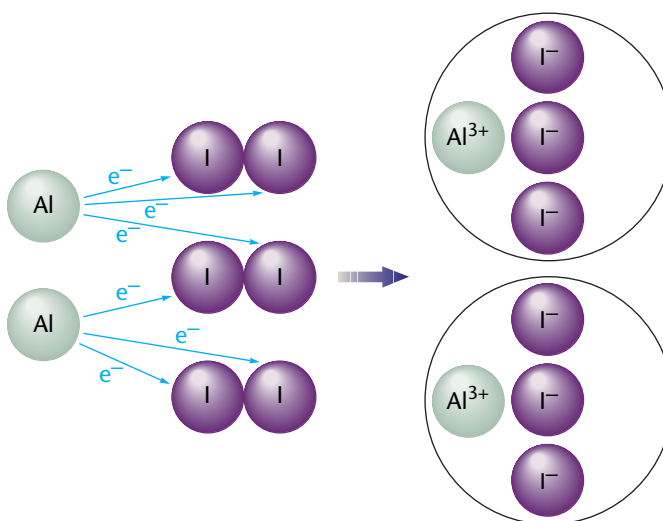
When powdered aluminum and iodine (shown in the foreground) are mixed (and a little water added), they react vigorously.

For each of the following reactions, show how electrons are gained and lost.

- $2\text{Al}(s) + 3\text{I}_2(s) \rightarrow 2\text{AlI}_3(s)$ (This reaction is shown in **Figure 8.6**. Note the purple “smoke,” which is excess I_2 being driven off by the heat.)
- $2\text{Cs}(s) + \text{F}_2(g) \rightarrow 2\text{CsF}(s)$

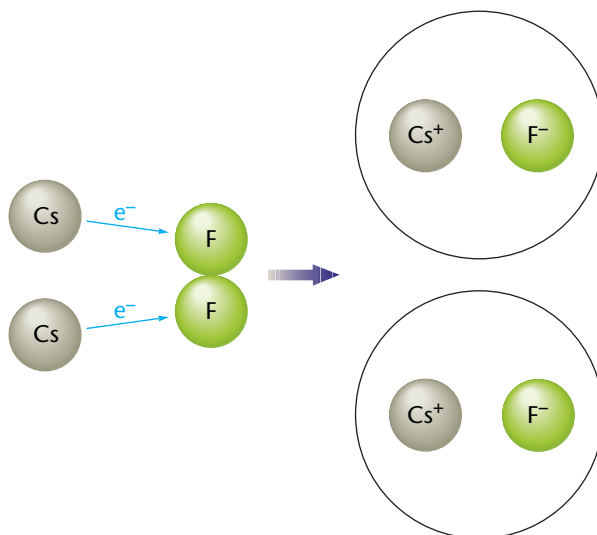
Solution

- In AlI_3 the ions are Al^{3+} and I^- (aluminum always forms Al^{3+} , and iodine always forms I^-). In $\text{Al}(s)$ the aluminum is present as uncharged atoms. Thus aluminum goes from Al to Al^{3+} by losing three electrons ($\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-$). In I_2 each iodine atom is uncharged. Thus each iodine atom goes from I to I^- by gaining one electron ($\text{I} + \text{e}^- \rightarrow \text{I}^-$). A schematic for this reaction is



- In CsF the ions present are Cs^+ and F^- . Cesium metal, $\text{Cs}(s)$, contains uncharged cesium atoms, and fluorine gas, $\text{F}_2(g)$, contains uncharged fluorine atoms. Thus in the reaction each cesium atom loses one

electron ($\text{Cs} \rightarrow \text{Cs}^+ + \text{e}^-$) and each fluorine atom gains one electron ($\text{F} + \text{e}^- \rightarrow \text{F}^-$). The schematic for this reaction is

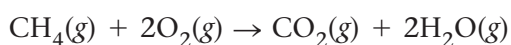


Practice Problem • Exercise 8.5

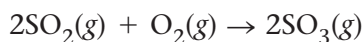
For each reaction, show how electrons are gained and lost.

- $2\text{Na}(s) + \text{Br}_2(l) \rightarrow 2\text{NaBr}(s)$
- $2\text{Ca}(s) + \text{O}_2(g) \rightarrow 2\text{CaO}(s)$

So far we have emphasized electron transfer (oxidation–reduction) reactions that involve a metal and a nonmetal. Electron transfer reactions can also take place between two nonmetals. We will not discuss these reactions in detail here. All we will say at this point is that one sure sign of an oxidation–reduction reaction between nonmetals is the presence of oxygen, $\text{O}_2(g)$, as a reactant or product. In fact, oxidation got its name from oxygen. Thus the reactions



and



are electron transfer reactions, even though it is not obvious at this point.

Let's Review

Characteristics of Oxidation–Reduction Reactions

- When a metal reacts with a nonmetal, an ionic compound is formed. The ions are formed when the metal transfers one or more electrons to the nonmetal, the metal atom becoming a cation and the nonmetal atom becoming an anion. *Therefore, a metal–nonmetal reaction can always be assumed to be an oxidation–reduction reaction, which involves electron transfer.*
- Two nonmetals can also undergo an oxidation–reduction reaction. At this point we can recognize these cases only by looking for O_2 as a reactant or product. When two nonmetals react, the compound formed is not ionic.

**What's in Your Water?****Materials**

- various water samples
- dropper bottle of silver nitrate solution

Procedure

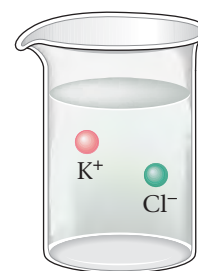
1. Obtain samples of water from your home faucet, bottled water, distilled water from school, water from the school drinking fountain, and a swimming pool (if available).
2. Get a dropper bottle of silver nitrate solution from your teacher. (**Caution:** Silver nitrate solution will stain your skin!). Add a few drops of the silver nitrate solution to your samples of water.

Results/Analysis

1. What do you notice when you add the silver nitrate solution to your water samples? Compare your results with your classmates.
2. Write the net ionic equation for any reaction.
3. What does this test tell you about the water samples?

**SECTION 8.2****REVIEW QUESTIONS**

1. What are the products in a reaction between an acid and a base? How can you tell that a reaction has occurred?
2. What is meant by the term *strong acid*? What is meant by the term *strong base*?
3. Draw a picture (showing the ions) for a strong acid solution of HNO_3 and a strong base solution of KOH in separate beakers.
4. What is an oxidation–reduction reaction? How can you identify this type of reaction?
5. Write the formulas for the strong acid and strong base that must have reacted to form the solution shown to the right. Write the chemical equation for this reaction.
6. Show how electrons are lost and gained in the reaction.
$$3\text{Mg}(s) + \text{N}_2(g) \rightarrow \text{Mg}_3\text{N}_2(s)$$

 $\text{KCl}(aq)$ **RESEARCH LINKS**
CLASSZONE.COM

Classifying Reactions

Objective

- To learn various classification schemes for reactions

A. Ways to Classify Reactions

So far in our study of chemistry we have seen many, many chemical reactions—and this is just Chapter 8. In the world around us and in our bodies, literally millions of chemical reactions are taking place. Obviously, we need a system for putting reactions into meaningful classes that will make them easier to remember and easier to understand.

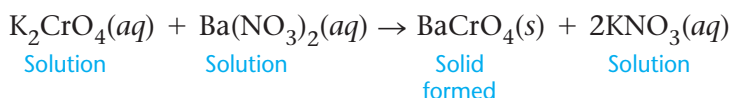
In Chapter 8 we have so far considered the following “driving forces” for chemical reactions:

- Formation of a solid
- Formation of water
- Transfer of electrons

Active Reading Question

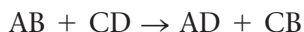
Write an example chemical equation for each of the driving forces listed.

Formation of a Solid In the following reaction, solid BaCrO_4 (a precipitate) is formed.



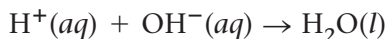
Because the *formation of a solid when two solutions are mixed* is called *precipitation*, we call this a *precipitation reaction*.

Notice in this reaction that two anions (NO_3^- and CrO_4^{2-}) are simply exchanged. Note that CrO_4^{2-} was originally associated with K^+ in K_2CrO_4 and that NO_3^- was associated with Ba^{2+} in $\text{Ba}(\text{NO}_3)_2$. In the products these associations are reversed. Because of this double exchange, we sometimes call this reaction a double-exchange reaction or **double-displacement reaction**. We might represent such a reaction as

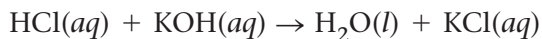


So we can classify a reaction such as this one as a precipitation reaction or as a double-displacement reaction. Either name is correct, but precipitation is more commonly used by chemists.

Formation of Water In this chapter we have also considered reactions in which water is formed when a strong acid is mixed with a strong base. All of these reactions had the same net ionic equation:



The H^+ ion comes from a strong acid, such as $\text{HCl}(aq)$ or $\text{HNO}_3(aq)$, and the origin of the OH^- ion is a strong base, such as $\text{NaOH}(aq)$ or $\text{KOH}(aq)$. An example is



We classify these reactions as **acid–base reactions**. You can recognize an acid–base reaction because it *involves an H^+ ion that ends up in the product water*.

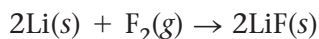
Key Terms

- Double-displacement reaction
- Acid–base reactions
- Single-replacement reaction
- Combustion reactions
- Synthesis (combination) reaction
- Decomposition reactions

Information

A precipitation reaction is also called a double-displacement reaction.

Transfer of Electrons The third driving force is electron transfer. We see evidence of this driving force particularly in the “desire” of a metal to donate electrons to nonmetals. An example is

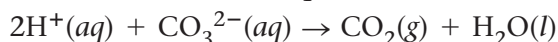


where each lithium atom loses one electron to form Li^+ , and each fluorine atom gains one electron to form the F^- ion. The process of electron transfer is also called oxidation–reduction. Thus we classify the above reaction as an *oxidation–reduction reaction*.

Formation of a Gas An additional driving force for chemical reactions that we have not yet discussed is *formation of a gas*. A reaction in aqueous solution that forms a gas (which escapes as bubbles) is pulled toward the products by this event. An example is the reaction

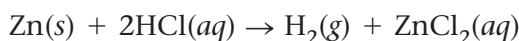


for which the net ionic equation is

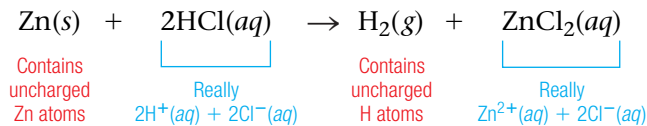


Note that this reaction forms carbon dioxide gas as well as water, so it illustrates two of the driving forces that we have considered. Because this reaction involves H^+ that ends up in the product water, we classify it as an acid–base reaction.

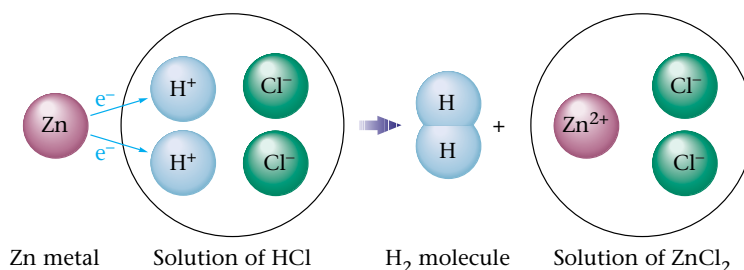
Consider another reaction that forms a gas:



How might we classify this reaction? A careful look at the reactants and products shows the following:

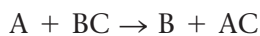


Note that in the reactant zinc metal, Zn exists as uncharged atoms, whereas in the product it exists as Zn^{2+} . Thus each Zn atom loses two electrons. Where have these electrons gone? They have been transferred to two H^+ ions to form H_2 . The schematic for this reaction is



This is an electron transfer process, so the reaction can be classified as an oxidation–reduction reaction.

Another way this reaction is sometimes classified is based on the fact that a *single* type of anion (Cl^-) has been exchanged between H^+ and Zn^{2+} . That is, Cl^- is originally associated with H^+ in HCl and ends up associated with Zn^{2+} in the product ZnCl_2 . We can call this a **single-replacement reaction** in contrast to double-displacement reactions, in which two types of anions are exchanged. We can represent a single replacement as



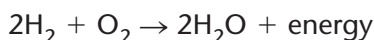
CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

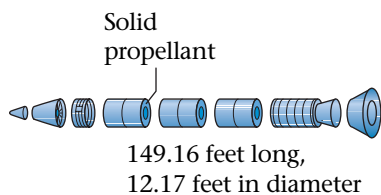
Launching the Space Shuttle

Launching into space a vehicle such as the space shuttle that weighs millions of pounds requires unimaginable quantities of energy—all furnished by oxidation–reduction reactions.

Notice that three cylindrical objects are attached to the shuttle orbiter. In the center is a tank about 28 feet in diameter and 154 feet long that contains liquid oxygen and liquid hydrogen (in separate compartments). These fuels are fed to the orbiter's main engines where they react to form water and release a huge quantity of energy.



Note that we can recognize this reaction as an oxidation–reduction reaction because O_2 is a reactant.



Two solid-fuel rockets about 12 feet in diameter and 150 feet long are also attached to the orbiter. Each rocket contains 1.1 million pounds of fuel: ammonium perchlorate (NH_4ClO_4) and powdered aluminum mixed with a binder (“glue”). Because the rockets are so large, they are built in segments and assembled at the launch site as shown above. Each segment is filled with the syrupy propellant, which then solidifies to a consistency much like that of a hard rubber eraser.

The oxidation–reduction reaction between the ammonium perchlorate and the aluminum is



It produces temperatures of about 5700 °F (3150 °C) and 3.3 million pounds of thrust in each rocket.

Oxidation–reduction reactions furnish the energy to launch the space shuttle.

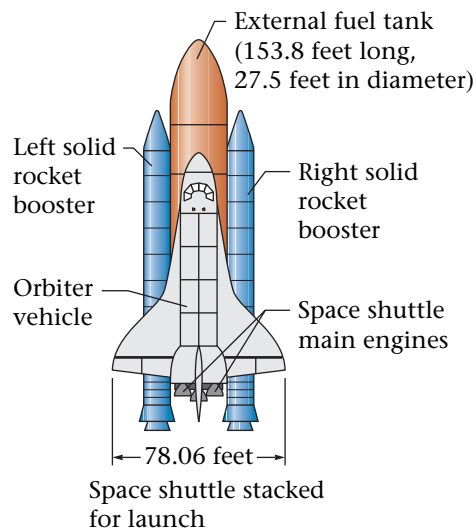


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Mixing the rocket propellant

B. Other Ways to Classify Reactions

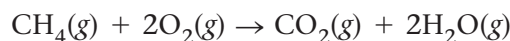
So far in this chapter we have classified chemical reactions in several ways. The most commonly used of these classifications are:

- Precipitation reactions
- Acid–base reactions
- Oxidation–reduction reactions

However, there are still other ways to classify reactions that you may encounter in your future studies of chemistry. We will consider several of these in this section.

Combustion Reactions

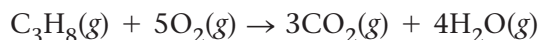
Many chemical reactions that involve oxygen produce energy (heat) so rapidly that a flame results. Such reactions are called **combustion reactions**. We have considered some of these reactions previously. For example, the methane in natural gas reacts with oxygen according to the following balanced equation:



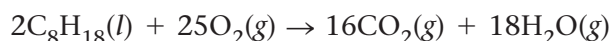
This reaction produces the flame of the common laboratory burner and is also used to heat most homes in the United States. Recall that we originally classified this reaction as an oxidation–reduction reaction. Thus we can say that the reaction of methane with oxygen is both an oxidation–reduction reaction and a combustion reaction. Combustion reactions, in fact, are a special class of oxidation–reduction reactions (see **Figure 8.7**).

There are many combustion reactions, most of which are used to provide heat or electricity for homes or businesses or energy for transportation. Some examples are:

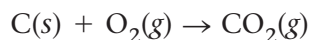
- Combustion of propane (used to heat some rural homes)



- Combustion of gasoline* (used to power cars and trucks)



- Combustion of coal* (used to generate electricity)

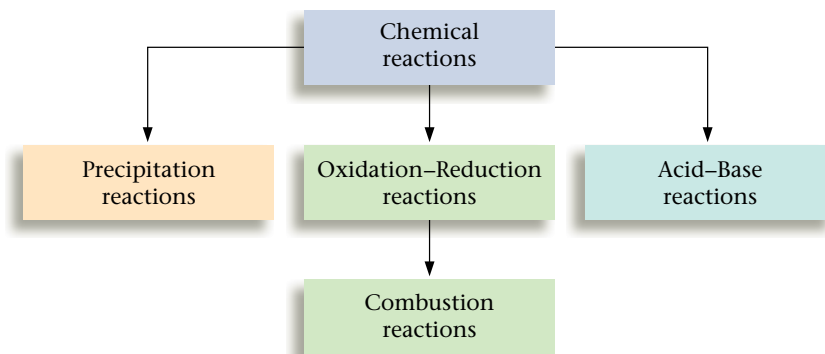


Combustion reaction

A chemical reaction involving oxygen as one of the reactants that produces enough heat so that a flame results

Figure 8.7

Classes of reactions—combustion reactions are a special type of oxidation–reduction reaction.

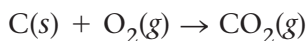
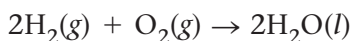


*This substance is really a complex mixture of compounds, but the reaction shown is representative of what takes place.

Synthesis (Combination) Reactions

One of the most important activities in chemistry is the synthesis of new compounds. Each of our lives has been greatly affected by synthetic compounds such as plastic, polyester, and aspirin. When a given compound is formed from simpler materials, we call this a **synthesis (combination) reaction**.

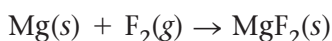
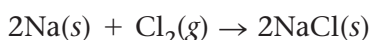
Examples of synthesis reactions starting with elements:



Other ways to classify these reactions:

- Oxidation–reduction \rightarrow involves oxygen
- Combustion \rightarrow produces flames

Examples of synthesis reactions that do not involve oxygen:



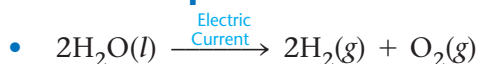
Another way to classify these reactions:

- Oxidation–reduction \rightarrow uncharged atoms become ions in compounds

We have seen that synthesis reactions in which the reactants are elements are oxidation–reduction reactions as well. In fact, we can think of these synthesis reactions as another subclass of the oxidation–reduction class of reactions.

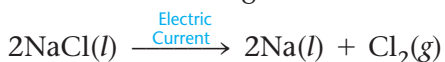
Decomposition Reactions

In many cases a compound can be broken down into simpler compounds or all the way to the component elements. This is usually accomplished by heating or by the application of an electric current. Such reactions are called **decomposition reactions**. Examples of decomposition reactions are



Because O_2 is involved in the first reaction, we recognize it as an oxidation–reduction reaction. In the second reaction, HgO , which contains Hg^{2+} and O^{2-} ions, is decomposed to the elements, which contain uncharged atoms. In this process each Hg^{2+} gains two electrons and each O^{2-} loses two electrons, so this is both a decomposition reaction and an oxidation–reduction reaction.

A decomposition reaction, in which a compound is broken down into its elements, is just the opposite of the synthesis (combination) reaction, in which elements combine to form the compound. For example, we have just discussed the synthesis of sodium chloride from its elements. Sodium chloride can be decomposed into its elements by melting it and passing an electric current through it:



Synthesis reaction (combination reaction)

A chemical reaction in which a compound is formed from simpler materials



Formation of the colorful plastics used in these zippers is an example of a synthesis reaction.

Decomposition reaction

A chemical reaction in which a compound is broken down into simpler compounds, or to the component elements

EXAMPLE 8.6

Classifying Reactions

Classify each of the following reactions in as many ways as possible.

- $2\text{K}(s) + \text{Cl}_2(g) \rightarrow 2\text{KCl}(s)$
- $\text{Fe}_2\text{O}_3(s) + 2\text{Al}(s) \rightarrow \text{Al}_2\text{O}_3(s) + 2\text{Fe}(s)$
- $2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s)$
- $\text{HNO}_3(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaNO}_3(aq)$
- $\text{KBr}(aq) + \text{AgNO}_3(aq) \rightarrow \text{AgBr}(s) + \text{KNO}_3(aq)$
- $\text{PbO}_2(s) \rightarrow \text{Pb}(s) + \text{O}_2(g)$

Solution

- This is both a synthesis reaction (elements combine to form a compound) and an oxidation–reduction reaction (uncharged potassium and chlorine atoms are changed to K^+ and Cl^- ions in KCl).
- This is an oxidation–reduction reaction. Iron is present in $\text{Fe}_2\text{O}_3(s)$ as Fe^{3+} ions and in elemental iron, $\text{Fe}(s)$, as uncharged atoms. So each Fe^{3+} must gain three electrons to form Fe. The reverse happens to aluminum, which is present initially as uncharged aluminum atoms, each of which loses three electrons to give Al^{3+} ions in Al_2O_3 . Note that this reaction might also be called a single-replacement reaction because O is switched from Fe to Al.
- This is both a synthesis reaction (elements combine to form a compound) and an oxidation–reduction reaction (each magnesium atom loses two electrons to give Mg^{2+} ions in MgO, and each oxygen atom gains two electrons to give O^{2-} in MgO).
- This is an acid–base reaction. It might also be called a double-displacement reaction because NO_3^- and OH^- “switch partners.”
- This is a precipitation reaction that might also be called a double-displacement reaction in which the anions Br^- and NO_3^- are exchanged.
- This is a decomposition reaction (a compound breaks down into elements). It also is an oxidation–reduction reaction, because the ions in PbO_2 (Pb^{4+} and O^{2-}) are changed to uncharged atoms in the elements $\text{Pb}(s)$ and $\text{O}_2(g)$. That is, electrons are transferred from O^{2-} to Pb^{4+} in the reaction.

Practice Problems • Exercise 8.6

Classify each of the following reactions in as many ways as possible.

- $4\text{NH}_3(g) + 5\text{O}_2(g) \rightarrow 4\text{NO}(g) + 6\text{H}_2\text{O}(g)$
- $\text{S}_8(s) + 8\text{O}_2(g) \rightarrow 8\text{SO}_2(g)$
- $2\text{Al}(s) + 3\text{Cl}_2(g) \rightarrow 2\text{AlCl}_3(s)$
- $2\text{AlN}(s) \rightarrow 2\text{Al}(s) + \text{N}_2(g)$
- $\text{BaCl}_2(aq) + \text{Na}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + 2\text{NaCl}(aq)$
- $2\text{Cs}(s) + \text{Br}_2(l) \rightarrow 2\text{CsBr}(s)$
- $\text{KOH}(aq) + \text{HCl}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{KCl}(aq)$
- $2\text{C}_2\text{H}_2(g) + 5\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 2\text{H}_2\text{O}(l)$

There are other schemes for classifying reactions that we have not considered. It should be apparent that many important reactions can be classified as oxidation–reduction reactions. As shown in **Figure 8.8**, various types of reactions can be viewed as subclasses of the overall oxidation–reduction category.

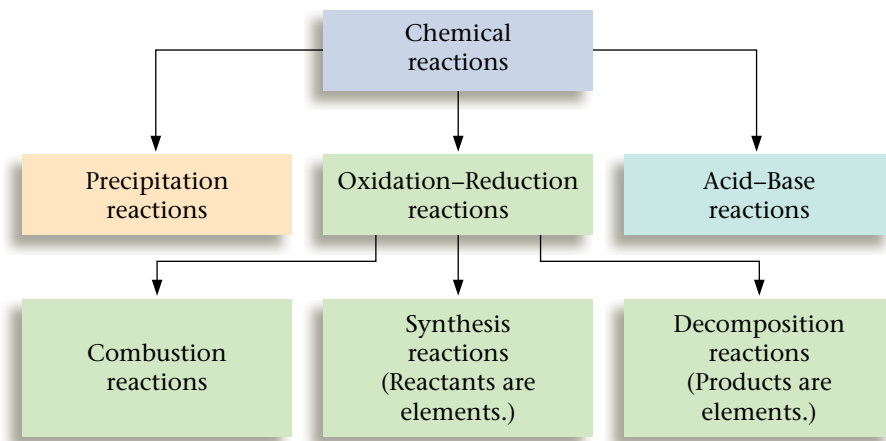


Figure 8.8
Summary of classes of reactions

CRITICAL THINKING ?

What if no ionic solids were soluble in water? Could reactions occur in aqueous solutions? Explain.



SECTION 8.3 REVIEW QUESTIONS

- What are the four driving forces for a chemical reaction?
- Why are precipitation reactions also called double displacement reactions?
- What are three other ways to classify oxidation–reduction reactions?
- Classify the following unbalanced equations as: precipitation, acid–base, or oxidation–reduction
 - $\text{Fe}(s) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Fe}_3(\text{SO}_4)_2(aq) + \text{H}_2(g)$
 - $\text{HClO}_4(aq) + \text{RbOH}(aq) \rightarrow \text{RbClO}_4(aq) + \text{H}_2\text{O}(l)$
 - $\text{K}_2\text{SO}_4(aq) + \text{CaCl}_2(aq) \rightarrow \text{KCl}(aq) + \text{CaSO}_4(s)$
 - $\text{Ni}(s) + \text{Cl}_2(g) \rightarrow \text{NiCl}_2(s)$
- One of the reactants in a combustion reaction must always be _____.
- Classify the reaction of hydrogen gas with oxygen gas to form water in three different ways.
- Balance the following equations and classify each in as many ways as possible.
 - $\text{Al}(s) + \text{O}_2(g) \rightarrow \text{Al}_2\text{O}_3(s)$
 - $\text{Ba}(\text{OH})_2(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + \text{H}_2\text{O}(l)$
 - $\text{H}_2\text{O}_2(aq) \rightarrow \text{H}_2\text{O}(l) + \text{O}_2(g)$



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Key Terms

8.1 Precipitation
Precipitate
Precipitation reaction
Strong electrolyte
Soluble solid
Insoluble solid (slightly soluble solid)
Molecular equation
Complete ionic equation
Spectator ions
Net ionic equation

8.2 Acid
Strong acids
Base
Strong bases
Salt
Oxidation–reduction reaction

8.3 Double-displacement reaction
Acid–base reactions
Single-replacement reaction
Combustion reactions
Synthesis (combination) reaction
Decomposition reactions

Key Ideas

8.1 Understanding Reactions in Aqueous Solution

- Four driving forces favor chemical change.
 - Formation of a solid
 - Formation of water
 - Transfer of electrons
 - Formation of a gas
- A reaction in which a solid forms is called a precipitation reaction.
- Solubility rules help predict what solid (if any) will form when solutions are mixed.
- Three types of equations are used to describe reactions in solution.
 - Molecular (formula) equation which shows the complete formulas of all reactants and products
 - Complete ionic equation in which all strong electrolytes are shown as ions
 - Net ionic equation which includes only those components of the solution that undergo a change
 - ◆ Spectator ions (those that remain unchanged) are not shown in the net ionic equation.

8.2 Other Reactions in Aqueous Solutions

- A strong acid is one in which virtually every molecule dissociates (ionizes) in water to an H^+ ion and an anion.
- A strong base is a metal hydroxide that is completely soluble in water, giving separate OH^- ions and cations.
- The products of the reaction of a strong acid and a strong base are water and a salt.
- Reactions between metals and nonmetals involve a transfer of electrons from the metal to the nonmetal which is called an oxidation–reduction reaction.

8.3 Classifying Reactions

- Reactions can be classified in various ways.
 - A synthesis reaction is one in which a compound forms from simpler substances, such as elements.
 - A decomposition reaction occurs when a compound is broken down into simpler substances.
 - A combustion reaction is an oxidation–reduction reaction that involves O_2 .
 - A precipitation reaction can also be classified as a double displacement reaction.

All exercises with **blue numbers** have answers in the back of this book.

8.1 Understanding Reactions in Aqueous Solutions

A. Predicting Whether a Reaction Will Occur

- Why is water an important solvent? Although you have not yet studied water in detail, can you think of some properties of water that make it so important?
- What is a “driving force”? What are some of the driving forces discussed in this section that tend to make reactions likely to occur? Can you think of any other possible driving forces?

B. Reactions in Which a Solid Forms

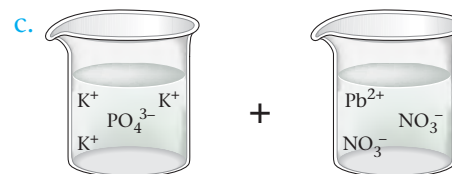
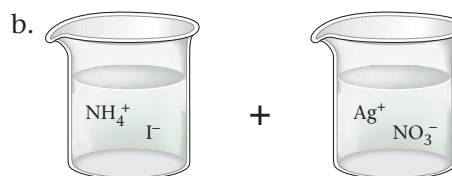
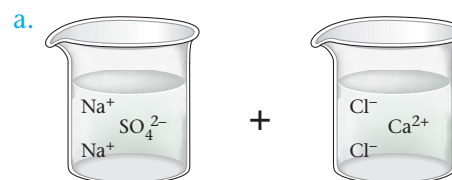
- When two solutions of ionic substances are mixed and a precipitate forms, what is the net charge of the precipitate? Why?
- What does it mean to say that the ions of an electrolyte behave independently of one another when the electrolyte is dissolved in water?
- What is meant by a *strong electrolyte*? Give two examples of substances that behave in solution as strong electrolytes.
- When aqueous solutions of sodium chloride, NaCl, and silver nitrate, AgNO₃, are mixed, a precipitate forms, but this precipitate is *not* sodium nitrate. What does this reaction tell you about the solubility of NaNO₃ in water?
- What do we mean when we say that a solid is “slightly” soluble in water? Is there a practical difference between *slightly* soluble and *insoluble*?
- On the basis of the general solubility rules given in Table 8.1, predict which of the following substances are likely to be soluble in water.

- barium nitrate
- potassium carbonate
- sodium sulfate
- copper(II) hydroxide
- mercury(I) chloride
- ammonium phosphate
- chromium(III) sulfide
- lead(II) sulfate

- On the basis of the general solubility rules given in Table 8.1, for each of the following compounds, indicate why the compound is *not* likely to be soluble in water. Indicate *which* of the solubility rules covers each substance’s particular situation.

- iron(III) hydroxide
- calcium carbonate
- cobalt(III) phosphate
- silver chloride

- On the basis of the general solubility rules given in Table 8.1, predict the identity of the precipitate that forms when aqueous solutions of the following substances are mixed. If no precipitate is likely, indicate which rules apply.



d. sodium hydroxide, NaOH, and iron(III) chloride, FeCl₃

e. potassium sulfate, K₂SO₄, and sodium nitrate, NaNO₃

f. sodium carbonate, Na₂CO₃, and barium nitrate, Ba(NO₃)₂

11. On the basis of the general solubility rules given in Table 8.1, write a balanced molecular equation for the precipitation reactions that take place when the following aqueous solutions are mixed. Underline the formula of the precipitate (solid) that forms. If no precipitation reaction is likely for the solutes given, so indicate.

- nitric acid, HNO_3 , and barium chloride, BaCl_2
- ammonium sulfide, $(\text{NH}_4)_2\text{S}$, and cobalt(II) chloride, CoCl_2
- sulfuric acid, H_2SO_4 , and lead(II) nitrate, $\text{Pb}(\text{NO}_3)_2$
- calcium chloride, CaCl_2 , and potassium carbonate, K_2CO_3
- sodium acetate, $\text{NaC}_2\text{H}_3\text{O}_2$, and ammonium nitrate, NH_4NO_3
- sodium phosphate, Na_3PO_4 , and chromium(III) chloride, CrCl_3

12. Balance each of the following equations that describe precipitation reactions.

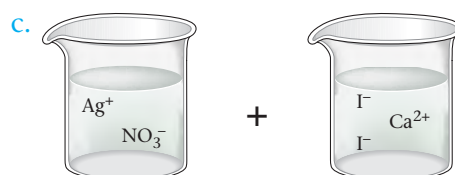
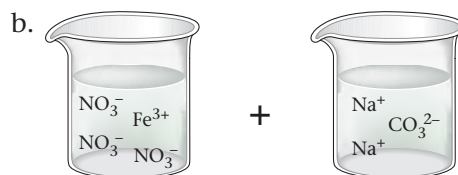
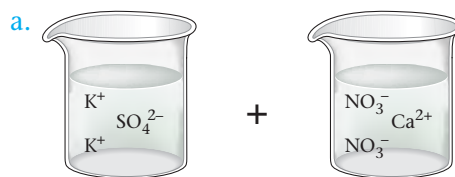
- $\text{AgNO}_3(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{Ag}_2\text{SO}_4(\text{s}) + \text{HNO}_3(\text{aq})$
- $\text{Ca}(\text{NO}_3)_2(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{CaSO}_4(\text{s}) + \text{HNO}_3(\text{aq})$
- $\text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{PbSO}_4(\text{s}) + \text{HNO}_3(\text{aq})$

13. For each of the following precipitation reactions, complete and balance the equation, indicating clearly which product is the precipitate.

- $(\text{NH}_4)_2\text{S}(\text{aq}) + \text{CoCl}_2(\text{aq}) \rightarrow ?$
- $\text{FeCl}_3(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow ?$
- $\text{CuSO}_4(\text{aq}) + \text{Na}_2\text{CO}_3(\text{aq}) \rightarrow ?$

C. Describing Reactions in Aqueous Solutions

14. Write balanced net ionic equations for the reactions that occur when the following aqueous solutions are mixed. If no reaction is likely to occur, so indicate.



- ammonium phosphate and cobalt(II) sulfate
- strontium chloride and mercury(I) nitrate
- barium bromide and lead(II) nitrate

15. In the laboratory, students often learn to analyze mixtures of the common positive and negative ions, separating and confirming the presence of the particular ions in the mixture. One of the first steps in such an analysis is to treat the mixture with hydrochloric acid, which precipitates and removes silver ion, lead(II) ion, and mercury(I) ion from the aqueous mixture as the insoluble chloride salts. Write balanced net ionic equations for the precipitation reactions of these three cations with chloride ion.

16. Many plants are poisonous because their stems and leaves contain oxalic acid, $\text{H}_2\text{C}_2\text{O}_4$, or sodium oxalate, $\text{Na}_2\text{C}_2\text{O}_4$; when ingested, these substances cause swelling of the respiratory tract and suffocation. A standard analysis for determining the amount of oxalate ion, $\text{C}_2\text{O}_4^{2-}$, in a sample is to precipitate this species as calcium oxalate, which is insoluble in water. Write the net ionic equation for the reaction between sodium oxalate and calcium chloride, CaCl_2 , in aqueous solution.

8.2 Other Reactions in Aqueous Solutions

A. Reactions That Form Water: Acids and Bases

17. What is meant by a *strong acid*? Are the strong acids also *strong electrolytes*? Explain.
18. What is meant by a *strong base*? Are the strong bases also *strong electrolytes*? Explain.
19. If 1000 NaOH units were dissolved in a sample of water, the NaOH would produce _____ Na⁺ ions and _____ OH⁻ ions.
20. Along with the three strong acids emphasized in the chapter (HCl, HNO₃, and H₂SO₄), hydrobromic acid, HBr, and perchloric acid, HClO₄, are also strong acids. Write equations for the dissociation of each of these additional strong acids in water.
21. Along with the strong bases NaOH and KOH discussed in this chapter, the hydroxide compounds of other Group 1 elements also behave as strong bases when dissolved in water. Write equations for RbOH and CsOH that show which ions form when they dissolve in water.
22. Complete and balance each of the following molecular equations for strong acid/strong base reactions; circle the formula of the salt produced in each.
- $\text{HCl}(aq) + \text{RbOH}(aq) \rightarrow ?$
 - $\text{HClO}_4(aq) + \text{NaOH}(aq) \rightarrow ?$
 - $\text{HBr}(aq) + \text{NaOH}(aq) \rightarrow ?$
 - $\text{H}_2\text{SO}_4(aq) + \text{CsOH}(aq) \rightarrow ?$
23. Below are indicated the formulas of some salts. Such salts could be formed by the reaction of the appropriate strong acid and strong base (with the other product of the reaction being, of course, water). For each salt, write an equation showing the formation of the salt from reaction of the appropriate strong acid and strong base.
- | | |
|------------------------------------|----------------------|
| a. Na ₂ SO ₄ | c. KClO ₄ |
| b. RbNO ₃ | d. KCl |

B. Reactions of Metals with Nonmetals (Oxidation-Reduction)

24. What do we mean when we say that the transfer of electrons can be the “driving force” for a reaction? Give an example of a reaction where this happens.
25. If atoms of a metallic element (such as sodium) react with atoms of a nonmetallic element (such as sulfur), which element loses electrons and which element gains them?

26. If potassium atoms were to react with atoms of the nonmetal sulfur, how many electrons would each potassium atom lose? How many electrons would each sulfur atom gain? How many potassium atoms would have to react to provide enough electrons for one sulfur atom? What charges would the resulting potassium and sulfur ions have?
27. For the reaction $2\text{Al}(s) + 3\text{Br}_2(l) \rightarrow 2\text{AlBr}_3(s)$, show how electrons are gained and lost by the atoms.
28. Balance each of the following oxidation-reduction chemical reactions.
- $\text{K}(s) + \text{F}_2(g) \rightarrow \text{KF}(s)$
 - $\text{K}(s) + \text{O}_2(g) \rightarrow \text{K}_2\text{O}(s)$
 - $\text{K}(s) + \text{N}_2(g) \rightarrow \text{K}_3\text{N}(s)$
 - $\text{K}(s) + \text{C}(s) \rightarrow \text{K}_4\text{C}(s)$
29. Balance each of the following oxidation-reduction chemical reactions.
- $\text{Fe}(s) + \text{S}(s) \rightarrow \text{Fe}_2\text{S}_3(s)$
 - $\text{Zn}(s) + \text{HNO}_3(aq) \rightarrow \text{Zn}(\text{NO}_3)_2(aq) + \text{H}_2(g)$
 - $\text{Sn}(s) + \text{O}_2(g) \rightarrow \text{SnO}(s)$
 - $\text{K}(s) + \text{H}_2(g) \rightarrow \text{KH}(s)$
 - $\text{Cs}(s) + \text{H}_2\text{O}(l) \rightarrow \text{CsOH}(s) + \text{H}_2(g)$

8.3 Classifying Reactions

A. Ways to Classify Reactions

30. What is a *double-displacement* reaction? What is a *single-replacement* reaction? Write balanced chemical equations showing two examples of each type.
31. Two “driving forces” for reactions discussed in this section are the formation of water in an acid-base reaction and the formation of a gaseous product. Write balanced chemical equations showing two examples of each type.

32. Identify each of the following unbalanced reaction equations as belonging to one or more of the following categories: precipitation, acid–base, or oxidation–reduction.

- $\text{K}_2\text{SO}_4(aq) + \text{Ba}(\text{NO}_3)_2(aq) \rightarrow \text{BaSO}_4(s) + \text{KNO}_3(aq)$
- $\text{HCl}(aq) + \text{Zn}(s) \rightarrow \text{H}_2(g) + \text{ZnCl}_2(aq)$
- $\text{HCl}(aq) + \text{AgNO}_3(aq) \rightarrow \text{HNO}_3(aq) + \text{AgCl}(s)$
- $\text{HCl}(aq) + \text{KOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{KCl}(aq)$
- $\text{Zn}(s) + \text{CuSO}_4(aq) \rightarrow \text{ZnSO}_4(aq) + \text{Cu}(s)$
- $\text{NaH}_2\text{PO}_4(aq) + \text{NaOH}(aq) \rightarrow \text{Na}_3\text{PO}_4(aq) + \text{H}_2\text{O}(l)$
- $\text{Ca}(\text{OH})_2(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{CaSO}_4(s) + \text{H}_2\text{O}(l)$
- $\text{ZnCl}_2(aq) + \text{Mg}(s) \rightarrow \text{Zn}(s) + \text{MgCl}_2(aq)$
- $\text{BaCl}_2(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{BaSO}_4(s) + \text{HCl}(aq)$

33. Identify each of the following unbalanced reaction equations as belonging to one or more of the following categories: precipitation, acid–base, or oxidation–reduction.

- $\text{H}_2\text{O}_2(aq) \rightarrow \text{H}_2\text{O}(l) + \text{O}_2(g)$
- $\text{H}_2\text{SO}_4(aq) + \text{Cu}(s) \rightarrow \text{CuSO}_4(aq) + \text{H}_2(g)$
- $\text{H}_2\text{SO}_4(aq) + \text{NaOH}(aq) \rightarrow \text{Na}_2\text{SO}_4(aq) + \text{H}_2\text{O}(l)$
- $\text{H}_2\text{SO}_4(aq) + \text{Ba}(\text{OH})_2(aq) \rightarrow \text{BaSO}_4(s) + \text{H}_2\text{O}(l)$
- $\text{AgNO}_3(aq) + \text{CuCl}_2(aq) \rightarrow \text{Cu}(\text{NO}_3)_2(aq) + \text{AgCl}(s)$
- $\text{KOH}(aq) + \text{CuSO}_4(aq) \rightarrow \text{Cu}(\text{OH})_2(s) + \text{K}_2\text{SO}_4(aq)$
- $\text{Cl}_2(g) + \text{F}_2(g) \rightarrow \text{ClF}(g)$
- $\text{NO}(g) + \text{O}_2(g) \rightarrow \text{NO}_2(g)$
- $\text{Ca}(\text{OH})_2(s) + \text{HNO}_3(aq) \rightarrow \text{Ca}(\text{NO}_3)_2(aq) + \text{H}_2\text{O}(l)$

B. Other Ways to Classify Reactions

34. What is a *synthesis* or *combination* reaction? Give an example. Can such reactions also be classified in other ways? Give an example of a synthesis reaction that is also a *combustion* reaction. Give an example of a synthesis reaction that is also an *oxidation–reduction* reaction, but which does not involve combustion.

35. What is a *decomposition* reaction? Give an example. Can such reactions also be classified in other ways?

36. Balance each of the following equations that describe combustion reactions.

- $\text{C}_2\text{H}_5\text{OH}(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)$
- $\text{C}_6\text{H}_{14}(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)$
- $\text{C}_6\text{H}_{12}(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)$

37. Balance each of the following equations that describe combustion reactions.

- $\text{C}_2\text{H}_6(g) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)$
- $\text{C}_2\text{H}_6\text{O}(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)$
- $\text{C}_2\text{H}_6\text{O}_2(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)$

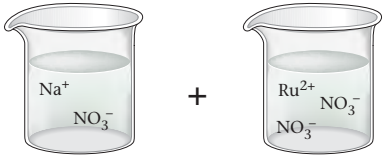
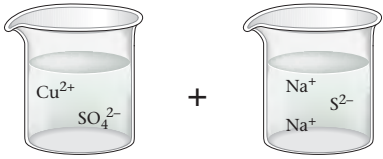
38. Balance each of the following equations that describe synthesis reactions.

- $\text{Co}(s) + \text{S}(s) \rightarrow \text{Co}_2\text{S}_3(s)$
- $\text{NO}(g) + \text{O}_2(g) \rightarrow \text{NO}_2(g)$
- $\text{FeO}(s) + \text{CO}_2(g) \rightarrow \text{FeCO}_3(s)$
- $\text{Al}(s) + \text{F}_2(g) \rightarrow \text{AlF}_3(s)$
- $\text{NH}_3(g) + \text{H}_2\text{CO}_3(aq) \rightarrow (\text{NH}_4)_2\text{CO}_3(s)$

39. Balance each of the following equations that describe decomposition reactions.

- $\text{NI}_3(s) \rightarrow \text{N}_2(g) + \text{I}_2(s)$
- $\text{BaCO}_3(s) \rightarrow \text{BaO}(s) + \text{CO}_2(g)$
- $\text{C}_6\text{H}_{12}\text{O}_6(s) \rightarrow \text{C}(s) + \text{H}_2\text{O}(g)$
- $\text{Cu}(\text{NH}_3)_4\text{SO}_4(s) \rightarrow \text{CuSO}_4(s) + \text{NH}_3(g)$
- $\text{NaN}_3(s) \rightarrow \text{Na}_3\text{N}(s) + \text{N}_2(g)$

Critical Thinking

40. Distinguish between the *molecular* equation, the *complete ionic* equation, and the *net ionic* equation for a reaction in solution. Which type of equation most clearly shows the substances that actually react with one another?
41. Using the general solubility rules given in Table 8.1, name three reactants that would form precipitates with each of the following ions in aqueous solution. Write the net ionic equation for each of your suggestions.
- chloride ion
 - calcium ion
 - iron(III) ion
 - sulfate ion
 - mercury(I) ion, Hg_2^{2+}
 - silver ion
42. Without first writing a full molecular or ionic equation, write the net ionic equations for any precipitation reactions that occur when aqueous solutions of the following compounds are mixed. If no reaction occurs, so indicate.
- 
 - 
 - lithium chloride and lead(II) nitrate
 - calcium nitrate and lithium carbonate
 - gold(III) chloride and sodium hydroxide

43. Complete and balance each of the following molecular equations for strong acid/strong base reactions. Underline the formula of the *salt* produced in each reaction.

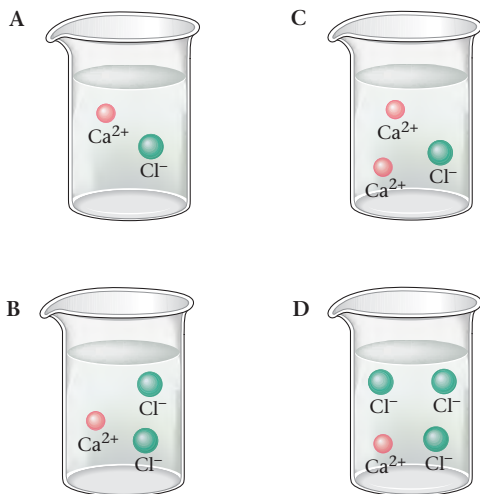
- $\text{HNO}_3(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow ?$
- $\text{H}_2\text{SO}_4(\text{aq}) + \text{Ba}(\text{OH})_2(\text{aq}) \rightarrow ?$
- $\text{HClO}_4(\text{aq}) + \text{NaOH}(\text{aq}) \rightarrow ?$
- $\text{HCl}(\text{aq}) + \text{Ca}(\text{OH})_2(\text{aq}) \rightarrow ?$

44. Copy and complete the following table with the formulas for the precipitates that would form. If no precipitate would form, write "No PPT."

	$\text{C}_2\text{H}_3\text{O}_2^-$	Cl^-	CO_3^{2-}	NO_3^-	OH^-	PO_4^{3-}	S^{2-}	SO_4^{2-}
Ag^+								
Ba^{2+}								
Ca^{2+}								
Fe^{3+}								
Hg_2^{2+}								
Na^+								
Ni^{2+}								
Pb^{2+}								

- 45.** On the basis of the general solubility rules given in Table 8.1, write a balanced molecular equation for the precipitation reactions that take place when the following aqueous solutions are mixed. Underline the formula of the precipitate (solid) that forms. If no precipitation reaction is likely for the reactants given, so indicate.
- silver nitrate and hydrochloric acid
 - copper(II) sulfate and ammonium carbonate
 - iron(II) sulfate and potassium carbonate
 - silver nitrate and potassium nitrate
 - lead(II) nitrate and lithium carbonate
 - tin(IV) chloride and sodium hydroxide
- 46.** For each of the following *unbalanced* molecular equations, write the corresponding *balanced net ionic equation* for the reaction.
- $\text{HCl}(aq) + \text{AgNO}_3(aq) \rightarrow \text{AgCl}(s) + \text{HNO}_3(aq)$
 - $\text{CaCl}_2(aq) + \text{Na}_3\text{PO}_4(aq) \rightarrow \text{Ca}_3(\text{PO}_4)_2(s) + \text{NaCl}(aq)$
 - $\text{Pb}(\text{NO}_3)_2(aq) + \text{BaCl}_2(aq) \rightarrow \text{PbCl}_2(s) + \text{Ba}(\text{NO}_3)_2(aq)$
 - $\text{FeCl}_3(aq) + \text{NaOH}(aq) \rightarrow \text{Fe}(\text{OH})_3(s) + \text{NaCl}(aq)$
- 47.** Write a balanced oxidation–reduction equation for the reaction of each of the metals below
- Ba K Mg Rb Ca Li
- with each of the nonmetals below.
- O_2 S Cl_2 N_2 Br_2
- 48.** For each of the following metals, how many electrons will the metal atoms lose when the metal reacts with a nonmetal?
- sodium
 - potassium
 - magnesium
 - barium
 - aluminum
- 49.** For each of the following nonmetals, how many electrons will each atom of the nonmetal gain in reacting with a metal?
- oxygen
 - fluorine
 - nitrogen
 - chlorine
 - sulfur
- 50.** Classify the reactions represented by the following unbalanced equations by as many methods as possible. Balance the equations.
- $\text{C}_3\text{H}_8\text{O}(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g)$
 - $\text{HCl}(aq) + \text{AgC}_2\text{H}_3\text{O}_2(aq) \rightarrow \text{AgCl}(s) + \text{HC}_2\text{H}_3\text{O}_2(aq)$
 - $\text{HCl}(aq) + \text{Al}(\text{OH})_3(s) \rightarrow \text{AlCl}_3(aq) + \text{H}_2\text{O}(l)$
 - $\text{H}_2\text{O}_2(aq) \rightarrow \text{H}_2\text{O}(l) + \text{O}_2(g)$
 - $\text{N}_2\text{H}_4(l) + \text{O}_2(g) \rightarrow \text{N}_2(g) + \text{H}_2\text{O}(g)$
- 51.** Classify each equation in one or more of the following categories:
- precipitation
 - acid–base
 - oxidation–reduction
 - combustion
 - synthesis
 - decomposition
- $\text{HNO}_3(aq) + \text{KOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{KNO}_3(aq)$
 - $\text{Ca}(\text{NO}_3)_2(aq) + \text{K}_2\text{SO}_4(aq) \rightarrow \text{CaSO}_4(s) + 2\text{KNO}_3(aq)$
 - $\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g)$
 - $2\text{H}_2\text{O}(l) \rightarrow 2\text{H}_2(g) + \text{O}_2(g)$

- 1 Which of the following is the best microscopic representation of an aqueous solution of calcium chloride?

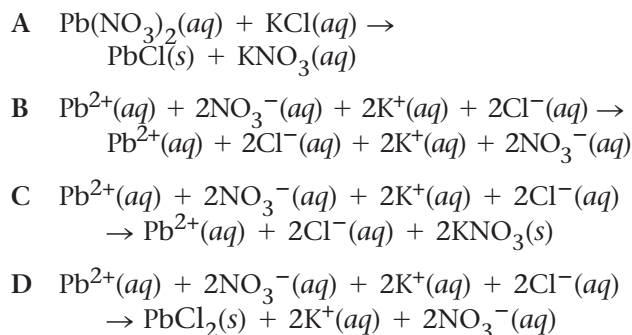


- 2 If two moles of sodium sulfate are dissolved in water, how many moles of ions are in solution?
- A 2
B 3
C 6
D 15

For questions 3–5, use the solubility rules (Table 8.1).

- 3 Which of the following would result in the production of a solid?
- A $\text{NaCl}(aq) + \text{KNO}_3(aq)$
B $\text{AgNO}_3(aq) + \text{NaNO}_3(aq)$
C $\text{Ba}(\text{NO}_3)_2(aq) + \text{K}_2\text{SO}_4(aq)$
D $\text{HCl}(aq) + \text{NaOH}(aq)$
- 4 When an aqueous solution of silver nitrate is mixed with an aqueous solution of potassium chloride, which are the spectator ions?
- A nitrate ions and chloride ions
B nitrate ions and potassium ions
C silver ions and chloride ions
D silver ions and potassium ions

- 5 Choose the correct balanced complete ionic equation for the reaction of an aqueous solution of lead(II) nitrate with an aqueous solution of potassium chloride.



- 6 Which of the following types of reactions always has water as a product?

A precipitation reaction
B oxidation–reduction reaction
C combustion reaction
D strong acid–strong base reaction

- 7 Which of the following types of reactions is *not* an oxidation–reduction reaction?

A precipitation reaction
B synthesis reaction
C combustion reaction
D decomposition reaction

- 8 Aluminum reacts with oxygen in the air as an oxidation–reduction reaction. Which of the following best describes what happens?

A Each aluminum atom gains three electrons and each oxygen atom loses two electrons.
B Each aluminum atom loses two electrons and each oxygen atom gains two electrons.
C Each aluminum atom loses three electrons and each oxygen atom gains two electrons.
D Each aluminum atom gains two electrons and each oxygen atom loses two electrons.

LOOKING AHEAD

9.1 Using Chemical Equations

- A. Information Given by Chemical Equations
- B. Mole–mole Relationships

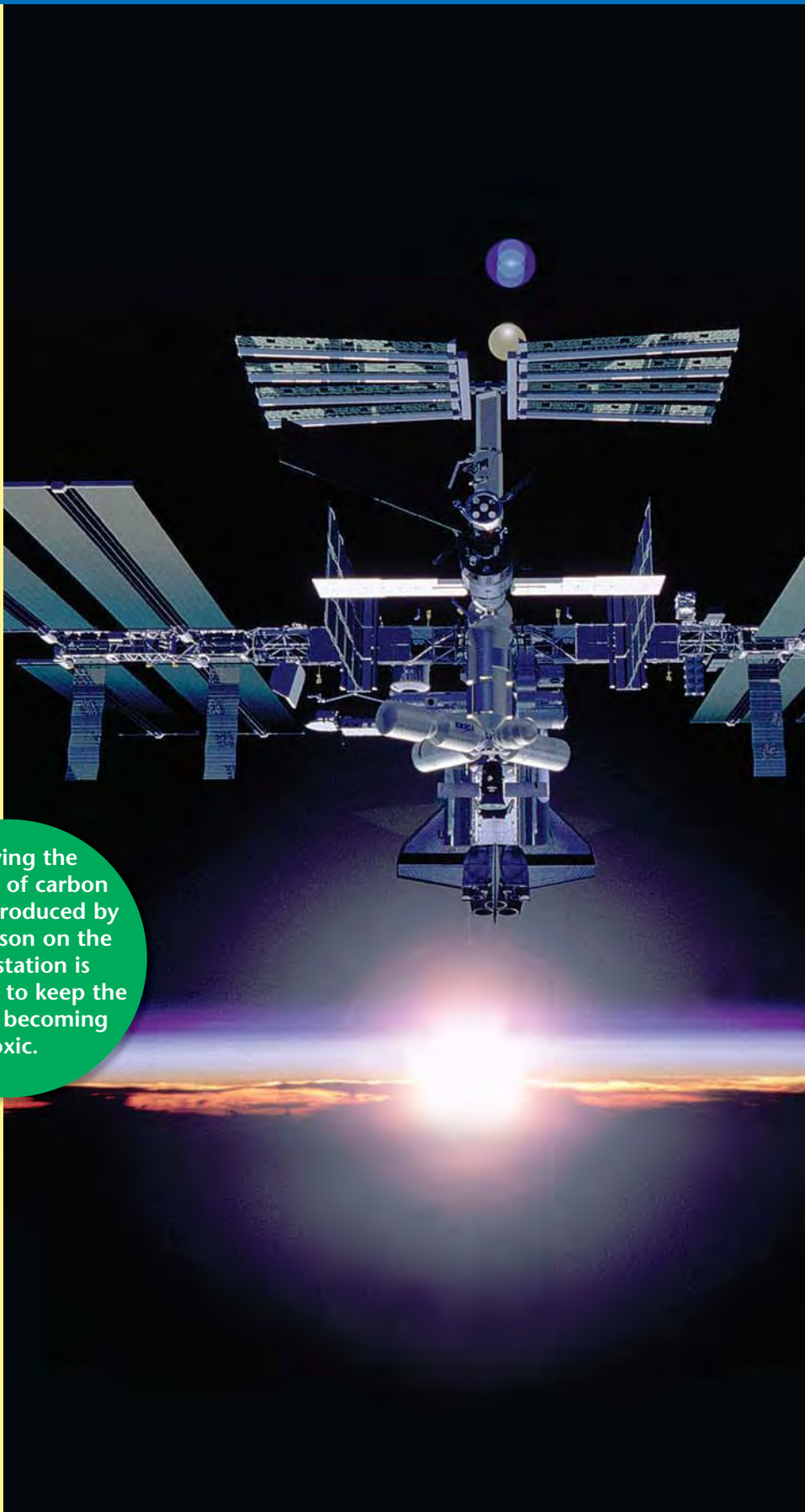
9.2 Using Chemical Equations to Calculate Mass

- A. Mass Calculations
- B. Mass Calculations Using Scientific Notation
- C. Mass Calculations: Comparing Two Reactions

9.3 Limiting Reactants and Percent Yield

- A. The Concept of Limiting Reactants
 - A Closer Look
- B. Calculations Involving a Limiting Reactant
- C. Percent Yield

Knowing the quantity of carbon dioxide produced by each person on the space station is important to keep the air from becoming toxic.



IN YOUR LIFE

Suppose that you work for a consumer advocate organization and you want to test a company's advertising claims about the effectiveness of its antacid. The company claims that its product neutralizes ten times as much stomach acid per tablet as its nearest competitor. How would you test the validity of this claim?

Or suppose that some day you go to work for a chemical company that makes methanol (methyl alcohol), a substance used as a starting material for the manufacture of products such as antifreeze and aviation fuels and as a fuel in the cars that race in the Indianapolis 500. You are working with an experienced chemist who is trying to improve the company's process for making methanol from the reaction of gaseous hydrogen with carbon monoxide gas. The first day on the job, you are instructed to order enough hydrogen and carbon monoxide to produce 6.0 kg of methanol in a test run. How would you determine how much carbon monoxide and hydrogen to order?

After you study this chapter, you will be able to answer these questions.

WHAT DO YOU KNOW? **Prereading Questions**

1. You are building model cars. Each car has four tires, two doors, and one body. If you have 16 tires, 10 doors, and 7 car bodies, how many model cars can you build? What is left over after you have built all of the cars?
2. Consider the reaction $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$. One mole of nitrogen gas requires how many moles of hydrogen gas? Two moles of nitrogen gas requires how many moles of hydrogen gas?
3. What is the molar mass of water?
4. Consider the reaction $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$. If there are two moles of each reactant, which reactant runs out first?
5. A certain recipe for chocolate chip cookies lists the yield as two dozen cookies. You make slightly larger cookies than recommended and make only 14 cookies. What percent of cookies did you make?

SECTION 9.1

Using Chemical Equations

Key Term

- Mole ratio

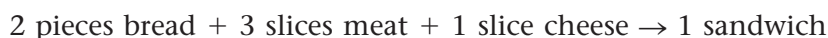
Objectives

- To understand the information given in a balanced equation
- To use a balanced equation to determine relationships between moles of reactant and products

A. Information Given by Chemical Equations

As we saw in the last two chapters, chemistry is really about reactions. Chemical changes involve rearrangements of atom groupings as one or more substances change to new substances. Recall that reactions are described by equations that give the identities of the reactants and products and show how much of each reactant and product participates in the reaction. It is the numbers (coefficients) in the balanced chemical equation that enable us to determine just how much product we can get from a given quantity of reactants.

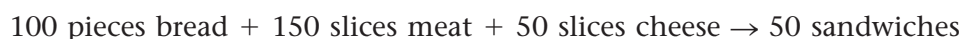
To explore this idea, consider a nonchemical analogy. Assume that you are in charge of making deli sandwiches at a local fast-food restaurant. A particular type of sandwich requires 2 pieces of bread, 3 slices of meat, and 1 slice of cheese. You might represent making this sandwich by the following equation:



Your boss sends you to the store to get enough ingredients to make 50 sandwiches. How do you figure out how much of each ingredient to buy? Because you need enough to make 50 sandwiches, you could multiply the preceding equation by 50.



That is,



Notice that the numbers 100:150:50 correspond to the ratio 2:3:1, which represents the original numbers of bread, meat, and cheese ingredients in the equation for making a sandwich. If you needed the ingredients for 73 sandwiches, it would be easy to use the original sandwich equation to figure out what quantity of ingredients you need.

The equation for a chemical reaction gives you the same type of information. It indicates the relative numbers of reactant and product molecules required for the reaction to take place. Using the equation permits us to determine the amounts of reactants needed to give a certain amount of product or to predict how much product we can make from a given quantity of reactants.

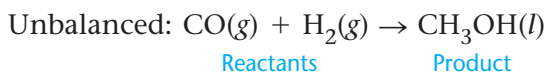
Active Reading Question

Why are we able to multiply coefficients in a chemical equation?
Why must we multiply each coefficient by the same amount?



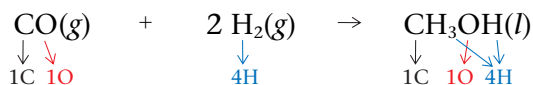
A worker making sandwiches in a deli

To illustrate how this idea works, consider the reaction between gaseous carbon monoxide and hydrogen to produce liquid methanol, $\text{CH}_3\text{OH}(l)$. The reactants and products are



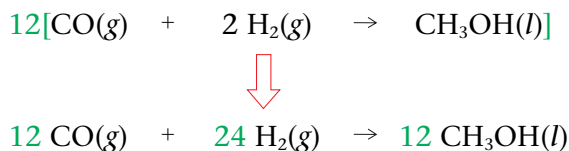
Because atoms are just rearranged (not created or destroyed) in a chemical reaction, we must always balance a chemical equation.

That is, we must choose coefficients that give the same number of each type of atom on both sides. Using the smallest set of integers that satisfies this condition gives the balanced equation



It is important to recognize that the coefficients in a balanced equation give the *relative* numbers of molecules.

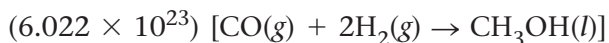
That is, we could multiply this balanced equation by any number and still have a balanced equation. For example, we could multiply by 12,



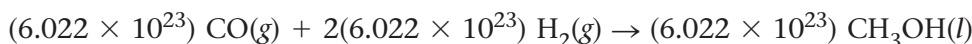
This is still a balanced equation (check to be sure). Because 12 represents a dozen, we could even describe the reaction in terms of dozens:



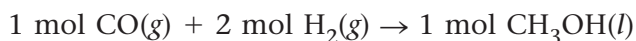
We could also multiply the original equation by a very large number, such as 6.022×10^{23} ,



which leads to the equation

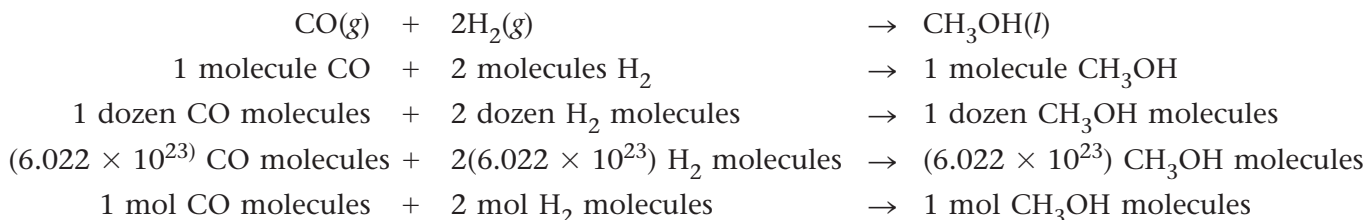


Just as 12 is called a dozen, chemists call 6.022×10^{23} a *mole* (abbreviated mol). Our equation, then, can be written in terms of moles:



Let's Review

Information Conveyed by the Balanced Equation for the Production of Methanol



Information

One mole is
 6.022×10^{23} units.

EXAMPLE 9.1

Relating Moles to Molecules in Chemical Equations



Propane is often used as a fuel for outdoor grills.

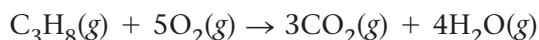
Propane, C_3H_8 , is a fuel commonly used for cooking on gas grills and for heating in rural areas where natural gas is unavailable. Propane reacts with oxygen gas to produce heat and the products carbon dioxide and water. This combustion reaction is represented by the unbalanced equation



Give the balanced equation for this reaction, and state the meaning of the equation in terms of numbers of molecules and moles of molecules.

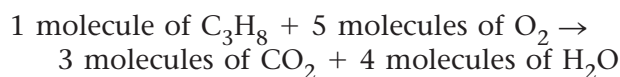
Solution

Using the techniques explained in Chapter 7, we can balance the equation.

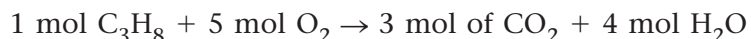


Check: 3 C, 8 H, 10 O \rightarrow 3 C, 8 H, 10 O

This equation can be interpreted in terms of molecules as follows:



Alternatively, it can be stated in terms of moles (of molecules):



HANDS-ON CHEMISTRY

• MINI-LAB •

The Nuts and Bolts of Stoichiometry

Materials

- a cup of nuts and bolts

Procedure

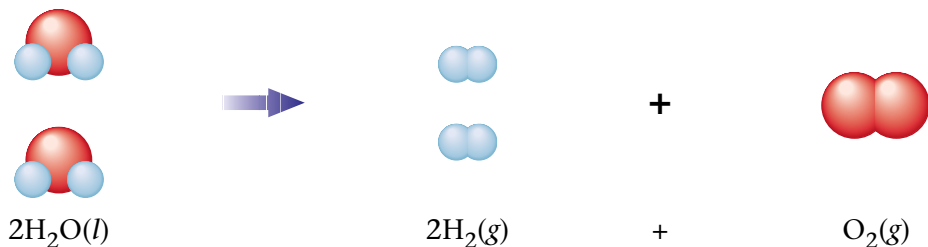
1. Obtain a cup of nuts and bolts from your teacher.
2. The nuts and bolts are the reactants. The product consists of 2 nuts on each bolt. Make as many products as possible.

Results/Analysis

1. Using N to symbolize a nut and B to symbolize a bolt, write an equation for the formation of the product. Pay attention to the difference between a subscript and a coefficient.
2. How many nuts did you have? How many bolts?
3. How many products could you make?
4. Which reactant (nut or bolt) was limiting? How did you make this determination?
5. The limiting reactant was the one that had (fewer/more) pieces.
6. An average mass of a bolt is 10.64 g and an average mass of a nut is 4.35 g. Suppose you are given "about 1500 g" of bolts and "about 1500 g" of nuts. Answer the following questions:
 - a. How many bolts are in "about 1500 g"? How many nuts are in "about 1500 g"?
 - b. Which reactant is limiting? Why is there a limiting reactant, given that you have equal masses of each?
 - c. The limiting reactant was the one that had (fewer/more) pieces. Compare this answer to your answer in question 5. What does it tell you?
 - d. What is the largest possible mass of product? How many products could you make?
 - e. What is the mass of the leftover reactant?

B. Mole–mole Relationships

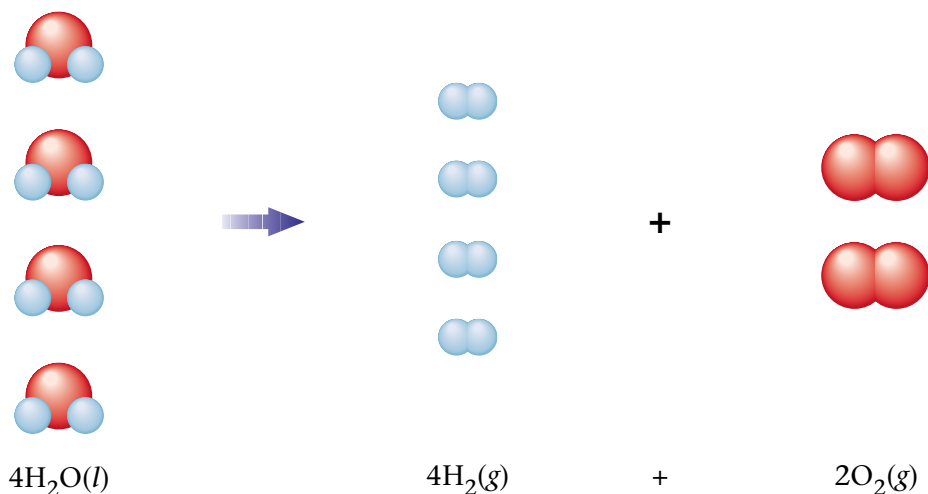
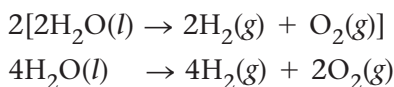
Now that we have discussed the meaning of a balanced chemical equation in terms of moles of reactants and products, we can use an equation to predict the moles of products that a given number of moles of reactants will yield. For example, consider the decomposition of water to give hydrogen and oxygen, which is represented by the following balanced equation:



This equation tells us that 2 mol of H_2O yields 2 mol of H_2 and 1 mol of O_2 .

Now suppose that we have 4 mol of water. If we decompose 4 mol of water, how many moles of products do we get?

One way to answer this question is to multiply the entire equation by 2; that will give us 4 mol of H_2O .



Now we can state that

4 mol of H_2O yields 4 mol of H_2 plus 2 mol of O_2

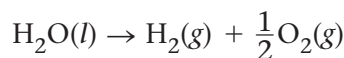
which answers the question of how many moles of products we get with 4 mol of H_2O .

Next, suppose that we decompose 5.8 mol of water. What numbers of moles of products are formed in this process?

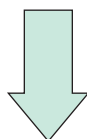
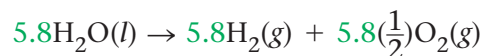
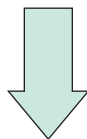
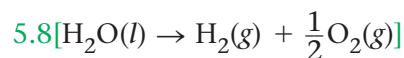


We could answer this question by rebalancing the chemical equation as follows:

- First, we divide *all coefficients* of the balanced equation by 2 to give



- Now, because we have 5.8 mol of H_2O , we multiply this equation by 5.8.



- Verify that this is a balanced equation.

Now we can state that

5.8 mol of H_2O yields 5.8 mol of H_2 plus 2.9 mol of O_2 .

This procedure of rebalancing the equation to obtain the number of moles involved in a particular situation always works, but it can be cumbersome. In Example 9.2 we will develop a more convenient procedure, which uses conversion factors, or **mole ratios**, based on the balanced chemical equation.

i nformation

This equation with noninteger coefficients makes sense only if the equation means moles (of molecules) of the various reactants and products.

Mole ratio

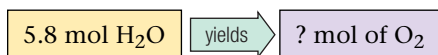
The ratio of moles of one substance to moles of another substance in a balanced chemical equation

Active Reading Question

What is meant by the term “mole ratio”? Give an example from an actual chemical equation.

EXAMPLE 9.2**Determining Mole Ratios**

What number of moles of O_2 will be produced by the decomposition of 5.8 mol of water?

Solution**Where do we want to go?****What do we know?**

- $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$
- 5.8 mol H_2O

How do we get there?

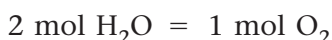
To answer this question, we need to know the relationship between moles of H_2O and moles of O_2 in the balanced equation (conventional form):



From this equation we can state that



which can be represented by the following equivalence statement:



We now want to use this equivalence statement to obtain the conversion factor (mole ratio) that we need. Because we want to go from moles of H_2O to moles of O_2 , we need the mole ratio

$$\frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O}$$

so that mol H_2O will cancel in the conversion from moles of H_2O to moles of O_2 .

$$5.8 \text{ mol } H_2O \times \frac{1 \text{ mol } O_2}{2 \text{ mol } H_2O} = 2.9 \text{ mol } O_2$$

So if we decompose 5.8 mol of H_2O , we will get 2.9 mol of O_2 .

Does it make sense?

Note that this is the same answer we obtained earlier when we rebalanced the equation to give $5.8 H_2O(l) \rightarrow 5.8 H_2(g) + 2.9 O_2(g)$.

We saw in Example 9.2 that to determine the moles of a product that can be formed from a specified number of moles of a reactant, we can use the balanced equation to obtain the appropriate mole ratio. We will now extend these ideas in Example 9.3.

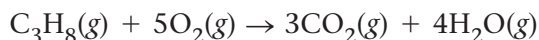
i n f o r m a t i o n

The statement $2 \text{ mol } H_2O = 1 \text{ mol } O_2$ is obviously not true in a literal sense, but it correctly expresses the chemical equivalence between H_2O and O_2 .

EXAMPLE 9.3

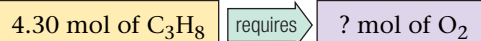
Using Mole Ratios in Calculations

Calculate the number of moles of oxygen required to react exactly with 4.30 mol of propane, C_3H_8 , in the reaction described by the following balanced equation:



Solution

Where do we want to go?



What do we know?

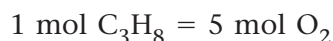
- $C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
- 4.30 mole C_3H_8

How do we get there?

To solve this problem, we need to consider the relationship between the reactants C_3H_8 and O_2 . Using the balanced equation, we find that

1 mol of C_3H_8 requires 5 mol of O_2

which can be represented by the equivalence statement



This leads to the required mole ratio

$$\frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8}$$

for converting from moles of C_3H_8 to moles of O_2 . We construct the conversion ratio this way so that mol C_3H_8 cancels:

$$4.30 \text{ mol } C_3H_8 \times \frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8} = 21.5 \text{ mol } O_2$$

We can now answer the original question:

4.30 mol of C_3H_8 requires 21.5 mol of O_2 .

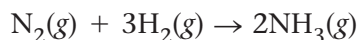
Practice Problem • Exercise 9.3

Calculate the moles of CO_2 formed when 4.30 mol of C_3H_8 reacts with the required 21.5 mol of O_2 .

Hint: Use the moles of C_3H_8 , and obtain the mole ratio between C_3H_8 and CO_2 from the balanced equation.

EXAMPLE 9.4**Using Mole Ratios**

Ammonia (NH_3) is used in huge quantities as a fertilizer. It is manufactured by combining nitrogen and hydrogen according to the following equation:



Calculate the number of moles of NH_3 that can be made from 1.30 mol $\text{H}_2(\text{g})$ reacting with excess $\text{N}_2(\text{g})$.

Solution**Where do we want to go?****What do we know?**

- $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$
- 1.30 mol H_2

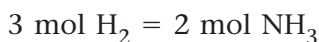
How do we get there?

The problem states that we have excess $\text{N}_2(\text{g})$, which means that we have more than is needed to react with all of the H_2 (1.30 mol).

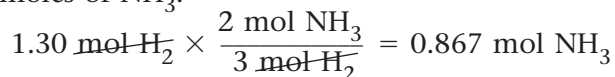
From the balanced equation we see that



which can be represented by the equivalence statement



This leads to the required mole ratio $\frac{2 \text{ mol } \text{NH}_3}{3 \text{ mol } \text{H}_2}$ for converting moles of H_2 to moles of NH_3 :

**Does it make sense?**

Notice that the number of moles of NH_3 that can be produced is less than 1.30. This answer makes sense because it takes 3 moles of H_2 to yield 2 moles of NH_3 . Thus the number of moles of NH_3 produced (0.867) should be smaller than the number of moles of H_2 consumed (1.30).

**SECTION 9.1****REVIEW QUESTIONS**

- What part of a chemical equation allows us to determine how much product we can get from a known amount of reactants?
- Why is it important to balance a chemical equation before proceeding to answer questions about it?
- Give at least three different ways to interpret the coefficient in a chemical equation.
- What number of moles of $\text{O}_2(\text{g})$ is required to react with 3.6 mol $\text{SO}_2(\text{g})$ in the following reaction?

$$2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})$$
- Butane (C_4H_{10}) reacts with oxygen in the air to form carbon dioxide and water. How many moles of carbon dioxide can be formed from 1.25 mol butane and excess oxygen?



RESEARCH LINKS
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SECTION 9.2

Using Chemical Equations to Calculate Mass

Key Term

- Stoichiometry

Objectives

- To learn to relate masses of reactant and products in a chemical reaction
- To perform mass calculations that involve scientific notation

A. Mass Calculations

Previously, we saw how to use the balanced equation for a reaction to calculate the numbers of moles of reactants and products for a particular case. However, moles represent numbers of molecules, and we cannot count molecules directly. In chemistry we count by weighing. Therefore, we need to review the procedures for converting between moles and masses and see how these procedures are applied to chemical calculations.

Let's consider the combustion of propane, C_3H_8 . Propane, when used as a fuel, reacts with oxygen to produce carbon dioxide and water according to the following unbalanced equation:



What mass of oxygen will be required to react exactly with 44.1 g of propane?

Where do we want to go?

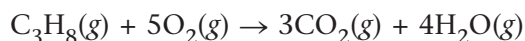


What do we know?

- $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
- 44.1 g propane

How do we get there?

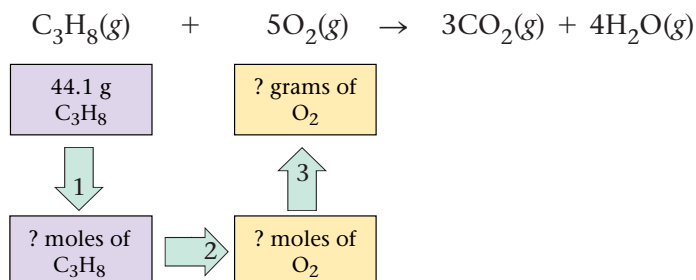
- Balance the equation:



- Our overall plan of attack is as follows:

- We are given the number of grams of propane, so we must convert to moles of propane (C_3H_8) because the balanced equation deals in moles rather than grams.
- Next, we can use the coefficients in the balanced equation to determine the moles of oxygen (O_2) required.
- Finally, we will use the molar mass of O_2 to calculate grams of oxygen.

We can sketch this strategy as follows:



Information

Always balance the equation for the reaction first.

Critical Thinking

Your lab partner has made the observation that we always measure the mass of chemicals but then use mole ratios to balance equations.

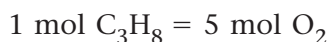
What if your lab partner decided to balance equations by using masses as coefficients? Is this possible? Why or why not?

1 The first question we must answer is, *How many moles of propane are present in 44.1 g of propane?*

The molar mass of propane is 44.09 g ($3 \times 12.01 + 8 \times 1.008$). The moles of propane present can be calculated as follows:

$$44.1 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} = 1.00 \text{ mol C}_3\text{H}_8$$

2 Next, we recognize that each mole of propane reacts with 5 mol of oxygen. This gives us the equivalence statement



from which we construct the mole ratio

$$\frac{5 \text{ mol O}_2}{1 \text{ mol C}_3\text{H}_8}$$

that we need to convert from moles of propane molecules to moles of oxygen molecules.

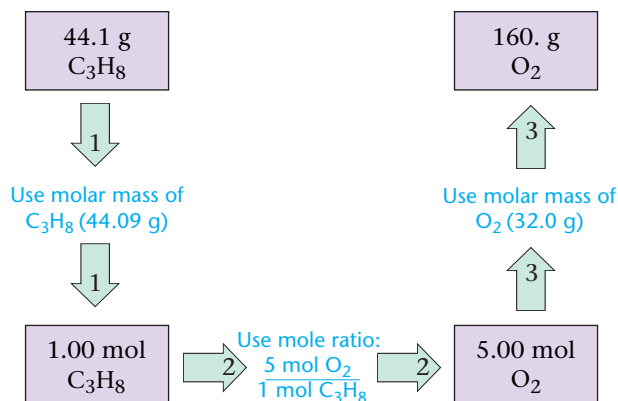
$$1.00 \text{ mol C}_3\text{H}_8 \times \frac{5 \text{ mol O}_2}{1 \text{ mol C}_3\text{H}_8} = 5.00 \text{ mol O}_2$$

Notice that the mole ratio is set up so that the moles of C_3H_8 cancel and the resulting units are moles of O_2 .

3 Because the original question asked for the *mass* of oxygen needed to react with 44.1 g of propane, we must convert the 5.00 mol of O_2 to grams, using the molar mass of O_2 ($32.00 = 2 \times 16.00$).

$$5.00 \text{ mol O}_2 \times \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} = 160. \text{ g O}_2$$

Therefore, 160 g of oxygen is required to burn 44.1 g of propane. The procedure we have followed is summarized at the right. This is a convenient way to make sure the final units are correct.



We can summarize this problem by writing out a "conversion string" that shows how the problem was done.

$$44.1 \text{ g C}_3\text{H}_8 \times \frac{1 \text{ mol C}_3\text{H}_8}{44.09 \text{ g C}_3\text{H}_8} \times \frac{5 \text{ mol O}_2}{1 \text{ mol C}_3\text{H}_8} \times \frac{32.0 \text{ g O}_2}{1 \text{ mol O}_2} = 160. \text{ g O}_2$$

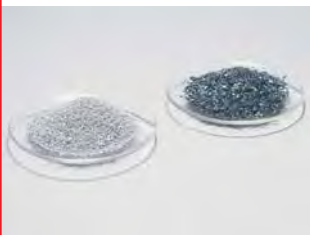
MATH



Use units as a check to see that you have used the correct conversion factors (mole ratios).

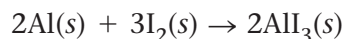
EXAMPLE 9.5

Mass Relationships in Chemical Reactions



Aluminum (left) and iodine (right) shown at the top, react vigorously to form aluminum iodide. The purple cloud results from excess iodine vaporized by the heat of the reaction.

Consider the reaction of powdered aluminum metal and finely ground iodine to produce aluminum iodide. The balanced equation for this vigorous chemical reaction is



Calculate the mass of $\text{I}_2(s)$ needed to just react with 35.0 g $\text{Al}(s)$.

Solution

Where do we want to go?

35.0 g $\text{Al}(s)$ requires ? g $\text{I}_2(s)$

What do we know?

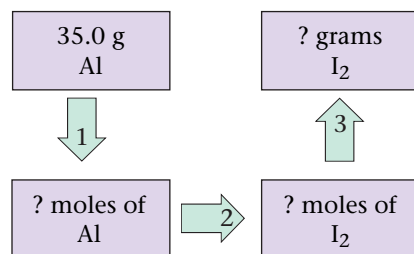
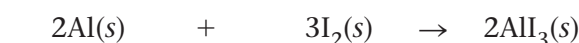
- $2\text{Al}(s) + 3\text{I}_2(s) \rightarrow 2\text{AlI}_3(s)$
- 35.0 g $\text{Al}(s)$

How do we get there?

Our plan

1. Convert 35.0 g Al to mol Al because the balanced equation deals in moles rather than grams.
2. Use the balanced equation to find the mol I_2 required to react with all of the Al.
3. Use the molar mass for iodine to calculate the mass of $\text{I}_2(s)$ required.

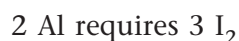
A schematic for this strategy is



- ➡ 1 Using the molar mass of aluminum (26.98 g), we will first calculate the moles of aluminum present in 35.0 g Al:

$$35.0 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.8 \text{ g Al}} = 1.30 \text{ mol Al}$$

- ➡ 2 From the balanced equation for the reaction, we see that



Thus, the required conversion factor from Al to I_2 is:

$$\frac{3 \text{ mol I}_2}{2 \text{ mol Al}}$$

Now we can determine the moles of I_2 required by 1.30 moles of Al:

$$1.30 \text{ mol Al} \times \frac{3 \text{ mol I}_2}{2 \text{ mol Al}} = 1.95 \text{ mol I}_2$$

3 We know the moles of I_2 required to react with the 1.30 mol of Al (35.0 g). Convert 1.95 mol of I_2 to grams so we will know how much to weigh out. We do so by using the molar mass of I_2 . The atomic mass of iodine is 126.9 g (for 1 mol of I atoms), so the molar mass of I_2 is

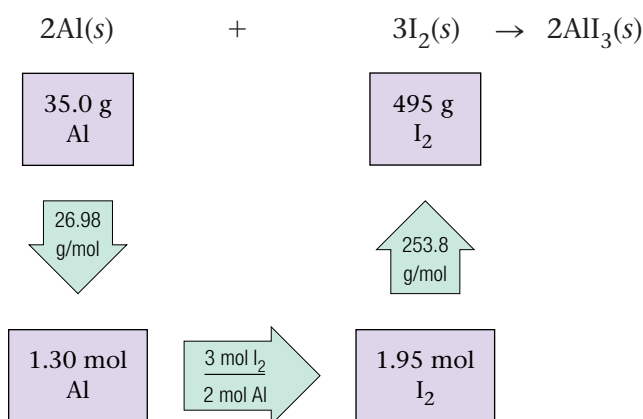
$$2 \times 126.9 \text{ g/mol} = 253.8 \text{ g/mol}$$

Now we convert the 1.95 mol of I_2 to grams of I_2 .

$$1.95 \text{ mol } I_2 \times \frac{253.8 \text{ g } I_2}{\text{mol } I_2} = 495 \text{ g } I_2$$

We have solved the problem. We need to weigh out 495 g of iodine (contains I_2 molecules) to react exactly with the 35.0 g of aluminum.

This problem can be summarized as follows:



Practice Problem • Exercise 9.5

Calculate the mass of $AlI_3(s)$ formed by the reaction of 35.0 g Al(s) with 495 g $I_2(s)$.

B. Mass Calculations Using Scientific Notation

So far in this chapter, we have spent considerable time “thinking through” the procedures for calculating the masses of reactants and products in chemical reactions.

Let's Review

Steps for Calculating the Masses of Reactants and Products in Chemical Reactions

Step 1 Balance the equation for the reaction.

Step 2 Convert the masses of reactants or products to moles.

Step 3 Use the balanced equation to set up the appropriate mole ratio(s).

Step 4 Use the mole ratio(s) to calculate the number of moles of the desired reactant or product.

Step 5 Convert from moles back to mass.

Stoichiometry

The process of using a balanced chemical equation to determine the relative masses of reactants and products involved in a reaction

The process of using a chemical equation to calculate the relative masses of reactants and products involved in a reaction is called **stoichiometry** (pronounced *stoy·kē·oim·etry*). Chemists say that the balanced equation for a chemical reaction describes the stoichiometry of the reaction.

EXAMPLE 9.6

Stoichiometric Calculations: Using Scientific Notation

Solid lithium hydroxide has been used in space vehicles to remove exhaled carbon dioxide from the living environment. The products are solid lithium carbonate and liquid water. What mass of gaseous carbon dioxide can 1.00×10^3 g of lithium hydroxide absorb?

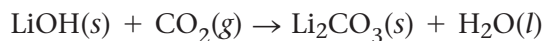
Solution

Where do we want to go?

1.00×10^3 g LiOH absorbs ?g CO₂

What do we know?

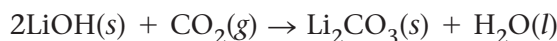
- We can write an equation from the description in the problem:



- 1.00×10^3 g LiOH

How do we get there?

Step 1 The balanced equation is



Step 2 We convert the given mass of LiOH to moles, using the molar mass of LiOH, which is $6.941 \text{ g} + 16.00 \text{ g} + 1.008 \text{ g} = 23.95 \text{ g}$.

$$1.00 \times 10^3 \text{ g LiOH} \times \frac{1 \text{ mol LiOH}}{23.95 \text{ g LiOH}} = 41.8 \text{ mol LiOH}$$

Step 3 The appropriate mole ratio is

$$\frac{1 \text{ mol CO}_2}{2 \text{ mol LiOH}}$$

Step 4 Using this mole ratio, we calculate the moles of CO₂ needed to react with the given mass of LiOH.

$$41.8 \text{ mol LiOH} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol LiOH}} = 20.9 \text{ mol CO}_2$$

Step 5 We calculate the mass of CO₂ by using its molar mass (44.01 g).

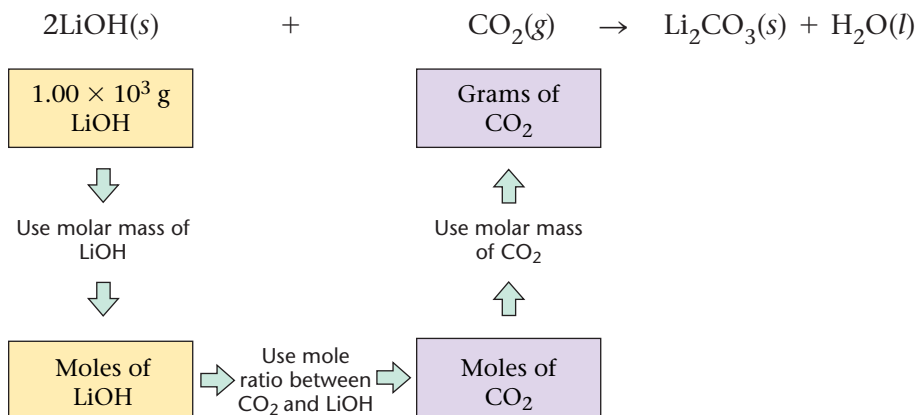
$$20.9 \text{ mol CO}_2 \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} = 920. \text{ g CO}_2 = 9.20 \times 10^2 \text{ g CO}_2$$

Thus, 1.00×10^3 g of LiOH(s) can absorb 920. g of CO₂(g).

Information

For a review of writing formulas of ionic compounds, see Chapter 4.

We can summarize this problem as follows:



The conversion string is

$$1.00 \times 10^3 \text{ g LiOH} \times \frac{1 \text{ mol LiOH}}{23.95 \text{ g LiOH}} \times \frac{1 \text{ mol CO}_2}{2 \text{ mol LiOH}} \times \frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2}$$

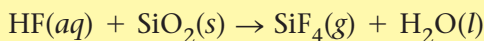
$$= 9.20 \times 10^2 \text{ g CO}_2$$

MATH

Carrying extra significant figures and rounding off only at the end gives an answer of 919 g CO₂.

Practice Problem • Exercise 9.6

Hydrofluoric acid, an aqueous solution containing dissolved hydrogen fluoride, is used to etch glass by reacting with the silica, SiO₂, in the glass to produce gaseous silicon tetrafluoride and liquid water. The unbalanced equation is



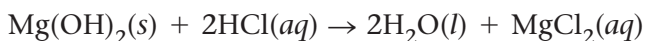
- Calculate the mass of hydrogen fluoride needed to react with 5.68 g of silica. *Hint:* Think carefully about this problem. What is the balanced equation for the reaction? What is given? What do you need to calculate? Sketch a map of the problem before you do the calculations.
- Calculate the mass of water produced in the reaction described in part a.

C. Mass Calculations: Comparing Two Reactions

Let's consider the relative effectiveness of two antacids to illustrate how chemical calculations can be important in daily life. Baking soda, NaHCO₃, is often used as an antacid. It neutralizes excess hydrochloric acid secreted by the stomach. The balanced equation for the reaction is



Milk of magnesia, which is an aqueous suspension of magnesium hydroxide, Mg(OH)₂, is also used as an antacid. The balanced equation for its reaction is



Which antacid can consume the most stomach acid, 1.00 g of NaHCO₃ or 1.00 g of Mg(OH)₂?

Where do we want to go?

The question we must ask for each antacid is, *How many moles of HCl will react with 1.00 g of each antacid?*



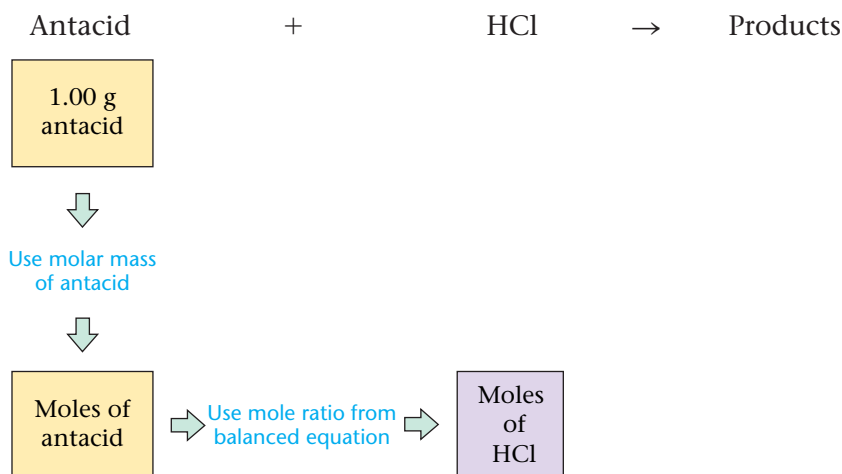
Two antacid tablets containing HCO₃⁻ dissolve to produce CO₂ gas.

What do we know?

- $\text{NaHCO}_3(s) + \text{HCl}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)$
- $\text{Mg}(\text{OH})_2(s) + 2 \text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + 2 \text{H}_2\text{O}(l)$
- 1.00 g $\text{Mg}(\text{OH})_2$
- 1.00 g NaHCO_3

How do we get there?

The antacid that reacts with the larger number of moles of HCl is more effective because it will neutralize more moles of acid. A schematic for this procedure is



Notice that in this case, we do not need to calculate how many grams of HCl react; we can answer the question with moles of HCl. We will now solve this problem for each antacid.

Step 1 Both of the equations are balanced, so we can proceed with the calculations.

For NaHCO_3 :

Step 2 Using the molar mass of NaHCO_3 , which is

$$22.99 \text{ g} + 1.008 \text{ g} + 12.01 \text{ g} + 3(16.00 \text{ g}) = 84.01 \text{ g},$$

we determine the moles of NaHCO_3 in 1.00 g of NaHCO_3 .

$$\begin{aligned} 1.00 \text{ g NaHCO}_3 &\times \frac{1 \text{ mol NaHCO}_3}{84.01 \text{ g NaHCO}_3} = 0.0119 \text{ mol NaHCO}_3 \\ &= 1.19 \times 10^{-2} \text{ mol NaHCO}_3 \end{aligned}$$

Step 3 Next, we determine the moles of HCl, using the mole ratio

$$\begin{aligned} &\frac{1 \text{ mol HCl}}{1 \text{ mol NaHCO}_3} \\ 1.19 \times 10^{-2} \text{ mol NaHCO}_3 &\times \frac{1 \text{ mol HCl}}{1 \text{ mol NaHCO}_3} \\ &= 1.19 \times 10^{-2} \text{ mol HCl} \end{aligned}$$

Step 4 Thus, 1.00 g of NaHCO_3 neutralizes 1.19×10^{-2} mol of HCl. We need to compare this to the number of moles of HCl that 1.00 g of $\text{Mg}(\text{OH})_2$ neutralizes.

For $\text{Mg}(\text{OH})_2$:

Step 2 Using the molar mass of $\text{Mg}(\text{OH})_2$, which is $24.31 \text{ g} + 2(16.00 \text{ g}) + 2(1.008 \text{ g}) = 58.33 \text{ g}$, we determine the moles of $\text{Mg}(\text{OH})_2$ in 1.00 g of $\text{Mg}(\text{OH})_2$.

$$1.00 \text{ g } \text{Mg}(\text{OH})_2 \times 1 \frac{\text{mol } \text{Mg}(\text{OH})_2}{58.33 \text{ g } \text{Mg}(\text{OH})_2} = 0.0171 \text{ mol } \text{Mg}(\text{OH})_2$$

$$= 1.71 \times 10^{-2} \text{ mol } \text{Mg}(\text{OH})_2$$

Step 3 To determine the moles of HCl that react with this amount of $\text{Mg}(\text{OH})_2$, we use the mole ratio $\frac{2 \text{ mol } \text{HCl}}{1 \text{ mol } \text{Mg}(\text{OH})_2}$.

$$1.71 \times 10^{-2} \text{ mol } \text{Mg}(\text{OH})_2 \times \frac{2 \text{ mol } \text{HCl}}{1 \text{ mol } \text{Mg}(\text{OH})_2}$$

$$= 3.42 \times 10^{-2} \text{ mol } \text{HCl}$$

Step 4 1.00 g of $\text{Mg}(\text{OH})_2$ neutralizes $3.42 \times 10^{-2} \text{ mol}$ of HCl . We have already calculated that 1.00 g of NaHCO_3 neutralizes only $1.19 \times 10^{-2} \text{ mol}$ of HCl . Therefore, $\text{Mg}(\text{OH})_2$ is a more effective antacid than NaHCO_3 on a mass basis.

Active Reading Question

In the problem above, we added equal masses (1.0 g) of each antacid. Why would one antacid consume more stomach acid than the other?



SECTION 9.2 REVIEW QUESTIONS

- Why do we need to convert mass to moles in stoichiometry problems?
- Solutions of sodium hydroxide cannot be kept for very long because they absorb carbon dioxide from the air, forming sodium carbonate. The unbalanced equation is

$$\text{NaOH}(aq) + \text{CO}_2(g) \rightarrow \text{Na}_2\text{CO}_3(aq) + \text{H}_2\text{O}(l)$$
 Calculate the number of grams of carbon dioxide that can be absorbed by complete reaction with a solution that contains 5.00 g of sodium hydroxide.
- Show how the steps in solving a stoichiometry problem are similar to a method that could be used in a supermarket to determine which of two products would be the best buy.
- You react 10.0 g of nitrogen gas with hydrogen gas according to the following reaction.

$$\text{N}_2(g) + 3\text{H}_2(g) \rightarrow 2\text{NH}_3(g)$$
 - What mass of hydrogen is required to completely react with 10.0 g sample of nitrogen gas?
 - What mass of ammonia is produced from 10.0 g of nitrogen gas and sufficient hydrogen gas?
- You react hydrogen gas with chlorine gas according to the following reaction.

$$\text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2\text{HCl}(g)$$
 What mass of $\text{HCl}(g)$ can be produced from $2.5 \times 10^3 \text{ g}$ of hydrogen gas with an excess of chlorine gas?
- Determine what mass of carbon monoxide and what mass of hydrogen are required to form 6.0 kg of methanol by the reaction

$$\text{CO}(g) + 2\text{H}_2(g) \rightarrow \text{CH}_3\text{OH}(l).$$

SECTION 9.3

Limiting Reactants and Percent Yield

Key Terms

- Limiting reactant (Limiting reagent)
- Theoretical yield
- Percent yield

Objectives

- To understand the concept of limiting reactants
- To learn to recognize the limiting reactant in a reaction
- To learn to use the limiting reactant to do stoichiometric calculations
- To learn to calculate percent yield

A. The Concept of Limiting Reactants

Earlier in this chapter, we discussed making sandwiches. Recall that the sandwich-making process could be described as follows:



In our earlier discussion, we always purchased the ingredients in the correct ratios so that we used all of the components, with nothing left over.

Now assume that you came to work one day and found the following quantities of ingredients:

20 slices of bread

24 slices of meat

12 slices of cheese

How many sandwiches can you make? What will be left over?

To solve this problem, let's see how many sandwiches we can make with each component.

Bread:

$$20 \text{ slices bread} \times \frac{1 \text{ sandwich}}{2 \text{ slices of bread}} = 10 \text{ sandwiches}$$

Meat:

$$24 \text{ slices meat} \times \frac{1 \text{ sandwich}}{3 \text{ slices of meat}} = 8 \text{ sandwiches}$$

Cheese:

$$12 \text{ slices cheese} \times \frac{1 \text{ sandwich}}{1 \text{ slice of cheese}} = 12 \text{ sandwiches}$$

How many sandwiches can you make? The answer is 8. When you run out of meat, you must stop making sandwiches. The meat is the limiting ingredient.

What do you have left over? Making 8 sandwiches requires 16 pieces of bread. You started with 20 pieces, so you have 4 pieces of bread left. You also used 8 pieces of cheese for the 8 sandwiches, so you have $12 - 8 = 4$ pieces of cheese left.

Active Reading Question

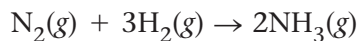
There are more slices of meat than of bread or cheese. So why does the meat limit the number of sandwiches you can make?

In this example, the ingredient present in the largest number (the meat) was actually the component that limited the number of sandwiches you could make. This situation arose because each sandwich required 3 slices of meat—more than the quantity required of any other ingredient.

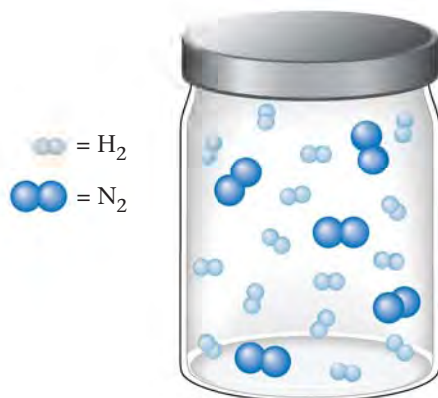


A Closer Look

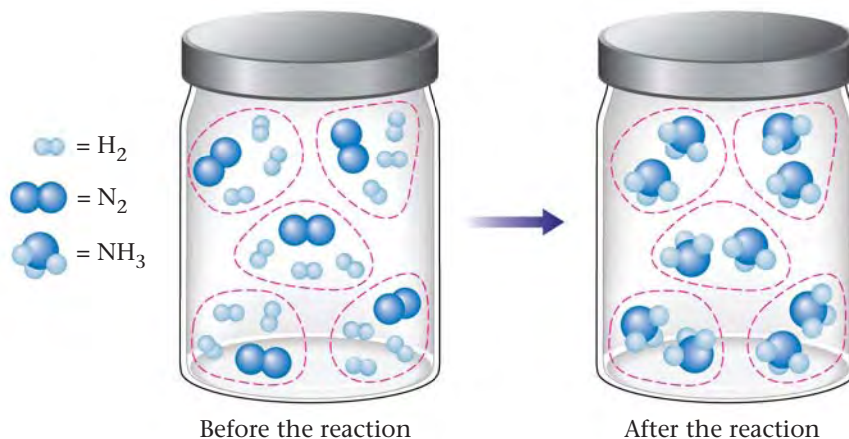
When molecules react with each other to form products, considerations very similar to those involved in making sandwiches arise. We can illustrate these ideas with the reaction of $\text{N}_2(\text{g})$ and $\text{H}_2(\text{g})$ to form $\text{NH}_3(\text{g})$:



Consider the following container of $\text{N}_2(\text{g})$ and $\text{H}_2(\text{g})$:



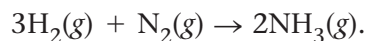
What will this container look like if the reaction between N_2 and H_2 proceeds to completion? To answer this question, you need to remember that each N_2 requires 3 H_2 molecules to form 2 NH_3 . To make things clear, we will circle groups of reactants:



In this case, the mixture of N_2 and H_2 contained just the number of molecules needed to form NH_3 with nothing left over. That is, the ratio of the number of H_2 molecules to N_2 molecules was

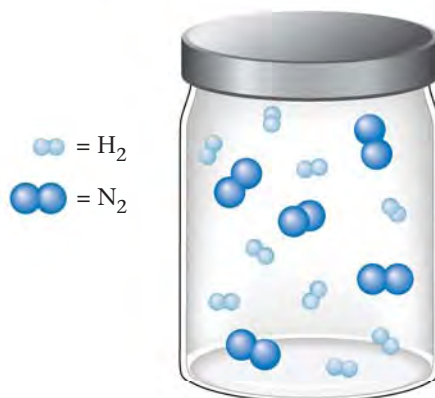
$$\frac{15 \text{ H}_2}{5 \text{ N}_2} = \frac{3 \text{ H}_2}{1 \text{ N}_2}$$

This ratio exactly matches the numbers in the balanced equation

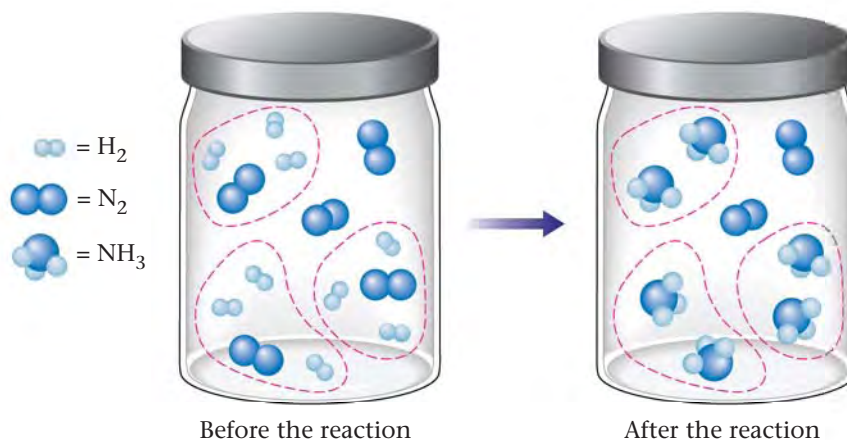


This type of mixture is called a *stoichiometric mixture*—one that contains the relative amounts of reactants that matches the numbers in the balanced equation. In this case all reactants will be consumed to form products.

Now consider another container of $\text{N}_2(\text{g})$ and $\text{H}_2(\text{g})$:



What will the container look like if the reaction between $\text{N}_2(\text{g})$ and $\text{H}_2(\text{g})$ proceeds to completion? Remember that each N_2 requires 3 H_2 . Circling groups of reactants we have



In this case, the hydrogen (H_2) is limiting. That is, the H_2 molecules are used up before all of the N_2 molecules are consumed. In this situation the amount of hydrogen limits the amount of product (ammonia) that can form—hydrogen is the limiting reactant. Some N_2 molecules are left over in this case because the reaction runs out of H_2 molecules first.

To determine how much product can be formed from a given mixture of reactants, we have to look for the reactant that is limiting—the one that runs out first and thus limits the amount of product that can form.

In some cases, the mixture of reactants might be stoichiometric—that is, all reactants run out at the same time. In general, however, you cannot assume that a given mixture of reactants is a stoichiometric mixture, so you must determine whether one of the reactants is limiting.

The reactant that runs out first and thus limits the amounts of products that can form is called the **limiting reactant (limiting reagent)**.

To this point, we have considered examples where the numbers of reactant molecules could be counted. In “real life” you can’t count the molecules directly—you can’t see them and, even if you could, there would be far too many to count. Instead, you must count by weighing. We must therefore explore how to find the limiting reactant, given the masses of the reactants.

Limiting reactant (limiting reagent)

The reactant that is completely used up when a reaction is run to completion

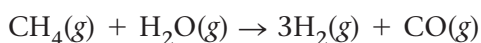
Active Reading Question

What is meant by the term “limiting reactant”?

B. Calculations Involving a Limiting Reactant

Manufacturers of cars, bicycles, and appliances order parts in the same proportion as they are used in their products. For example, auto manufacturers order four times as many wheels as engines and bicycle manufacturers order twice as many pedals as seats. Likewise, when chemicals are mixed together so that they can undergo a reaction, they are often mixed in stoichiometric quantities—that is, in exactly the correct amounts so that all reactants “run out” (are used up) at the same time.

Let’s consider the production of hydrogen for use in the manufacture of ammonia. Ammonia, a very important fertilizer itself and a starting material for other fertilizers, is made by combining nitrogen from the air with hydrogen. The hydrogen for this process is produced by the reaction of methane with water according to the balanced equation



Let’s consider the question, *What mass of water is required to react exactly with 249 g of methane?*

Where do we want to go?

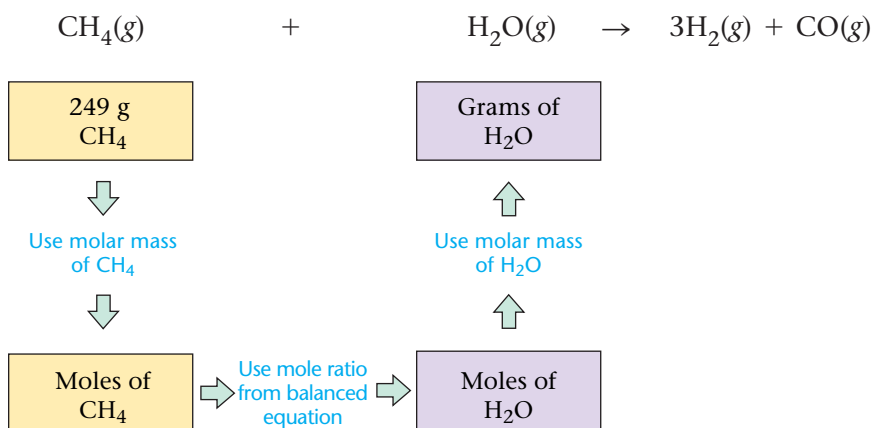
How much water will just use up all of the 249 g of methane, leaving no water or methane remaining?

What do we know?

- $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow 3\text{H}_2(\text{g}) + \text{CO}(\text{g})$
- 249 g CH_4

How do we get there?

Drawing a map of the problem is helpful.



Ammonia being dissolved in irrigation water to provide fertilizer for a field of corn

Step 1 The equation is already balanced.

Step 2 We first convert the mass of CH_4 to moles, using the molar mass of CH_4 (16.04 g/mol).

$$249 \text{ g } \cancel{\text{CH}_4} \times \frac{1 \text{ mol } \text{CH}_4}{16.04 \text{ g } \cancel{\text{CH}_4}} = 15.5 \text{ mol } \text{CH}_4$$

Step 3 Because in the balanced equation 1 mol of CH_4 reacts with 1 mol of H_2O , we have

$$15.5 \text{ mol } \cancel{\text{CH}_4} \times \frac{1 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \cancel{\text{CH}_4}} = 15.5 \text{ mol } \text{H}_2\text{O}$$

Step 4 Therefore, 15.5 mol of H_2O will react exactly with the given mass of CH_4 . Converting 15.5 mol of H_2O to grams of H_2O (molar mass = 18.02 g/mol) gives

$$15.5 \text{ mol } \cancel{\text{H}_2\text{O}} \times \frac{18.02 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \cancel{\text{H}_2\text{O}}} = 279 \text{ g } \text{H}_2\text{O}$$

This result means that if 249 g of methane is mixed with 279 g of water, both reactants will “run out” at the same time. The reactants have been mixed in stoichiometric quantities.

information

The reactant that is consumed first limits the amounts of products that can form.

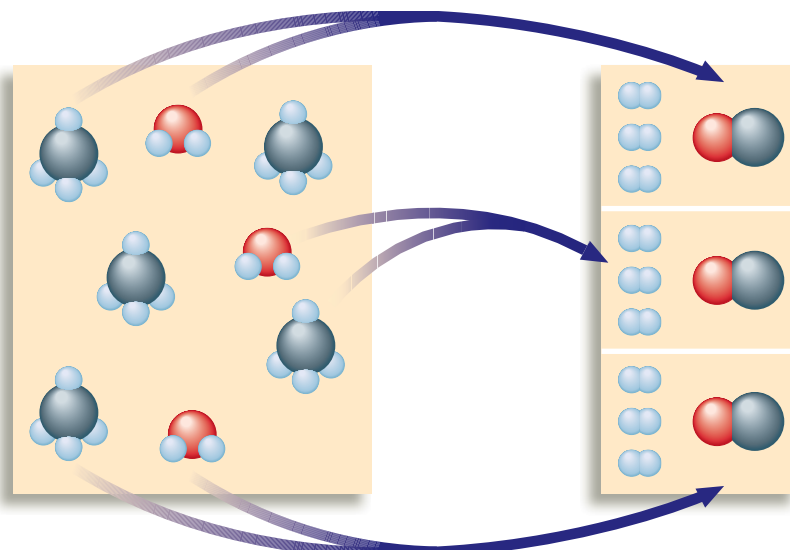
If, on the other hand, 249 g of methane is mixed with 300 g of water, the methane will be consumed before the water runs out. The water will be in *excess*. In this case, the quantity of products formed will be determined by the quantity of methane present. After the methane is consumed, no more products can be formed, even though some water still remains. In this situation, the amount of methane *limits* the amount of products that can be formed, and it is the limiting reactant.

In any stoichiometry problem, where reactants are not mixed in stoichiometric quantities, it is essential to determine which reactant is limiting to calculate correctly the amounts of products that will be formed. This concept is illustrated in **Figure 9.1**. Note from this figure that because there are fewer water molecules than CH_4 molecules, the water is consumed first. After the water molecules are gone, no more products can form. So in this case water is the limiting reactant.

Figure 9.1

A mixture of 5 CH_4 and 3 H_2O molecules undergoes the reaction $\text{CH}_4(g) + \text{H}_2\text{O}(g) \rightarrow 3\text{H}_2(g) + \text{CO}(g)$.

Note that the H_2O molecules are used up first, leaving two CH_4 molecules unreacted.



EXAMPLE 9.7**Stoichiometric Calculations: Identifying the Limiting Reactant**

Suppose that 25.0 kg (2.50×10^4 g) of nitrogen gas and 5.00 kg (5.00×10^3 g) of hydrogen gas are mixed and reacted to form ammonia. Calculate the mass of ammonia produced when this reaction is run to completion.

Solution**Where do we want to go?**

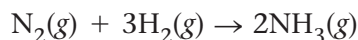
How much ammonia will be produced when the reaction is complete?

What do we know?

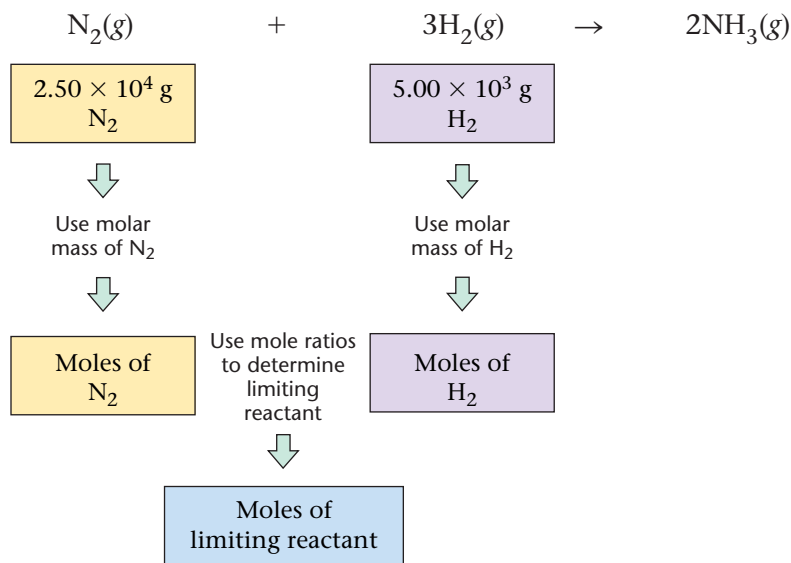
- $\text{N}_2(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{NH}_3(\text{g})$
- 2.5×10^4 g N_2
- 5.00×10^3 g H_2

How do we get there?

Step 1 The balanced equation for this reaction is



This problem is different from the others we have done so far in that we are mixing *specified amounts of two reactants*. To know how much product forms, we must determine which reactant is consumed first. That is, we must determine which is the limiting reactant in this experiment. To do so, we must add a step to our normal procedure. We can map this process as follows:



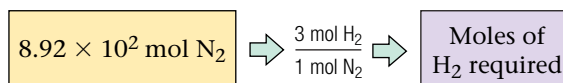
We will use the moles of the limiting reactant to calculate the moles and then the grams of the product.

Step 2 We first calculate the moles of the two reactants present:

$$2.50 \times 10^4 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} = 8.92 \times 10^2 \text{ mol N}_2$$

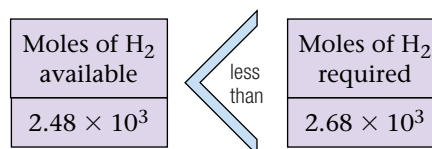
$$5.00 \times 10^3 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} = 2.48 \times 10^3 \text{ mol H}_2$$

Step 3 Now we must determine which reactant is limiting (will be consumed first). We have 8.92×10^2 mol of N_2 . Let's determine *how many moles of H_2 are required to react with this much N_2* . Because 1 mol of N_2 reacts with 3 mol of H_2 , the number of moles of H_2 we need to react completely with 8.92×10^2 mol of N_2 is determined as follows:



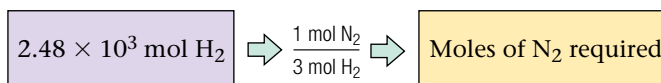
$$8.92 \times 10^2 \text{ mol N}_2 \times \frac{3 \text{ mol H}_2}{1 \text{ mol N}_2} = 2.68 \times 10^3 \text{ mol H}_2$$

Is N_2 or H_2 the limiting reactant? The answer comes from the comparison



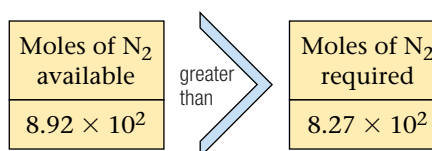
We see that 8.92×10^2 mol of N_2 requires 2.68×10^3 mol of H_2 to react completely. However, only 2.48×10^3 mol of H_2 is present. This means that the hydrogen will be consumed before the nitrogen runs out, so hydrogen is the *limiting reactant* in this particular situation.

Note that in our effort to determine the limiting reactant, we could have started instead with the given amount of hydrogen and calculated the moles of nitrogen required.



$$2.48 \times 10^3 \text{ mol H}_2 \times \frac{1 \text{ mol N}_2}{3 \text{ mol H}_2} = 8.27 \times 10^2 \text{ mol N}_2$$

Thus, 2.48×10^3 mol of H_2 requires 8.27×10^2 mol of N_2 . Because 8.92×10^2 mol of N_2 is actually present, the nitrogen is in excess.



If nitrogen is in excess, hydrogen will “run out” first; again we find that hydrogen limits the amount of ammonia formed.

H₂ is limiting reactant.

N₂ is in excess.

Information

Always check to see which, if any, reactant is limiting when you are given the amounts of two or more reactants.

Step 4 Because the moles of H₂ present are limiting, we must use this quantity to determine the moles of NH₃ that can form.

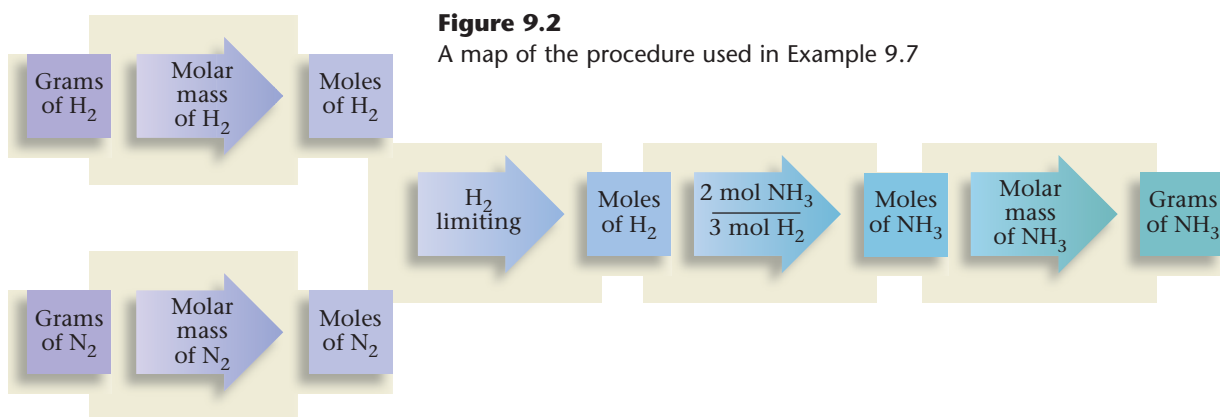
$$2.48 \times 10^3 \text{ mol H}_2 \times \frac{2 \text{ mol NH}_3}{3 \text{ mol H}_2} = 1.65 \times 10^3 \text{ mol NH}_3$$

Step 5 Next, we convert moles of NH₃ to mass of NH₃.

$$1.65 \times 10^3 \text{ mol NH}_3 \times \frac{17.03 \text{ g NH}_3}{1 \text{ mol NH}_3} = 2.81 \times 10^4 \text{ g NH}_3 \\ = 28.1 \text{ kg NH}_3$$

Therefore, 25.0 kg of N₂ and 5.00 kg of H₂ can form 28.1 kg of NH₃.

The strategy used in Example 9.7 is summarized in **Figure 9.2**.



Let's Review

Steps for Solving Stoichiometry Problems Involving Limiting Reactants

- Step 1** Write and balance the equation for the reaction.
- Step 2** Convert known masses of reactants to moles.
- Step 3** Using the numbers of moles of reactants and the appropriate mole ratios, determine which reactant is limiting.
- Step 4** Using the amount of the limiting reactant and the appropriate mole ratios, compute the number of moles of the desired product.
- Step 5** Convert from moles of product to grams of product, using the molar mass (if this is required by the problem).

EXAMPLE 9.8

Stoichiometric Calculations: Reactions Involving the Masses of Two Reactants



Copper(II) oxide reacting with ammonia in a heated tube

Nitrogen gas can be prepared by passing gaseous ammonia over solid copper(II) oxide at high temperatures. The other products of the reaction are solid copper and water vapor. How many grams of N_2 are formed when 18.1 g of NH_3 is reacted with 90.4 g of CuO ?

Solution

Where do we want to go?

How much nitrogen will be produced when the reaction is complete?

What do we know?

- $NH_3(g) + CuO(s) \rightarrow N_2(g) + Cu(s) + H_2O(g)$
- 18.1 g NH_3
- 90.4 g CuO

How do we get there?

Step 1 Balance the equation:



Step 2 From the masses of reactants available, we must compute the moles of NH_3 (molar mass = 17.03 g) and of CuO (molar mass = 79.55 g).

$$18.1 \text{ g } \cancel{NH_3} \times \frac{1 \text{ mol } NH_3}{17.03 \text{ g } \cancel{NH_3}} = 1.06 \text{ mol } NH_3$$

$$90.4 \text{ g } \cancel{CuO} \times \frac{1 \text{ mol } CuO}{79.55 \text{ g } \cancel{CuO}} = 1.14 \text{ mol } CuO$$

Step 3 To determine which reactant is limiting, we use the mole ratio between CuO and NH_3 to determine how CuO is required.

$$1.06 \text{ mol } \cancel{NH_3} \times \frac{3 \text{ mol } CuO}{2 \text{ mol } \cancel{NH_3}} = 1.59 \text{ mol } CuO$$

Then we compare how much CuO we have with how much of it we need.

Moles of CuO available	\leftarrow less than \rightarrow	Moles of CuO needed to react with all the NH_3
1.14		1.59

Therefore, 1.59 mol of CuO is required to react with 1.06 mol of NH_3 , but only 1.14 mol of CuO is actually present. So the amount of CuO is limiting; CuO will run out before NH_3 does.

Note that CuO is limiting even though the original mass of CuO was much greater than the original mass of NH_3 .

Step 4 CuO is the limiting reactant, so we must use the amount of CuO in calculating the amount of N₂ formed. Using the mole ratio between CuO and N₂ from the balanced equation, we have

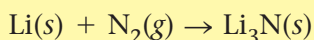
$$1.14 \text{ mol CuO} \times \frac{1 \text{ mol N}_2}{3 \text{ mol CuO}} = 0.380 \text{ mol N}_2$$

Step 5 Using the molar mass of N₂ (28.02), we can now calculate the mass of N₂ produced.

$$0.380 \text{ mol N}_2 \times \frac{28.02 \text{ g N}_2}{1 \text{ mol N}_2} = 10.6 \text{ g N}_2$$

Practice Problem • Exercise 9.8

Lithium nitride, an ionic compound containing the Li⁺ and N³⁻ ions, is prepared by the reaction of lithium metal and nitrogen gas. Calculate the mass of lithium nitride formed from 56.0 g of nitrogen gas and 56.0 g of lithium in the unbalanced reaction:



Active Reading Question

In the example, how can CuO be limiting if we started with a greater mass of CuO than NH₃?

C. Percent Yield

Previously, we calculated the amount of product formed when specified amounts of reactants were mixed together. In doing these calculations, we used the fact that the amount of product is controlled by the limiting reactant. Products stop forming when one reactant runs out.

The amount of product calculated in this way is called the **theoretical yield** of that product. It is the amount of product predicted from the amounts of reactants used. For instance, in Example 9.8, 10.6 g of nitrogen represents the theoretical yield. This is the *maximum amount* of nitrogen that can be produced from the quantities of reactants used. Actually, however, the amount of product predicted (the theoretical yield) is seldom obtained. One reason for this is the presence of side reactions (other reactions that consume one or more of the reactants or products).

The *actual yield* of product, which is the amount of product *actually obtained*, is often compared to the theoretical yield. This comparison, usually expressed as a percent, is called the **percent yield**.

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \text{percent yield}$$

For example, *if* the reaction considered in Example 9.8 *actually* gave 6.63 g of nitrogen instead of the *predicted* 10.6 g, the percent yield of nitrogen would be

$$\frac{6.63 \text{ g N}_2}{10.6 \text{ g N}_2} \times 100\% = 62.5\%$$

Theoretical yield

The maximum amount of a given product that can be formed when the limiting reactant is completely consumed

Percent yield

The actual yield of a product as the percentage of the theoretical yield

information

Percent yield is important as an indicator of the efficiency of a particular reaction.

Active Reading Question

Which do we calculate by using stoichiometry, the actual or the theoretical yield?

EXAMPLE 9.9

Stoichiometric Calculations: Determining Percent Yield

Methanol can be produced by the reaction between carbon monoxide and hydrogen. Let's consider this process again. Suppose 68.5 kg (6.85×10^4 g) of CO(g) is reacted with 8.60 kg (8.60×10^3 g) of H₂(g).

MATH



$$\text{Percent} = \frac{\text{Part}}{\text{Whole}} \times 100\%$$

- Calculate the theoretical yield of methanol.
- If 3.57×10^4 g of CH₃OH is actually produced, what is the percent yield of methanol?

Solution

Where do we want to go?

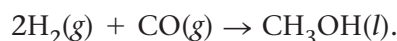
- theoretical yield of CH₃OH
- percent yield of CH₃OH

What do we know?

- CO(g) + H₂(g) → CH₃OH(l)
- 8.60×10^3 g H₂
- 6.85×10^4 g CO
- 3.57×10^4 g CH₃OH produced

How do we get there?

Step 1 The balanced equation is



Step 2 Next, we calculate the moles of reactants.

$$6.85 \times 10^4 \text{ g CO} \times \frac{1 \text{ mol CO}}{28.01 \text{ g CO}} = 2.45 \times 10^3 \text{ mol CO}$$

$$8.60 \times 10^3 \text{ g H}_2 \times \frac{1 \text{ mol H}_2}{2.016 \text{ g H}_2} = 4.27 \times 10^3 \text{ mol H}_2$$

Step 3 Now we determine which reactant is limiting. Using the mole ratio between CO and H₂ from the balanced equation, we have

$$2.45 \times 10^3 \text{ mol CO} \times \frac{2 \text{ mol H}_2}{1 \text{ mol CO}} = 4.90 \times 10^3 \text{ mol H}_2$$

Moles of H ₂ present	less than	Moles of H ₂ needed to react with all the CO
4.27×10^3		4.90×10^3

We see that 2.45×10^3 mol of CO requires 4.90×10^3 mol of H_2 . Because only 4.27×10^3 mol of H_2 is actually present, H_2 is limiting.

Step 4 We must therefore use the amount of H_2 and the mole ratio between H_2 and CH_3OH to determine the maximum amount of methanol that can be produced in the reaction.

$$4.27 \times 10^3 \text{ mol } H_2 \times \frac{1 \text{ mol } CH_3OH}{2 \text{ mol } H_2} = 2.14 \times 10^3 \text{ mol } CH_3$$

This represents the theoretical yield in moles.

Step 5 Using the molar mass of CH_3OH (32.04 g), we can calculate the theoretical yield in grams.

$$2.14 \times 10^3 \text{ mol } CH_3OH \times \frac{32.04 \text{ g } CH_3OH}{1 \text{ mol } CH_3OH} = 6.86 \times 10^4 \text{ g } CH_3OH$$

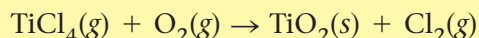
So, from the amounts of reactants given, the maximum amount of CH_3OH that can be formed is 6.85×10^4 g. This is the *theoretical yield*.

The percent yield is

$$\begin{aligned} & \frac{\text{Actual yield (grams)}}{\text{Theoretical yield (grams)}} \times 100\% \\ &= \frac{3.57 \times 10^4 \text{ g } CH_3OH}{6.86 \times 10^4 \text{ g } CH_3OH} \times 100\% \\ &= 52.0\% \end{aligned}$$

Practice Problem • Exercise 9.9

Titanium(IV) oxide is a white compound used as a coloring pigment. In fact, the page you are now reading is white because of the presence of this compound in the paper. Solid titanium(IV) oxide can be prepared by reacting gaseous titanium(IV) chloride with oxygen gas. A second product of this reaction is chlorine gas.



- Suppose that 6.71×10^3 g of titanium(IV) chloride is reacted with 2.45×10^3 g of oxygen. Calculate the maximum mass of titanium(IV) oxide that can form.
- If the percent yield of TiO_2 is 75%, what mass is actually formed?

CHEMISTRY IN YOUR WORLD

Consumer Connection

The “Golden” Dollar

The dollar bill is very popular with U.S. consumers, but it has some problems—it has an average life of only 18 months and can be very frustrating to use in vending machines.



To solve these problems a dollar coin—the Susan B. Anthony dollar—was minted in 1979 and 1980. This coin never became popular with consumers (although 900 million are now in circulation) because it too closely resembled a quarter.



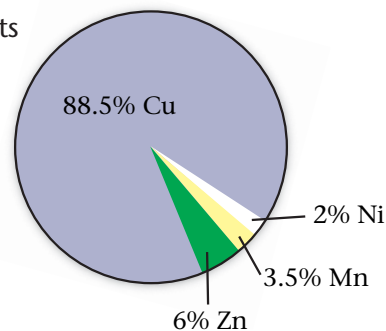
In the late 1990s, Congress decided to try again. The United States Dollar Coin Act of 1997 called for a new dollar coin to be golden in color, have a distinctive edge, and be the same size as the Anthony dollar. The idea was to mint a coin that looks and feels different from the quarter but that is enough like the Susan B. Anthony dollar to work in vending machines.

The new dollar coin, released in 2000, depicts Sacagawea, the young Native American (Shoshone) woman who guided the Lewis and Clark expedition, and her son, Jean Baptiste, on the front and a flying eagle on the back. The coin has the same color and luster as 14-carat gold, although it contains none of that precious metal.



The Sacagawea dollar is 88.5% copper, 6.0% zinc, 3.5% manganese, and 2% nickel. The copper core accounts for half of the coin's total thickness, with manganese brass, the surrounding golden-colored alloy of copper, zinc, manganese, and nickel making up the rest.

The coin costs 12 cents to manufacture, so the U.S. government makes 88 cents of every dollar minted. Such a deal!



SECTION 9.3 REVIEW QUESTIONS

- How do you recognize a stoichiometric mixture in a chemical reaction?
- Suppose that a lemonade recipe calls for 1 cup of sugar for every 6 lemons. You have 12 lemons and 3 cups of sugar. Which ingredient is limiting? Why?
- How is a limiting reactant problem different from other stoichiometry problems? (What is your clue that it is a limiting reactant problem?)
- What is the difference between theoretical yield and the actual yield? Why is there a difference?
- You react natural gas (methane, CH_4) with chlorine gas to produce carbon tetrachloride and hydrogen gas according to the following reaction:
$$\text{CH}_4(g) + 2\text{Cl}_2(g) \rightarrow \text{CCl}_4(l) + 2\text{H}_2(g)$$
Consider reacting 25.0 g of each reactant and answer the following.
 - Which reactant is limiting?
 - What mass of carbon tetrachloride should be produced (assuming 100% yield)?
 - Determine the mass of leftover reactant.

Key Terms

9.1 Mole ratio

9.2 Stoichiometry

9.3 Limiting reactant
Theoretical yield
Percent yield

Key Ideas

9.1 Using Chemical Equations

- A balanced chemical equation gives relative numbers (or moles) of reactant and product molecules that participate in a chemical reaction.
- Stoichiometric calculations involve using a balanced chemical equation to determine the amounts of reactants needed or products formed in a reaction.
- To convert between moles of reactants and moles of products, we use mole ratios derived from the balanced chemical equation.

9.2 Using Chemical Equations to Calculate Mass

- To calculate masses from the moles of reactants needed or products formed, we can use the molar masses of substances for finding the masses (g) needed or formed.

9.3 Limiting Reactants and Percent Yield

- Often, reactants in a chemical reaction are not present in stoichiometric quantities (they do not “run out” at the same time).
 - In this case, we must determine which reactant runs out first and thus limits the amount of products that can form—this is called the limiting reactant.
- The actual yield (amount produced) of a reaction is usually less than the maximum expected (theoretical yield).
- The actual yield is often expressed as a percentage of the theoretical yield

$$\text{percent yield} = \frac{\text{actual yield (g)}}{\text{theoretical yield (g)}} \times 100\%$$

All exercises with **blue numbers** have answers in the back of this book.

9.1 Using Chemical Equations

A. Information Given by Chemical Equations

- Although *mass* is a property of matter we can conveniently measure in the laboratory, the coefficients of a balanced chemical equation are *not* directly interpreted on the basis of mass. Explain why.
- Explain why, in the balanced chemical equation $C + O_2 \rightarrow CO_2$, we know that 1 g of C will *not* react exactly with 1 g of O_2 .
- For each of the following reactions, give the balanced equation for the reaction and state the meaning of the equation in terms of numbers of *individual molecules* and in terms of *moles of molecules*.
 - $NO(g) + O_2(g) \rightarrow NO_2(g)$
 - $AgC_2H_3O_2(aq) + CuSO_4(aq) \rightarrow Ag_2SO_4(s) + Cu(C_2H_3O_2)_2(aq)$
 - $PCl_3(l) + H_2O(l) \rightarrow H_3PO_3(l) + HCl(g)$
 - $C_2H_6(g) + Cl_2(g) \rightarrow C_2H_5Cl(g) + HCl(g)$
- For each of the following reactions, give the balanced chemical equation for the reaction and state the meaning of the equation in terms of *individual molecules* and in terms of *moles of molecules*.
 - $MnO_2(s) + Al(s) \rightarrow Mn(s) + Al_2O_3(s)$
 - $B_2O_3(s) + CaF_2(s) \rightarrow BF_3(g) + CaO(s)$
 - $NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + NO(g)$
 - $C_6H_6(l) + H_2(g) \rightarrow C_6H_{12}(l)$

B. Mole-mole Relationships

- True or false? For the reaction represented by the chemical equation

$$2H_2O_2(aq) \rightarrow 2H_2O(l) + O_2(g)$$
 if 2.0 g of hydrogen peroxide decomposes, then 2.0 g of water and 1.0 g of oxygen gas will be produced.
- Consider the balanced equation

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$
 What is the mole ratio that would enable you to calculate the number of moles of oxygen needed to react exactly with a given number of moles of $CH_4(g)$? What mole ratios would you use to calculate how many moles of each product form from a given number of moles of CH_4 ?

- Consider the unbalanced chemical equation



Balance the equation. Identify the mole ratios that you would use to calculate the number of moles of each product that would form for a given number of moles of silver reacting.

- For each of the following *unbalanced* chemical equations, calculate how many *moles of each product* would be produced by the complete conversion of 0.125 mol of the reactant indicated in boldface. State clearly the mole ratio used for the conversion.
 - $FeO(s) + C(s) \rightarrow Fe(l) + CO_2(g)$
 - $Cl_2(g) + KI(aq) \rightarrow KCl(aq) + I_2(s)$
 - $Na_2B_4O_7(s) + H_2SO_4(aq) + H_2O(l) \rightarrow H_3BO_3(s) + Na_2SO_4(aq)$
 - $CaC_2(s) + H_2O(l) \rightarrow Ca(OH)_2(s) + C_2H_2(g)$
- For each of the following balanced chemical equations, calculate how many *moles* and how many *grams* of each product would be produced by the complete conversion of 0.50 mol of the reactant indicated in boldface. State clearly the mole ratio used for each conversion.
 - $NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$
 - $CH_4(g) + 4S(s) \rightarrow CS_2(l) + 2H_2S(g)$
 - $PCl_3(l) + 3H_2O(l) \rightarrow H_3PO_3(aq) + 3HCl(aq)$
 - $NaOH(s) + CO_2(g) \rightarrow NaHCO_3(s)$
- For each of the following *unbalanced* equations, indicate how many *moles* of the *first product* are produced if 0.625 mol of the *second product* forms. State clearly the mole ratio used for each conversion.
 - $KO_2(s) + H_2O(l) \rightarrow O_2(g) + KOH(s)$
 - $SeO_2(g) + H_2Se(g) \rightarrow Se(s) + H_2O(g)$
 - $CH_3CH_2OH(l) + O_2(g) \rightarrow CH_3CHO(aq) + H_2O(l)$
 - $Fe_2O_3(s) + Al(s) \rightarrow Fe(l) + Al_2O_3(s)$

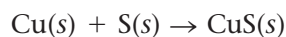
9.2 Using Chemical Equations to Calculate Mass

A. Mass Calculations

- What quantity serves as the conversion factor between the mass of a sample and the number of moles the sample contains?
- What does it mean to say that the balanced chemical equation for a reaction describes the *stoichiometry* of the reaction?

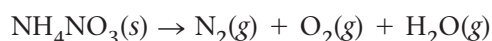
- 13.** Using the average atomic masses given inside the front cover of the text, calculate the number of *moles* of each substance contained in the following *masses*.
- 2.01×10^{-2} g of silver
 - 45.2 mg of ammonium sulfide
 - 61.7 μg of uranium
 - 5.23 kg of sulfur dioxide
 - 272 g of iron(III) nitrate
 - 12.7 mg of iron(II) sulfate
 - 6.91×10^3 g of lithium hydroxide
- 14.** Using the average atomic masses given inside the front cover of the text, calculate the *mass in grams* of each of the following samples.
- 2.21×10^{-4} mol of calcium carbonate
 - 2.75 mol of helium
 - 0.00975 mol of oxygen gas
 - 7.21×10^{-3} mol of carbon dioxide
 - 0.835 mol of iron(II) sulfide
 - 4.01 mol of potassium hydroxide
 - 0.0219 mol of hydrogen gas
- 15.** For each of the following *unbalanced* equations, calculate how many moles of the *second* reactant would be required to react completely with exactly 25.0 g of the *first* reactant. Indicate clearly the mole ratio used for each conversion.
- $\text{Mg}(s) + \text{CuCl}_2(aq) \rightarrow \text{MgCl}_2(aq) + \text{Cu}(s)$
 - $\text{AgNO}_3(aq) + \text{NiCl}_2(aq) \rightarrow \text{AgCl}(s) + \text{Ni}(\text{NO}_3)_2(aq)$
 - $\text{NaHSO}_3(aq) + \text{NaOH}(aq) \rightarrow \text{Na}_2\text{SO}_3(aq) + \text{H}_2\text{O}(l)$
 - $\text{KHCO}_3(aq) + \text{HCl}(aq) \rightarrow \text{KCl}(aq) + \text{H}_2\text{O}(l) + \text{CO}_2(g)$
- 16.** For each of the following *unbalanced* equations, calculate how many *milligrams of each product* would be produced by complete reaction of 10.0 mg of the reactant indicated in boldface. Indicate clearly the mole ratio used for the conversion.
- $\text{FeSO}_4(aq) + \text{K}_2\text{CO}_3(aq) \rightarrow \text{FeCO}_3(s) + \text{K}_2\text{SO}_4(aq)$
 - $\text{Cr}(s) + \text{SnCl}_4(l) \rightarrow \text{CrCl}_3(s) + \text{Sn}(s)$
 - $\text{Fe}(s) + \text{S}_8(s) \rightarrow \text{Fe}_2\text{S}_3(s)$
 - $\text{Ag}(s) + \text{HNO}_3(aq) \rightarrow \text{AgNO}_3(aq) + \text{H}_2\text{O}(l) + \text{NO}(g)$
- 17.** Although mixtures of hydrogen and oxygen are highly explosive, pure elemental hydrogen gas itself burns quietly in air with a pale blue flame, producing water vapor.
- $$2\text{H}_2(g) + \text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(g)$$
- Calculate the mass (in grams) of water vapor produced when 56.0 g of pure hydrogen gas burns in air.
- 18.** Given the information in Problem 17, calculate the mass of oxygen gas that would be necessary to burn 0.0275 mol of hydrogen gas.
- 19.** When elemental carbon is burned in the open atmosphere, with plenty of oxygen gas present, the product is carbon dioxide.
- $$\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g)$$
- However, when the amount of oxygen present during the burning of the carbon is restricted, carbon monoxide is more likely to result.
- $$2\text{C}(s) + \text{O}_2(g) \rightarrow 2\text{CO}(g)$$
- What mass of each product is expected when a 5.00-g sample of pure carbon is burned under each of these conditions?
- 20.** Although we usually think of substances as “burning” only in oxygen gas, the process of rapid oxidation to produce a flame may also take place in other strongly oxidizing gases. For example, when iron is heated and placed in pure chlorine gas, the iron “burns” according to the following (unbalanced) reaction:
- $$\text{Fe}(s) + \text{Cl}_2(g) \rightarrow \text{FeCl}_3(s)$$
- How many milligrams of iron(III) chloride result when 15.5 mg of iron is reacted with an excess of chlorine gas?
- 21.** Small quantities of oxygen gas can be generated in the laboratory by the decomposition of hydrogen peroxide. The unbalanced equation for the reaction is
- $$\text{H}_2\text{O}_2(aq) \rightarrow \text{H}_2\text{O}(l) + \text{O}_2(g)$$
- Calculate the mass of oxygen produced when 10.00 g of hydrogen peroxide decomposes.

22. An introductory chemistry experiment involves heating finely divided copper metal with sulfur to determine the proportions in which the elements react to form copper(II) sulfide. The experiment works well because any excess sulfur beyond that required to react with the copper may be simply boiled away from the reaction container.



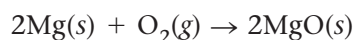
If 1.25 g of copper is heated with an excess of sulfur, how many grams of sulfur will react?

23. Ammonium nitrate has been used as a high explosive because it is unstable and decomposes into several gaseous substances. The rapid expansion of the gaseous substances produces the explosive force.



Calculate the mass of each product gas if 1.25 g of ammonium nitrate reacts.

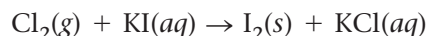
24. Magnesium metal burns in oxygen with an intensely bright white flame. The balanced equation for this reaction is



How many grams of MgO(s) are produced by complete reaction of 1.25 g of magnesium metal?

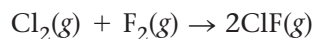
B. Mass Calculations Using Scientific Notation

25. If chlorine gas is bubbled through a potassium iodide solution, elemental iodine is produced. The *unbalanced* equation is



Calculate the mass of iodine produced when 4.50×10^3 g of chlorine gas is bubbled through an excess amount of potassium iodide solution.

26. Elemental fluorine and chlorine gases are very reactive. For example, they react with each other to form chlorine monofluoride.



Calculate the mass of chlorine gas required to produce 5.00×10^{-3} g of chlorine monofluoride given an excess of fluorine gas.

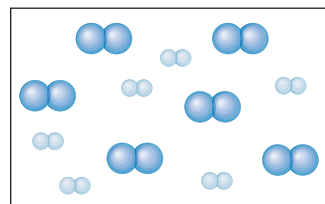
C. Mass Calculations: Comparing Two Reactions

27. Both propane (C_3H_8) and butane (C_4H_{10}) react with oxygen gas to form carbon dioxide and water. If you have equal masses of each, which will require a greater mass of oxygen to react?
28. Methane (CH_4) reacts with oxygen in the air to produce carbon dioxide and water. Ammonia (NH_3) reacts with oxygen in the air to produce nitrogen monoxide and water. What mass of ammonia is needed to react with excess oxygen to produce the same amount of water as 1.00 g of methane reacting with excess oxygen?

9.3 Limiting Reactants and Percent Yield

A. The Concept of Limiting Reactants

29. What is the *limiting reactant* for a process? Why does a reaction stop when the limiting reactant is consumed, even though there may be plenty of the other reactants present?
30. Nitrogen (N_2) and hydrogen (H_2) react to form ammonia (NH_3). Consider the mixture of N_2 (●●) and H_2 (●●) in a closed container as illustrated below:

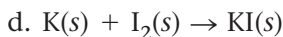
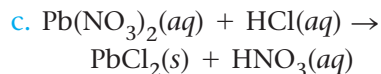
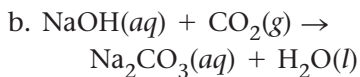
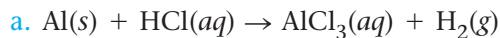


Assuming the reaction goes to completion, draw a representation of the product mixture. Explain how you arrived at this representation.

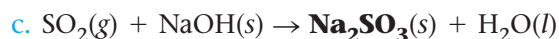
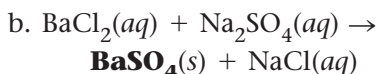
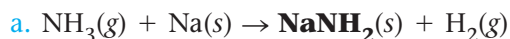
B. Calculations Involving a Limiting Reactant

31. Explain how one determines which reactant in a process is the limiting reactant. Does this depend only on the masses of the reactant present? Is the mole ratio in which the reactants combine involved?
32. How does the maximum *yield* of products for a reaction depend on the limiting reactant?
33. What does it mean to say that a reactant is present “in excess” in a process? Can the *limiting reactant* be present in excess? Does the presence of an excess of a reactant affect the mass of products expected for a reaction?

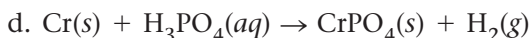
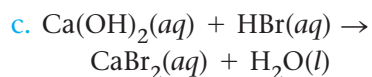
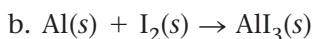
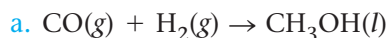
34. For each of the following *unbalanced* chemical equations, suppose that exactly 15.0 g of *each* reactant is taken. Determine which reactant is limiting, and calculate what mass of each product is expected. (Assume that the limiting reactant is completely consumed.)



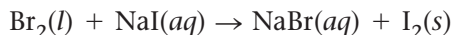
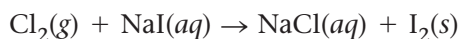
35. For each of the following *unbalanced* chemical equations, suppose that exactly 50.0 g of *each* reactant is taken. Determine which reactant is limiting, and calculate what mass of the product in boldface is expected. (Assume that the limiting reactant is completely consumed.)



36. For each of the following *unbalanced* chemical equations, suppose that 10.0 mg of *each* reactant is taken. Show by calculation which reactant is limiting. Calculate the mass of each product that is expected.

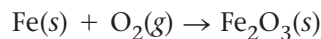


37. The more reactive halogen elements are able to replace the less reactive halogens from their compounds.



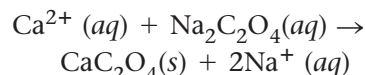
Suppose that separate solutions, each containing 25.0 g of NaI, are available. If 5.00 g of Cl_2 gas is bubbled into one NaI solution, and 5.00 g of liquid bromine is added to the other, calculate the number of grams of elemental iodine produced in each case.

38. If steel wool (iron) is heated until it glows and is placed in a bottle containing pure oxygen, the iron reacts spectacularly to produce iron(III) oxide.



If 1.25 g of iron is heated and placed in a bottle containing 0.0204 mol of oxygen gas, what mass of iron(III) oxide is produced?

39. One method for chemical analysis involves finding some reagent that will precipitate the species of interest. The mass of the precipitate is then used to determine what mass of the species of interest was present in the original sample. For example, calcium ion can be precipitated from solution by addition of sodium oxalate. The balanced equation is



Suppose that a solution is known to contain approximately 15 g of calcium ion. Show by calculation whether the addition of a solution containing 15 g of sodium oxalate will precipitate all of the calcium from the sample.

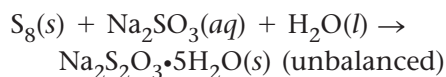
C. Percent Yield

40. What is the *actual yield* of a reaction? What is the *percent yield* of a reaction? How do the actual yield and the percent yield differ from the theoretical yield?

41. The text explains that one reason why the actual yield for a reaction may be less than the theoretical yield is side reactions. Suggest some other reasons why the percent yield for a reaction might not be 100%.

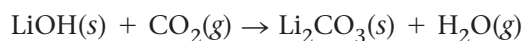
42. According to his prelaboratory theoretical yield calculations, a student's experiment should have produced 1.44 g of magnesium oxide. When he weighed his product after reaction, only 1.23 g of magnesium oxide was present. What is the student's percent yield?

- 43.** The compound sodium thiosulfate pentahydrate, $\text{Na}_2\text{S}_2\text{O}_3 \cdot 5\text{H}_2\text{O}$, is important commercially to the photography business as “hypo,” because it has the ability to dissolve unreacted silver salts from photographic film during development. Sodium thiosulfate pentahydrate can be produced by boiling elemental sulfur in an aqueous solution of sodium sulfite.



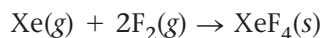
What is the theoretical yield of sodium thiosulfate pentahydrate when 3.25 g of sulfur is boiled with 13.1 g of sodium sulfite? Sodium thiosulfate pentahydrate is very soluble in water. What is the percent yield of the synthesis if a student doing this experiment is able to isolate (collect) only 5.26 g of the product?

- 44.** Alkali metal hydroxides are sometimes used to “scrub” excess carbon dioxide from the air in closed spaces (such as submarines and spacecraft). For example, lithium hydroxide reacts with carbon dioxide according to the unbalanced chemical equation



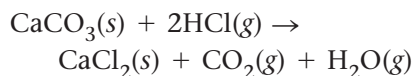
Suppose that a lithium hydroxide canister contains 155 g of $\text{LiOH}(\text{s})$. What mass of $\text{CO}_2(\text{g})$ will the canister be able to absorb? If it is found that after 24 hours of use the canister has absorbed 102 g of carbon dioxide, what percentage of its capacity has been reached?

- 45.** Although they were formerly called the inert gases, at least the heavier elements of Group 8 do form relatively stable compounds. For example, xenon combines directly with elemental fluorine at elevated temperatures in the presence of a nickel catalyst.



What is the theoretical mass of xenon tetrafluoride that should form when 130. g of xenon is reacted with 100. g of F_2 ? What is the percent yield if only 145 g of XeF_4 is actually isolated?

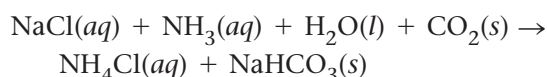
- 46.** Anhydrous calcium chloride, CaCl_2 , is frequently used in the laboratory as a drying agent for solvents because it absorbs 6 mol of water molecules for every mole of CaCl_2 used (forming a stable solid hydrated salt, $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$). Calcium chloride is typically prepared by treating calcium carbonate with hydrogen chloride gas.



A large amount of heat is generated by this reaction, so the water produced from the reaction is usually driven off as steam. Some liquid water may remain, however, and it may dissolve some of the desired calcium chloride. What is the percent yield if 155 g of calcium carbonate is treated with 250. g of anhydrous hydrogen chloride and only 142 g of CaCl_2 is obtained?

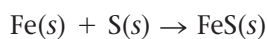
Critical Thinking

- 47.** One process for the commercial production of baking soda (sodium hydrogen carbonate) involves the following reaction, in which the carbon dioxide is used in its solid form (“dry ice”) both to serve as a source of reactant and to cool the reaction system to a temperature low enough for the sodium hydrogen carbonate to precipitate:



Because they are relatively cheap, sodium chloride and water are typically present in excess. What is the expected yield of NaHCO_3 when one performs such a synthesis using 10.0 g of ammonia and 15.0 g of dry ice, with an excess of NaCl and water?

- 48.** A favorite demonstration among chemistry teachers, performed to show that the properties of a compound differ from those of its constituent elements, involves iron filings and powdered sulfur. If the teacher takes samples of iron and sulfur and simply mixes them together, the two elements can be separated from each other with a magnet (iron is attracted to a magnet, sulfur is not). If the teacher then combines and *heats* the mixture of iron and sulfur, a reaction takes place and the elements combine to form iron(II) sulfide (which is not attracted by a magnet).

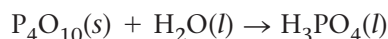
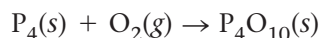


Suppose that 5.25 g of iron filings is combined with 12.7 g of sulfur. What is the theoretical yield of iron(II) sulfide?

49. When the sugar glucose, $C_6H_{12}O_6$, is burned in air, carbon dioxide and water vapor are produced. Write the balanced chemical equation for this process, and calculate the theoretical yield of carbon dioxide when 1.00 g of glucose is burned completely.
50. The traditional method of analyzing the amount of chloride ion present in a sample was to dissolve the sample in water and then slowly add a solution of silver nitrate. Silver chloride is very insoluble in water, and by adding a slight excess of silver nitrate, it is possible to effectively remove all chloride ion from the sample.
- $$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$
- Suppose that a 1.054-g sample is known to contain 10.3% chloride ion by mass. What mass of silver nitrate must be used to completely precipitate the chloride ion from the sample? What mass of silver chloride will be obtained?
51. Consider the balanced equation
- $$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$
- What mole ratio enables you to calculate the number of moles of oxygen needed to react exactly with a given number of moles of $C_3H_8(g)$? What mole ratios enable you to calculate how many moles of each product form from a given number of moles of C_3H_8 ?
52. For each of the following balanced reactions, calculate how many *moles of each product* would be produced by complete conversion of 0.50 mol of the reactant indicated in boldface. Indicate clearly the mole ratio used for the conversion.
- $2H_2O_2(l)$** $\rightarrow 2H_2O(l) + O_2(g)$
 - $2KClO_3(s)$** $\rightarrow 2KCl(s) + 3O_2(g)$
 - $2Al(s)$** $+ 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
 - $C_3H_8(g)$** $+ 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$
53. For each of the following balanced equations, indicate how many *moles of the product* could be produced by complete reaction of 1.00 g of the reactant indicated in boldface. Indicate clearly the mole ratio used for the conversion.
- $NH_3(g)$** $+ HCl(g) \rightarrow NH_4Cl(s)$
 - $CaO(s)$** $+ CO_2(g) \rightarrow CaCO_3(s)$
 - $4Na(s)$** $+ O_2(g) \rightarrow 2Na_2O(s)$
 - $2P(s)$** $+ 3Cl_2(g) \rightarrow 2PCl_3(l)$
54. If sodium peroxide is added to water, elemental oxygen gas is generated:
- $$Na_2O_2(s) + H_2O(l) \rightarrow NaOH(aq) + O_2(g)$$
- Suppose that 3.25 g of sodium peroxide is added to a large excess of water. What mass of oxygen gas will be produced?
55. The gaseous hydrocarbon acetylene, C_2H_2 , is used in welders' torches because of the large amount of heat released when acetylene burns with oxygen.
- $$2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(g)$$
- How many grams of oxygen gas are needed for the complete combustion of 150 g of acetylene?
56. For each of the following *unbalanced* chemical equations, suppose that 25.0 g of each reactant is taken. Show by calculation which reactant is limiting. Calculate the theoretical yield in grams of the product in boldface.
- $C_2H_5OH(l) + O_2(g) \rightarrow \mathbf{CO_2(g)} + H_2O(l)$
 - $N_2(g) + O_2(g) \rightarrow \mathbf{NO(g)}$
 - $NaClO_2(aq) + Cl_2(g) \rightarrow ClO_2(g) + \mathbf{NaCl(aq)}$
 - $H_2(g) + N_2(g) \rightarrow \mathbf{NH_3(g)}$
57. Hydrazine, N_2H_4 , emits a large quantity of energy when it reacts with oxygen, which has led to hydrazine's use as a fuel for rockets:
- $$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$$
- How many moles of each of the gaseous products are produced when 20.0 g of pure hydrazine is ignited in the presence of 20.0 g of pure oxygen? How many grams of each product are produced?
58. Before going to lab, a student read in her lab manual that the percent yield for a difficult reaction to be studied was likely to be only 40.% of the theoretical yield. The student's prelab stoichiometric calculations predict that the theoretical yield should be 12.5 g. What is the student's actual yield likely to be?
59. When a certain element whose formula is X_4 combines with HCl, the results are XCl_3 and hydrogen gas. Write a balanced equation for the reaction. When 24.0 g of hydrogen gas results from such a reaction, it is noted that 248 g of X_4 are consumed. Identify element X.

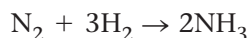
60. When a 5.00-g sample of element X reacts completely with a 15.0-g sample of element Y, compound XY is formed. When a 3.00-g sample of element X reacts with an 18.0-g sample of element Z, compound XZ₃ is formed. The molar mass of Y is 60.0 g/mol. Find the molar masses of X and Z.

61. Phosphoric acid, H₃PO₄, can be synthesized from phosphorus, oxygen, and water according to the following reactions:



If you start with 20.0 g of phosphorus, 30.0 g of oxygen gas, and 15.0 g of water, what is the maximum mass of phosphoric acid that can be formed?

62. Which of the following reaction mixtures would produce the greatest amount of product, assuming all went to completion? Each involves the reaction symbolized by the equation:



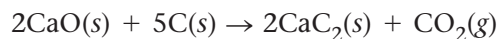
- 1 mole of N₂ and 3 moles of H₂
- 2 moles of N₂ and 3 moles of H₂
- 3 moles of N₂ and 3 moles of H₂
- Each would produce the same amount of product.

63. Which of the following statements is always true concerning a reaction represented by the following **balanced** chemical equation?



- If we react equal masses of C₂H₆ and O₂, there is no limiting reactant.
- If we react an equal number of moles of C₂H₆ and O₂, there is no limiting reactant.
- If we react a greater mass of C₂H₆ than of O₂, then O₂ must be limiting.
- If we react a greater mass of O₂ than of C₂H₆, then C₂H₆ must be limiting.

64. Determine the limiting reactant and mass of left-over reactant when a 50.0 g sample of CaO is reacted with 50.0 g of carbon according to the balanced equation



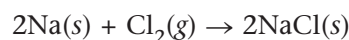
65. A sample of 13.97 g KClO₃ is decomposed according to the unbalanced equation



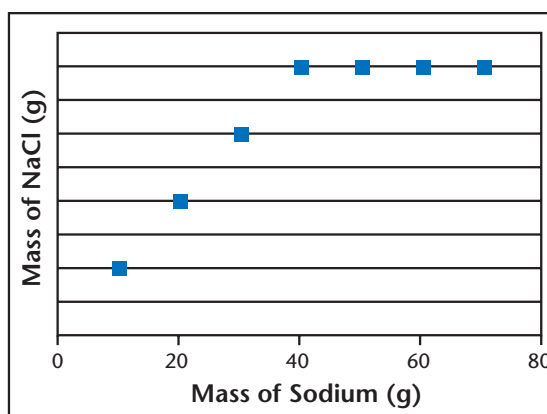
After the reaction 6.23 g KCl is collected. What is the percent yield of your experiment?

66. True or false? The limiting reactant is always the reactant that is present in the smallest amount (with respect to mass). Provide mathematical support for your answer using the chemical equation N₂ + 3H₂ → 2NH₃.

67. You have seven closed containers each with equal masses of chlorine gas (Cl₂). You add 10.0 g of sodium to the first sample, 20.0 g of sodium to the second sample, and so on (adding 70.0 g of sodium to the seventh sample). Sodium and chlorine react to form sodium chloride according to the equation



After each reaction is complete, you collect and measure the amount of sodium chloride formed. A graph of your results is shown below.



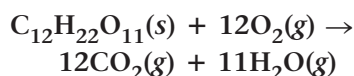
Answer the following questions:

- Explain the shape of the graph.
- Calculate the mass of NaCl formed when 20.0 g of sodium is used.
- Calculate the mass of Cl₂ in each container.
- Calculate the mass of NaCl formed when 50.0 g of sodium is used.
- Identify the leftover reactant and determine its mass for parts b and d above.

1 Which of the following statements concerning balancing equations is **false**?

- A There must always be a coefficient of 1 in an equation balanced in standard form.
- B The ratio of coefficients in a balanced equation is much more meaningful than an individual coefficient in a balanced equation.
- C The total number of atoms must be the same on the reactant side and the product side of the balanced equation.
- D When there are two products in an equation, the order in which they are written does not matter.

2 When table sugar is burned in the air, carbon dioxide and water vapor are products as shown by the following balanced equation:



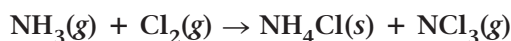
How many moles of oxygen gas are required to react completely with 2.0 moles of sugar?

- A 12
- B 24
- C 30
- D 35

3 When ethane (C_2H_6) is reacted with oxygen in the air, the products are carbon dioxide and water. This process requires _____ mole(s) of oxygen for every one mole of ethane.

- A 1
- B 2.5
- C 3.5
- D 7

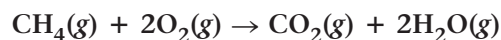
4 Consider the reaction represented by the following **unbalanced** chemical equation



What mass of NH_4Cl can be produced from 10.0 g of NH_3 and an excess of Cl_2 ?

- A 160.5 g
- B 53.5 g
- C 35.3 g
- D 23.6 g

5 The flame of the Bunsen burner is caused by the reaction of methane (CH_4) and oxygen according to the following balanced equation



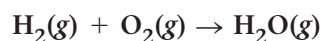
If 5.00 g of methane is burned, what mass of water can be produced? Assume the reaction goes to completion.

- A 5.62 g
- B 10.0 g
- C 11.2 g
- D 18.0 g

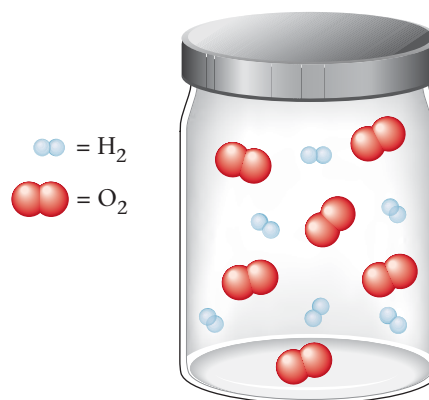
6 The maximum amount of a product that can be formed given the amounts of reactants used is called

- A the actual yield.
- B the theoretical yield.
- C the percent yield.
- D the limiting reactant.

7 Hydrogen gas (H_2) and oxygen gas (O_2) react to form water vapor according to the following unbalanced chemical equation:



Consider the mixture of H_2 and O_2 in a closed container as illustrated:



Draw a molecular level picture of the product mixture, including any left over reactant.

LOOKING AHEAD

10.1 Energy, Temperature, and Heat

- A. The Nature of Energy
- B. Temperature and Heat
- C. Exothermic and Endothermic Processes

10.2 The Flow of Energy


- A. Thermodynamics
- B. Measuring Energy Changes

10.3 Energy and Chemical Reactions

- A. Thermochemistry (Enthalpy)
 - Calorimetry
- B. Hess's Law
 - Characteristics of Enthalpy Changes

10.4 Using Energy in the Real World

- A. Quality Versus Quantity of Energy
- B. Energy and Our World
 - Petroleum and Natural Gas
 - Coal
 - Effects of Carbon Dioxide on Climate
 - New Energy Sources
- C. Energy as a Driving Force
 - Entropy



Wind is an alternative source of energy for producing electricity.

IN YOUR LIFE

Energy is the essence of our very existence as individuals and as a society. The food that we eat furnishes the energy to live, work, and play, just as the coal and oil consumed by manufacturing and transportation systems power our modern industrialized civilization.

Huge quantities of carbon-based fossil fuels have been available for the taking. This abundance of fuels has led to a world society with a voracious appetite for energy, consuming millions of barrels of petroleum every day. We are now dangerously dependent on the dwindling supplies of oil, and this dependence is an important source of tension among nations in today's world. In an incredibly short time we have moved from a period of ample and cheap supplies of petroleum to one of high prices and uncertain supplies. If our present standard of living is to be maintained, we must find alternatives to petroleum. To do this, we need to know the relationship between chemistry and energy, which we explore in this chapter.

**WHAT DO YOU KNOW?** **Prereading Questions**

1. What first comes to mind when you hear the term *energy*?
2. What do calories measure?
3. List three sources of energy.
4. Have you heard of the term *entropy*? What does it mean?

SECTION 10.1

Energy, Temperature, and Heat

Key Terms

- Energy
- Potential energy
- Kinetic energy
- Law of conservation of energy
- Work
- State function
- Temperature
- Heat
- System
- Surroundings
- Exothermic
- Endothermic

Objectives

- To understand the general properties of energy
- To understand the concepts of temperature and heat
- To understand the direction of energy flow as heat

A. The Nature of Energy

Although energy is a familiar concept, it is difficult to define precisely. For our purposes we will define **energy** as *the ability to do work or produce heat*. We will define these terms below.

Energy can be classified as either potential or kinetic energy. **Potential energy** is energy due to position or composition. For example, water behind a dam has potential energy that can be converted to work when the water flows down through turbines, thereby creating electricity. Attractive and repulsive forces also lead to potential energy. The energy released when gasoline is burned results from differences in attractive forces between the nuclei and electrons in the reactants and products. The **kinetic energy** of an object is energy due to the motion of the object and depends on the mass of the object m and its velocity v : $KE = \frac{1}{2}mv^2$.



Energy

The ability to do work or to produce heat

Potential energy

Energy due to position or composition

Kinetic energy

Energy due to the motion of the object

Law of conservation of energy

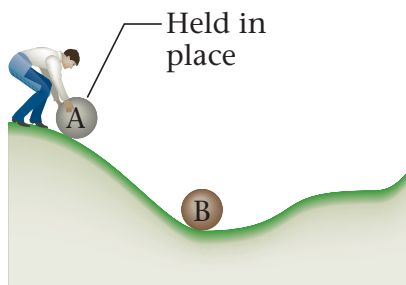
Energy cannot be created or destroyed, but it can be converted from one form to another.

Active Reading Question

Explain the difference between kinetic energy and potential energy.

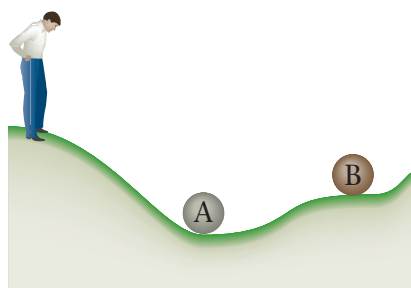
One of the most important characteristics of energy is that it is conserved. The **law of conservation of energy** states *that energy can be converted from one form to another but can be neither created nor destroyed*. That is, the energy of the universe is constant.

Although the energy of the universe is constant, it can be converted from one form to another. Consider the two balls as shown.



Ball A, because of its initially higher position, has more potential energy than ball B.

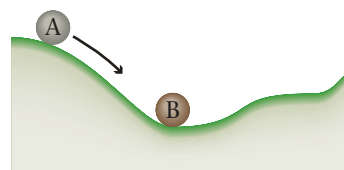
When ball A is released, it moves down the hill and strikes ball B, giving this arrangement.



What has happened in going from the initial to the final arrangement? The potential energy of A has decreased because its position was lowered. However, this energy cannot disappear. Where is the energy lost by A?

Initially, the potential energy of A is changed to kinetic energy as the ball rolls down the hill. Part of this energy is transferred to B, causing it to be raised to a higher final position. Thus the potential energy of B has been increased, which means that **work** (force acting over a distance) has been performed on B. Because the final position of B is lower than the original position of A, however, some of the energy is still unaccounted for. Both balls in their final positions are at rest, so the missing energy cannot be attributed to their motions. What has happened to the remaining energy?

The answer lies in the interaction between the hill's surface and ball A. As ball A rolls down the hill, some of its kinetic energy is transferred to the surface of the hill as heat.



Work

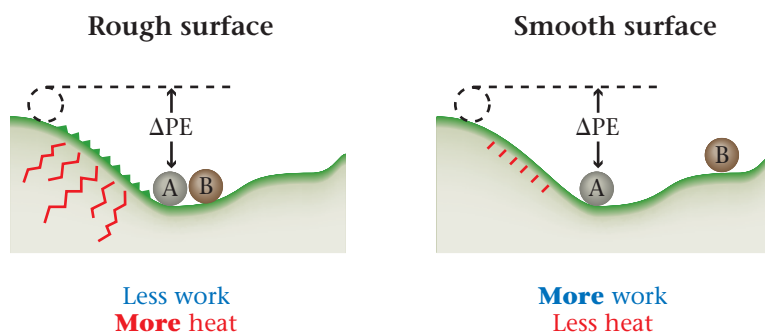
Force acting over a distance

CRITICAL THINKING ?

What if energy were not conserved? How would it affect our lives?

This transfer of energy is called *frictional heating*. The temperature of the hill increases very slightly as the ball rolls down. Thus the energy stored in A in its original position (potential energy) is distributed to B through work and to the surface of the hill by heat.

Imagine that we perform this same experiment several times, varying the surface of the hill from very smooth to very rough.



In rolling to the bottom of the hill, A always loses the same amount of energy because its position always changes by exactly the same amount. The way that this energy transfer is divided between work and heat, however, depends on the specific conditions—the *pathway*.

For example,

- On rough surface
 - Energy of A \longrightarrow frictional heating
 - No work is done on B.
- On smooth surface
 - Energy of A \longrightarrow frictional heating
 - Energy of A \longrightarrow Work done on B.
(B moves to a higher level absorbing energy.)

Regardless of the condition of the hill's surface, the *total energy* transferred will be constant, although the amounts of heat and work will differ.

Energy change is independent of the pathway, whereas work and heat are both dependent on the pathway.

State function

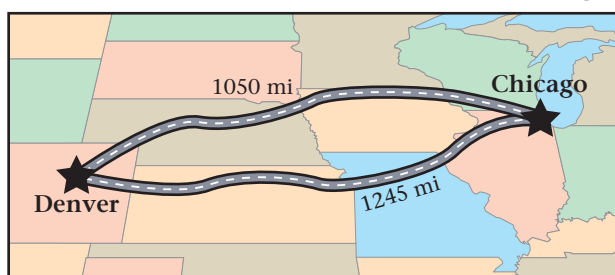
A property of a system that does not depend on the pathway

This brings us to a very important idea, the state function. A **state function** is a property of the system that changes independently of its pathway. Let's consider a nonchemical example. Suppose you are traveling from Chicago to Denver. Which of the following are state functions?

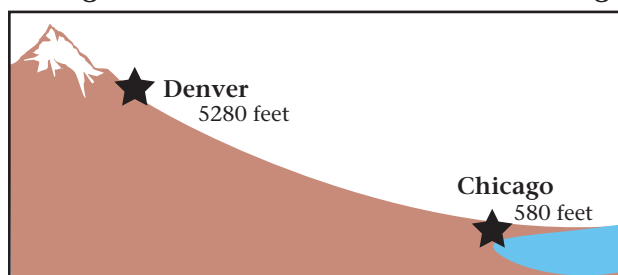
Distance traveled or Change in elevation

Because the distance traveled depends on the route taken (that is, the *pathway* between Chicago and Denver), it is *not* a state function.

Distance traveled from Denver to Chicago



Change in elevation from Denver to Chicago



On the other hand, the change in elevation depends only on the difference between Denver's elevation (5280 ft) and Chicago's elevation (580 ft). It does not depend on the route taken between the two cities.

We can also learn about state functions from the ball example. Because ball A always goes from its initial position on the hill to the bottom of the hill, its energy change is always the same, regardless of whether the hill is smooth or bumpy. This energy is a state function—a given change in energy is independent of the pathway of the process. In contrast, work and heat are *not* state functions. For a given change in the position of A, a smooth hill produces more work and less heat than a rough hill does. That is, for a given change in the position of A, the change in energy is always the same (state function) but the way the resulting energy is distributed as heat or work depends on the nature of the hill's surface (heat and work are not state functions).

B. Temperature and Heat

What does the temperature of a substance tell us about that substance? Put another way, how is warm water different from cold water? The answer lies in the motions of the water molecules. **Temperature** is a *measure of the random motions of the components of a substance*. That is, the H_2O molecules in warm water are moving around more rapidly than the H_2O molecules in cold water.

Consider an experiment in which we place 1.00 kg of hot water ($90.^\circ\text{C}$) next to 1.00 kg of cold water ($10.^\circ\text{C}$) in an insulated box. The water samples are separated from each other by a thin metal plate (see **Figure 10.1**). You already know what will happen: the hot water will cool down and the cold water will warm up.

Assuming that no energy is lost to the air, can we determine the final temperature of the two samples of water? Let's consider how to think about this problem.

First picture what is happening. Remember that the H_2O molecules in the hot water are moving faster than those in the cold water (see **Figure 10.2**). As a result, energy will be transferred through the metal wall from the hot water to the cold water. This energy transfer will cause the H_2O molecules in the hot water to slow down and the H_2O molecules in the cold water to speed up.

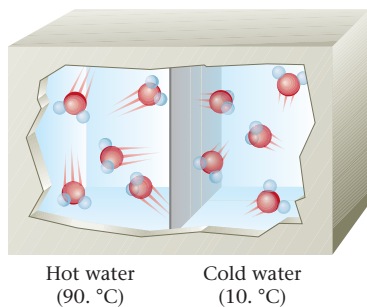


Figure 10.2

The H_2O molecules in hot water have much greater random motions than the H_2O molecules in cold water.

Thus we have a transfer of energy from the hot water to the cold water. This flow of energy is called heat. **Heat** can be defined as a *flow of energy due to a temperature difference*. What will eventually happen? The two water samples will reach the same temperature (see **Figure 10.3**). At this point, how does the energy lost by the hot water compare to the energy gained by the cold water? They must be the same (remember that energy is conserved).

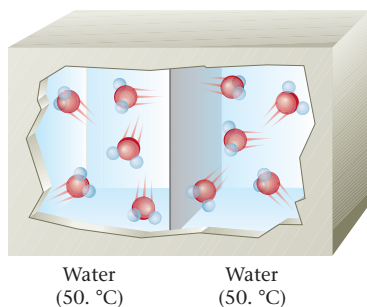


Figure 10.3

The water samples now have the same temperature ($50.^\circ\text{C}$) and have the same random motions.

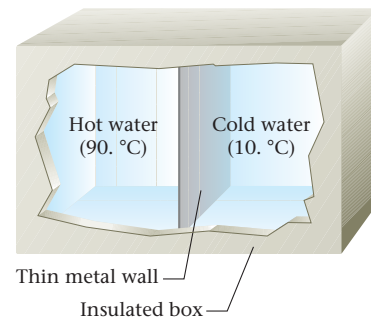


Figure 10.1

Equal masses of hot water and cold water separated by a thin metal wall in an insulated box

Temperature

Measure of the random motions (average kinetic energy) of the components of a substance

Heat

Flow of energy due to a temperature difference

Top Ten Lowest Temperatures Recorded by State

Temperature ($^\circ\text{C}$)	State	Date
-62.2	Alaska	1/23/71
-56.7	Montana	1/20/54
-56.1	Utah	2/1/85
-54.4	Wyoming	2/9/33
-51.7	Colorado	2/1/85
-51.1	Idaho	1/18/43
-51.1	North Dakota	2/15/36
-51.1	Minnesota	2/2/96
-50.0	South Dakota	2/17/36
-47.8	Oregon	2/10/33

We conclude that the final temperature is the average of the original temperatures:

$$T_{\text{final}} = \frac{T_{\text{initial}}^{\text{hot}} + T_{\text{initial}}^{\text{cold}}}{2} = \frac{90. \text{ }^{\circ}\text{C} + 10. \text{ }^{\circ}\text{C}}{2} = 50. \text{ }^{\circ}\text{C}$$

For the hot water, the temperature change is

$$\text{Change in temperature (hot)} = \Delta T_{\text{hot}} = 90. \text{ }^{\circ}\text{C} - 50. \text{ }^{\circ}\text{C} = 40. \text{ }^{\circ}\text{C}$$

The temperature change for the cold water is

$$\text{Change in temperature (cold)} = \Delta T_{\text{cold}} = 50. \text{ }^{\circ}\text{C} - 10. \text{ }^{\circ}\text{C} = 40. \text{ }^{\circ}\text{C}$$

In this example, the masses of hot water and cold water are equal. If they were unequal, this problem would be more complicated.

Let's Review

Temperature and Heat

- Temperature is a measure of the random motions of the components of an object.
- Heat is a *flow* of energy due to a temperature difference.
- The random motions of the components of an object constitute the *thermal energy* of that object.
- The flow of energy called heat is the way in which thermal energy is transferred from a hot object to a colder object.

C. Exothermic and Endothermic Processes

Next we will consider the energy changes that accompany chemical reactions. To explore this idea, let's consider the striking and burning of a match. Energy is clearly released through heat as the match burns. To discuss this reaction, we divide the universe into two parts: the system and the surroundings. The **system** is the part of the universe on which we wish to focus attention; the **surroundings** include everything else in the universe. In this case we define the system as the reactants and products of the reaction. The surroundings consist of the air in the room and anything else other than the reactants and products.

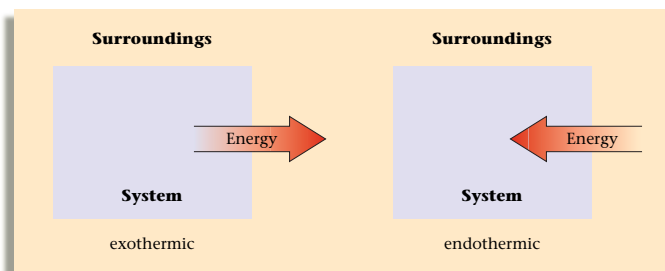
When a process results in the evolution of heat, it is said to be **exothermic** (*exo-* is a prefix meaning "out of"); that is, energy flows *out of the system*. For example, in the combustion of a match, energy flows out of the system as heat. Processes that absorb energy from the surroundings are said to be **endothermic**. When the heat flow moves *into a system*, the process is endothermic. Boiling water to form steam is a common endothermic process.

Exothermic

Refers to a process in which energy (as heat) flows out of the system into the surroundings

Endothermic

Refers to a process in which energy (as heat) flows from the surroundings into the system



Active Reading Question

Define *exothermic* and *endothermic* using the term *surroundings* instead of *system*.

Where does the energy, released as heat, come from in an exothermic reaction? The answer lies in the difference in potential energies between the products and the reactants. Which has lower potential energy, the reactants or the products? We know that total energy is conserved and that energy flows from the system into the surroundings in an exothermic reaction. Thus *the energy gained by the surroundings must be equal to the energy lost by the system*. In the combustion of a match, the burned match has lost potential energy (in this case potential energy stored in the bonds of the reactants), which was transferred through heat to the surroundings (see **Figure 10.4**). The heat flow into the surroundings results from a lowering of the potential energy of the reaction system. *In any exothermic reaction, some of the potential energy stored in the chemical bonds is converted to thermal energy (random kinetic energy) via heat.*



A burning match releases energy.

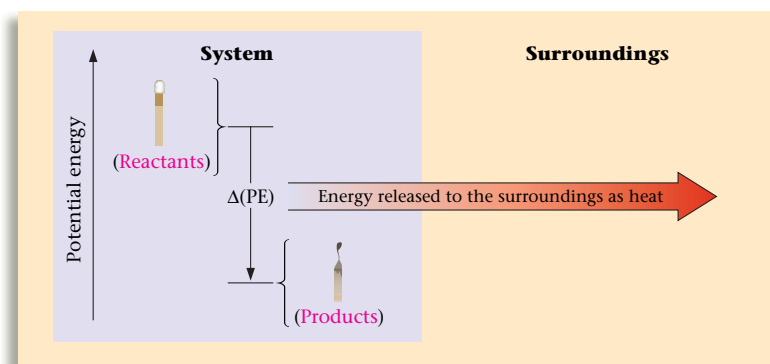
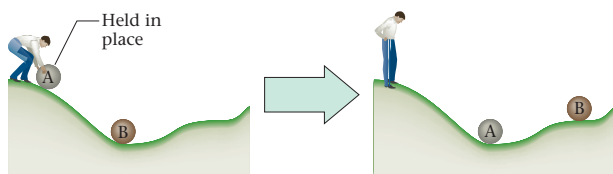


Figure 10.4
The energy changes accompanying the burning of a match

SECTION 10.1 REVIEW QUESTIONS

- 1 Explain why energy is a state function, and why heat and work are not.
- 2 What is the significance of the law of conservation of energy?
- 3 Explain the difference between potential energy and kinetic energy.
- 4 Write a paragraph explaining what happens to the balls using the terms *potential energy*, *kinetic energy*, *work*, and *heat*.
- 5 What is the difference between temperature and heat?
- 6 Provide an example of an endothermic process and an exothermic process different from the examples in the book.
- 7 Draw a picture of an endothermic reaction similar to the one for an exothermic reaction shown in Figure 10.4.



SECTION 10.2

The Flow of Energy

Key Terms

- Thermodynamics
- First law of thermodynamics
- Internal energy
- Calorie
- Joule

Objectives

- To understand how energy flow affects internal energy
- To understand how heat is measured

A. Thermodynamics

The study of energy is called **thermodynamics**. The law of conservation of energy is often called the **first law of thermodynamics**.

The energy of the universe is constant.

The **internal energy**, E , of a system can be defined most precisely as the sum of the kinetic and potential energies of all “particles” in the system. The internal energy of a system can be changed by a flow of work, heat, or both.

That is,

$$\Delta E = q + w$$

where

Δ (delta) means a change in the function that follows,

q represents heat,

w represents work.

Thermodynamic quantities always consist of two parts:

- a *number*, giving the magnitude of the change
- a *sign*, indicating the direction of the flow. *The sign reflects the system's point of view.*

For example, if a quantity of energy flows *into* the system via heat (an endothermic process), q is equal to $+x$, where the *positive* sign indicates that the *system's energy is increasing*. On the other hand, when energy flows *out* of the system via heat (an exothermic process), q is equal to $-x$, where the *negative* sign indicates that the *system's energy is decreasing*.

Thermodynamics

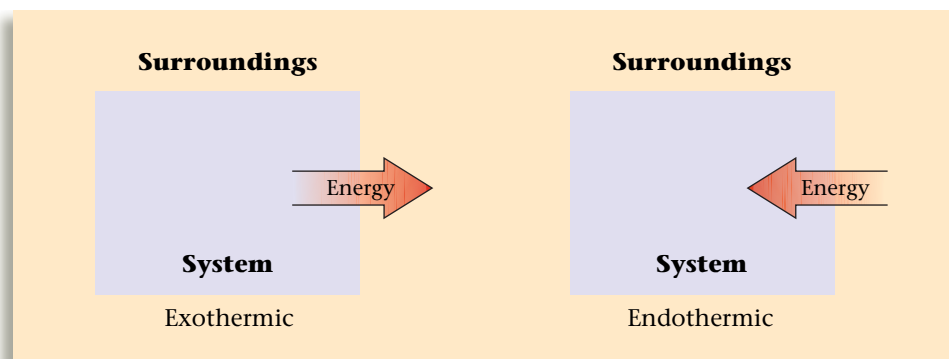
The study of energy

First law of thermodynamics

The energy of the universe is constant.

Internal energy

Sum of the kinetic and potential energies of all particles in the system



- $q = -x$
- negative $q \rightarrow$ system's energy decreases
- $\Delta E < 0$

- $q = +x$
- positive $q \rightarrow$ system's energy increases
- $\Delta E > 0$

In this text the same conventions apply to the flow of work. If the system does work on the surroundings (energy flows out of the system), w is negative. If the surroundings do work on the system (energy flows into the system), w is positive. We define work from the system's point of view to be consistent for all thermodynamic quantities. That is, in this convention the signs of both q and w reflect what happens to the system; thus we use $\Delta E = q + w$.

B. Measuring Energy Changes

Earlier in this chapter we saw that when we heat a substance to a higher temperature, we increase the motions of the components of the substance—that is, we increase the thermal energy of the substance. Different materials respond differently to being heated. To explore this idea we need to introduce the common units of energy: the *calorie* and the *joule* (pronounced *jewel*).

In the metric system the **calorie** is defined as the amount of energy (heat) required to raise the temperature of one gram of water by one Celsius degree. The “calorie” with which you are probably familiar is used to measure the energy content of food and is actually a kilocalorie (1000 calories), written with a capital C (Calorie) to distinguish it from the calorie used in chemistry. The **joule** (an SI unit) can be most conveniently defined in terms of the calorie:

$$1 \text{ calorie} = 4.184 \text{ joules}$$

or using the normal abbreviations

$$1 \text{ cal} = 4.184 \text{ J}$$

You need to be able to convert between calories and joules. We will consider that conversion process in Example 10.1.

EXAMPLE 10.1

Converting Calories to Joules

Express 60.1 cal of energy in units of joules.

Solution

Where do we want to go?



What do we know?

- 60.1 cal of energy
- 1 cal = 4.184 J

How do we get there?

The conversion factor needed is $\frac{4.184 \text{ J}}{1 \text{ cal}}$. The result is $(60.1 \text{ cal}) \left(\frac{4.184 \text{ J}}{1 \text{ cal}} \right) = 251 \text{ J}$.

Does it make sense?

Note that the 1 in the denominator is an exact number by definition and so does not limit the number of significant figures. The answer 251 J has the same number of significant figures as the original quantity (60.1 cal).

Practice Problem • Exercise 10.1

How many calories of energy correspond to 28.4 J?



Diet drinks are now labeled as “low joule” instead of “low calorie” in European countries.

Calorie

The amount of energy needed to raise the temperature of one gram of water by one Celsius degree

Now think about heating a substance from one temperature to another. How does the amount of substance heated affect the energy required? In 2 g of water there are twice as many molecules as in 1 g of water. It takes twice as much energy to change the temperature of 2 g of water by 1 °C, because we must change the motions of twice as many molecules in a 2-g sample as in a 1-g sample. Also, as we would expect, it takes twice as much energy to raise the temperature of a given sample of water by 2 degrees as it does to raise the temperature by 1 degree.

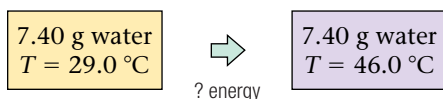
EXAMPLE 10.2

Calculating Energy Requirements

Determine the amount of energy (heat) in joules required to raise the temperature of 7.40 g water from 29.0 °C to 46.0 °C.

Solution

Where do we want to go?



What do we know?

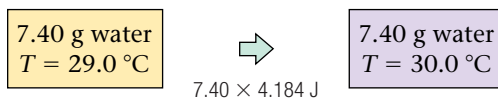
- Mass of water = 7.40 g
- Final temperature = 46.0 °C
- Initial temperature = 29.0 °C
- 1 cal = 4.184 J

How do we get there?

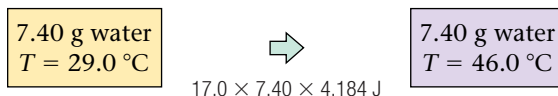
From the discussion in the text, we know that 4.184 J of energy is required to raise the temperature of *one* gram of water by *one* Celsius degree.



Because in our case we have 7.40 g of water instead of 1.00 g, it will take $7.40 \times 4.184\text{ J}$ to raise the temperature by one degree.



However, we want to raise the temperature of our sample of water by more than 1 °C. In fact, the temperature change required is from 29.0 °C to 46.0 °C. This is change of 17.0 °C ($46.0\text{ °C} - 29.0\text{ °C} = 17.0\text{ °C}$). Thus we will have to supply 17.0 times the energy necessary to raise the temperature of 7.40 g of water by 1 °C.



MATH



The result you will get on your calculator is 526.3472, which rounds to 526.

This calculation is summarized as follows:

$$4.184 \frac{\text{J}}{\text{g} \cdot \text{°C}} \times 7.40 \text{ g} \times 17.0 \text{ °C} = 526 \text{ J}$$

Energy per gram of water per degree of temperature × Actual grams of water × Actual temperature change = Energy required

Does it make sense?

The original measurements for mass and temperature have 3 significant figures which means our answer should have 3 significant figures. We are asked for the energy (heat) required. The units in our answer are J, which are energy units.

Practice Problem • Exercise 10.2

Calculate the joules of energy required to heat 454 g of water from 5.4 °C to 98.6 °C.

Let's Review

The energy (heat) required to change the temperature of a substance depends on:

- The amount of substance being heated (number of grams)
- The temperature change (number of degrees)

Information

Each state of a substance has a different specific heat.

There is, however, another important factor: the identity of the substance.

Different substances respond differently to being heated. We have seen that 4.184 J of energy raises the temperature of 1 g of water 1 °C. In contrast, this same amount of energy applied to 1 g of gold raises its temperature by approximately 32 °C! The point is that some substances require relatively large amounts of energy to change their temperatures, whereas others require relatively little. Chemists describe this difference by saying that substances have different heat capacities. *The amount of energy required to change the temperature of one gram of a substance by one Celsius degree* is called its **specific heat capacity** or, more commonly, its *specific heat*.

The specific heat capacities for several substances are listed in **Table 10.1**. You can see from the table that the specific heat capacity for water is very high compared to those of the other substances listed. This is why lakes and seas are much slower to respond to cooling or heating than are the surrounding land masses.

Table 10.1

The Specific Heat Capacities of Some Common Substances

Substance	Specific Heat Capacity (J/g °C)
water (l)* (liquid)	4.184
water (s) (ice)	2.03
water (g) (steam)	2.0
aluminum (s)	0.89
iron (s)	0.45
mercury (l)	0.14
carbon (s)	0.71
silver (s)	0.24
gold (s)	0.13
copper (s)	0.385

*The symbols (s), (l), and (g) indicate the solid, liquid, and gaseous states, respectively.

Specific heat capacity

The amount of energy needed to raise the temperature of one gram of a substance by one Celsius degree

Active Reading Question

Do metals have relatively high or low heat capacities? When in your life have you noticed this?

EXAMPLE 10.3

Calculations Involving Specific Heat Capacity

What quantity of energy (in joules and in calories) is required to heat a piece of iron weighing 1.3 g from 25 °C to 46 °C?

Solution

Where do we want to go?



What do we know?

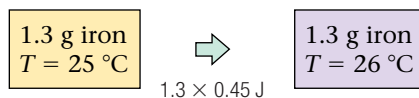
- Mass of iron sample = 1.3 g
- Initial temperature = 25 °C
- Final temperature = 46 °C
- Specific heat of iron = 0.45 J/g °C
- 1 cal = 4.184 J

How do we get there?

Our task is to determine how much energy is required (J and cal). It takes 0.45 J to raise the temperature of a 1-g piece of iron by 1 °C.



In this case our sample is 1.3 g, so $1.3 \times 0.45\text{ J}$ is required for *each* degree of temperature increase.



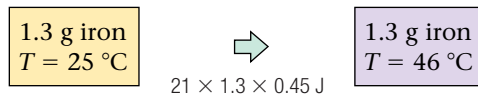
MATH



The result you will get on your calculator is 12.285, which rounds to 12.

Because the temperature increase is 21 °C ($46\text{ }^{\circ}\text{C} - 25\text{ }^{\circ}\text{C} = 21\text{ }^{\circ}\text{C}$), the total amount of energy required is

$$0.45 \frac{\text{J}}{\text{g}^{\circ}\text{C}} \times 1.3\text{ g} \times 21\text{ }^{\circ}\text{C} = 12\text{ J}$$



To calculate this energy in calories, we construct the appropriate conversion factor. We want to change from joules to calories, so cal must be in the numerator and J in the denominator, where it cancels:

$$12\text{ J} \times = \frac{1\text{ cal}}{4.184\text{ J}} = 2.9\text{ cal}$$

Does it make sense?

The final units of J and cal are energy units as required. The answers have 2 significant figures because all of the measurements are given with 2 significant figures.

Practice Problem • Exercise 10.3

A 5.63-g sample of solid gold is heated from 21 °C to 32 °C. How much energy (in joules and calories) is required?

CHEMISTRY IN YOUR WORLD

Connection to Biology

Nature Has Hot Plants

The voodoo lily is a beautiful and seductive plant. The exotic-looking lily features an elaborate reproductive mechanism—a purple spike that can reach nearly 3 feet in length and is cloaked by a hoodlike leaf. But approach to the plant reveals bad news—it smells terrible!

Despite its antisocial odor, this putrid plant has fascinated biologists for many years because of its ability to generate heat. At the peak of its metabolic activity, the plant's blossom can be as much as 15 °C above its surrounding temperature. To generate this much heat, the metabolic rate of the plant must be close to that of a flying hummingbird!

What's the purpose of this intense heat production? For a plant faced with limited food supplies in the very competitive tropical climate where it grows, heat production seems like a great waste of energy. The answer to this mystery is that the voodoo lily is pollinated mainly by carrion-loving insects. Thus the lily prepares a malodorous mixture of chemicals

characteristic of rotting meat, which it then “cooks” off into the surrounding air to attract flesh-feeding beetles and flies. Then, once the insects enter the pollination chamber, the high temperatures there (as high as 110 °F) cause the insects to remain very active to better carry out their pollination duties.

The voodoo lily is only one of many thermogenic (heat-producing) plants. These plants are of special interest to biologists because they provide opportunities to study metabolic reactions that are quite subtle in “normal” plants.

Titan Arum is reputedly the largest flower in the world.



Note that in Example 10.3, to calculate the energy (heat) required, we took the product of the specific heat capacity, the sample size in grams, and the change in temperature in Celsius degrees.

$$\boxed{\text{Energy (heat) required (Q)}} = \boxed{\text{Specific heat capacity (s)}} \times \boxed{\text{Mass (m) in grams of sample}} \times \boxed{\text{Changes in temperature } (\Delta T) \text{ in } ^\circ\text{C}}$$

We can represent this by the following equation:

$$Q = s \times m \times \Delta T$$

where

Q = energy (heat) required

s = specific heat capacity

m = mass of the sample in grams

ΔT = change in temperature in Celsius degrees

This equation always applies when a substance is being heated (or cooled) and no change of state occurs. Before you begin to use this equation, however, make sure you understand what it means.

MATH



The symbol Δ (the Greek letter delta) is shorthand for “change in.”

EXAMPLE 10.4

Specific Heat Capacity Calculations: Using the Equation

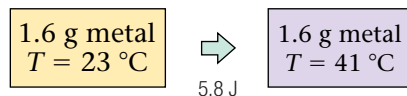
A 1.6-g sample of a metal that has the appearance of gold requires 5.8 J of energy to change its temperature from 23 °C to 41 °C. Is the metal pure gold?

Solution

Where do we want to go?

Is the metal sample pure gold?

What do we know?



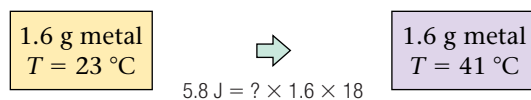
$$\Delta T = 41\text{ }^{\circ}\text{C} - 23\text{ }^{\circ}\text{C} = 18\text{ }^{\circ}\text{C}$$

How do we get there?

We can calculate the value of the specific heat capacity for the metal and compare this value to the one for gold given in Table 10.1. We know that

$$Q = s \times m \times \Delta T$$

or, pictorially,



When we divide both sides of the equation

$$Q = s \times m \times \Delta T$$

by $m \times \Delta T$, we get

$$\frac{Q}{m \times \Delta T} = \frac{s \times m \times \Delta T}{m \times \Delta T}$$

$$\frac{Q}{m \times \Delta T} = s$$

Now, we can calculate the value of s .

$$Q = \text{energy (heat) required} = 5.8\text{ J}$$

$$m = \text{mass of the sample} = 1.6\text{ g}$$

$$\Delta T = \text{change in temperature} = 18\text{ }^{\circ}\text{C} \quad (41\text{ }^{\circ}\text{C} - 23\text{ }^{\circ}\text{C} = 18\text{ }^{\circ}\text{C})$$

Thus

$$s = \frac{Q}{m \times \Delta T} = \frac{5.8\text{ J}}{(1.6\text{ g})(18\text{ }^{\circ}\text{C})} = 0.20\text{ J/g }^{\circ}\text{C}$$

As shown in Table 10.1, the specific heat capacity for gold is 0.13 J/g °C. Thus the metal must not be pure gold.

Practice Problem • Exercise 10.4

A 2.8-g sample of pure metal requires 10.1 J of energy to change its temperature from 21 °C to 36 °C. What is this metal? (Use Table 10.1.)

MATH



The result you will get on your calculator is 0.2013889, which rounds to 0.20.

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Firewalking: Magic or Science?

For millennia people have been amazed at the ability of Eastern mystics to walk across beds of glowing coals without any apparent discomfort. Even in the United States, thousands of people have performed feats of firewalking as part of motivational seminars. How can this be possible? Do firewalkers have supernatural powers?

Actually, there are good scientific explanations of why firewalking is possible. First, human tissue is mainly composed of water, which has a relatively large specific heat capacity. This means that a large amount of energy must be transferred from the coals to change significantly the temperature of the feet. During the brief contact between feet and coals involved in firewalking, there is relatively little time for energy flow, so the feet do not reach a high enough temperature to cause damage.

Also, although the surface of the coals has a very high temperature, the red-hot layer is very thin. Therefore, the quantity of energy available to heat the feet is smaller than might be expected. (Don't try this on your own!)



A group of firewalkers in Japan

SECTION 10.2

REVIEW QUESTIONS

- In our study of thermodynamics we look at changes in the internal energy of systems. To do this we consider thermodynamic quantities to consist of two parts. What are they? Why do we take the system's point of view?
- How is a food calorie different from the calorie we use in chemistry?
- Copy and complete the following table.
- Suppose you wanted to heat a mug of water (250 mL) from room temperature 25 °C to 100 °C to make a cup of tea. How much energy (in units of calories and joules) would you need from your microwave?
- We apply the same amount of energy to 10.0-g samples of aluminum, iron, and silver which begin at the same temperature. Rank the metals from highest to lowest temperature after the heat is applied. (See Table 10.1.)
- Determine the final temperature of a 10.0-g aluminum block originally at 25.0 °C if you apply 435 J of energy to it. (See Table 10.1.)

Mass of water	Change in temperature	Energy (cal)	Energy (J)
100.0 g	24.2 °C		
75.0 g		1.29×10^3 cal	
	18.4 °C		59.6 J

SECTION 10.3

Energy and Chemical Reactions

Key Terms

- Enthalpy
- Calorimeter
- Hess's law

Objectives

- To consider the heat (enthalpy) of chemical reactions
- To understand Hess's law

A. Thermochemistry (Enthalpy)

We have seen that some reactions are exothermic (produce heat energy) and other reactions are endothermic (absorb heat energy). Chemists also like to know exactly how much energy is produced or absorbed by a given reaction. To make that process more convenient, we have invented a special energy function called **enthalpy**, which is designated by H . For a reaction occurring under conditions of constant pressure, the change in enthalpy (ΔH) is equal to the energy that flows as heat. That is,

$$\Delta H_p = \text{heat}$$

where the subscript p indicates that the process has occurred under conditions of constant pressure and Δ means "a change in." Thus the enthalpy change for a reaction (that occurs at constant pressure) is the same as the heat for that reaction.

Enthalpy

At constant pressure, a change in enthalpy (ΔH) is equal to the energy flow as heat.

EXAMPLE 10.5

Enthalpy

When 1 mol of methane (CH_4) is burned at constant pressure, 890 kJ of energy is released as heat. Calculate ΔH for a process in which a 5.8-g sample of methane is burned at constant pressure.

Solution

Where do we want to go?



What do we know?

- 1 mol CH_4
- molar mass (CH_4) = $12.01 \text{ g/mol} + 4(1.008 \text{ g/mol}) = 16.04 \text{ g/mol}$
- constant pressure
- energy released per mol = 890 kJ
- $\Delta H = -890 \text{ kJ/mol } (\text{CH}_4) = q_p$

How do we get there?

The actual value can be calculated as follows:

$$5.8 \text{ g } \cancel{\text{CH}_4} \times \frac{1 \text{ mol } \text{CH}_4}{16.04 \text{ g } \cancel{\text{CH}_4}} = 0.36 \text{ mol } \text{CH}_4$$

$$0.36 \text{ mol } \cancel{\text{CH}_4} \times \frac{-890 \text{ kJ}}{\text{mol } \cancel{\text{CH}_4}} = -320 \text{ kJ}$$

Thus, when a 5.8-g sample of CH_4 is burned at constant pressure,

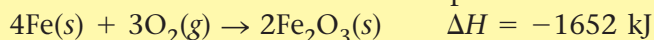
$$\Delta H = \text{heat flow} = -320 \text{ kJ}$$

Does it make sense?

Since the mass of CH_4 burned is smaller than 1 mol, less than 890 kJ will be released as heat. The answer has 2 significant figures as required by the given quantities.

Practice Problem • Exercise 10.5

The reaction that occurs in the heat packs used to treat sports injuries is



How much heat is released when 1.00 g of $\text{Fe}(s)$ is reacted with excess $\text{O}_2(g)$?

Calorimetry

A **calorimeter** (see **Figure 10.5**) is a device used to determine the heat associated with a chemical reaction. The reaction is run in the calorimeter and the temperature change of the calorimeter is observed. Knowing the temperature change that occurs in the calorimeter and the heat capacity of the calorimeter enables us to calculate the heat energy released or absorbed by the reaction. Thus we can determine ΔH for the reaction.

Once we have measured the ΔH values for various reactions, we can use these data to *calculate* the ΔH values of other reactions. We will see how to carry out these calculations in the next section.

B. Hess's Law

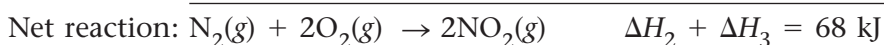
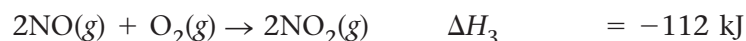
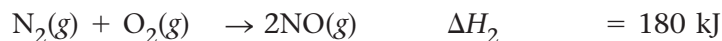
One of the most important characteristics of enthalpy is that it is a state function. That is, the change in enthalpy for a given process is independent of the pathway for the process.

In going from a particular set of reactants to a particular set of products, the change in enthalpy is the same whether the reaction takes place in one step or in a series of steps.

This principle, which is known as **Hess's law**, can be illustrated by examining the oxidation of nitrogen to produce nitrogen dioxide. The overall reaction can be written in one step, where the enthalpy change is represented by ΔH_1 .



This reaction can also be carried out in two distinct steps, with the enthalpy changes being designated as ΔH_2 and ΔH_3 :



Note that the sum of the two steps gives the net, or overall, reaction and that

$$\Delta H_1 = \Delta H_2 + \Delta H_3 = 68 \text{ kJ}$$

Calorimeter

A device used to determine the heat associated with a chemical or physical change

Hess's law

The change in enthalpy in going from a given set of reactants to a given set of products does not depend on the number of steps in the reaction.

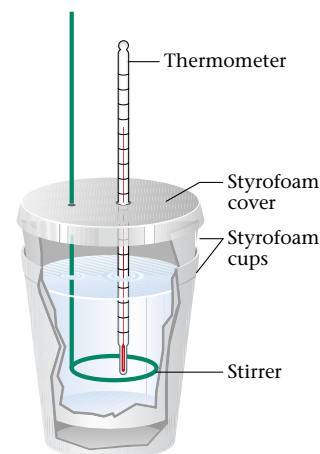


Figure 10.5

A coffee-cup calorimeter made of two Styrofoam cups

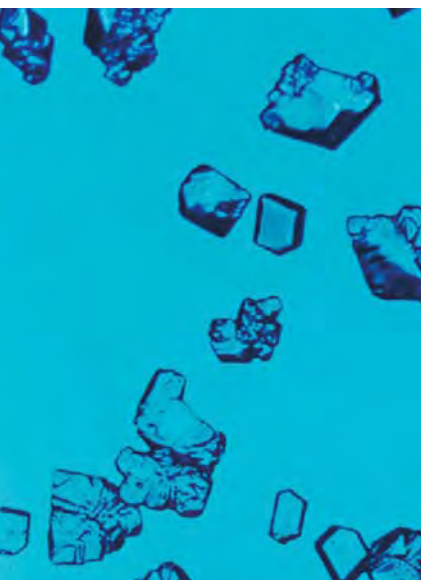
The importance of Hess's law is that it allows us to *calculate* heats of reaction that might be difficult or inconvenient to measure directly in a calorimeter.

Characteristics of Enthalpy Changes

To use Hess's law to compute enthalpy changes for reactions, it is important to understand two characteristics of ΔH for a reaction:

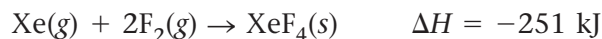
Characteristics of ΔH

- If a reaction is reversed, the sign of ΔH is also reversed.
- The magnitude of ΔH is directly proportional to the quantities of reactants and products in a reaction. If the coefficients in a balanced reaction are multiplied by an integer, the value of ΔH is multiplied by the same integer.



Crystals of xenon tetrafluoride, the first reported binary compound containing a noble gas element

Both these rules follow in a straightforward way from the properties of enthalpy changes. The first rule can be explained by recalling that the *sign* of ΔH indicates the *direction* of the heat flow at constant pressure. If the direction of the reaction is reversed, the direction of the heat flow also will be reversed. To see this, consider the preparation of xenon tetrafluoride, which was the first binary compound made from a noble gas:

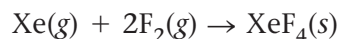


This reaction is exothermic, and 251 kJ of energy flows into the surroundings as heat. On the other hand, if the colorless XeF_4 crystals are decomposed into the elements, according to the equation

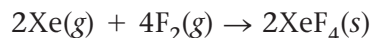


the opposite energy flow occurs because 251 kJ of energy must be added to the system to produce this endothermic reaction. Thus, for this reaction, $\Delta H = +251 \text{ kJ}$.

The second rule comes from the fact that ΔH depends on the amount of substances reacting. For example, since 251 kJ of energy is evolved for the reaction



then for a preparation involving twice the quantities of reactants and products, or



twice as much heat would be evolved:

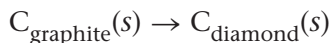
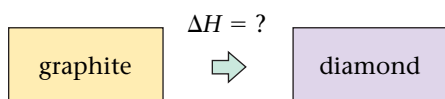
$$\Delta H = 2(-251 \text{ kJ}) = -502 \text{ kJ}$$

Active Reading Question

Why is the sign of ΔH reversed if the reaction is reversed?

EXAMPLE 10.6**Hess's Law**

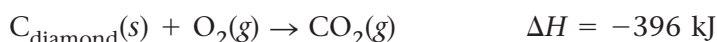
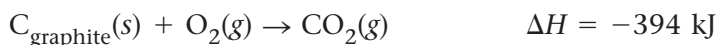
Two forms of carbon are graphite, the soft, black, slippery material used in “lead” pencils and as a lubricant for locks, and diamond, the brilliant, hard gemstone. Using the enthalpies of combustion for graphite (-394 kJ/mol) and diamond (-396 kJ/mol), calculate ΔH for the conversion of graphite to diamond:

**Solution****Where do we want to go?****What do we know?**

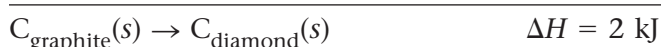
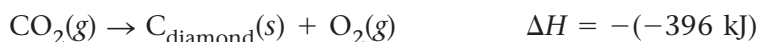
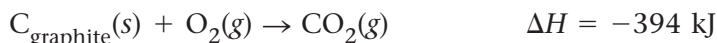
- $\Delta H_{\text{graphite}}^{\text{combustion}} = -394$ kJ/mol
- $\Delta H_{\text{diamond}}^{\text{combustion}} = -396$ kJ/mol

How do we get there?

The combustion reactions are



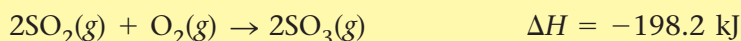
Note that if we reverse the second reaction (which means we must change the sign of ΔH) and sum the two reactions, we obtain the desired reaction:



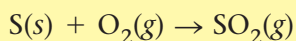
Thus 2 kJ of energy is required to change 1 mol graphite to diamond. This process is endothermic.

Practice Problem • Exercise 10.6

From the following information



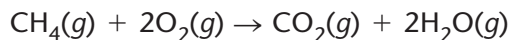
calculate ΔH for the reaction:



CELEBRITY CHEMICAL

Methane (CH₄)

Methane is the main component of natural gas, a valuable fossil fuel. It is such a good fuel because the combustion of methane with oxygen



produces 55 kJ of energy per gram of methane. Natural gas, which is associated with petroleum deposits and contains as much as 97% methane, originated from the decomposition of plants in ancient forests that became buried in natural geological processes.

Although the methane in natural gas represents a tremendous source of energy for our civilization, an even more abundant source of methane lies in the depths of the ocean. The U.S. Geological Survey estimates that 320,000 trillion cubic feet of methane is trapped in the deep ocean near the United States. This amount is 200 times the amount of methane contained in the natural gas deposits in the United States. In the ocean, the

methane is trapped in cavities formed by water molecules that are arranged very much like the water molecules in ice. These structures are called methane hydrates.

Although extraction of methane from the ocean floor offers tremendous potential benefits, it also carries risks. Methane is a “greenhouse gas”—its presence in the atmosphere helps to trap the heat from the sun. As a result, any accidental release of the methane from the ocean could produce serious warming of the earth’s climate. As usual, environmental trade-offs accompany human activities.

Flaming pieces of methane hydrate

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SECTION 10.3

REVIEW QUESTIONS

- 1 If enthalpy is the heat for a reaction, it must have a sign as well as a magnitude. What sign should the enthalpy for an exothermic reaction have? Why?
- 2 Suppose you ran a chemical reaction in the calorimeters shown in Figure 10.5. If the temperature of the solution goes from 27 °C to 36 °C for a sample of mass 5.0 g, how would you determine the energy produced by the reaction?
- 3 What is Hess’s law and why is it useful?
- 4 The enthalpy of combustion of solid carbon to form carbon dioxide is -393.7 kJ/mol carbon, and the enthalpy of combustion of carbon monoxide to form carbon dioxide is -283.3 kJ/mol CO. Using these data calculate the change in enthalpy for the reaction
$$2\text{C}(s) + \text{O}_2(g) \rightarrow 2\text{CO}(g)$$
- 5 Given the following data
$$\text{N}_2(g) + 2\text{O}_2(g) \rightarrow 2\text{NO}_2(g) \quad \Delta H = 67.7 \text{ kJ}$$
$$\text{N}_2(g) + 2\text{O}_2(g) \rightarrow \text{N}_2\text{O}_4(g) \quad \Delta H = 9.7 \text{ kJ}$$
calculate ΔH for $2\text{NO}_2(g) \rightarrow \text{N}_2\text{O}_4(g)$.

Using Energy in the Real World

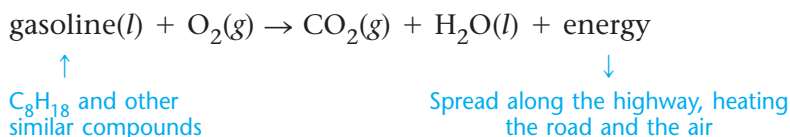
Objectives

- To understand how the quality of energy changes as it is used
- To consider the energy resources of our world
- To understand energy as a driving force for natural processes

A. Quality Versus Quantity of Energy

One of the most important characteristics of energy is that it is conserved. Thus the total energy content of the universe will always be what it is now. If that is the case, why are we concerned about energy? For example, why should we worry about conserving our petroleum supply? Surprisingly, the “energy crisis” is not about the *quantity* of energy, but rather about the *quality* of energy.

To understand this idea, consider an automobile trip from Chicago to Denver. Along the way you would put gasoline into the car to get to Denver. What happens to that energy? The energy stored in the bonds of the gasoline and of the oxygen that reacts with it is changed to thermal energy, which is spread along the highway to Denver. The total quantity of energy remains the same as before the trip but the energy concentrated in the gasoline becomes widely distributed in the environment:



Which energy is easier to use to do work: the concentrated energy in the gasoline or the thermal energy spread from Chicago to Denver? Of course, the energy concentrated in the gasoline is more convenient to use. This example illustrates a very important general principle:

When we use energy to do work, we degrade its usefulness.

In other words, when we use energy the *quality* of that energy (its ease of use) is lowered.

In summary,



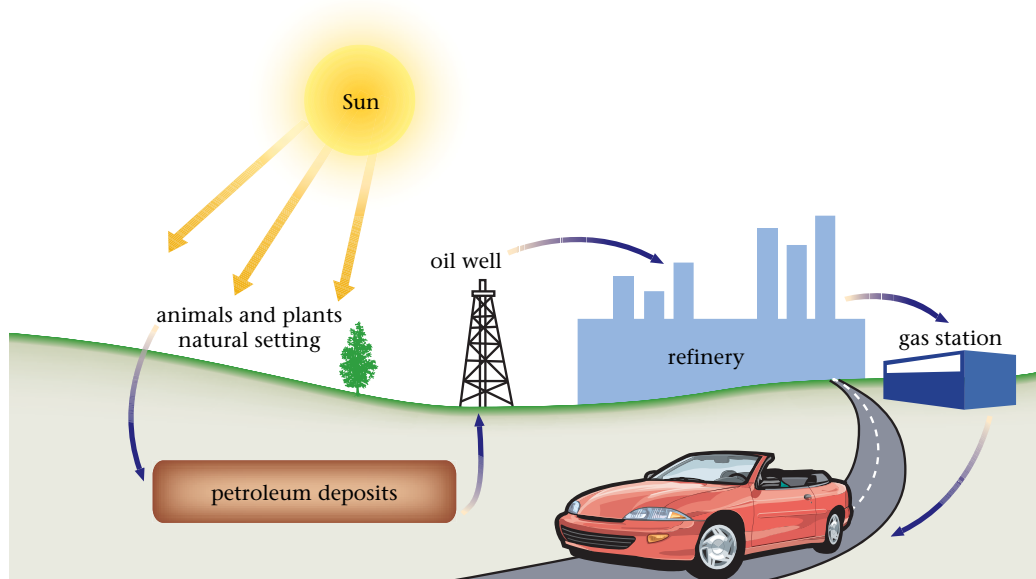
You may have heard someone mention the “heat death” of the universe. Eventually (many eons from now), all energy will be spread evenly throughout the universe and everything will be at the same temperature. At this point it will no longer be possible to do any work. The universe will be “dead.”

Key Terms

- Fossil fuels
- Petroleum
- Natural gas
- Coal
- Greenhouse effect
- Energy spread
- Matter spread
- Entropy
- Second law of thermodynamics

We don't have to worry about the heat death of the universe anytime soon, of course, but we do need to think about conserving "quality" energy supplies. The energy stored in petroleum molecules got there over millions of years through plants and simple animals absorbing energy from the sun and using this energy to construct molecules. As these organisms died and became buried, natural processes changed them into the petroleum deposits we now access for our supplies of gasoline and natural gas.

Petroleum is highly valuable because it furnishes a convenient, concentrated source of energy. Unfortunately, we are using this fuel at a much faster rate than natural processes can replace it, so we are looking for new sources of energy. The most logical energy source is the sun. *Solar energy* refers to using the sun's energy directly to do productive work in our society. We will discuss energy supplies next.



Fossil fuel

Fuel that consists of carbon-based molecules derived from the decomposition of once-living organisms

B. Energy and Our World

Woody plants, coal, petroleum, and natural gas provide a vast resource of energy that originally came from the sun. By the process of photosynthesis, plants store energy that can be claimed by burning the plants themselves or the decay products that have been converted over millions of years to **fossil fuels**. Although the United States currently depends heavily on petroleum for energy, this dependency is a relatively recent phenomenon, as shown in **Figure 10.6**. Now we will discuss some sources of energy and their effects on the environment.

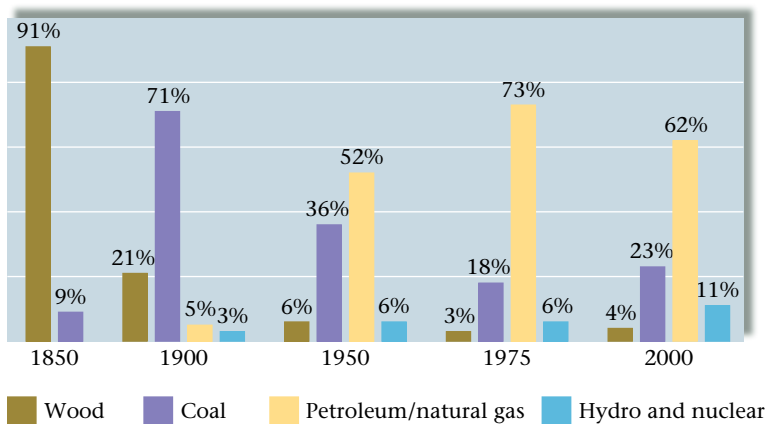


Figure 10.6

Energy sources used in the United States



Petroleum and Natural Gas

Although how they were produced is not completely understood, petroleum and natural gas were most likely formed from the remains of marine organisms that lived approximately 500 million years ago. **Petroleum** is a thick, dark liquid composed mostly of compounds called *hydrocarbons* that contain carbon and hydrogen. (Carbon is unique among elements in the extent to which it can bond to itself to form chains of various lengths.) **Table 10.2** gives the formulas and names for several common hydrocarbons. **Natural gas**, usually associated with petroleum deposits, consists mostly of methane, but it also contains significant amounts of ethane, propane, and butane.

The composition of petroleum varies somewhat, but it includes mostly hydrocarbons having chains that contain from 5 to more than 25 carbons. To be used efficiently, the petroleum must be separated into fractions by boiling. The lighter molecules (having the lowest boiling points) can be boiled off, leaving the heavier ones behind. The commercial uses of various petroleum fractions are shown in **Table 10.3**.

Petroleum

A thick, dark liquid composed mostly of hydrocarbon compounds

Natural gas

A gaseous fossil fuel, mostly consisting of methane (with some ethane, propane, and butane) and usually associated with petroleum deposits

Table 10.2

Names and Formulas for Some Common Hydrocarbons

Formula	Name
CH_4	Methane
C_2H_6	Ethane
C_3H_8	Propane
C_4H_{10}	Butane
C_5H_{12}	Pentane
C_6H_{14}	Hexane
C_7H_{16}	Heptane
C_8H_{18}	Octane

Table 10.3

Uses of the Various Petroleum Fractions

Petroleum Fraction in Terms of Numbers of Carbon Atoms	Major Uses
$\text{C}_5\text{--C}_{10}$	Gasoline
$\text{C}_{10}\text{--C}_{18}$	Kerosene Jet fuel
$\text{C}_{15}\text{--C}_{25}$	Diesel fuel Heating oil Lubricating oil
$>\text{C}_{25}$	Asphalt

History of Gasoline



First "practical" automobile

1850

1860

1859

First oil well (Titusville, PA)

1870

Kerosene from petroleum used for lamps—gasoline discarded

1880

1890

1900

1910

1913

First cracking process to convert oil to gasoline

1920

1922

Tetraethyl lead added to gasoline to prevent knocking

1930

1937

Catalytic cracking of low grade fuels to "high test" gasoline

1940

1950

1960

1970

1975

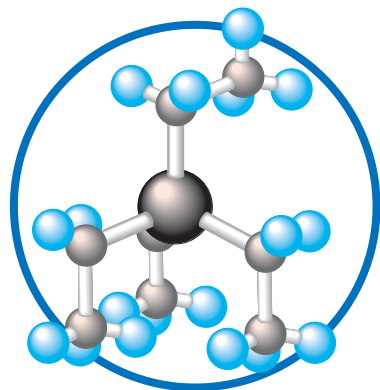
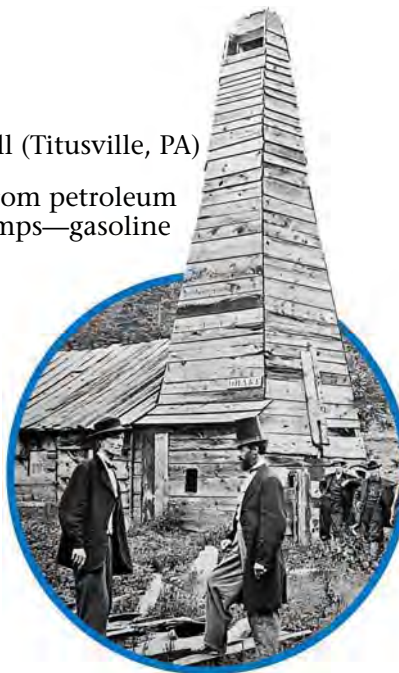
Catalytic converters on cars

1980

1975–1990

Leaded gas phased out

1990





Coal being strip-mined

Coal

Coal was formed from the remains of plants that were buried and subjected to high pressure and heat over long periods of time. Plant materials have a high content of cellulose, a complex molecule whose empirical formula is CH_2O but whose molar mass is approximately 500,000 g/mol. After the plants and trees that grew on the earth at various times and places died and were buried, chemical changes gradually lowered the oxygen and hydrogen content of the cellulose molecules. Coal “matures” through four stages: lignite, subbituminous, bituminous, and anthracite. Each stage has a higher carbon-to-oxygen and carbon-to-hydrogen ratio; that is, the relative carbon content gradually increases. Typical elemental compositions of the various coals are given in **Table 10.4**. The energy available from the combustion of a given mass of coal increases as the carbon content increases. Anthracite is the most valuable coal, and lignite is the least valuable.

Coal is an important and plentiful fuel in the United States, currently furnishing approximately 20% of our energy. As the supply of petroleum decreases, the share of the energy supply from coal could eventually increase to as high as 30%. However, coal is expensive and dangerous to mine underground, and the strip-mining of fertile farmland in the Midwest or of scenic land in the West causes obvious problems. In addition, the burning of coal, especially high-sulfur coal, yields air pollutants such as sulfur dioxide, which, in turn, can lead to acid rain. However, even if coal were pure carbon, the carbon dioxide produced when it was burned would still have significant effects on the earth’s climate.

Effects of Carbon Dioxide on Climate

The earth receives a tremendous quantity of radiant energy from the sun, about 30% of which is reflected back into space by the earth’s atmosphere. The remaining energy passes through the atmosphere to the earth’s surface. Some of this energy is absorbed by plants for photosynthesis and some by the oceans to evaporate water, but most of it is absorbed by soil, rocks, and water, increasing the temperature of the earth’s surface. This energy is, in turn, radiated from the heated surface mainly as *infrared radiation*, often called *heat radiation*.

Coal

A solid fossil fuel mostly consisting of carbon

Table 10.4

Element Composition of Various Types of Coal

Type of Coal	Mass Percent of Each Element				
	C	H	O	N	S
Lignite	71	4	23	1	1
Subbituminous	77	5	16	1	1
Bituminous	80	6	8	1	5
Anthracite	92	3	3	1	1

DID YOU KNOW

Coal has variable composition depending on both its age and location.

Greenhouse effect

The process by which an atmosphere warms a planet

Greenhouse Effect The atmosphere, like window glass, is transparent to visible light but does not allow all the infrared radiation to pass back into space. Molecules in the atmosphere, principally H_2O and CO_2 , strongly absorb infrared radiation and radiate it back toward the earth, as shown in **Figure 10.7**. A net amount of thermal energy is retained by the earth's atmosphere, causing the earth to be much warmer than it would be without its atmosphere. In a way, the atmosphere acts like the glass of a greenhouse, which is transparent to visible light but absorbs infrared radiation, thus raising the temperature inside the building. This **greenhouse effect** is seen even more spectacularly on Venus, where the dense atmosphere is thought to be responsible for the high surface temperature of that planet.

Thus the temperature of the earth's surface is controlled to a significant extent by the carbon dioxide and water content of the atmosphere. The effect of atmospheric moisture (humidity) is readily apparent in the Midwest, for example. In summer, when the humidity is high, the heat of the sun is retained well into the night, giving very high nighttime temperatures. In winter, the coldest temperatures always occur on clear nights, when the low humidity allows efficient radiation of energy back into space.

DID YOU KNOW

The average temperature of the earth's surface is 298 K. It would be 255 K without the "greenhouse gases."

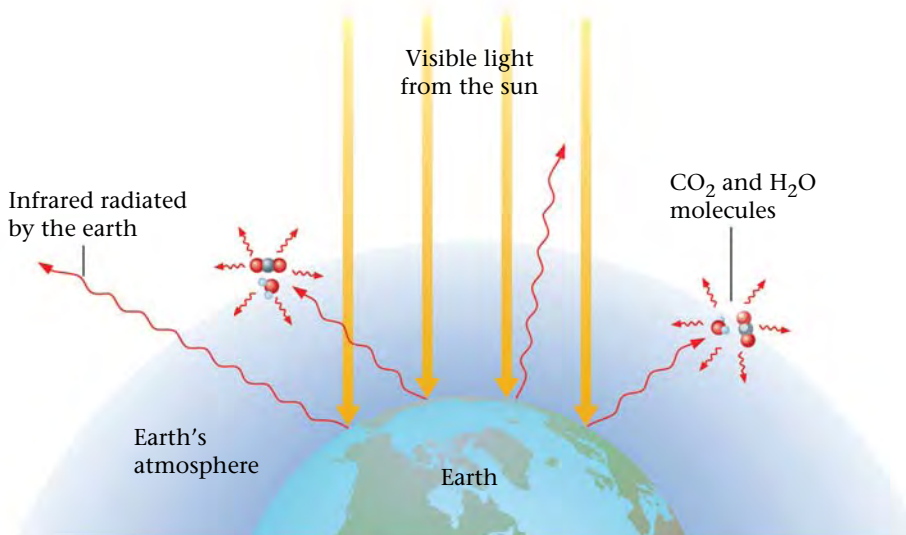


Active Reading Question

How does carbon dioxide act as a "greenhouse gas"?

Figure 10.7

The earth's atmosphere is transparent to visible light from the sun. This visible light strikes the earth, and part of it is changed to infrared radiation. The infrared radiation from the earth's surface is strongly absorbed by CO_2 , H_2O , and other molecules present in smaller amounts (for example, CH_4 and N_2O) in the atmosphere. In effect, the atmosphere traps some of the energy, acting like the glass in a greenhouse and keeping the earth warmer than it would otherwise be.



Atmospheric CO_2 The atmosphere's water content is controlled by the water cycle (evaporation and precipitation), and the average has remained constant over the years. However, as fossil fuels have been used more extensively, the carbon dioxide concentration has increased—up about 20% from 1880 to the present. Projections indicate that the carbon dioxide content of the atmosphere may be double in the twenty-first century what it was in 1880. This trend *could* increase the earth's average temperature by as much as $10\text{ }^\circ\text{C}$, causing dramatic changes in climate and greatly affecting the growth of food crops.

How well can we predict the long-term effects of carbon dioxide? Because weather has been studied for a period of time that is minuscule compared with the age of the earth, the factors that control the earth's climate in the long range are not clearly understood. For example, we do not understand what causes the earth's periodic ice ages. So it is difficult to estimate the effects of the increasing carbon dioxide levels.

Although the exact relationship between the carbon dioxide concentration in the atmosphere and the earth's temperature is not known at present, one thing is clear: The increase in the atmospheric concentration of carbon dioxide is quite dramatic (see **Figure 10.8**). We must consider the implications of this increase as we consider our future energy needs.

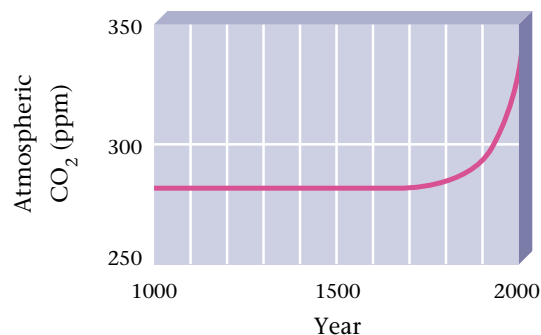


Figure 10.8

The atmospheric CO₂ concentration over the past 1000 years, based on ice core data and direct readings (since 1958). Note the dramatic increase in the past 100 years.

CHEMISTRY IN YOUR WORLD

Consumer Connection

Veggie Power?

Gasoline usage is as high as ever, and world petroleum supplies will eventually run out. One possible alternative to petroleum as a source of fuels and lubricants is vegetable oil—the same vegetable oil we now use to cook french fries. Researchers believe that the oils from soybeans, corn, canola, and sunflowers all have the potential to be used in cars as well as on salads.

The use of vegetable oil for fuel is not a new idea. Rudolf Diesel reportedly used peanut oil to run one of his engines at the Paris Exposition in 1900. In addition, ethyl alcohol has been used widely as a fuel in South America and as a fuel additive in the United States.

Biodiesel, a fuel made from vegetable oil, has some real advantages over regular diesel fuel. Biodiesel produces fewer pollutants. Also, because vegetable oils have no sulfur, there is no noxious sulfur dioxide in the exhaust gases. Biodiesel can run in existing engines with little modification. In addition, it is much more biodegradable than petroleum-based fuels, so spills cause less environmental damage.

Of course, biodiesel has some drawbacks. Biodiesel produces more nitrogen oxides in the exhaust than conventional diesel fuel and is less stable in storage. It can leave more gummy deposits in engines and

engines must be “winterized” by removing components that tend to solidify at low temperatures.

The best solution may be to use biodiesel as an additive to regular diesel fuel. One such fuel is known as B20 because it is 20% biodiesel and 80% conventional diesel fuel. B20 is especially attractive because of the higher lubricating ability of vegetable oils, which reduces diesel engine wear.

Vegetable oils are also being considered as replacements for motor oils and hydraulic fluids. Veggie oil fuels and lubricants seem to have a growing market as petroleum supplies decrease and as environmental laws become more severe. In Germany's Black Forest region, for example, environmental protection laws require that farm equipment use only vegetable oil fuels and lubricants. In the near future there may be veggie oil in your garage as well as in your kitchen.

This promotion bus both advertises biodiesel and demonstrates its usefulness.

Adapted from “Fill ‘Er Up . . . with Veggie Oil,” by Corinna Wu, as appeared in Science News, Vol 154, December 5, 1998, p. 364.

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New Energy Sources

CRITICAL THINKING ?

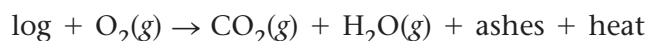
A government study concludes that burning fossil fuels to power our automobiles causes too much pollution.

What if Congress decided that all cars and trucks must be powered by batteries? Would this change solve the air pollution problems caused by transportation?

As we search for the energy sources of the future, we need to consider economic, climatic, and supply factors. There are several potential energy sources: the sun (solar), nuclear processes (fission and fusion), biomass (plants), wind, and synthetic fuels. Direct use of the sun's radiant energy to heat our homes and run our factories and transportation systems seems a sensible long-term goal. But what can we do now? Conservation of fossil fuels is one obvious step, but substitutes for fossil fuels also must be found. There is much research going on now to solve this problem.

C. Energy as a Driving Force

A major goal of science is to understand why things happen as they do. In particular, we are interested in the driving forces of nature. Why do things occur in a particular direction? For example, consider a log that has burned in a fireplace, producing ashes and heat energy. If you are sitting in front of the fireplace, you would be very surprised to see the ashes begin to absorb heat from the air and reconstruct themselves into the log. It just doesn't happen. That is, the process that always occurs is

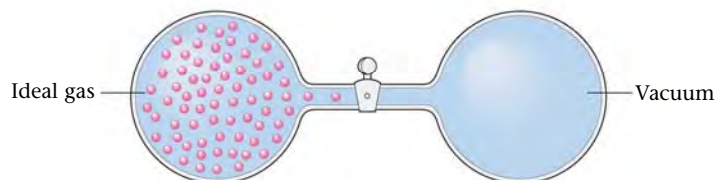


The reverse of this process

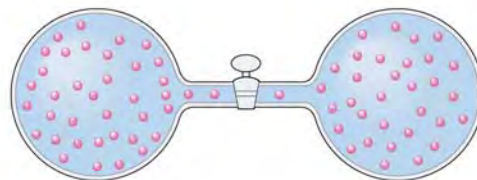


never happens.

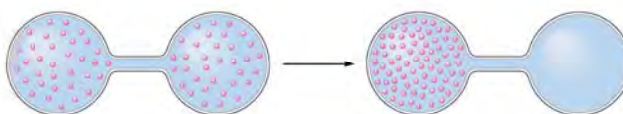
Consider another example. A gas is trapped in one end of a vessel as shown below.



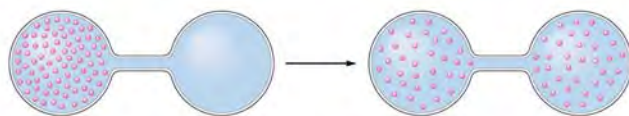
When the valve is opened, what always happens? The gas spreads evenly throughout the entire container.



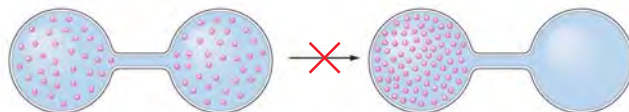
You would be very surprised to see the following process occur spontaneously:



So, why does this process



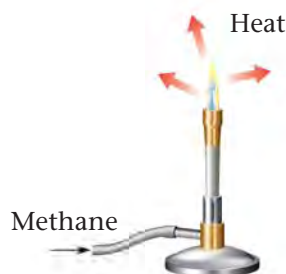
occur spontaneously but the reverse process never occur?



In many years of analyzing these and many other processes, scientists have discovered two very important driving forces:

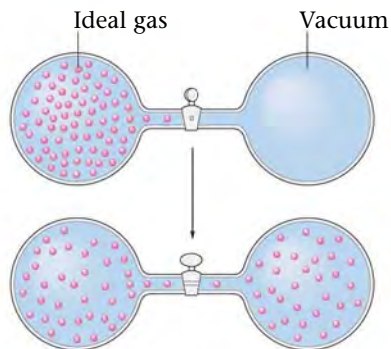
- Energy spread
- Matter spread

Energy spread means that in a given process, concentrated energy is dispersed widely. This distribution happens every time an exothermic process occurs. For example, when a Bunsen burner burns, the energy stored in the fuel (natural gas—mostly methane) is dispersed into the surrounding air:



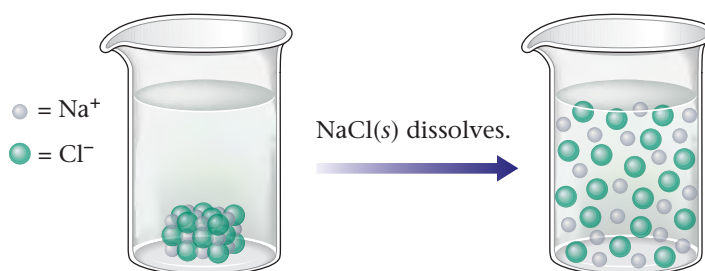
The energy that flows into the surroundings through heat increases the thermal motions of the molecules in the surroundings. In other words, this process increases the random motions of the molecules in the surroundings. *This always happens in every exothermic process.*

Matter spread means exactly what it says: the molecules of a substance are spread out and occupy a larger volume.



After looking at thousands of processes, scientists have concluded that these two factors are the important driving forces that cause events to occur. That is, processes are favored if they involve energy spread and matter spread.

Do these driving forces ever occur in opposition? Yes, they do—in many, many processes. For example, consider ordinary table salt dissolving in water.



This process occurs spontaneously. You observe it every time you add salt to water to cook potatoes or pasta. Surprisingly, dissolving salt in water is *endothermic*. This process seems to go in the wrong direction—it involves energy concentration, not energy spread. Why does the salt dissolve? Because of matter spread. The Na^+ and Cl^- that are closely packed in the solid NaCl become spread around randomly in a much larger volume in the resulting solution. Salt dissolves in water because the favorable matter spread overcomes an unfavorable energy change.

Entropy

Entropy is a function we have invented to keep track of the natural tendency for the components of the universe to become disordered. Entropy (designated by the letter S) is a measure of disorder or randomness. As randomness increases, S increases. Which has lower entropy, solid water (ice) or gaseous water (steam)? Remember that ice contains closely packed, ordered H_2O molecules, and steam has widely dispersed, randomly moving H_2O molecules (see **Figure 10.9**). Thus ice has more order and a lower value of S .

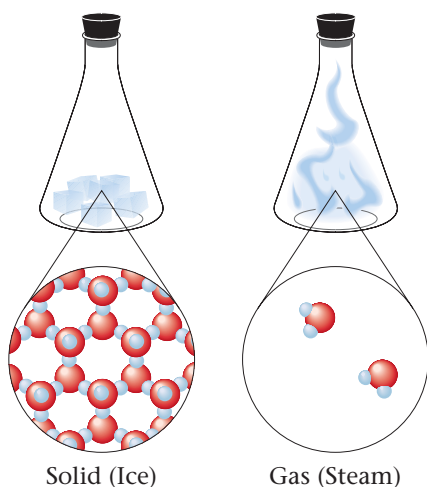


Figure 10.9
Comparing the entropies of ice and steam

Active Reading Question

What is the driving force for water to freeze below $0\text{ }^\circ\text{C}$? What is the driving force for ice to melt above $0\text{ }^\circ\text{C}$?

CHEMISTRY IN YOUR WORLD

Chemical Impact

Farming the Wind

In the Midwest the wind blows across fields of corn, soybeans, wheat, and wind turbines—wind turbines? It turns out that the wind that seems to blow almost continuously across the plains is now becoming the latest cash crop.

There is plenty of untapped wind power in the United States. Wind mappers rate regions on a scale of 1 to 6 (with 6 being the best) to indicate the quality of the wind resource. Wind farms are now being developed in areas rated from 4 to 6. The farmers who own the land welcome the increased income derived from the wind blowing across their land. Economists estimate that each acre devoted to wind turbines can pay royalties to the farmers of as much as \$8000 per year, or many times the revenue from growing corn on that same land. Globally, wind generation of electricity has nearly quadrupled in the last five years and is expected to increase by about 60% per year in the United States. The economic feasibility of wind-generated electricity has greatly improved in the last 30 years as the wind turbines have become more efficient. Today's

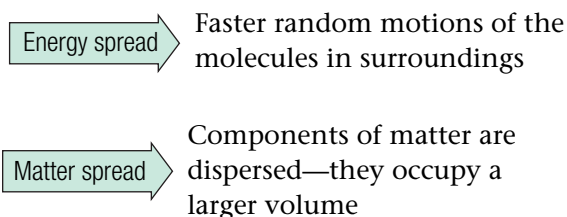
turbines can produce electricity that costs about the same as that from other sources. The most impressive thing about wind power is the magnitude of the supply. According to the American Wind Energy Association in Washington, D.C., the wind power potential in the United States is comparable to or larger than the energy resources under the sands of Saudi Arabia.

Within a few years wind power could be a major source of electricity. There could be a fresh wind blowing across the energy landscape of the United States in the near future.

This State Line Wind Project along the Oregon-Washington border uses approximately 399 wind turbines to create enough electricity to power about 70,000 households.

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What do you suppose happens to the disorder of the universe as energy spread and matter spread occur during a process?



It seems clear that both energy spread and matter spread lead to greater entropy (greater disorder) in the universe. This idea leads to a very important conclusion that is summarized in the **second law of thermodynamics**:

The entropy of the universe is always increasing.

That is, all processes that occur in the universe lead to a net increase in the disorder of the universe. As the universe “runs,” it is always heading toward more disorder. We are plunging slowly but inevitably toward total randomness—the heat death of the universe. But don't despair, it will not happen soon.

CRITICAL THINKING

What if the first law of thermodynamics were true, but the second law were not? How would the world be different?

**Distribution of Wealth****Materials**

- 20 pennies
- Graph paper

Procedure

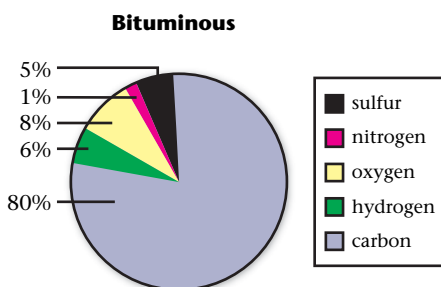
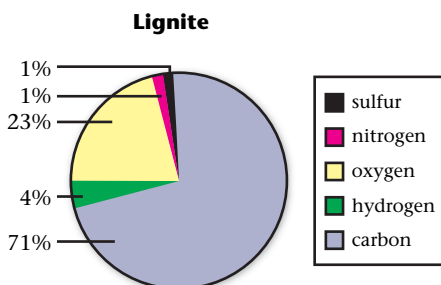
1. Use 20 pennies. Flip each penny and record the number of heads (head count). Do this 30 times.
2. Determine the frequency of each head count by counting the number of times each different head count appears.

Results/Analysis

1. Construct a bar graph of frequency versus head count.
2. Obtain data from the other groups in your class and make a bar graph of frequency versus head count for all of the data.
3. How can these graphs be related to probability?

**SECTION 10.4****REVIEW QUESTIONS**

- 1 If energy is conserved how can there be an “energy crisis”?
- 2 Use the charts below to answer the following questions.
- 3 What does the “heat death” of the universe have to do with quality of energy?
- 4 What are three uses of petroleum fractions?
- 5 What is the greenhouse effect and what are the key molecules which cause it?
- 6 What driving force must be predominant in order for an endothermic reaction to occur? Why?
- 7 What is meant by the term *entropy*? Which has higher entropy, a solid or a gas of a particular compound at the same temperature? Why?



- a. Which type of coal contains the larger percent mass of carbon?
- b. Which type of coal contains the larger percent mass of sulfur?
- c. Which element is present in equal amounts in both types of carbon?

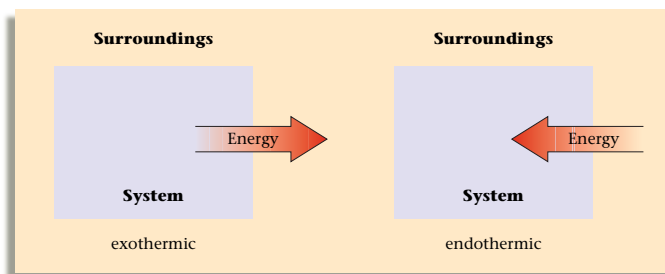
Key Terms

10.1 Energy
Potential energy
Kinetic energy
Law of conservation of energy
Work
State function
Temperature
Heat
System
Surroundings
Exothermic
Endothermic

Key Ideas

10.1 Energy, Temperature, and Heat

- Energy is conserved.
- The law of conservation of energy states that energy is neither created or destroyed in any process.
- In a process, energy can be changed from one form to another but the amount of energy remains constant.
- Thermodynamics is the study of energy and its changes.
- Energy is classified as one of the following:
 - Kinetic energy—energy due to the motion of an object
 - Potential energy—energy due to the position or composition of an object
- Some functions, called state functions, depend only on the beginning and final states of the system, not on the specific pathway followed.
 - Energy is a state function.
 - Heat and work are not state functions.
- Temperature indicates the vigor of the random motions of the components of that substance.
- Thermal energy is the sum of the energy produced by the random motions of the components.
- Heat is a flow of energy between two objects due to a temperature difference between the objects.
 - An exothermic process is one in which energy as heat flows out of the system into the surroundings.
 - An endothermic process is one in which energy as heat flows into the system from the surroundings.
 - The common units for heat are calories and joules.



10.2 Thermodynamics

First law of thermodynamics

Internal energy

Calorie

Joule

10.3 Enthalpy

Calorimeter

Hess's law

10.4 Fossil fuels

Petroleum

Natural gas

Coal

Greenhouse effect

Energy spread

Matter spread

Entropy

Second law of thermodynamics

10.2 The Flow of Energy

- Internal energy (E) is the sum of the kinetic and potential energy associated with an object.
 - Internal energy (E) can be changed by two types of energy flow:
 - ◆ Heat (q)
 - ◆ Work (w)
 - $\Delta E = q + w$
- Specific heat capacity is the energy required to change the temperature of a mass of one gram of a substance by one Celsius degree.

10.3 Energy and Chemical Reactions

- For a process carried out at constant pressure, the change in enthalpy (ΔH) of that process is equal to the heat.

Characteristics of ΔH

- If a reaction is reversed, the sign of ΔH is also reversed.
- The magnitude of ΔH is directly proportional to the quantities of reactants and products in a reaction. If the coefficients in a balanced reaction are multiplied by an integer, the value of ΔH is multiplied by the same integer.

- Hess's law enables the calculation of the heat for a given reaction from known heats of related reactions.
- A calorimeter is a device used to measure the heat associated with a given chemical reaction.

10.4 Using Energy in the Real World

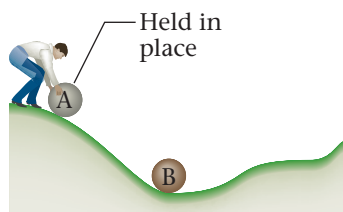
- Although energy is conserved in every process, the quality (usefulness) of the energy decreases in every real process.
- Natural processes occur in the direction that leads to an increase in the disorder (increase in entropy) of the universe.
 - The principal driving forces for natural processes can be described in terms of "matter spread" and "energy spread."
- Our world has many sources of energy. The use of this energy affects the environment in various ways.

All exercises with **blue numbers** have answers in the back of this book.

10.1 Energy, Temperature, and Heat

A. The Nature of Energy

1. Explain the difference between kinetic and potential energy.
2. Why isn't all energy available for work?
3. Explain what is meant by the term *state function*. Provide examples of state functions.
4. Which ball has the higher potential energy? Explain.



5. The law of conservation of energy means that energy is a state function. Explain why.

B. Temperature and Heat

6. Explain the differences among heat, temperature, and thermal energy.
7. Do each of the following depend on the amount of substance you have? Explain.
 - a. temperature
 - b. thermal energy
8. Provide a molecular-level explanation of why the temperatures of a cold soft drink and hot coffee in the same room will eventually be the same.
9. In which case is more heat involved: mixing 100.0-g samples of 90 °C water and 80 °C water or mixing 100.0-g samples of 60 °C water and 10 °C water? Assume no heat is lost to the environment.
10. If 100.0 g of water at 90 °C is added to 50.0 g of water at 10 °C, estimate the final temperature of the water. Explain your reasoning.

C. Exothermic and Endothermic Processes

11. What is meant by potential energy in a chemical reaction? Where is it located?
12. What does it mean for a chemical to have a low potential energy?
13. In an endothermic reaction, do the reactants or the products have the lower potential energy?

14. Are the following processes exothermic or endothermic?
 - a. When solid KBr is dissolved in water, the solution gets colder.
 - b. Natural gas (CH₄) is burned in a furnace.
 - c. When concentrated sulfuric acid is added to water, the solution gets very hot.
 - d. Water is boiled in a teakettle.

10.2 The Flow of Energy

A. Thermodynamics

15. What does it mean when energy is reported with a negative sign? A positive sign?
16. What does it mean when work is reported with a negative sign? A positive sign?
17. In thermodynamics, the chemist takes the system's point of view. Explain.
18. How do engineers define *energy* and *work* differently from chemists? Why do engineers take this approach?
19. Calculate ΔE for each of the following cases.
 - a. $q = +51 \text{ kJ}$, $w = -15 \text{ kJ}$
 - b. $q = +100. \text{ kJ}$, $w = -65 \text{ kJ}$
 - c. $q = -65 \text{ kJ}$, $w = -20. \text{ kJ}$
20. A gas absorbs 45 kJ of heat and does 29 kJ of work. Calculate ΔE .
21. A system releases 125 kJ of heat, and 104 kJ of work is done on it. Calculate ΔE .

B. Measuring Energy Changes

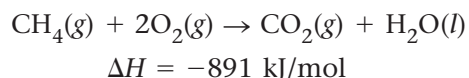
22. Describe what happens to the molecules in a sample of ice as the sample is slowly heated until it liquefies and then vaporizes.
23. Metallic substances tend to have (lower/higher) specific heat capacities than nonmetallic substances.
24. If it takes 526 J of energy to warm 7.40 g of water by 17 °C, how much energy would be needed to warm 7.40 g of water by 55 °C?
25. Convert the following numbers of calories or kilocalories into joules or kilojoules.
 - a. 7845 cal
 - b. $4.55 \times 10^4 \text{ cal}$
 - c. 62.142 kcal
 - d. 43,024 cal

- 26.** If 7.24 kJ of heat is applied to a 952-g block of metal, the temperature increases by 10.7 °C. Calculate the specific heat capacity of the metal in J/g °C.
- 27.** Three 75.0-g samples of copper, silver, and gold are available. Each of these samples is initially at 24.0 °C, and then 2.00 kJ of heat is applied to each sample. Which sample will end up at the highest temperature?
- 28.** A 35.2-g sample of metal X requires 1251 J of energy to heat the sample by 25.0 °C. Calculate the specific heat capacity of this metal.

10.3 Energy and Chemical Reactions

A. Thermochemistry (Enthalpy)

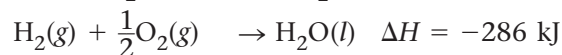
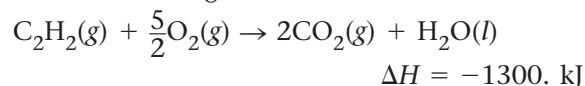
- 29.** The equation for the fermentation of glucose to alcohol and carbon dioxide is
- $$\text{C}_6\text{H}_{12}\text{O}_6(aq) \rightarrow 2\text{C}_2\text{H}_5\text{OH}(aq) + 2\text{CO}_2(g)$$
- The enthalpy change for the reaction is -67 kJ. Is the reaction endothermic or exothermic? Is energy, in the form of heat, absorbed or released as the reaction occurs?
- 30.** For the reaction
- $$\text{S}(s) + \text{O}_2(g) \rightarrow \text{SO}_2(g) \quad \Delta H = -296 \text{ kJ/mol}$$
- How much heat is released when 275 g of sulfur is burned in excess oxygen?
 - How much heat is released when 25 mol of sulfur is burned in excess oxygen?
 - How much heat is released when 150. g of sulfur dioxide is produced?
- 31.** Calculate the enthalpy change when 1.00 g of methane is burned in excess oxygen according to the reaction



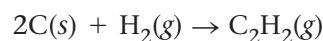
B. Hess's Law

- 32.** Given the following data:
- $$\text{S}(s) + \frac{3}{2}\text{O}_2(g) \rightarrow \text{SO}_3(g) \quad \Delta H = -395.2 \text{ kJ}$$
- $$2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g) \quad \Delta H = -198.2 \text{ kJ}$$
- calculate ΔH for the reaction
- $$\text{S}(s) + \text{O}_2(g) \rightarrow \text{SO}_2(g)$$

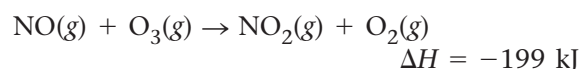
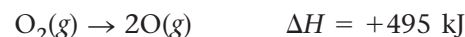
- 33.** Given the following data:



calculate ΔH for the reaction



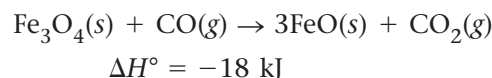
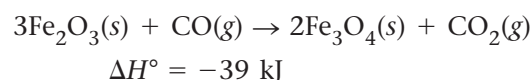
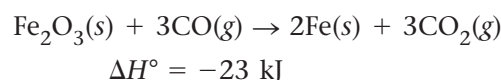
- 34.** Given the following data:



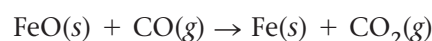
calculate ΔH for the reaction



- 35.** Given the following data:



calculate ΔH° for the reaction



10.4 Using Energy in the Real World

A. Quality Versus Quantity of Energy

- 36.** What is the difference between the quality of energy and the quantity of energy? Which is decreasing?
- 37.** Why can no work be done when everything in the universe is at the same temperature?

B. Energy and Our World

- 38.** What was the advantage of using tetraethyl lead in gasoline? What were two disadvantages?
- 39.** Why do we need some greenhouse gases? What is the problem with having too much of the greenhouse gases?
- 40.** Which energy sources used in the United States have declined the most in the last 150 years? Which have increased the most?

C. Energy as a Driving Force

41. Why can't the first law of thermodynamics explain why a ball doesn't spontaneously roll up a hill?
42. What is meant by the term *driving force*?

Critical Thinking

45. Consider a sample of *steam* (water in the gaseous state) at 150 °C. Describe what happens to the molecules in the sample as the sample is slowly cooled until it liquefies and then solidifies.
46. Convert the following numbers of kilojoules into kilocalories. (Remember: kilo means 1000.)
- 462.4 kJ
 - 18.28 kJ
 - 1.014 kJ
 - 190.5 kJ
47. Perform the indicated conversions.
- 45.62 kcal into kilojoules
 - 72.94 kJ into kilocalories
 - 2.751 kJ into calories
 - 5.721 kcal into joules
48. Calculate the amount of energy required (in calories) to heat 145 g of water from 22.3 °C to 75.0 °C.
49. It takes 1.25 kJ of energy to heat a certain sample of pure silver from 12.0 °C to 15.2 °C. Calculate the mass of the sample of silver.
50. If 50. J of heat is applied to 10. g of iron, by how much will the temperature of the iron increase? (See Table 10.1.)
51. The specific heat capacity of gold is 0.13 J/g °C. Calculate the specific heat capacity of gold in cal/g °C.
52. Calculate the amount of energy required (in joules) to heat 2.5 kg of water from 18.5 °C to 55.0 °C.
53. If 10. J of heat is applied to 5.0-g samples of each of the substances listed in Table 10.1, which substance's temperature will increase the most? Which substance's temperature will increase the least?
43. What is the driving force of an exothermic reaction—matter spread or energy spread? Explain.
44. Exothermic reactions have a driving force. Nevertheless, water melting into a liquid is endothermic and this process occurs at room conditions. Explain why.
54. A 50.0-g sample of water at 100. °C is poured into a 50.0-g sample of water at 25 °C. What will be the final temperature of the water?
55. A 25.0-g sample of pure iron at 85 °C is dropped into 75 g of water at 20. °C. What is the final temperature of the water-iron mixture?
56. If it takes 4.5 J of energy to warm 5.0 g of aluminum from 25 °C to a certain higher temperature, then it will take _____ J to warm 10. g of aluminum over the same temperature interval.
57. For each of the substances listed in Table 10.1, calculate the quantity of heat required to heat 150. g of the substance by 11.2 °C.
58. Suppose you have 10.0-g samples of each of the substances listed in Table 10.1 and that 1.00 kJ of heat is applied to each of these samples. By what amount would the temperature of each sample be raised?
59. Calculate ΔE for each of the following.
- $q = -47 \text{ kJ}$, $w = +88 \text{ kJ}$
 - $q = +82 \text{ kJ}$, $w = +47 \text{ kJ}$
 - $q = +47 \text{ kJ}$, $w = 0$
 - In which of these cases do the surroundings do work on the system?
60. Are the following processes exothermic or endothermic?
- the combustion of gasoline in a car engine
 - water condensing on a cold pipe
 - $\text{CO}_2(s) \rightarrow \text{CO}_2(g)$
 - $\text{F}_2(g) \rightarrow 2\text{F}(g)$

61. The overall reaction in commercial heat packs can be represented as



- How much heat is released when 4.00 mol iron is reacted with excess O_2 ?
- How much heat is released when 1.00 mol Fe_2O_3 is produced?
- How much heat is released when 1.00 g iron is reacted with excess O_2 ?
- How much heat is released when 10.0 g Fe and 2.00 g O_2 are reacted?

62. Consider the following equations:



Suppose the first equation is reversed and multiplied by $\frac{1}{6}$, the second and third equations are divided by 2, and the three adjusted equations are added. What is the net reaction and what is the overall heat of this reaction?

63. It has been determined that the body can generate 5500 kJ of energy during one hour of strenuous exercise. Perspiration is the body's mechanism for eliminating this heat. How many grams and how many liters of water would have to be evaporated through perspiration to rid the body of the heat generated during two hours of exercise? (The heat of vaporization of water is 40.6 kJ/mol.)

64. One way to lose weight is to exercise. Walking briskly at 4.0 miles per hour for an hour consumes about 400 kcal of energy. How many miles would you have to walk at 4.0 miles per hour to lose one pound of body fat? One gram of body fat is equivalent to 7.7 kcal of energy. There are 454 g in 1 lb.

65. You are given a metal and asked to determine its identity. You are to do this by determining the specific heat of the metal. You place the metal in a boiling water bath for a few minutes and then transfer the metal to a 100.0 g sample of water at a measured temperature. You then record the highest temperature of the water. The data are given in the table below.

	Data
Mass water	100.0 g
Initial temperature of the water	21.31 °C
Final temperature of the water	24.80 °C
Mass metal	50.0 g

- Given that the specific heat capacity of water is 4.184 J/g °C, calculate the amount of heat that is absorbed by the water.
- How does the heat absorbed by the water compare to the heat lost by the metal?
- What is the initial temperature of the "hot" metal? What is the final temperature of the metal?
- Calculate the specific heat capacity of the metal.
- Given the following specific heat capacities, determine the identity of your metal.

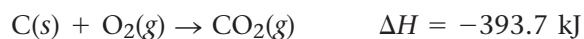
Tin: 0.227 J/g °C

Zinc: 0.388 J/g °C

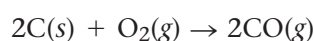
Aluminum: 0.891 J/g °C

- 1 Energy of position is also called
- kinetic energy
 - potential energy
 - heat energy
 - enthalpy
- 2 Temperature is
- a measure of the potential energy stored in a substance.
 - the same as heat.
 - a measure of the random motions of the particles in a system.
 - a state function.
- 3 Of energy, enthalpy, heat, and work, how many are state functions?
- 1
 - 2
 - 3
 - 4
- 4 Which of the following processes is exothermic?
- Candle wax melting
 - A puddle evaporating
 - Dry ice (solid carbon dioxide) subliming to form gaseous carbon dioxide
 - Water freezing to form ice
- 5 Metals generally have relatively _____ specific heat capacities, meaning they respond rather _____ to cooling or heating.
- low, quickly
 - low, slowly
 - high, slowly
 - high, quickly
- 6 Calculate the heat change when 45.0 g of water cools from 45 °C to 22 °C. The specific heat capacity of water is 4.184 J/g °C.
- 1.04 kJ
 - 4.33 kJ
 - 4.33 kJ
 - 1.04 kJ

- 7 Given the following data:

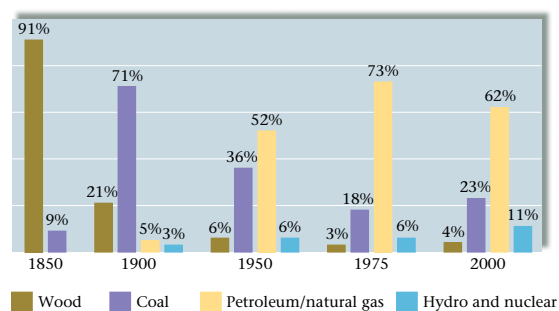


calculate ΔH for the reaction



- 172.9 kJ
 - 172.9 kJ
 - 960.3 kJ
 - 220.8 kJ
- 8 Why is there an energy crisis in the world?
- Because the law of conservation of energy states that energy cannot be created nor destroyed.
 - Because the quantity of energy in the world is always decreasing.
 - Because the quality of energy in the world is always decreasing.
 - Because energy is a state function.
- 9 Given that the releasing of heat is a driving force for reactions, why do endothermic reactions take place at all?
- Because in an endothermic reaction, heat is released.
 - In order for an endothermic reaction to occur, matter spread must occur as well.
 - Endothermic reactions are always spontaneous.
 - Endothermic reactions never occur naturally.

For questions 10 and 11, use the following graph on energy usage.



- 10 Which source of energy has had the largest increase in usage since 1900?
- 11 Which source of energy is no longer used to a great extent?

Modern Atomic Theory

LOOKING AHEAD

11.1 Atoms and Energy

- A. Rutherford's Atom
- B. Energy and Light
- C. Emission of Energy by Atoms

11.2 The Hydrogen Atom

- A. The Energy Levels of Hydrogen
- B. The Bohr Model of the Atom
- C. The Wave Mechanical Model of the Atom

11.3 Atomic Orbitals


- A. The Hydrogen Orbitals
 - Hydrogen Energy Levels
 - Hydrogen Orbitals
- B. The Wave Mechanical Model: Further Development

11.4 Electron Configurations and Atomic Properties

- A. Electron Arrangements in the First 18 Atoms on the Periodic Table
- B. Electron Configurations and the Periodic Table
 - Summary of the Wave Mechanical Model and Valence-Electron Configurations
- C. Atomic Properties and the Periodic Table
 - Metals and Nonmetals
 - Atomic Size
 - Ionization Energies

The colors in this fireworks display are the result of electron transfers between atomic energy levels.



IN YOUR LIFE

A fireworks display gives us a dramatic example of chemistry in action. Chemical reactions are used to provide the energy to propel the fireworks into the air, and other reactions are used for the explosions in the air. Of course, what most people remember are the beautiful and various colors that go along with the smoke and the sounds. But where do these colors come from?

It turns out that many are due to the positively charged ions (cations) of ionic solids (salts). For example, sodium salts give us yellow colors, strontium salts give us red colors, and barium salts give us green colors. But why do these ions display colors at all? And why do they display different colors? The answers have to do with the unique arrangement of the electrons in a given atom.

Until now we have dealt with rather simple models of the atom. For example, Dalton's model did not include electrons at all. In order to answer our questions about the colors of fireworks, we need to develop a more complicated model of the atom, and especially of the electrons in the atom. We have not yet looked deeply at the structure of an atom, but that is what this chapter is all about.

WHAT DO YOU KNOW? **Prereading Questions**

1. What particles make up an atom?
2. Recall Rutherford's experiment from Chapter 3. What was the experiment and what did it show?
3. What does the term *quantum leap* mean?

SECTION 11.1

Atoms and Energy

Key Terms

- Electromagnetic radiation
- Wavelength
- Frequency
- Photons

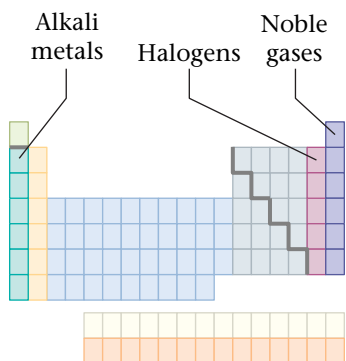
Objectives

- To describe Rutherford's model of the atom
- To explore the nature of electromagnetic radiation
- To see how atoms emit light

The concept of atoms is a very useful one. It explains many important observations, such as why compounds always have the same composition (a specific compound always contains the same types and numbers of atoms) and how chemical reactions occur (they involve a rearrangement of atoms).

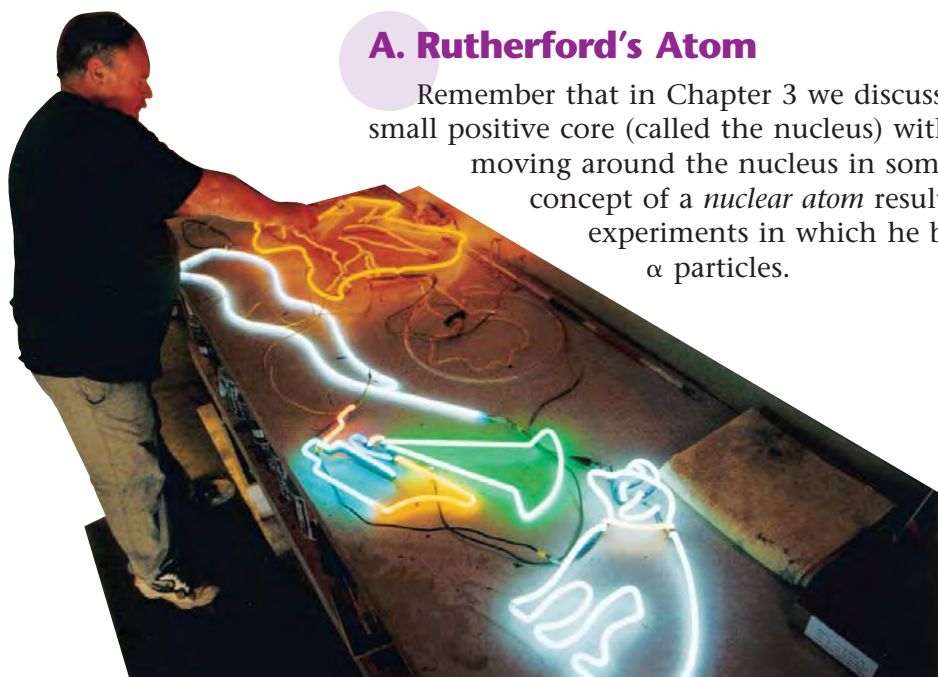
Once chemists came to “believe” in atoms, logical questions followed: What are atoms like? What is the structure of an atom? In Chapter 3 we learned to picture the atom with a positively charged nucleus composed of protons and neutrons at its center and electrons moving around the nucleus in a space very large compared to the size of the nucleus.

In this chapter, we will look at atomic structure in more detail. In particular, we will develop a picture of the electron arrangements in atoms—a picture that allows us to account for the chemistry of the various elements. Recall from our discussion of the periodic table in Chapter 3 that, although atoms exhibit a great variety of characteristics, certain elements can be grouped together because they behave similarly. For example, fluorine, chlorine, bromine, and iodine (the halogens) show great chemical similarities. Likewise lithium, sodium, potassium, rubidium, and cesium (the alkali metals) exhibit many similar properties, and helium, neon, argon, krypton, xenon, and radon (the noble gases) are all very nonreactive. Although the members of each of these groups of elements show great similarity *within* the group, the differences in behavior *between* groups are striking. In this chapter we will see that it is the way the electrons are arranged in various atoms that accounts for these facts. However, before we examine atomic structure, we must consider the nature of electromagnetic radiation, which plays a central role in the study of the atom's behavior.



A. Rutherford's Atom

Remember that in Chapter 3 we discussed the idea that an atom has a small positive core (called the nucleus) with negatively charged electrons moving around the nucleus in some way (see **Figure 11.1**). This concept of a *nuclear atom* resulted from Ernest Rutherford's experiments in which he bombarded metal foil with α particles.



Neon signmaker and artist Jess Baird shows off a few of the items he has made in his Weatherford, Texas, shop.

Let's Review

Nuclear Model of the Atom

- The atom has a small dense nucleus which
 - is positively charged
 - contains protons (+1 charge)
 - contains neutrons (no charge)
- The remainder of the atom
 - Is mostly empty space
 - Contains electrons (−1 charge)

A major question left unanswered by Rutherford's work was, "What are the electrons doing?" That is, how are the electrons arranged and how do they move? Rutherford suggested that electrons might revolve around the nucleus like the planets revolve around the sun in our solar system. He couldn't explain, however, why the negative electrons aren't attracted into the positive nucleus, causing the atom to collapse.

At this point it became clear that more observations of the properties of atoms were needed to understand the structure of the atom more fully. To help us understand these observations, we need to discuss the nature of light and how it transmits energy.

B. Energy and Light

If you hold your hand a few inches from a brightly glowing light bulb, what do you feel? Your hand gets warm. The "light" from the bulb somehow transmits energy to your hand. The same thing happens if you move close to the glowing embers of wood in a fireplace—you receive energy that makes you feel warm. The energy you feel from the sun is a similar example.

In all three of these instances, energy is being transmitted from one place to another by light—more properly called **electromagnetic radiation**. Many kinds of electromagnetic radiation exist. Examples include

- X rays used to make images of bones
- "white" light from a light bulb
- microwaves used to cook hot dogs and other food
- radio waves that transmit voices and music.

How do these various types of electromagnetic radiation differ from one another? To answer this question we need to talk about waves. To explore the characteristics of waves, let's think about ocean waves. In **Figure 11.2** a seagull is shown floating on the ocean and being raised and lowered by the motion of the water surface as waves pass by. Notice that the gull just moves up and down as the waves pass—it is not moved forward. A particular wave is characterized by three properties: *wavelength*, *frequency*, and *speed*.

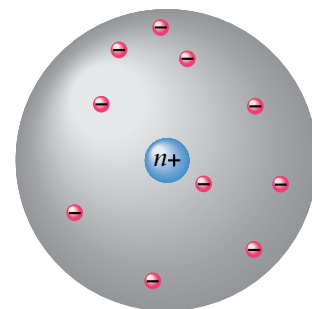


Figure 11.1

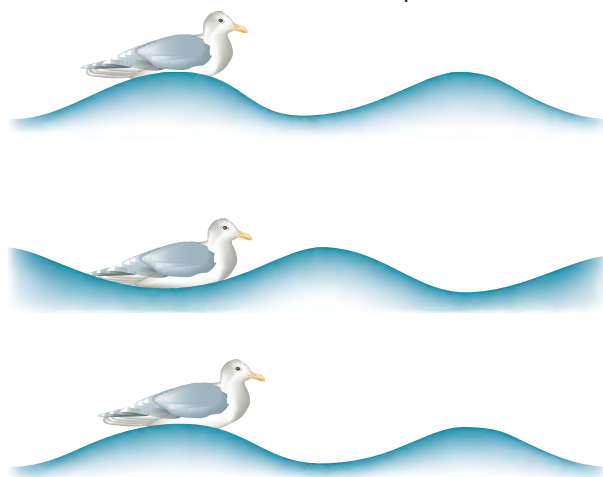
The Rutherford atom. The nuclear charge ($n+$) is balanced by the presence of n electrons moving in some way around the nucleus.

Electromagnetic radiation

Radiant energy that exhibits wavelike behavior and travels through space at the speed of light in a vacuum

Figure 11.2

A seagull floating on the ocean moves up and down as waves pass.



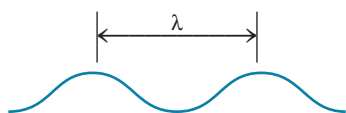


Figure 11.3

The wavelength of a wave is the distance between peaks.

The **wavelength** (symbolized by the Greek letter lambda, λ) is the distance between two consecutive wave peaks (see **Figure 11.3**). The **frequency** of the wave (symbolized by the Greek letter nu, ν) indicates how many wave peaks pass a certain point per a given time period. This idea can best be understood by thinking about how many times the seagull in Figure 11.2 goes up and down per minute. The *speed* of a wave indicates how fast a given peak travels through the water.

Active Reading Question

Describe the relationship between *wavelength* and *frequency*.

Wavelength

The distance between two consecutive peaks or troughs in a wave

Frequency

The number of waves (cycles) per second that pass a given point in space

Photon

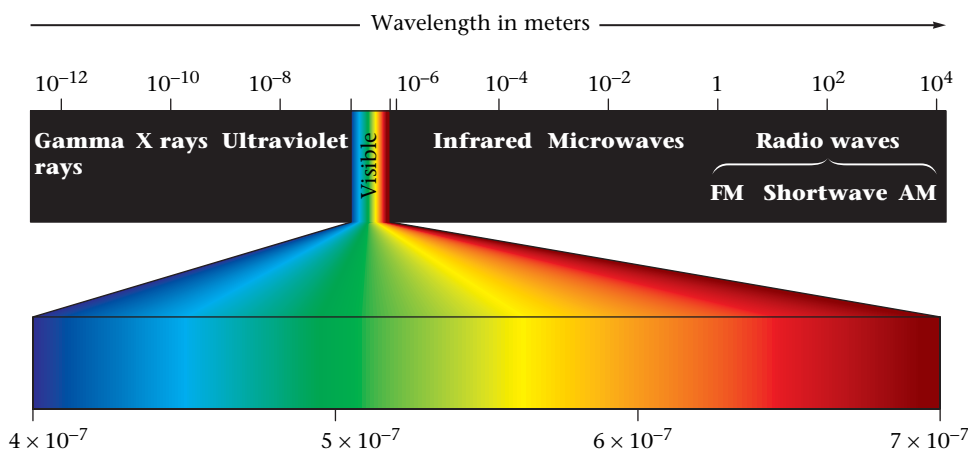
A particle of electromagnetic radiation

Although it is more difficult to picture than water waves, light (electromagnetic radiation) also travels as waves. The various types of electromagnetic radiation (X rays, microwaves, and so on) differ in their wavelengths. The classes of electromagnetic radiation are shown in **Figure 11.4**. Notice that X rays have very short wavelengths, whereas radiowaves have very long wavelengths.

Radiation provides an important means of energy transfer. For example, the energy from the sun reaches the earth mainly in the forms of visible and ultraviolet radiation. The glowing coals of a fireplace transmit heat energy by infrared radiation. In a microwave oven, the water molecules in food absorb microwave radiation, which increases their motions; this energy is then transferred to other types of molecules by collisions, thus increasing the food's temperature.

Figure 11.4

The different wavelengths of electromagnetic radiation

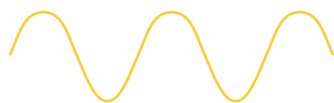


Active Reading Question

Which has the longer wavelength: blue light or red light?

Dual Nature of Light Thus we visualize electromagnetic radiation (“light”) as a wave that carries energy through space. Sometimes, however, light doesn’t behave as though it were a wave. That is, electromagnetic radiation can sometimes have properties that are characteristic of particles. (You will learn more about this idea in a physics course.) Another way to think of a beam of light traveling through space, then, is as a stream of tiny packets of energy called **photons**.

What is the exact nature of light? Does it consist of waves or is it a stream of particles of energy? It seems to be both (see **Figure 11.5**). This situation is often referred to as the wave-particle nature of light.



Light as a wave



Light as a stream of photons
(packets of energy)

Figure 11.5

Electromagnetic radiation (a beam of light) can be pictured in two ways: as a wave and as a stream of individual packets of energy called photons.

CHEMISTRY IN YOUR WORLD

Connection to Biology

Plants Seeing Red

Vegetable growers have long used plastic sheeting (mulch) in the rows between plants to retain moisture, retard weeds, and provide warmth for the roots of young plants. It is now becoming apparent that the color of the plastic used has a significant effect on the plants. How can this be? Why does the color of their mulch matter to plants?

Michael J. Kasperbauer of the Department of Agriculture's Coastal Plains Soil, Water, and Plant Research Laboratory in Florence, South Carolina, has spent most of his 40-year career studying the responses of plants to various colors of light. Plants use proteins called phytochromes to sense light in the red (640–670 nm) and the far red (700–750 nm) ranges. Although far red light is not photosynthetically active—it does not provide energy for plant growth—this light gives plants important information about their environment. For example, green leaves reflect a lot of light in the far red region. Therefore, when a given plant's phytochromes sense a high ratio of far red to red light wavelengths, the plant knows that it has many neighbors—many other plants around it are reflecting red light. Because these neighbors are competitors for the life-giving light from the sun—their leaves will shade neighboring plants—a plant sensing this situation tends to direct its growth above ground, producing a taller, thinner structure that can compete more successfully for sunlight. In fact, Kasperbauer and his colleagues have

found that by using a red plastic mulch they can fool tomatoes into “thinking” they are crowded, leading to faster seedling growth and eventually producing earlier, larger fruit.

Since the opposite effect should benefit root crops, the Department of Agriculture scientists have grown turnips in soil covered by an orange mulch. These turnips proved much bigger than those mulched with black or red plastic, presumably because of increased reflection of red light by the orange plastic. The increased red light signaled no significant competition for light from other plants, encouraging growth of roots rather than above-ground foliage.

Besides affecting the plant's structure, the type of reflected light influences the nature of the waxy coating on the leaves and the taste of the plant product. Surprisingly, the color of the reflected light also seems to affect the plant's response to insect damage.

This research shows that plants are very sensitive to the type of red light that bathes them. Our plants may benefit from rose-colored glasses as much as we do.

Red plastic mulch being used in an experimental plot

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Making Waves**Materials**

- Graph paper

Procedure

1. On a sheet of graph paper, draw three waves with a small wavelength (3–4 cm).
2. Draw two more waves, the first with wavelength 2 times the original wave and the second with wavelength 4 times the original wave.

3. Order the wavelengths from lowest to highest frequency.
4. Order the wavelengths from lowest to highest energy.

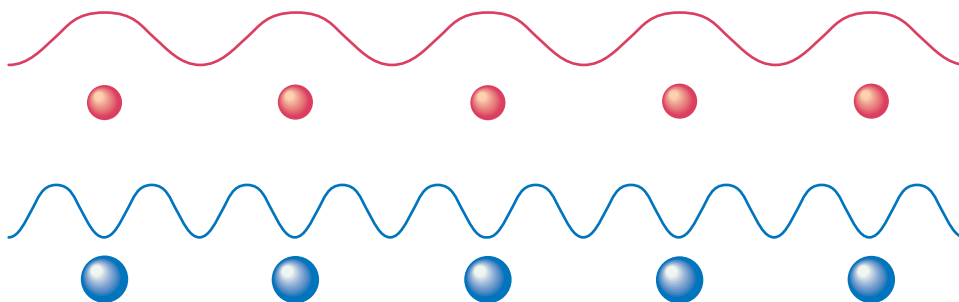
Results/Analysis

1. Explain your answers for steps 2 and 3.

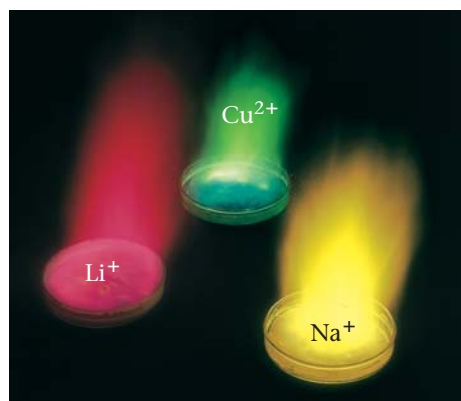
Different wavelengths of electromagnetic radiation carry different amounts of energy. For example, the photons that correspond to red light carry less energy than the photons that correspond to blue light. In general, the longer the wavelength of light, the lower the energy of its photons (see **Figure 11.6**).

Figure 11.6

A photon of red light (relatively long wavelength) carries less energy than a photon of blue light (relatively short wavelength) does.

**C. Emission of Energy by Atoms**

Consider the results of the experiment shown in **Figure 11.7**. This experiment is run by dissolving compounds containing the Li^+ ion, the Cu^{2+} ion, and the Na^+ ion in separate dishes containing methyl alcohol (with a little water added to help dissolve the compounds). The solutions are then set on fire. The brilliant colors that result are shown below.



As we will see in more detail in the next section, the colors of these flames result from atoms in these solutions releasing energy by emitting visible light of specific wavelengths (that is, specific colors). The heat from the flame causes the atoms to absorb energy; we say that the atoms become *excited*.

Some of this excess energy is then released in the form of light. The atom moves to a lower energy state as it emits a photon of light.

Lithium emits red light because its energy change corresponds to photons of red light (see **Figure 11.7**). Copper emits green light because it undergoes a different energy change than lithium; the energy change for copper corresponds to the energy of a photon of green light. Likewise, the energy change for sodium corresponds to a photon with a yellow–orange color.

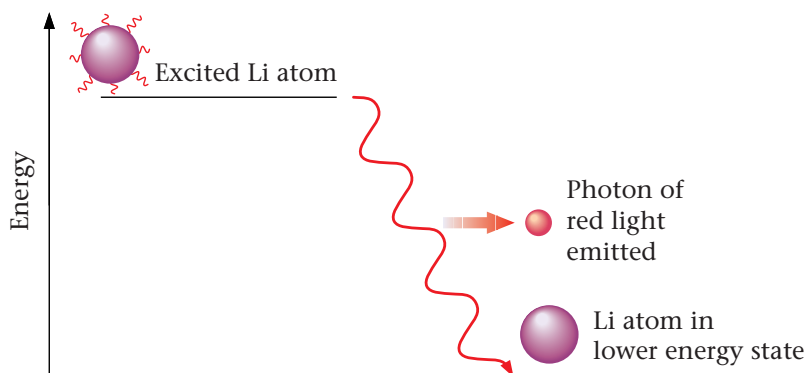


Figure 11.7

An excited lithium atom emitting a photon of red light to drop to a lower energy state

Active Reading Question

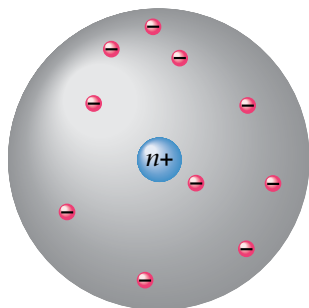
Why do we see colors when salt solutions in methanol are burned?

To summarize, we have the following situation. When atoms receive energy from some source—when they become excited—they can release this energy by emitting light. The emitted energy is carried away by a photon. Thus the energy of the photon corresponds exactly to the energy change experienced by the emitting atom. High-energy photons correspond to short-wavelength light and low-energy photons correspond to long-wavelength light. The photons of red light therefore carry less energy than the photons of blue light because red light has a longer wavelength than blue light does.

SECTION 11.1

REVIEW QUESTIONS

- Copy and label the atom showing the locations of the *electrons*, *protons*, *neutrons*, and *nucleus*.



- What is wrong with Rutherford's model of the atom? Why did it need to be modified?

- What is *electromagnetic radiation*? Provide three examples.
- How is the frequency of a wave different from its speed?
- What is a *photon*?
- What is the relationship between wavelength of light and the energy of its photons?



SECTION 11.2

The Hydrogen Atom

Key Terms

- Quantized
- Wave mechanical model

Objectives

- To understand how the emission spectrum of hydrogen demonstrates the quantized nature of energy
- To learn about Bohr's model of the hydrogen atom
- To understand how the electron's position is represented in the wave mechanical model

A. The Energy Levels of Hydrogen

As we learned in the last section, an atom with excess energy is said to be in an *excited state*. An excited atom can release some or all of its excess energy by emitting a photon (a “particle” of electromagnetic radiation) and thus move to a lower energy state. The lowest possible energy state of an atom is called its *ground state*.

We can learn a great deal about the energy states of hydrogen atoms by observing the photons they emit.

Different wavelengths of light carry different amounts of energy per photon.

Information

An atom can lose energy by *emitting* a photon.

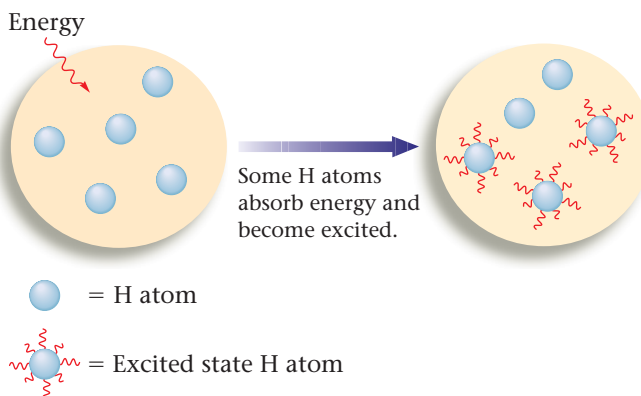


Each photon of blue light carries a larger quantity of energy than a photon of red light.

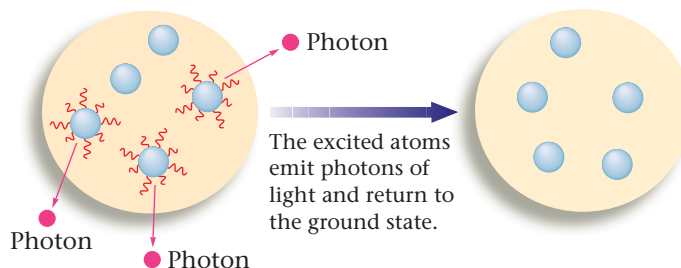


Recall that a beam of red light has lower-energy photons than a beam of blue light.

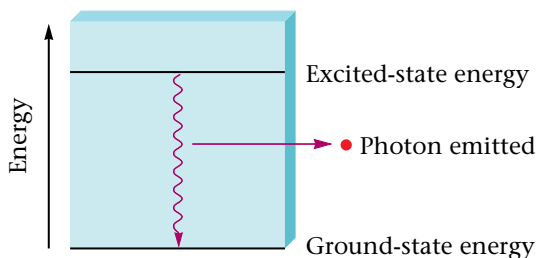
When a hydrogen atom absorbs energy from some outside source, it uses this energy to enter an excited state.



It can release this excess energy (go back to a lower state) by emitting a photon of light.



We can picture this process in terms of the energy-level diagram.



Information

A particular color (wavelength) of light carries a particular amount of energy per photon.

The energy contained in the photon corresponds to the change in energy that the atom experiences in going from the excited state to the lower state.

Consider the following experiment. Suppose we take a sample of H atoms and put a lot of energy into the system. When we study the photons of visible light emitted, we see only certain colors (see **Figure 11.8**). That is, *only certain types of photons* are produced. We don't see all colors, which would add up to give "white light;" we see only selected colors. This is a very significant result. Let's discuss carefully what it means.



Figure 11.8

When excited hydrogen atoms return to lower energy states, they emit photons of certain energies, and thus certain colors. Shown here are the colors and wavelengths (in nanometers) of the photons in the visible region that are emitted by excited hydrogen atoms.

Quantized Energy Levels Because only certain photons are emitted, we know that only certain energy changes are occurring (see **Figure 11.9**). This means that the hydrogen atom must have *certain discrete energy levels* (see **Figure 11.10**). Excited hydrogen atoms *always* emit photons with the same discrete colors (wavelengths) as those shown in Figure 11.8. They *never* emit photons with energies (colors) in between those shown.

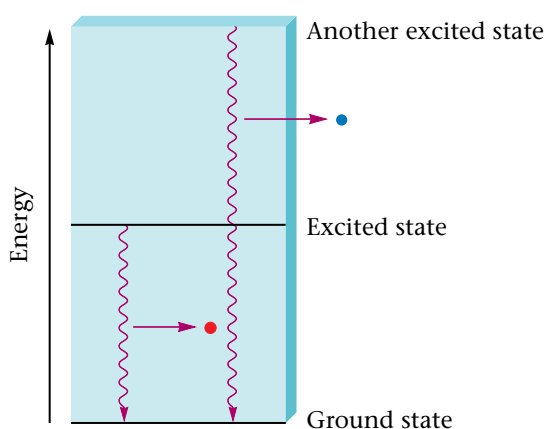


Figure 11.9

Hydrogen atoms have several excited-state energy levels. The color of the photon emitted depends on the energy change that produces it. A larger energy change may correspond to a blue photon, whereas a smaller change may produce a red photon.

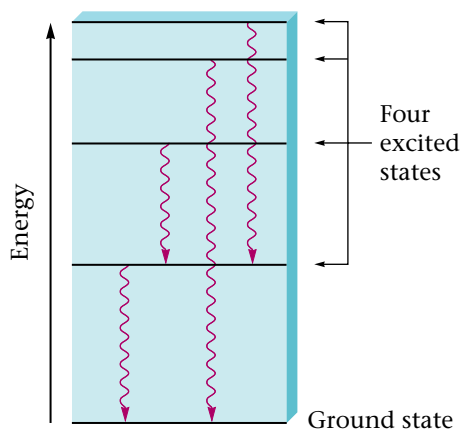


Figure 11.10

Each photon emitted by an excited hydrogen atom corresponds to a particular energy change in the hydrogen atom. In this diagram the horizontal lines represent discrete energy levels present in the hydrogen atom. A given H atom can exist in any of these energy states and can undergo energy changes to the ground state as well as to other excited states.

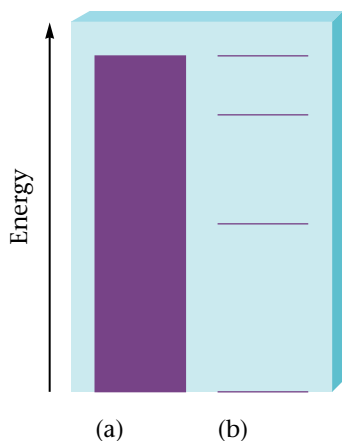


Figure 11.11

(a) Continuous energy levels. Any energy value is allowed.
 (b) Discrete (quantized) energy levels. Only certain energy states are allowed.

So we can conclude that all hydrogen atoms have the same set of discrete energy levels. We say the energy levels of hydrogen are **quantized**. That is, only *certain values are allowed*. Scientists have found that the energy levels of *all* atoms are quantized.

Active Reading Question

What does the term *quantized* mean?

The quantized nature of the energy levels in atoms was a surprise when scientists discovered it. It had been assumed previously that an atom could exist at any energy level. That is, everyone had assumed that atoms could have a continuous set of energy levels rather than only certain discrete values (see **Figure 11.11**). A useful analogy here is the contrast between the elevations allowed by a ramp, which vary continuously, and those allowed by a set of steps, which are discrete (see **Figure 11.12**). The discovery of the quantized nature of energy has radically changed our view of the atom, as we will see in the next few sections.

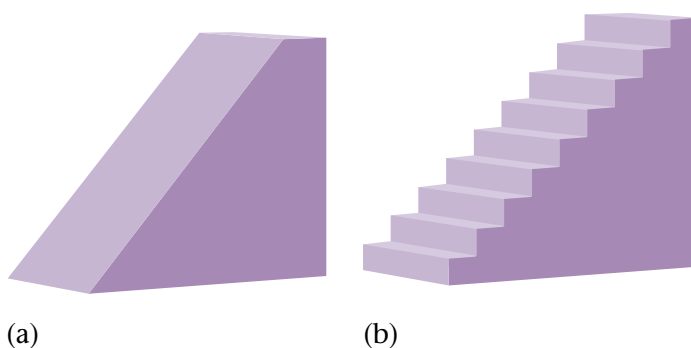


Figure 11.12

The difference between continuous and quantized energy levels can be illustrated by comparing a flight of stairs with a ramp. (a) A ramp varies continuously in elevation. (b) A flight of stairs allows only certain elevations; the elevations are quantized.

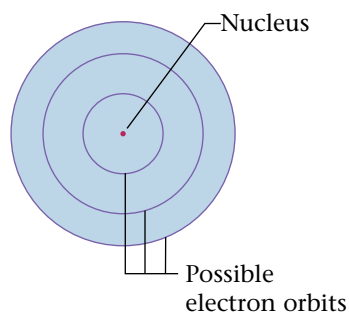
B. The Bohr Model of the Atom

In 1911 at the age of twenty-five, Niels Bohr received his Ph.D. in physics. He was convinced that the atom could be pictured as a small positive nucleus with electrons orbiting around it as the planets orbit the sun.

Over the next two years, Bohr constructed a model of the hydrogen atom with quantized energy levels that agreed with the hydrogen emission results we have just discussed. He pictured the electron moving in circular orbits corresponding to the various allowed energy levels. Bohr suggested that the electron could jump to a different orbit by absorbing or emitting a photon of light with exactly the correct energy content. Thus, in the Bohr atom, the energy levels in the hydrogen atom represented certain allowed circular orbits.

DID YOU KNOW

Although Bohr's model is consistent with the energy levels for hydrogen, it is fundamentally incorrect.



CHEMISTRY EXPLORERS

Niels Hendrik David Bohr • 1885–1962

Niels Hendrik David Bohr as a boy lived in the shadow of his younger brother Harald, who played on the 1908 Danish Olympic Soccer Team and later became a distinguished mathematician. In school, Bohr received his poorest marks in composition and struggled with writing his entire life. In fact, he wrote so poorly that he was forced to dictate his Ph.D. thesis to his mother. He is one of the very few people who felt the need to write rough drafts of postcards. Nevertheless, Bohr was a brilliant physicist. After receiving his Ph.D. in Denmark, he constructed a quantum model for the hydrogen atom by the time he was 27. Even though his model later proved to be incorrect, Bohr remained a central figure in the drive to understand the atom. He was awarded the Nobel Prize in Physics in 1922.

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At first Bohr's model appeared very promising. It fit the hydrogen atom very well. However, when this model was applied to atoms other than hydrogen, it did not work. In fact, further experiments showed that the Bohr model is fundamentally incorrect. Although the Bohr model paved the way for later theories, it is important to realize that the current theory of atomic structure is not the same as the Bohr model. Electrons do *not* move around the nucleus in circular orbits like planets orbiting the sun. Surprisingly, as we shall see later in this chapter, we do not know exactly how the electrons move in an atom.

C. The Wave Mechanical Model of the Atom

By the mid-1920s it had become apparent that the Bohr model was incorrect. Scientists needed to pursue a totally new approach. Two young physicists, Louis Victor de Broglie from France and Erwin Schrödinger from Austria, suggested that because light seems to have both wave and particle characteristics (it behaves simultaneously as a wave and as a stream of particles), the electron might also exhibit both of these characteristics. Although everyone had assumed that the electron was a tiny particle, these scientists said it might be useful to find out whether it could be described as a wave.

When Schrödinger carried out a mathematical analysis based on this idea, he found that it led to a new model for the hydrogen atom that seemed to apply equally well to other atoms—something Bohr's model failed to do. We will now explore a general picture of this model, which is called the **wave mechanical model** of the atom.

Orbitals In the Bohr model, the electron was assumed to move in circular orbits. In the wave mechanical model, on the other hand, the electron states are described by orbitals. *Orbitals are nothing like orbits.* To approximate the idea of an orbital, picture a single male firefly in a room in the center of which is suspended an open vial of a chemical which attracts fireflies. The room is extremely dark and in one corner there is a camera with an open shutter. Every time the firefly “flashes,” the camera records a pinpoint of light and thus the firefly's position in the room at that particular moment. The firefly senses the attractant and, as you can imagine, it spends a lot of time at the vial or close to it. However, now and then the insect flies randomly around the room.

DID YOU KNOW

There is now much experimental evidence suggesting that all matter exhibits both wave and particle properties.

CRITICAL THINKING ?

We now have evidence that electron energy levels in atoms are quantized.

What if energy levels in atoms were not quantized? What are some differences we would notice?



Louis Victor de Broglie

When the film is taken out of the camera and developed, the picture will probably look like this. Because a picture is brightest where the film has been exposed to the most light, the color intensity at any given point tells us how often the firefly visited that point in the room. Notice that, as we might expect, the firefly spent the most time near the room's center.

Now suppose you are watching the firefly in the dark room. You see it flash at a given point far from the center of the room. Where do you expect to see it next? There is really no way to be sure. The firefly's flight path is not precisely predictable. However, if you had seen the time-exposure picture of the firefly's activities, you would have some idea of where to look next. Your best chance would be to look toward the center of the room. The picture suggests there is the highest probability (the highest odds, the greatest likelihood) of finding the firefly at any particular moment near the center of the room. You *can't be sure* the firefly will fly toward the center of the room, but it *probably* will. So the time-exposure picture is a kind of "probability map" of the firefly's flight pattern.

According to the wave mechanical model, the electron in the hydrogen atom can be pictured as being something like this firefly. Schrödinger found that he could not precisely describe the electron's path. His mathematics enabled him only to predict the probabilities of finding the electron at given points in space around the nucleus. In its ground state the hydrogen electron has a probability map like that shown here.

The more intense the color at a particular point, the more probable it is that the electron will be found at that point at a given instant. The model gives *no information about when* the electron occupies a certain point in space or *how it moves*. In fact, we have good reasons to believe that we can *never know* the details of electron motion, no matter how sophisticated our models may become. But we do feel confident that the electron *does not* orbit the nucleus in circles as Bohr suggested.



SECTION 11.2

REVIEW QUESTIONS

- 1 In Figure 11.8 there are four different colored lines. You already know that hydrogen has only one electron. How can we get four lines from one electron?
- 2 What evidence leads us to believe that energy levels in the hydrogen atom are quantized?
- 3 Explain the terms *ground state* and *excited state*. When a photon of energy is absorbed by an atom, does the electron go from the ground state to an excited state or from an excited state to the ground state?
- 4 What is wrong with the Bohr model of the atom?
- 5 How does the wave mechanical model of the atom differ from Bohr's model?
- 6 What does the wave mechanical model of the atom tell us about how the electron moves around the nucleus?



Atomic Orbitals

Objectives

- To learn about the shapes of the s , p and d orbitals
- To review the energy levels and orbitals of the wave mechanical model of the atom
- To learn about electron spin

A. The Hydrogen Orbitals

The probability map for the hydrogen electron on the previous page is called an **orbital**. Although the probability of finding the electron decreases at greater distances from the nucleus, the probability of finding it at even great distances from the nucleus never becomes exactly zero. A useful analogy might be the lack of a sharp boundary between the earth's atmosphere and "outer space." The atmosphere fades away gradually, but there are always a few molecules present. Because the edge of an orbital is "fuzzy," an orbital does not have an exactly defined size. Chemists arbitrarily define its size as the sphere that contains 90% of the total electron probability. This definition means that the electron spends 90% of the time inside this surface and 10% somewhere outside this surface. (Note that we are *not* saying the electron travels only on the *surface* of the sphere.) The orbital represented in the diagram to the right is named the $1s$ orbital, and it describes the hydrogen electron's lowest energy state (the ground state).

We saw earlier that the hydrogen atom can absorb energy to transfer the electron to a higher energy state (an excited state). In terms of the obsolete Bohr model, this meant the electron was transferred to an orbit with a larger radius. In the wave mechanical model, these higher energy states correspond to different kinds of orbitals with different shapes.

Hydrogen Energy Levels

At this point we need to stop and consider how the hydrogen atom is organized. Remember, we showed earlier that the hydrogen atom has discrete energy levels. We call these levels **principal energy levels** and label them with whole numbers (see **Figure 11.13**). Next we find that each of these levels is subdivided into **sublevels**. The following analogy should help you understand this. Picture an inverted triangle (see **Figure 11.14**). We divide the principal levels into various numbers of sublevels. Principal level 1 consists of one sublevel, principal level 2 has two sublevels, principal level 3 has three sublevels, and principal level 4 has four sublevels.

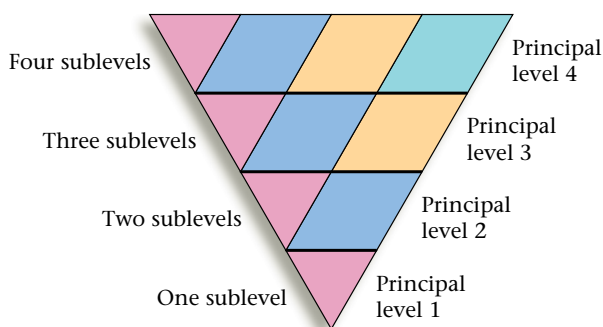


Figure 11.14

An illustration of how principal levels can be divided into sublevels

Key Terms

- Orbital
- Principal energy levels
- Sublevels
- Pauli exclusion principle

Orbital

The three-dimensional region in which there is a high probability of finding an electron in an atom

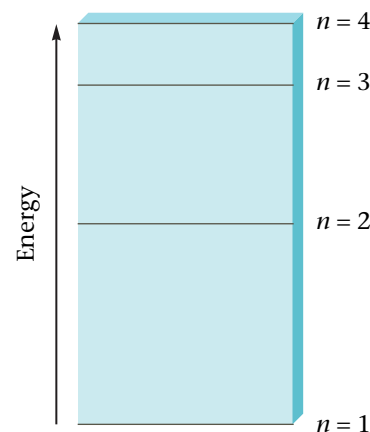


Figure 11.13

The first four principal energy levels in the hydrogen atom. Each level is assigned an integer, n .

Like our triangle, the principal energy levels in the hydrogen atom contain sublevels. As we will see presently, these sublevels contain spaces for the electron that we call orbitals. Principal energy level 1 consists of just one sublevel, or one type of orbital. This orbital is spherical in shape. We label this orbital 1s. The number 1 is for the principal energy level, and s is a shorthand way to label a particular sublevel (type of orbital).

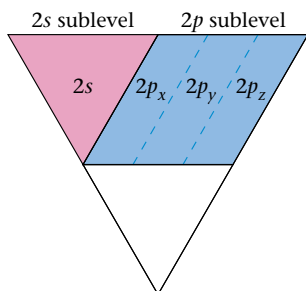
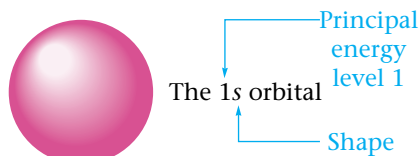


Figure 11.15
Principal level 2 shown divided into the 2s and 2p sublevels



Principal energy level 2 has two sublevels. (Note the correspondence between the principal energy level number and the number of sublevels.) These sublevels are labeled 2s and 2p. The 2s sublevel consists of one orbital (called the 2s), and the 2p sublevel consists of three orbitals (called $2p_x$, $2p_y$, and $2p_z$). Let's return to the inverted triangle to illustrate this. **Figure 11.15** shows principal level 2 divided into the sublevels 2s and 2p (which is subdivided into $2p_x$, $2p_y$, and $2p_z$). The orbitals have the shapes shown in **Figures 11.16** and **11.17**. The 2s orbital is spherical like the 1s orbital but larger in size (see Figure 11.16). The three 2p orbitals are not spherical but have two "lobes." These orbitals are shown in Figure 11.17 both as electron probability maps and as surfaces that contain 90% of the total electron probability. Notice that the label x, y, or z on a given 2p orbital tells along which axis the lobes of that orbital are directed.

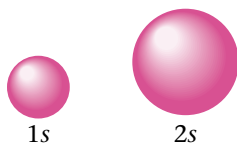


Figure 11.16
The relative sizes of the 1s and 2s orbitals of hydrogen

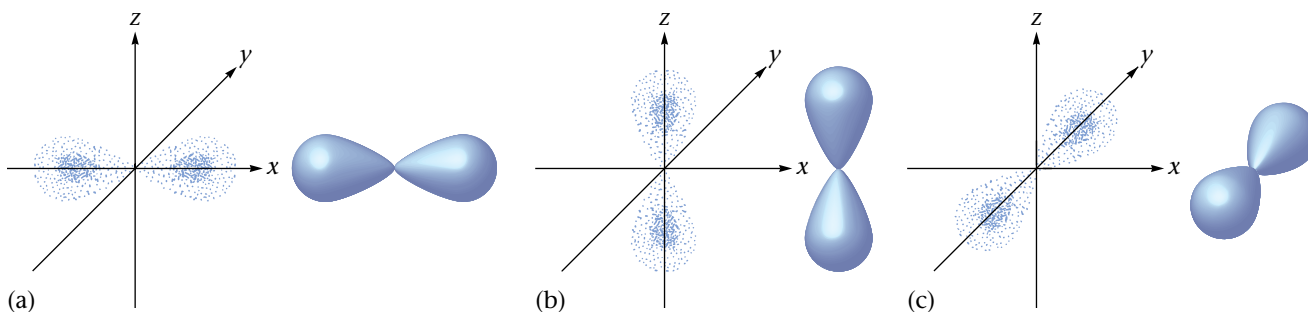
Active Reading Question

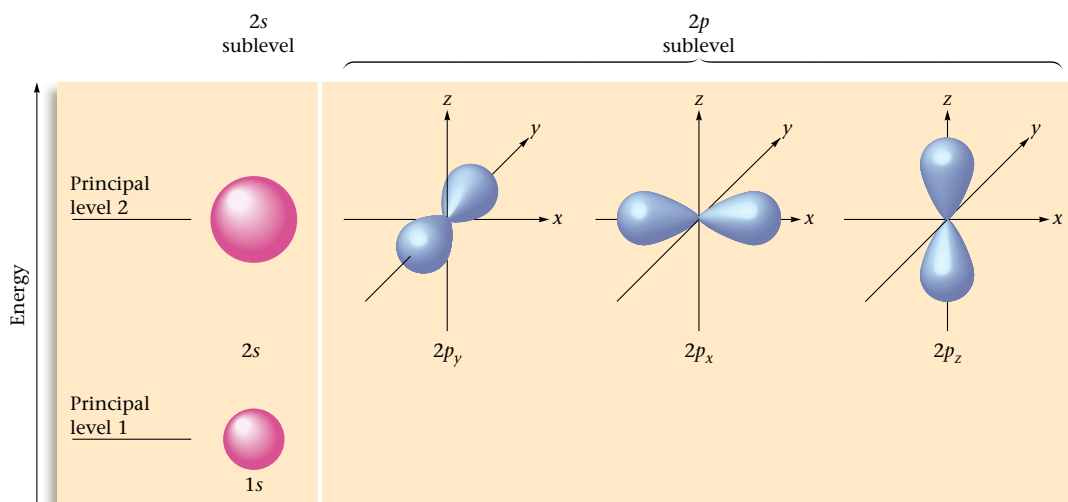
What is the relationship between the principal level and the number of sublevels?

What we have learned so far about the hydrogen atom is summarized in **Figure 11.18**.

- Principal energy level 1 has one sublevel, which contains the 1s orbital.
- Principal energy level 2 contains two sublevels, one of which contains the 2s orbital and one of which contains the 2p orbitals (three of them).
- Note that each orbital is designated by a symbol or label.

Figure 11.17
The three 2p orbitals: (a) $2p_x$, (b) $2p_z$, (c) $2p_y$. The x, y, or z label indicates along which axis the two lobes are directed. Each orbital is shown both as a probability map and as a surface that encloses 90% of the electron probability.



**Figure 11.18**

A diagram of principal energy levels 1 and 2 showing the shapes of orbitals that compose the sublevels

Let's Review

Orbital Labels

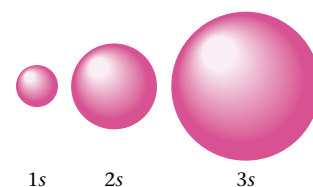
- The number tells the principal energy level.
- The letter tells the shape. The letter *s* means a spherical orbital; the letter *p* means a two-lobed orbital. The *x*, *y*, or *z* subscript on a *p* orbital label tells along which of the coordinate axes the two lobes lie.

One important characteristic of orbitals is that as the level number increases, the average distance of the electron in that orbital from the nucleus also increases. That is, when the hydrogen electron is in the 1*s* orbital (the ground state), it spends most of its time much closer to the nucleus than when it occupies the 2*s* orbital (an excited state).

Hydrogen Orbitals

You may be wondering at this point why hydrogen, which has only one electron, has more than one orbital. It is best to think of an orbital as a *potential space* for an electron. The hydrogen electron can occupy only a single orbital at a time, but the other orbitals are still available should the electron be transferred into one of them. For example, when a hydrogen atom is in its ground state (lowest possible energy state), the electron is in the 1*s* orbital. By adding the correct amount of energy (for example, a specific photon of light), we can excite the electron to the 2*s* orbital or to one of the 2*p* orbitals.

So far we have discussed only two of hydrogen's energy levels. There are many others. For example, level 3 has three sublevels (see Figure 11.14), which we label 3*s*, 3*p*, and 3*d*. The 3*s* sublevel contains a single 3*s* orbital, a spherical orbital larger than 1*s* and 2*s* (see **Figure 11.19**). Sublevel 3*p* contains three orbitals: 3*p_x*, 3*p_y*, and 3*p_z*, which are shaped like the 2*p* orbitals except that they are larger. The 3*d* sublevel contains five 3*d* orbitals with the shapes and labels shown in **Figure 11.20**. (You do not need to memorize the 3*d* orbital shapes and labels. They are shown for completeness.)

**Figure 11.19**

The relative sizes of the spherical 1*s*, 2*s*, and 3*s* orbitals of hydrogen

Information

Think of orbitals as ways of dividing up the space around a nucleus.

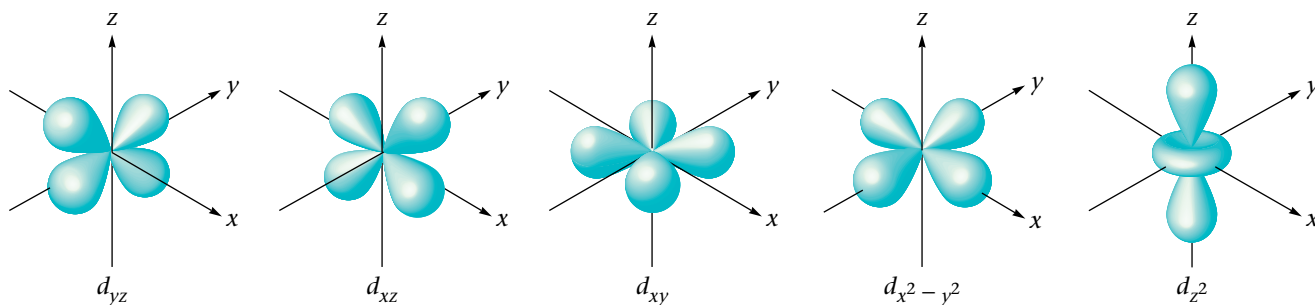


Figure 11.20
The shapes and labels of the five 3d orbitals

Notice as you compare levels 1, 2, and 3 that a new type of orbital (sub-level) is added in each principal energy level. (Recall that the p orbitals are added in level 2 and the d orbitals in level 3.) This makes sense because in going farther out from the nucleus, there is more space available and thus room for more orbitals.

It might help you to understand that the number of orbitals increases with the principal energy level if you think of a sports stadium. Picture a tennis court with circular rows of seats surrounding it. The farther from the court a row of seats is, the more seats it contains because the circle is larger. Orbitals divide up the space around a nucleus somewhat like the seats in this stadium. The greater the distance from the nucleus, the more space there is and the more orbitals we find.

The pattern of increasing numbers of orbitals continues with level 4. Level 4 has four sublevels labeled $4s$, $4p$, $4d$, and $4f$. The $4s$ sublevel has a single $4s$ orbital. The $4p$ sublevel contains three orbitals ($4p_x$, $4p_y$, and $4p_z$). The $4d$ sublevel has five $4d$ orbitals. The $4f$ sublevel has seven $4f$ orbitals.

The $4s$, $4p$, and $4d$ orbitals have the same shapes as the earlier s , p , and d orbitals, respectively, but are larger. We will not be concerned here with the shapes of the f orbitals.

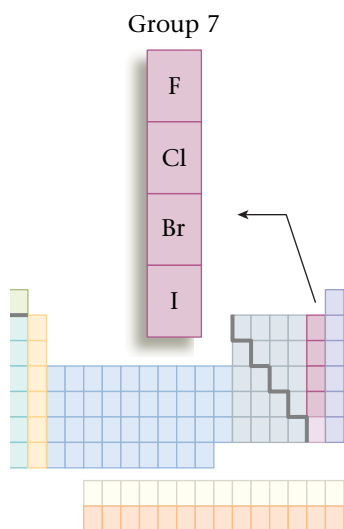


A circular tennis stadium

B. The Wave Mechanical Model: Further Development

A model for the atom is of little use if it does not apply to all atoms. The Bohr model was discarded because it could be applied only to hydrogen. The wave mechanical model can be applied to all atoms in basically the same form as we have just used it for hydrogen. In fact, the major triumph of this model is its ability to explain the periodic table of the elements. Recall that the elements on the periodic table are arranged in vertical groups, which contain elements that typically show similar chemical properties. For example, the halogens shown to the left are chemically similar. The wave mechanical model of the atom allows us to explain, based on electron arrangements, why these similarities occur. We will see later how this is done.

Atoms Beyond Hydrogen Remember that an atom has as many electrons as it has protons to give it a zero overall charge. Therefore, all atoms beyond hydrogen have more than one electron. Before we can consider the atoms beyond hydrogen, we must describe one more property of electrons that determines how they can be arranged in an atom's orbitals. This property is spin. Each electron appears to be spinning as a top spins on



its axis. Like the top, an electron can spin in only one of two directions. We often represent spin with an arrow: \uparrow or \downarrow . One arrow represents the electron spinning in the one direction, and the other represents the electron spinning in the opposite direction. For our purposes, what is most important about electron spin is that two electrons must have *opposite* spins to occupy the same orbital. That is, two electrons that have the same spin cannot occupy the same orbital.

Pauli exclusion principle: An atomic orbital can hold a maximum of two electrons, and those two electrons must have opposite spins.

Pauli exclusion principle

In a given atom, no two electrons can occupy the same atomic orbital and have the same spin.

Let's Review

Principal Components of the Wave Mechanical Model of the Atom

- Atoms have a series of energy levels called **principal energy levels**, which are designated by whole numbers symbolized by n ; n can equal 1, 2, 3, 4, . . . Level 1 corresponds to $n = 1$, level 2 corresponds to $n = 2$, and so on.
- The energy of the level increases as the value of n increases.
- Each principal energy level contains one or more *types* of orbitals, called **sublevels**.
- The number of sublevels present in a given principal energy level equals n . For example, level 1 contains one sublevel (1s); level 2 contains two sublevels (two types of orbitals), the 2s orbital and the three 2p orbitals; and so on. These are summarized in the following table. The number of each type of orbital is shown in parentheses.

n Sublevels (Types of Orbitals) Present

1	1s(1)
2	2s(1) 2p(3)
3	3s(1) 3p(3) 3d(5)
4	4s(1) 4p(3) 4d(5) 4f(7)

- The n value is always used to label the orbitals of a given principal level and is followed by a letter that indicates the type (shape) of the orbital. For example, the designation 3p means an orbital in level 3 that has two lobes (a p orbital always has two lobes).
- An orbital can be empty or it can contain one or two electrons, but never more than two. If two electrons occupy the same orbital, they must have opposite spins.
- The shape of an orbital does not indicate the details of electron movement. It indicates the probability distribution for an electron residing in that orbital.

EXAMPLE 11.1

Understanding the Wave Mechanical Model of the Atom

Indicate whether each of the following statements about atomic structure is true or false.

- An s orbital is always spherical in shape.
- The $2s$ orbital is the same size as the $3s$ orbital.
- The number of lobes on a p orbital increases as n increases. That is, a $3p$ orbital has more lobes than a $2p$ orbital.
- Level 1 has one s orbital, level 2 has two s orbitals, level 3 has three s orbitals, and so on.
- The electron path is indicated by the surface of the orbital.

Solution



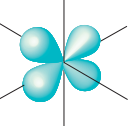
- True. The size of the sphere increases as n increases, but the shape is always spherical.
- False. The $3s$ orbital is larger (the electron is farther from the nucleus on average) than the $2s$ orbital.
- False. A p orbital always has two lobes.
- False. Each principal energy level has only one s orbital.
- False. The electron is *somewhere inside* the orbital surface 90% of the time. The electron does not move around *on* this surface.

Practice Problem • Exercise 11.1

Define the following terms.

- Bohr orbits
- orbitals
- orbital size
- sublevel

SECTION 11.3 REVIEW QUESTIONS

- What is the difference between an orbit and an orbital in atomic theory?
- Draw Figure 11.14 and fill in the different types of sublevels for each principal energy levels.
- Label each of the following orbitals as s , p , or d .
 - 
 - 
 - 
- Tell how many orbitals are found in each type of sublevel: s , p , d , f .
- How do the shapes of s orbitals in different energy levels of a hydrogen atom compare? How do the sizes of these orbitals compare?
- What is the Pauli exclusion principle and how does it help us determine where an electron is within the atom?



Electron Configurations and Atomic Properties

Objectives

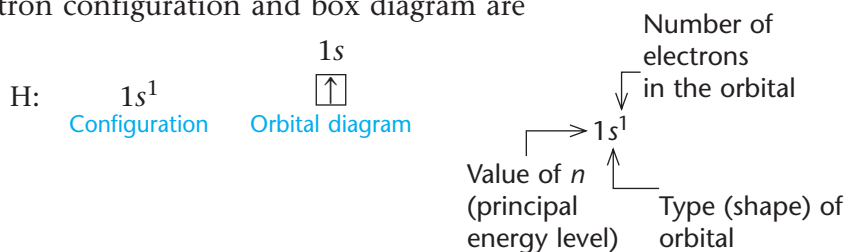
- To understand how the principal energy levels fill with electrons in atoms beyond hydrogen
- To learn about valence electrons and core electrons
- To learn about the electron configurations of atoms with $Z \leq 18$
- To understand the general trends in properties in the periodic table

A. Electron Arrangements in the First 18 Atoms on the Periodic Table

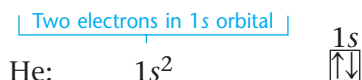
We will now describe the electron arrangements in atoms with $Z = 1$ to $Z = 18$ by placing electrons in the various orbitals in the principal energy levels, starting with $n = 1$, and then continuing with $n = 2$, $n = 3$, and so on. For the first 18 elements, the individual sublevels fill in the following order: $1s$, then $2s$, then $2p$, then $3s$, and then $3p$.

The most attractive orbital to an electron in an atom is always the $1s$, because in this orbital the negatively charged electron is closer to the positively charged nucleus than in any other orbital. That is, the $1s$ orbital involves the space around the nucleus that is closest to the nucleus. As n increases, the orbital becomes larger—the electron, on average, occupies space farther from the nucleus.

Electron Configurations In its ground state, hydrogen has its lone electron in the $1s$ orbital. This is commonly represented in two ways. First, we say that hydrogen has the electron arrangement, or **electron configuration**, $1s^1$. This means there is one electron in the $1s$ orbital. We can also represent this configuration by using an **orbital diagram**, also called a **box diagram**, in which orbitals are represented by boxes grouped by sublevel with small arrows indicating the electrons. For *hydrogen*, the electron configuration and box diagram are



The arrow represents an electron spinning in a particular direction. The next element is *helium*, $Z = 2$. It has two protons in its nucleus and so has two electrons. Because the $1s$ orbital is the most desirable, both electrons go there but with opposite spins. For helium, the electron configuration and box diagram are



The opposite electron spins are shown by the opposing arrows in the box.

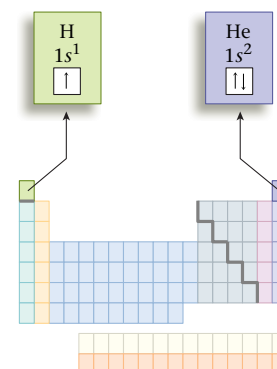
Lithium ($Z = 3$) has three electrons, two of which go into the $1s$ orbital. That is, two electrons fill that orbital. The $1s$ orbital is the only orbital for

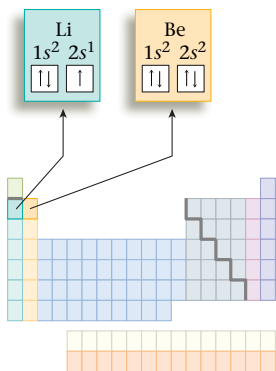
Key Terms

- Electron configuration
- Orbital diagram (box diagram)
- Valence electrons
- Core electrons
- Lanthanide series
- Actinide series
- Main-group elements (representative elements)
- Metals
- Nonmetals
- Metalloids
- Atomic size
- Ionization energy

Electron configuration

The arrangement of electrons in an atom





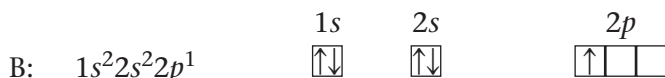
$n = 1$, so the third electron must occupy an orbital with $n = 2$ —in this case the $2s$ orbital. This gives a $1s^2 2s^1$ configuration. The electron configuration and box diagram are



The next element, *beryllium*, has four electrons, which occupy the $1s$ and $2s$ orbitals with opposite spins.



Boron has five electrons, four of which occupy the $1s$ and $2s$ orbitals. The fifth electron goes into the second type of orbital with $n = 2$, one of the $2p$ orbitals.



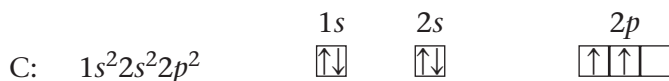
Because all the $2p$ orbitals have the same energy, it does not matter which $2p$ orbital the electron occupies.

Carbon, the next element, has six electrons: two electrons occupy the $1s$ orbital, two occupy the $2s$ orbital, and two occupy $2p$ orbitals. There are three $2p$ orbitals, so each of the mutually repulsive electrons occupies a different $2p$ orbital. For reasons we will not consider, in the separate $2p$ orbitals the electrons have the same spin.

The configuration for carbon could be written $1s^2 2s^2 2p^1 2p^1$ to indicate that the electrons occupy separate $2p$ orbitals. However, the configuration is usually given as $1s^2 2s^2 2p^2$, and it is understood that the electrons are in different $2p$ orbitals.

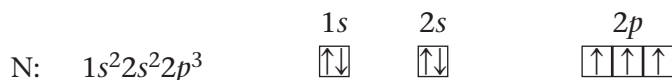
Information

Because electrons have the same charge, they repel each other.

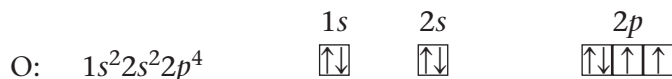


Note the like spins for the unpaired electrons in the $2p$ orbitals.

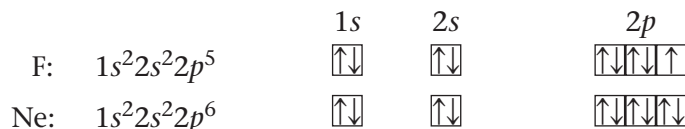
The configuration for *nitrogen*, which has seven electrons, is $1s^2 2s^2 2p^3$. The three electrons in $2p$ orbitals occupy separate orbitals and have like spins.



The configuration for *oxygen*, which has eight electrons, is $1s^2 2s^2 2p^4$. One of the $2p$ orbitals is now occupied by a pair of electrons with opposite spins, as required by the Pauli exclusion principle.



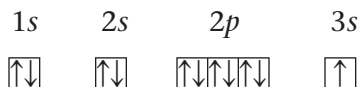
The electron configurations and orbital diagrams for *fluorine* (nine electrons) and *neon* (ten electrons) are



With neon, the orbitals with $n = 1$ and $n = 2$ are completely filled.

For *sodium*, which has eleven electrons, the first ten electrons occupy the $1s$, $2s$, and $2p$ orbitals, and the eleventh electron must occupy the first orbital with $n = 3$, the $3s$ orbital. The electron configuration for sodium is $1s^2 2s^2 2p^6 3s^1$. To avoid writing the inner-level electrons, we often abbreviate the configuration $1s^2 2s^2 2p^6 3s^1$ as $[\text{Ne}]3s^1$, where $[\text{Ne}]$ represents the electron configuration of neon, $1s^2 2s^2 2p^6$.

The orbital diagram for sodium is



The next element, *magnesium*, $Z = 12$, has the electron configuration $1s^2 2s^2 2p^6 3s^2$, or $[\text{Ne}]3s^2$.

The next six elements, *aluminum* through *argon*, have electron configurations obtained by filling the $3p$ orbitals one electron at a time. **Figure 11.21** summarizes the electron configurations of the first eighteen elements by giving the number of electrons in the type of orbital (sublevel) occupied last.

H $1s^1$									He $1s^2$
Li $2s^1$	Be $2s^2$			B $2p^1$	C $2p^2$	N $2p^3$	O $2p^4$	F $2p^5$	Ne $2p^6$
Na $3s^1$	Mg $3s^2$			Al $3p^1$	Si $3p^2$	P $3p^3$	S $3p^4$	Cl $3p^5$	Ar $3p^6$

Figure 11.21

The electron configurations in the sublevel last occupied for the first eighteen elements

EXAMPLE 11.2

Writing Orbital Diagrams

Write the orbital diagram for magnesium.

Solution

Magnesium ($Z = 12$) has twelve electrons that are placed successively in the $1s$, $2s$, $2p$, and $3s$ orbitals to give the electron configuration $1s^2 2s^2 2p^6 3s^2$. The orbital diagram is



Only occupied orbitals are shown here.

Practice Problem • Exercise 11.2

Write the complete electron configuration and the orbital diagram for each of the elements aluminum through argon.

Valence electron

The electron in the outermost principal energy level of an atom

Core electron

An inner electron; an electron not in the outermost principal energy level of an atom

Classifying Electrons At this point it is useful to introduce the concept of **valence electrons**—that is, *the electrons in the outermost (highest) principal energy level of an atom*. For example, nitrogen, which has the electron configuration $1s^22s^22p^3$, has electrons in principal levels 1 and 2. Therefore, level 2 (which has $2s$ and $2p$ sublevels) is the valence level of nitrogen, and the $2s$ and $2p$ electrons are the valence electrons. For the sodium atom (electron configuration $1s^22s^22p^63s^1$, or $[\text{Ne}]3s^1$) the valence electron is the electron in the $3s$ orbital, because in this case principal energy level 3 is the outermost level that contains an electron. The valence electrons are the most important electrons to chemists because, being the outermost electrons, they are the ones involved when atoms attach to each other (form bonds), as we will see in the next chapter. The inner electrons, which are known as **core electrons**, are not involved in bonding atoms to each other.

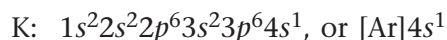
Note in Figure 11.21 that a very important pattern is developing: except for helium, *the atoms of elements in the same group (vertical column of the periodic table) have the same number of electrons in a given type of orbital (sublevel), except that the orbitals are in different principal energy levels*. Remember that the elements were originally organized into groups on the periodic table on the basis of similarities in chemical properties. Now we understand the reason behind these groupings.

Elements with the same valence electron arrangement show very similar chemical behavior.

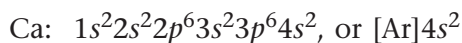
B. Electron Configurations and the Periodic Table

We have seen that we can describe the atoms beyond hydrogen by simply filling the atomic orbitals starting with level $n = 1$ and working outward in order. This works fine until we reach the element *potassium* ($Z = 19$), which is the next element after argon. Because the $3p$ orbitals are fully occupied in argon, we might expect the next electron to go into a $3d$ orbital (recall that for $n = 3$ the sublevels are $3s$, $3p$, and $3d$). However, experiments show that the chemical properties of potassium are very similar to those of lithium and sodium.

Because we have learned to associate similar chemical properties with similar valence-electron arrangements, we predict that the valence-electron configuration for potassium is $4s^1$, resembling sodium ($3s^1$) and lithium ($2s^1$). That is, we expect the last electron in potassium to occupy the $4s$ orbital instead of one of the $3d$ orbitals. This means that the principal energy level 4 begins to fill before level 3 has been completed. This conclusion is confirmed by many types of experiments. So the electron configuration of potassium is



The next element is *calcium*, with an additional electron that also occupies the $4s$ orbital.



The $4s$ orbital is now full.

After calcium the next electrons go into the $3d$ orbitals to complete principal energy level 3. The elements that correspond to filling the $3d$ orbitals are called transition metals. Then the $4p$ orbitals fill. **Figure 11.22** gives partial electron configurations for the elements potassium through krypton.

Information

[Ar] is shorthand for $1s^22s^22p^63s^23p^6$.

Figure 11.22

Partial electron configurations for the elements potassium through krypton. The transition metals shown in green (scandium through zinc) have the general configuration $[\text{Ar}]4s^23d^n$, except for chromium and copper.

K $4s^1$	Ca $4s^2$	Sc $3d^1$	Ti $3d^2$	V $3d^3$	Cr $4s^13d^5$	Mn $3d^5$	Fe $3d^6$	Co $3d^7$	Ni $3d^8$	Cu $4s^13d^{10}$	Zn $3d^{10}$	Ga $4p^1$	Ge $4p^2$	As $4p^3$	Se $4p^4$	Br $4p^5$	Kr $4p^6$

Note from Figure 11.22 that all of the transition metals have the general configuration $[\text{Ar}]4s^23d^n$ except chromium ($4s^13d^5$) and copper ($4s^13d^{10}$). The reasons for these exceptions are complex and will not be discussed here.

Instead of continuing to consider the elements individually, we will now look at the overall relationship between the periodic table and orbital filling. **Figure 11.23** shows which type of orbital is filling in each area of the periodic table.

Periods	1	Groups										8
1	1s											1s
2	2s							2p				
3	3s							3p				
4	4s			3d			4p					
5	5s			4d			5p					
6	6s	La			5d			6p				
7	7s	Ac			6d							
					4f						Lanthanide series	
					5f						Actinide series	

Information

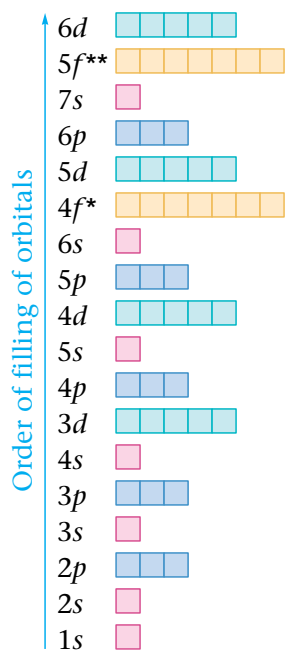
- The $(n + 1)s$ orbital fills before the nd orbitals fill.
- Lanthanides are elements in which the $4f$ orbitals are being filled.
- Actinides are elements in which the $5f$ orbitals are being filled.

Figure 11.23

The orbitals being filled for elements in various parts of the periodic table. Note that in going along a horizontal row (a period), the $(n + 1)s$ orbital fills before the nd orbital. The group label indicates the number of valence electrons (the number of s plus the number of p electrons in the highest occupied principal energy level) for the elements in each group.

Rules for Orbital Filling

1. In a principal energy level that has d orbitals, the s orbital from the *next* level fills before the d orbitals in the current level. That is, the $(n + 1)s$ orbitals always fill before the nd orbitals. For example, the $5s$ orbitals fill for rubidium and strontium before the $4d$ orbitals fill for the second row of transition metals (yttrium through cadmium).
2. After lanthanum, which has the electron configuration $[\text{Xe}]6s^25d^1$, a group of fourteen elements called the **lanthanide series**, or the lanthanides, occurs. This series of elements corresponds to the filling of the seven $4f$ orbitals.
3. After actinium, which has the configuration $[\text{Rn}]7s^26d^1$, a group of fourteen elements called the **actinide series**, or the actinides, occurs. This series corresponds to the filling of the seven $5f$ orbitals.
4. Except for helium, the group numbers indicate the sum of electrons in the ns and np orbitals in the highest principal energy level that contains electrons (where n is the number that indicates a particular principal energy level). These electrons are the valence electrons, the electrons in the outermost principal energy level of a given atom.



To help you further understand the connection between orbital filling and the periodic table, **Figure 11.24** shows the orbitals in the order in which they fill.

A periodic table is almost always available to you. If you understand the relationship between the electron configuration of an element and its position on the periodic table, you can figure out the expected electron configuration of any atom.

Active Reading Question

How does an electron configuration for an atom relate to the atom's placement on the periodic table?

Figure 11.24

A box diagram showing the order in which orbitals fill to produce the atoms in the periodic table. Each box can hold two electrons.

*After the 6s orbital is full, one electron goes into a 5d orbital. This corresponds to the element lanthanum ($[\text{Xe}]6s^25d^1$). After lanthanum, the 4f orbitals fill with electrons.

**After the 7s orbital is full, one electron goes into 6d. This is actinium ($[\text{Rn}]7s^26d^1$). The 5f orbitals then fill.

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

A Magnetic Moment

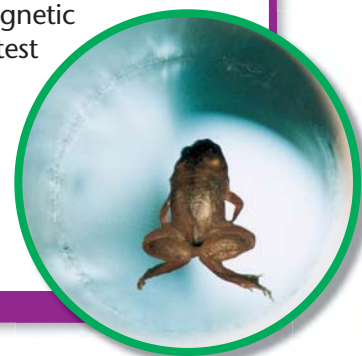
An anesthetized frog lies in the hollow core of an electromagnet. As the current in the coils of the magnet is increased, the frog magically rises and floats in midair (see photo). How can this happen? Is the electromagnet an antigravity machine? In fact, there is no magic going on here. This phenomenon demonstrates the magnetic properties of all matter. We know that iron magnets attract and repel each other depending on their relative orientations. Is a frog magnetic like a piece of iron? If a frog lands on a steel manhole cover, will it be trapped there by magnetic attractions? Of course not. The magnetism of the frog, as with most objects, shows up only in the presence of a strong inducing magnetic field. In other words, the powerful electromagnet surrounding the frog in the experiment described above *induces* a magnetic field in the frog that opposes the inducing field. The opposing magnetic field in the frog repels the inducing field, and the frog lifts up until the magnetic force is balanced by the gravitational pull on its body. The frog then “floats” in air.

How can a frog be magnetic if it is not made of iron? It's the electrons. Frogs are composed of cells containing many kinds of molecules. Of course, these molecules are made of atoms—carbon atoms,

nitrogen atoms, oxygen atoms, and other types. Each of these atoms contains electrons that are moving around the atomic nuclei. When these electrons sense a strong magnetic field, they respond by moving in a fashion that produces magnetic fields aligned to oppose the inducing field. This phenomenon is called *diamagnetism*.

All substances, animate and inanimate, because they are made of atoms, exhibit diamagnetism. Andre Geim and his colleagues at the University of Nijmegen, the Netherlands, have levitated frogs, grasshoppers, plants, and water droplets, among other objects. Geim says that, given a large enough electromagnet, even humans can be levitated. He notes, however, that constructing a magnet strong enough to float a human would be very expensive, and he sees no point in it. Geim does point out that inducing weightlessness with magnetic fields may be a good way to pretest experiments on weightlessness intended as research for future space flights—to see if the ideas fly as well as the objects.

A live frog being levitated in a magnetic field



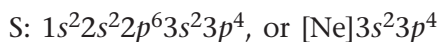
EXAMPLE 11.3

Determining Electron Configurations

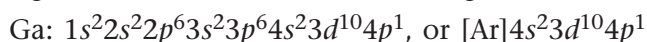
Using the periodic table inside the back cover of the text, give the electron configurations for sulfur (S), gallium (Ga), hafnium (Hf), and radium (Ra).

Solution

Sulfur is element 16 and resides in Period 3, where the $3p$ orbitals are being filled (see **Figure 11.25**). Because sulfur is the fourth among the “ $3p$ elements,” it must have four $3p$ electrons. Sulfur’s electron configuration is



Gallium is element 31 in Period 4 just after the transition metals (see **Figure 11.25**). It is the first element in the “ $4p$ series” and has a $4p^1$ arrangement. Gallium’s electron configuration is

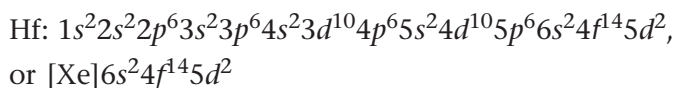


	Groups										
	1	2				3	4	5	6	7	8
Periods	1s	2s				3s	4s	5s	6s	7s	1s
2	2s							2p			
3	3s						3p	S			
4	4s				3d				Ga	4p	
5	5s				4d					5p	
6	6s	Hf				5d				6p	
7	7s	Ra				6d					

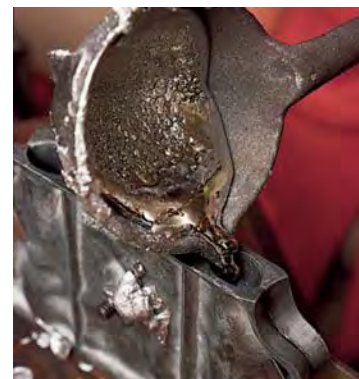
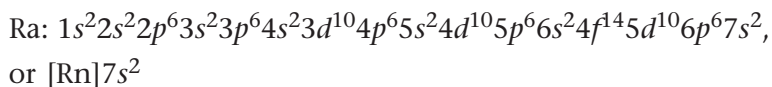
Figure 11.25

The positions of the elements considered in Example 11.3

Hafnium is element 72 and is found in Period 6, as shown in **Figure 11.25**. Note that it occurs just after the lanthanide series (see **Figure 11.23**). Thus the $4f$ orbitals are already filled. Hafnium is the second member of the $5d$ transition series and has two $5d$ electrons. Its electron configuration is



Radium is element 88 and is in Period 7 (and Group 2), as shown in **Figure 11.25**. Thus radium has two electrons in the $7s$ orbital, and its electron configuration is



Molten lead being poured into a mold to make toy soldiers

Practice Problem • Exercise 11.3

Using the periodic table inside the back cover of the text, predict the electron configurations for fluorine, silicon, cesium, lead, and iodine. If you have trouble, use **Figure 11.23**.

CRITICAL THINKING ?

Summary of the Wave Mechanical Model and Valence-Electron Configurations

You have learned that each orbital can hold two electrons and that this pattern is evident in the periodic table.

What if each orbital could hold three electrons? How would it change the appearance of the periodic table? For example, what would be the atomic numbers of the noble gases?

The concepts we have discussed in this chapter are very important. They allow us to make sense of a good deal of chemistry. When it was first observed that elements with similar properties occur periodically as the atomic number increases, chemists wondered why. Now we have an explanation. The wave mechanical model pictures the electrons in an atom as arranged in orbitals, with each orbital capable of holding two electrons. As we build up the atoms, the same types of orbitals recur in going from one principal energy level to another. This means that particular valence-electron configurations recur periodically.

For reasons we will explore in the next chapter, elements with a particular type of valence configuration all show very similar chemical behavior. Thus groups of elements, such as the alkali metals, show similar chemistry because all the elements in that group have the same type of valence-electron arrangement. This concept, which explains so much chemistry, is the greatest contribution of the wave mechanical model to modern chemistry.

For reference, the valence-electron configurations for all the elements are shown on the periodic table in **Figure 11.26**.

Period number, highest occupied electron level	Representative Elements		<i>d</i> -Transition Elements										Representative Elements					Noble Gases								
	1A <i>ns</i> ¹	2A <i>ns</i> ²											3A <i>ns</i> ² <i>np</i> ¹	4A <i>ns</i> ² <i>np</i> ²	5A <i>ns</i> ² <i>np</i> ³	6A <i>ns</i> ² <i>np</i> ⁴	7A <i>ns</i> ² <i>np</i> ⁵	8A <i>ns</i> ² <i>np</i> ⁶								
1	1 H <i>1s</i> ¹	2 He <i>1s</i> ²																								
2	3 Li <i>2s</i> ¹	4 Be <i>2s</i> ²											5 B <i>2s</i> ² <i>2p</i> ¹	6 C <i>2s</i> ² <i>2p</i> ²	7 N <i>2s</i> ² <i>2p</i> ³	8 O <i>2s</i> ² <i>2p</i> ⁴	9 F <i>2s</i> ² <i>2p</i> ⁵	10 Ne <i>2s</i> ² <i>2p</i> ⁶								
3	11 Na <i>3s</i> ¹	12 Mg <i>3s</i> ²											13 Al <i>3s</i> ² <i>3p</i> ¹	14 Si <i>3s</i> ² <i>3p</i> ²	15 P <i>3s</i> ² <i>3p</i> ³	16 S <i>3s</i> ² <i>3p</i> ⁴	17 Cl <i>3s</i> ² <i>3p</i> ⁵	18 Ar <i>3s</i> ² <i>3p</i> ⁶								
4	19 K <i>4s</i> ¹	20 Ca <i>4s</i> ²	21 Sc <i>4s</i> ² <i>3d</i> ¹	22 Ti <i>4s</i> ² <i>3d</i> ²	23 V <i>4s</i> ² <i>3d</i> ³	24 Cr <i>4s</i> ¹ <i>3d</i> ⁵	25 Mn <i>4s</i> ² <i>3d</i> ⁵	26 Fe <i>4s</i> ² <i>3d</i> ⁶	27 Co <i>4s</i> ² <i>3d</i> ⁷	28 Ni <i>4s</i> ² <i>3d</i> ⁸	29 Cu <i>4s</i> ¹ <i>3d</i> ¹⁰	30 Zn <i>4s</i> ² <i>3d</i> ¹⁰	31 Ga <i>4s</i> ² <i>4p</i> ¹	32 Ge <i>4s</i> ² <i>4p</i> ²	33 As <i>4s</i> ² <i>4p</i> ³	34 Se <i>4s</i> ² <i>4p</i> ⁴	35 Br <i>4s</i> ² <i>4p</i> ⁵	36 Kr <i>4s</i> ² <i>4p</i> ⁶								
5	37 Rb <i>5s</i> ¹	38 Sr <i>5s</i> ²	39 Y <i>5s</i> ² <i>4d</i> ¹	40 Zr <i>5s</i> ² <i>4d</i> ²	41 Nb <i>5s</i> ¹ <i>4d</i> ⁴	42 Mo <i>5s</i> ¹ <i>4d</i> ⁵	43 Tc <i>5s</i> ¹ <i>4d</i> ⁶	44 Ru <i>5s</i> ¹ <i>4d</i> ⁷	45 Rh <i>5s</i> ¹ <i>4d</i> ⁸	46 Pd <i>5s</i> ⁰ <i>4d</i> ¹⁰	47 Ag <i>5s</i> ¹ <i>4d</i> ¹⁰	48 Cd <i>5s</i> ⁰ <i>4d</i> ¹⁰	49 In <i>5s</i> ² <i>5p</i> ¹	50 Sn <i>5s</i> ² <i>5p</i> ²	51 Sb <i>5s</i> ² <i>5p</i> ³	52 Te <i>5s</i> ² <i>5p</i> ⁴	53 I <i>5s</i> ² <i>5p</i> ⁵	54 Xe <i>5s</i> ² <i>5p</i> ⁶								
6	55 Cs <i>6s</i> ¹	56 Ba <i>6s</i> ²	57 La* <i>6s</i> ² <i>5d</i> ¹	72 Hf <i>6s</i> ² <i>5d</i> ²	73 Ta <i>6s</i> ² <i>5d</i> ³	74 W <i>6s</i> ² <i>5d</i> ⁴	75 Re <i>6s</i> ² <i>5d</i> ⁵	76 Os <i>6s</i> ² <i>5d</i> ⁶	77 Ir <i>6s</i> ² <i>5d</i> ⁷	78 Pt <i>6s</i> ¹ <i>5d</i> ⁹	79 Au <i>6s</i> ¹ <i>5d</i> ¹⁰	80 Hg <i>6s</i> ² <i>5d</i> ¹⁰	81 Tl <i>6s</i> ² <i>6p</i> ¹	82 Pb <i>6s</i> ² <i>6p</i> ²	83 Bi <i>6s</i> ² <i>6p</i> ³	84 Po <i>6s</i> ² <i>6p</i> ⁴	85 At <i>6s</i> ² <i>6p</i> ⁵	86 Rn <i>6s</i> ² <i>6p</i> ⁶								
7	87 Fr <i>7s</i> ¹	88 Ra <i>7s</i> ²	89 Ac** <i>7s</i> ² <i>6d</i> ¹	104 Rf <i>7s</i> ² <i>6d</i> ²	105 Db <i>7s</i> ² <i>6d</i> ³	106 Sg <i>7s</i> ² <i>6d</i> ⁴	107 Bh <i>7s</i> ² <i>6d</i> ⁵	108 Hs <i>7s</i> ² <i>6d</i> ⁶	109 Mt <i>7s</i> ² <i>6d</i> ⁷	110 Ds <i>7s</i> ² <i>6d</i> ⁸	111 Rg <i>7s</i> ¹ <i>6d</i> ¹⁰	112 Uub <i>7s</i> ² <i>6d</i> ¹⁰	113 Uut <i>7s</i> ² <i>6p</i> ¹	114 Uuq <i>7s</i> ² <i>6p</i> ²	115 Uup <i>7s</i> ² <i>6p</i> ³											
			<i>f</i> -Transition Elements																							
			*Lanthanides																							
			58 Ce <i>6s</i> ² <i>4f</i> ¹ <i>5d</i> ¹	59 Pr <i>6s</i> ² <i>4f</i> ³ <i>5d</i> ⁰	60 Nd <i>6s</i> ² <i>4f</i> ⁴ <i>5d</i> ⁰	61 Pm <i>6s</i> ² <i>4f</i> ⁵ <i>5d</i> ⁰	62 Sm <i>6s</i> ² <i>4f</i> ⁶ <i>5d</i> ⁰	63 Eu <i>6s</i> ² <i>4f</i> ⁷ <i>5d</i> ⁰	64 Gd <i>6s</i> ² <i>4f</i> ⁷ <i>5d</i> ¹	65 Tb <i>6s</i> ² <i>4f</i> ⁹ <i>5d</i> ⁰	66 Dy <i>6s</i> ² <i>4f</i> ¹⁰ <i>5d</i> ⁰	67 Ho <i>6s</i> ² <i>4f</i> ¹¹ <i>5d</i> ⁰	68 Er <i>6s</i> ² <i>4f</i> ¹² <i>5d</i> ⁰	69 Tm <i>6s</i> ² <i>4f</i> ¹³ <i>5d</i> ⁰	70 Yb <i>6s</i> ² <i>4f</i> ¹⁴ <i>5d</i> ⁰	71 Lu <i>6s</i> ² <i>4f</i> ¹⁴ <i>5d</i> ¹										
			**Actinides																							
			90 Th <i>7s</i> ² <i>5f</i> ⁰ <i>6d</i> ²	91 Pa <i>7s</i> ² <i>5f</i> ² <i>6d</i> ¹	92 U <i>7s</i> ² <i>5f</i> ³ <i>6d</i> ¹	93 Np <i>7s</i> ² <i>5f</i> ⁴ <i>6d</i> ¹	94 Pu <i>7s</i> ² <i>5f</i> ⁶ <i>6d</i> ⁰	95 Am <i>7s</i> ² <i>5f</i> ⁷ <i>6d</i> ⁰	96 Cm <i>7s</i> ² <i>5f</i> ⁷ <i>6d</i> ¹	97 Bk <i>7s</i> ² <i>5f</i> ⁹ <i>6d</i> ⁰	98 Cf <i>7s</i> ² <i>5f</i> ¹⁰ <i>6d</i> ⁰	99 Es <i>7s</i> ² <i>5f</i> ¹¹ <i>6d</i> ⁰	100 Fm <i>7s</i> ² <i>5f</i> ¹² <i>6d</i> ⁰	101 Md <i>7s</i> ² <i>5f</i> ¹³ <i>6d</i> ⁰	102 No <i>7s</i> ² <i>5f</i> ¹⁴ <i>6d</i> ⁰	103 Lr <i>7s</i> ² <i>5f</i> ¹⁴ <i>6d</i> ¹										

Figure 11.26

The periodic table with atomic symbols, atomic numbers, and partial electron configurations

Periodic Table and Electron Configurations

1. The group labels for Groups 1, 2, 3, 4, 5, 6, 7, and 8 indicate the *total number* of valence electrons for the atoms in these groups. For example, all the elements in Group 5 have the configuration ns^2np^3 . (Any *d* electrons present are always in the next lower principal energy level than the valence electrons and so are not counted as valence electrons.)
2. The elements in Groups 1, 2, 3, 4, 5, 6, 7, and 8 are often called the **main-group elements**, or **representative elements**. Remember that every member of a given group (except for helium) has the same valence-electron configuration, except that the electrons are in different principal energy levels.

Information

The group label gives the total number of valence electrons for that group.

C. Atomic Properties and the Periodic Table

With all of this talk about electron probability and orbitals, we must not lose sight of the fact that chemistry is still fundamentally a science based on the observed properties of substances. We know that wood burns, steel rusts, plants grow, sugar tastes sweet, and so on because we *observe* these phenomena. The atomic theory is an attempt to help us understand why these things occur. If we understand why, we can hope to better control the chemical events that are so crucial in our daily lives.

In the next chapter we will see how our ideas about atomic structure help us understand how and why atoms combine to form compounds. As we explore this topic, and as we use theories to explain other types of chemical behavior later in the text, it is important that we distinguish the observation (steel rusts) from the attempts to explain why the observed event occurs (theories). The observations remain the same over the decades, but the theories (our explanations) change as we gain a clearer understanding of how nature operates. A good example of this is the replacement of the Bohr model for atoms by the wave mechanical model.

Because the observed behavior of matter lies at the heart of chemistry, you need to understand thoroughly the characteristic properties of the



Gold leaf being applied to the dome of the courthouse in Huntington, West Virginia

HANDS-ON CHEMISTRY

• MINI-LAB •

Which Element Is It?

Materials

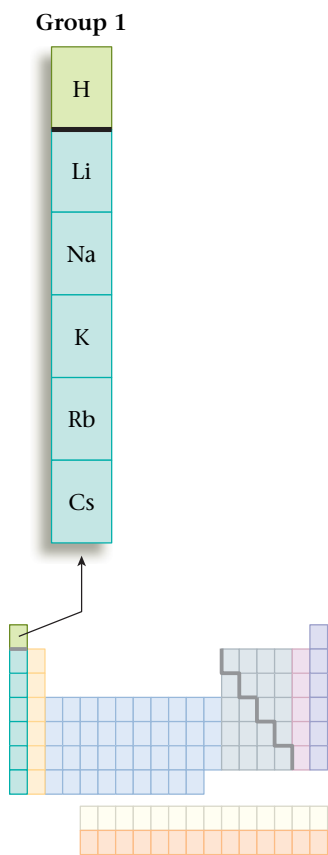
- *World of Chemistry* textbook

Procedure

1. Work in a group of three or four students.
2. Have one person write down one of the representative elements without showing it to the other group members.
3. Take turns asking this person yes/no questions about the element. The person who chose the element must answer these questions (he or she can look through Chapter 3 or 11 for the information).
4. If the element is not identified after each person has asked three questions, the person choosing the element “wins” and identifies the element.
5. Continue until each group member has chosen an element.

Results/Analysis

1. Think about which questions were best to determine the element. What was it about these questions that made them good? How is the periodic table set up so that these questions worked well?



various elements and the trends (systematic variations) that occur in those properties. To that end, we will now consider some especially important properties of atoms and see how they vary, horizontally and vertically, on the periodic table.

Metals and Nonmetals

The most fundamental classification of the chemical elements is into metals and nonmetals. **Metals** typically have the following physical properties: a lustrous appearance, the ability to change shape without breaking (they can be pulled into a wire or pounded into a thin sheet), and excellent conductivity of heat and electricity. **Nonmetals** typically do not have these physical properties, although there are some exceptions. (For example, solid iodine is lustrous; the graphite form of carbon is an excellent conductor of electricity; and the diamond form of carbon is an excellent conductor of heat.) However, it is the *chemical* differences between metals and nonmetals that interest us the most:

Metals tend to lose electrons to form positive ions, and nonmetals tend to gain electrons to form negative ions.

When a metal and a nonmetal react, a transfer of one or more electrons from the metal to the nonmetal often occurs.

Most of the elements are classed as metals, as is shown in **Figure 11.27**. Note that the metals are found on the left side and at the center of the periodic table. The relatively few nonmetals are in the upper right corner of the table. A few elements exhibit both metallic and nonmetallic behavior; they are classed as **metalloids** or semimetals.

It is important to understand that simply being classified as a metal does not mean that an element behaves exactly like all other metals. For example, some metals can lose one or more electrons much more easily than others. In particular, cesium can give up its outermost electron (a 6s electron) more easily than can lithium (a 2s electron). In fact, for the alkali metals (Group 1) the ease of giving up an electron varies as follows:

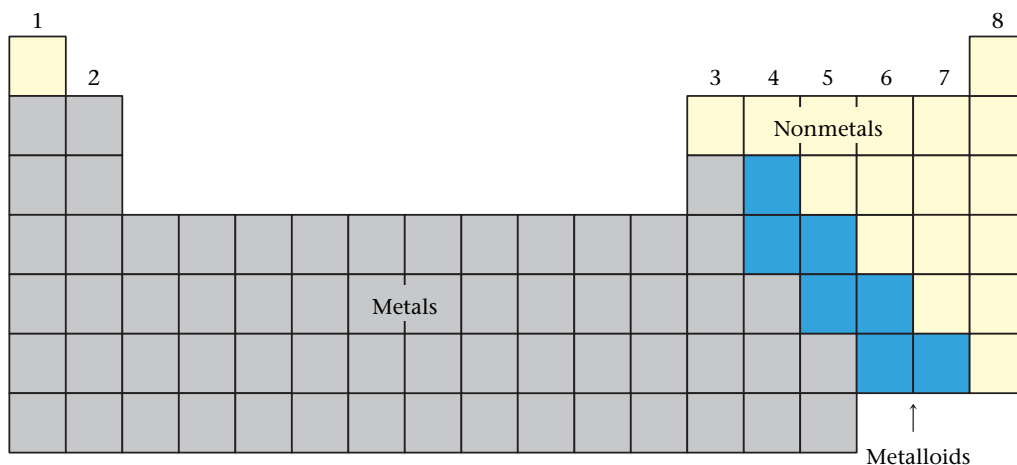
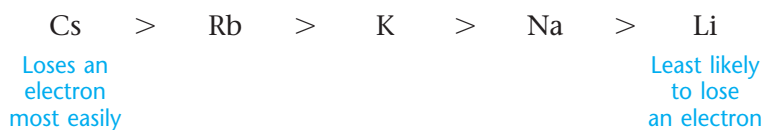


Figure 11.27
The classification of elements as metals, nonmetals, and metalloids

Note that as we go down the group, the metals become more likely to lose an electron. This makes sense because as we go down the group, the electron being removed resides, on average, farther and farther from the nucleus. That is, the 6s electron lost from Cs is much farther from the attractive positive nucleus—and so is easier to remove—than the 2s electron that must be removed from a lithium atom.

The same trend is also seen in the Group 2 metals (alkaline earth metals): the farther down in the group the metal resides, the more likely it is to lose an electron.

Just as metals vary somewhat in their properties, so do nonmetals. In general, the elements that can most effectively pull electrons from metals occur in the upper right corner of the periodic table.

As a general rule, we can say that the

- most chemically active metals appear in the lower left region of the periodic table
- most chemically active nonmetals appear in the upper right region
- properties of the semimetals, or metalloids, lie between the metals and the nonmetals

Atomic Size

The sizes of atoms vary as shown in **Figure 11.28**. Notice that atoms get larger as we go down a group on the periodic table and that they get smaller as we go from left to right across a period.

We can understand the increase in size that we observe as we go down a group by remembering that as the principal energy level increases, the average distance of the electrons from the nucleus also increases. So atoms get bigger as electrons are added to larger principal energy levels.

Group 2

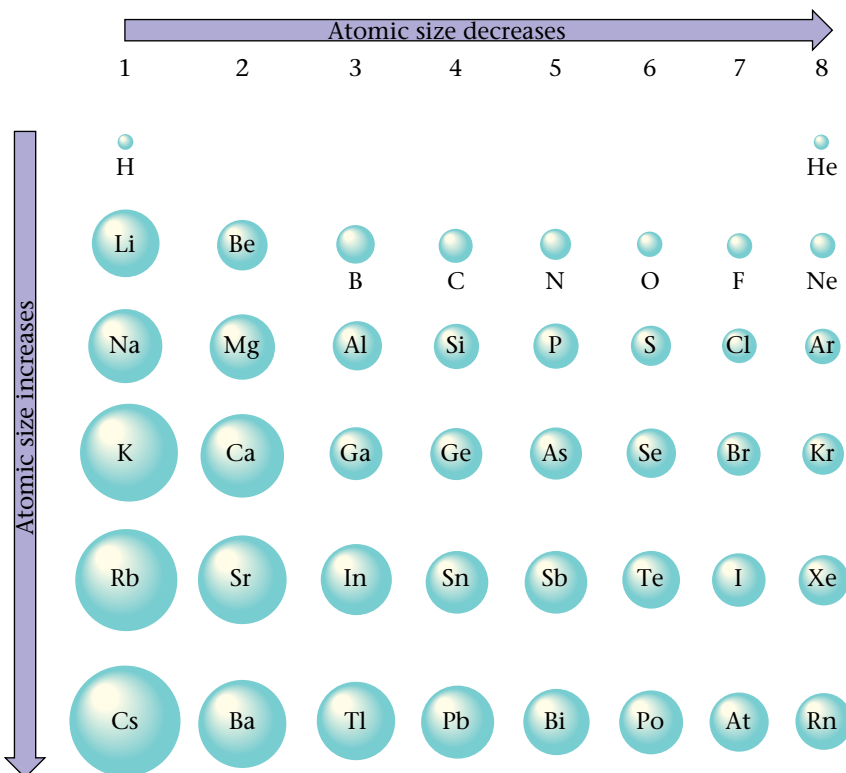
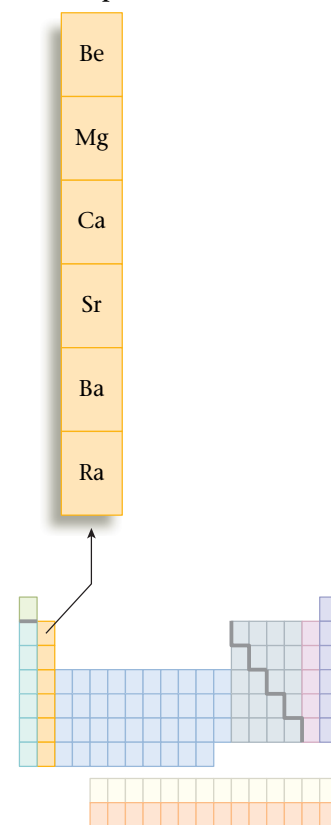


Figure 11.28

Relative atomic sizes for selected atoms. Note that atomic size increases down a group and decreases across a period.

Explaining the decrease in **atomic size** across a period requires some thought about the atoms in a given row (period) of the periodic table. Recall that the atoms in a particular period all have their outermost electrons in a given principal energy level. That is, the atoms in Period 1 have their outer electrons in the 1s orbital (principal energy level 1), the atoms in Period 2 have their outermost electrons in principal energy level 2 (2s and 2p orbitals), and so on (see Figure 11.23). Because all the orbitals in a given principal energy level are expected to be the same size, we might expect the atoms in a given period to be the same size. However, remember that the number of protons in the nucleus increases as we move from atom to atom in the period.

The resulting increase in positive charge on the nucleus tends to pull the electrons closer to the nucleus. So instead of remaining the same size across a period as electrons are added in a given principal energy level, the atoms get smaller as the electron “cloud” is drawn in by the increasing nuclear charge.

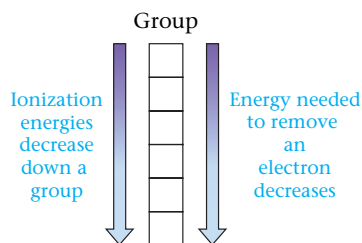
Ionization Energies

The **ionization energy** of an atom is the energy required to remove an electron from an individual atom in the gas phase:

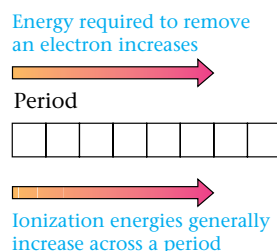


As we have noted, the most characteristic chemical property of a metal atom is losing electrons to nonmetals. Another way of saying this is that *metals have relatively low ionization energies*—a relatively small amount of energy is needed to remove an electron from a typical metal.

Recall that metals at the bottom of a group lose electrons more easily than those at the top. In other words, ionization energies tend to decrease in going from the top to the bottom of a group.



In contrast to metals, nonmetals have relatively large ionization energies. Nonmetals tend to gain, not lose, electrons. Recall that metals appear on the left side of the periodic table and nonmetals appear on the right side. Thus it is not surprising that ionization energies tend to increase from left to right across a given period on the periodic table.



Ionization energy

The amount of energy required to remove an electron from a gaseous atom or ion

In general, the elements in the

- Lower left region of the periodic table have the lowest ionization energies (are the most chemically active metals)
- Upper right region of the periodic table have the highest ionization energies (are the most chemically active nonmetals).

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Fireworks

The art of using mixtures of chemicals to produce explosives is an ancient one. Black powder—a mixture of potassium nitrate, charcoal, and sulfur—was being used in China well before A.D. 1000, and it has been used through the centuries in military explosives, in construction blasting, and for fireworks.

Before the nineteenth century, fireworks were confined mainly to rockets and loud bangs. Orange and yellow colors came from the presence of charcoal and iron filings. However, with the great advances in chemistry in the nineteenth century, new compounds found their way into fireworks. Salts of copper, strontium, and barium added brilliant colors. Magnesium and aluminum metals gave a dazzling white light.

How do fireworks produce their brilliant colors and loud bangs? Actually, only a handful of different chemicals are responsible for most of the spectacular effects. To produce the noise and flashes, an oxidizer (something with a strong affinity for electrons) is reacted with a metal such as magnesium or aluminum mixed with sulfur. The resulting reaction produces a brilliant flash, which is due to the aluminum or magnesium burning, and a loud report is produced by the rapidly expanding gases. For a color effect, an element with a colored flame is included.

Yellow colors in fireworks are due to sodium. Strontium salts give the red color familiar from highway safety flares. Barium salts give a green color.

Although you might think that the chemistry of fireworks is simple, achieving the vivid white flashes and the brilliant colors requires complex combinations of chemicals. For example, because the white flashes

produce high flame temperatures, the colors tend to wash out. Another problem arises from the use of sodium salts. Because sodium produces an extremely bright yellow color, sodium salts cannot be used when other colors are desired. In short, the manufacture of fireworks that produce the desired effects and are also safe to handle requires very careful selection of chemicals.*

*The chemical mixtures in fireworks are very dangerous. *Do not experiment with chemicals on your own.*

Brightly colored fireworks



CELEBRITY CHEMICAL

Hydrogen (H₂)

Hydrogen is the most abundant element in the universe, being the major constituent of all stars. Most of the hydrogen found on earth is in the combined state—in compounds with other elements. The amount of elemental hydrogen (H₂) found on earth is actually very small. The atmosphere contains only 0.00005% H₂ gas. Therefore, when H₂ gas is required to produce a substance such as NH₃ in the reaction



the H₂ must be obtained by decomposing a compound such as water.

Hydrogen is a colorless, odorless, and tasteless gas that reacts explosively with oxygen gas. In fact, the energetic reaction of H₂(g) with O₂(g) makes it a good candidate as a fuel. Hydrogen actually delivers four times as much energy per gram as does gasoline. The other advantage to using

hydrogen as a fuel is that the combustion product is water, which is not harmful to the environment. Currently, hydrogen is too expensive to produce to be competitive as a fuel compared to gasoline or natural gas. However, if cheaper ways are found to produce it, hydrogen could prove a valuable fuel in the future.

Liquid hydrogen is used to power the space shuttle.



SECTION 11.4 REVIEW QUESTIONS

- Write the orbital diagram for the elements listed below:
Li $1s^2 2s^1$
C $1s^2 2s^2 2p^2$
Mg $1s^2 2s^2 2p^6 3s^2$
- What is the difference between a valence electron and a core electron? Select an element from row 3 of the periodic table and label both a valence electron and a core electron using its electron configuration.
- Elements in vertical columns (groups) show similar chemical behavior. How are their electron configurations similar?
- For each of the following elements, refer to its position on the periodic table and write its valence shell electron configuration.
 - hafnium, $Z = 72$
 - radium, $Z = 88$
 - antimony, $Z = 51$
 - lead, $Z = 82$
- What chemical property distinguishes a metal from a nonmetal?
- How can the property in question 5 be explained by using ionization energy trends?
- Explain the general trends of atomic size and ionization energies across rows and down columns of the periodic table.



Key Terms

11.1 Electromagnetic radiation

Wavelength

Frequency

Photons

11.2 Quantized

Wave mechanical model

11.3 Orbital

Principal energy levels

Sublevels

Pauli exclusion principle

11.4 Electron configuration

Orbital diagram (box diagram)

Valence electrons

Core electrons

Lanthanide series

Actinide series

Main-group elements
(representative elements)

Metals

Nonmetals

Metalloids

Atomic size

Ionization energy

Key Ideas

11.1 Atoms and Energy

- Rutherford's atom consists of a tiny, dense nucleus at the center and electrons which occupy most of the volume of the atom.
- Electromagnetic radiation
 - Characterized by its wavelength and frequency
 - Can be thought of as a stream of packets of energy called photons
 - Atoms can gain energy by absorbing a photon or lose energy by emitting a photon.

11.2 The Hydrogen Atom

- The hydrogen atom can emit only certain energies as it changes from a higher to a lower energy.
- Hydrogen has quantized energy levels.

11.3 Atomic Orbitals

- The Bohr model assumed electrons travel around the nucleus in circular orbits which is incorrect.
- The wave mechanical model assumes the electron has both particle and wave properties and describes electrons as occupying orbitals.
 - The orbitals are different from the Bohr orbits.
 - Probability maps indicate the likelihood of finding the electron at a given point in space.
 - The size of an atom can be described by a surface that contains 90% of the total electron probability.

11.4 Electron Configurations and Atomic Properties

- Atomic energy levels are broken down into principal levels (n) which contain various numbers of sublevels.
 - The sublevels represent various types of orbitals (s, p, d, f) which have different shapes.
 - The number of sublevels increases as n increases.
- A given atom has Z protons in its nucleus and Z electrons surrounding the nucleus.
- The electrons occupy atomic orbitals starting with the lowest energy (the orbital closest to the nucleus).
- The Pauli exclusion principle states that an orbital can hold only two electrons with opposite spins.
- The electrons in the highest energy level are called valence electrons.
- The electron arrangement for a given atom explains its position on the periodic table.
- Atomic size generally increases down a group of the periodic table and decreases across a period.
- Ionization energy generally decreases down a group and increases across a period.

All exercises with **blue numbers** have answers in the back of this book.

11.1 Atoms and Energy

A. Rutherford's Atom

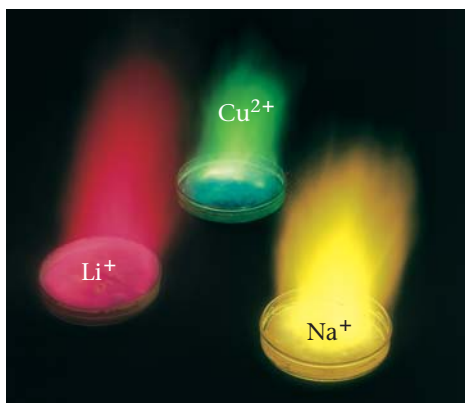
- Describe the experiment that allowed Rutherford to conclude that the atom consists of a relatively heavy positively charged nucleus surrounded by electrons.
- Sketch a picture of the model of the atom that Rutherford developed. What questions were left unanswered by Rutherford's model?

B. Energy and Light

- What do we mean by *electromagnetic radiation*? Give several examples of different sorts of electromagnetic radiation.
- How are the different types of electromagnetic radiation similar? How do they differ?
- What does the *wavelength* of electromagnetic radiation represent? Sketch a representation of a wave and indicate the *wavelength* on your graph. How does the *wavelength* of electromagnetic radiation differ from its *frequency*?
- What do we mean by the *speed* of electromagnetic radiation? How do the *frequency* and the *speed* of electromagnetic radiation differ?
- What is a "packet" of electromagnetic energy called?
- What is meant by the *wave-particle* nature of light?

C. Emission of Energy by Atoms

- Do the colors of flame tests result from taking in energy or releasing energy?
- In which case are higher-energy photons released, with Li^+ or with Cu^{2+} ? How do you know?



11.2 The Hydrogen Atom

A. The Energy Levels of Hydrogen

- What does it mean to say that an atom is in an "excited state"?
- When an atom in an excited state returns to its ground state, what happens to the excess energy of the atom?
- Describe briefly why the study of electromagnetic radiation has been important to our understanding of the arrangement of electrons in atoms.
- What does it mean to say that the hydrogen atom has *discrete energy levels*? How is this fact reflected in the radiation that excited hydrogen atoms emit?
- What experimental evidence do scientists have that the energy levels of hydrogen are *quantized*?
- What is meant by the *ground state* of an atom?

B. The Bohr Model of the Atom

- What are the essential points of Bohr's theory of the structure of the hydrogen atom?
- According to Bohr, what types of motions do electrons have in an atom, and what happens when energy is applied to the atom?
- How does the Bohr theory account for the observed phenomenon of the emission of discrete wavelengths of light by excited atoms?
- Why was Bohr's theory for the hydrogen atom initially accepted, and why was it ultimately discarded?

C. The Wave Mechanical Model of the Atom

- What major assumption (that was analogous to what had already been demonstrated for electromagnetic radiation) did de Broglie and Schrödinger make about the motion of tiny particles?
- Discuss briefly the difference between an orbit (as described by Bohr for hydrogen) and an orbital (as described by the more modern, wave mechanical picture of the atom).
- We cannot *exactly* specify the location of an electron in an atom but can discuss only where an electron is *most likely* to be at any given time. How does the concept of an orbital show this?

11.3 Atomic Orbitals

A. The Hydrogen Orbitals

- 24.** Why are the orbitals of the hydrogen atom described as “probability maps”? Why are the edges of the hydrogen orbitals sometimes drawn to appear “fuzzy”?
- 25.** When we draw a picture of an orbital, we are indicating that the probability of finding the electron within this region of space is greater than 90%. Why is this probability never 100%?
- 26.** What are the differences between the 2s orbital and the 1s orbital of hydrogen? How are they similar?
- 27.** How is the *distance* of an electron from the nucleus related to the *principal energy level*? Does a *higher* principal energy level mean that the electron is *closer* to or *farther* away from the nucleus?
- 28.** If an electron moves from the 1s orbital to the 2s orbital, its energy (increases/decreases).
- 29.** Although a hydrogen atom has only one electron, the hydrogen atom possesses a complete set of available orbitals. What purpose do these additional orbitals serve?

B. The Wave Mechanical Model: Further Development

- 30.** What is the *Pauli exclusion principle*? How many electrons can occupy an orbital, according to this principle? Why?
- 31.** How does the *energy* of a principal energy level depend on the value of *n*? Does a higher value of *n* mean a higher or lower energy?
- 32.** The number of sublevels in a principal energy level (increases/decreases) as *n* increases.
- 33.** Which of the following orbital designations is (are) not correct?
- 2p
 - 1d
 - 3f
 - 4s

11.4 Electron Configurations and Atomic Properties

A. Electron Arrangements in the First 18 Atoms on the Periodic Table

- 34.** Which orbital is the *first* to be filled in any atom? Why?

- 35.** Which electrons of an atom are the *valence* electrons? Why are these electrons especially important?
- 36.** How are the electron arrangements in a given group (vertical column) of the periodic table related? How is this relationship shown in the properties of the elements in the given group?
- 37.** Write the full electron configuration ($1s^22s^2$, etc.) for each of the following elements.
- strontium, $Z = 38$
 - zinc, $Z = 30$
 - helium, $Z = 2$
 - bromine, $Z = 35$
- 38.** Write the full electron configuration ($1s^22s^2$, etc.) for each of the following elements.
- calcium, $Z = 20$
 - potassium, $Z = 19$
 - fluorine, $Z = 9$
 - krypton, $Z = 36$
- 39.** Write the complete orbital diagram for each of the following elements, using boxes to represent orbitals and arrows to represent electrons.
- aluminum, $Z = 13$
 - phosphorus, $Z = 15$
 - bromine, $Z = 35$
 - argon, $Z = 18$
- 40.** How many valence electrons does each of the following atoms possess?
- sodium, $Z = 11$
 - calcium, $Z = 20$
 - iodine, $Z = 53$
 - nitrogen, $Z = 7$

B. Electron Configurations and the Periodic Table

- 41.** Why do we believe that the valence electrons of calcium and potassium reside in the 4s orbital rather than in the 3d orbital?
- 42.** Would you expect the valence electrons of rubidium and strontium to reside in the 5s or the 4d orbitals? Why?

- 43.** Using the symbol of the previous noble gas to indicate the core electrons, write the valence-electron configuration for each of the following elements.
- calcium, $Z = 20$
 - francium, $Z = 87$
 - yttrium, $Z = 39$
 - cerium, $Z = 58$
- 44.** Using the symbol of the previous noble gas to indicate the core electrons, write the valence-electron configuration for each of the following elements.
- phosphorus, $Z = 15$
 - chlorine, $Z = 17$
 - magnesium, $Z = 12$
 - zinc, $Z = 30$
- 45.** How many $4d$ electrons are found in each of the following elements?
- yttrium, $Z = 39$
 - zirconium, $Z = 40$
 - strontium, $Z = 38$
 - cadmium, $Z = 48$
- 46.** For each of the following elements, indicate which set of orbitals is being filled last.
- plutonium, $Z = 94$
 - nobelium, $Z = 102$
 - praseodymium, $Z = 59$
 - radon, $Z = 86$
- 47.** Write the shorthand valence-electron configuration of each of the following elements, basing your answer on the element's location on the periodic table.
- uranium, $Z = 92$
 - manganese, $Z = 25$
 - mercury, $Z = 80$
 - francium, $Z = 87$

C. Atomic Properties and the Periodic Table

- 48.** What types of ions do the metals and the nonmetallic elements form? Do the metals lose or gain electrons in doing this? Do the nonmetallic elements gain or lose electrons in doing this?
- 49.** Give some similarities that exist among the elements of Group 1.
- 50.** Give some similarities that exist among the elements of Group 7.
- 51.** Where are the most nonmetallic elements located on the periodic table? Why do these elements pull electrons from metallic elements so effectively during a reaction?
- 52.** Why do the metallic elements of a given period (horizontal row) typically have much lower ionization energies than do the nonmetallic elements of the same period?
- 53.** Explain why the atoms of the elements at the bottom of a given group (vertical column) of the periodic table are *larger* than the atoms of the elements at the top of the same group.
- 54.** Though all the elements in a given period (horizontal row) of the periodic table have their valence electrons in the same types of orbitals, the sizes of the atoms decrease from left to right within a period. Explain why.
- 55.** In each of the following sets of elements, which element shows the least ability to gain or lose electrons?
- Cs, Rb, Na
 - Ba, Ca, Be
 - F, Cl, Br
 - O, Te, S
- 56.** In each of the following sets of elements, which element would be expected to have the highest ionization energy?
- Cs, K, Li
 - Ba, Sr, Ca
 - I, Br, Cl
 - Mg, Si, S
- 57.** Arrange the following sets of elements in order of increasing atomic size.
- Sn, Xe, Rb, Sr
 - Rn, He, Xe, Kr
 - Pb, Ba, Cs, At

Critical Thinking

58. The distance in meters between two consecutive peaks (or troughs) in a wave is called the _____.
59. The speed at which electromagnetic radiation moves through a vacuum is called the _____.
60. The energy levels of hydrogen (and other atoms) are _____, which means that only certain values of energy are allowed.
61. In the modern theory of the atom, a(n) _____ represents a region of space in which there is a high probability of finding an electron.
62. Without referring to your textbook or a periodic table, write the full electron configuration, the orbital box diagram, and the noble gas short-hand configuration for the elements with the following atomic numbers.
- $Z = 19$
 - $Z = 22$
 - $Z = 14$
 - $Z = 26$
 - $Z = 30$
63. Write the general valence configuration (for example, ns^1 for Group 1) for the group in which each of the following elements is found.
- barium, $Z = 56$
 - bromine, $Z = 35$
 - tellurium, $Z = 52$
 - potassium, $Z = 19$
 - sulfur, $Z = 16$
64. How many valence electrons does each of the following atoms have?
- titanium, $Z = 22$
 - iodine, $Z = 53$
 - radium, $Z = 88$
 - manganese, $Z = 25$
65. How do we know that the energy levels of the hydrogen atom are not *continuous*, as physicists originally assumed?
66. Why do the two electrons in the $2p$ sublevel of carbon occupy *different* $2p$ orbitals?
67. Write the full electron configuration ($1s^22s^2$, etc.) for each of the following elements.
- bromine, $Z = 35$
 - xenon, $Z = 54$
 - barium, $Z = 56$
 - selenium, $Z = 34$
68. Metals have relatively (low/high) ionization energies, whereas nonmetals have relatively (high/low) ionization energies.
69. In each of the following sets of elements, indicate which element has the smallest atomic size.
- Ba, Ca, Ra
 - P, Si, Al
 - Rb, Cs, K
70. A $2+$ ion of a particular element has an atomic mass of 203 and 123 neutrons in its nucleus.
- What is the nuclear charge in this element?
 - What are the symbol and name of this element?
 - How many valence electrons does a neutral atom of this element have?
 - How many levels of electrons are occupied when this element is in its ground state?

- 71.** Given elements with the following electron configurations, specify which meet the following conditions.
- i.** $1s^2 2s^2 2p^6 3s^2$
 - ii.** $1s^2 2s^2 2p^6 3s^1$
 - iii.** $1s^2 2s^2 2p^6$
 - iv.** $1s^2 2s^2 2p^4$
 - v.** $1s^2 2s^2 2p^3$
 - vi.** $1s^2 2s^2 2p^1$
- a. Which will normally form a negative ion in solution?
 - b. Which will have the lowest ionization energy?
 - c. Which would be the most reactive metal?
 - d. Which has the greatest number of unpaired electrons?
 - e. Which metal would combine with oxygen in a one-to-one ratio?
 - f. Predict the formula for the compound that would form between (ii) and (v).
 - g. Predict the formula for the compound that would form between (i) and (v).
 - h. List the six elements in order of increasing ionization energy.

72. Identify the following elements.

- a. the most reactive nonmetal (not a noble gas) on the periodic table
- b. the smallest lanthanide
- c. the largest metalloid
- d. the metal with the lowest ionization energy
- e. the least reactive halogen
- f. the least reactive member of the fifth period
- g. the actinide with the highest nuclear charge

73. Suppose that element 120 has just been discovered. Identify four chemical or physical properties expected of this element.



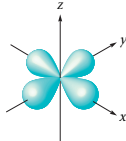
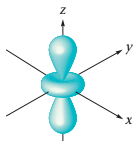
74. Suppose that in another universe there is a completely different set of elements from the ones we know. The inhabitants assign symbols of A, B, C, and so on in order of increasing atomic number, using all of the letters of our alphabet. They find that the following elements closely resemble one another in terms of their properties:

A, C, G, K, Q, and Y

B, F, J, P, and X

V is its own family

Sketch a periodic table, similar to ours, using this information.

- 1 The form of electromagnetic radiation that has less energy per photon than infrared rays but more energy per photon than radio waves is
- A microwaves C gamma rays
B ultraviolet D X rays
- 2 How many electrons can be contained in all of the orbitals with a principal energy level of 3 ($n = 3$)?
- A 8 C 18
B 10 D 32
- 3 In general, the rows on the periodic table correspond to _____ and the columns numbered 1A, 2A, . . . 8A correspond to _____.
- A the orbital of the valence electrons; the number of total electrons
B the energy level of the valence electrons; the number of valence electrons
C the energy level for the core electrons; the orbital of the valence electrons
D the number of valence electrons; the energy level of the valence electrons
- 4 An orbital is best described as
- A the pathway for an electron.
B a region of probability of finding an electron.
C the space in which only valence electrons exist.
D a physical structure which holds the electrons.
- 5 Which of the following is true about the trend of ionization energy across a row on the periodic table?
- A Ionization energy generally increases to the right because the number of protons increases from left to right.
B Ionization energy generally decreases to the right because the number of electrons increases from left to right.
C Ionization energy generally decreases to the right because the number of protons increases from left to right.
D Ionization energy is constant since all atoms in a row have the same valence energy level.
- 6 Which of the following is a representation of a p orbital?
- A  C 
B  D 
- 7 For the atoms Li, N, F, and Na, which of the following is the correct order from smallest to largest atomic radius?
- A Na, F, N, Li
B Na, Li, N, F
C N, F, Na, Li
D F, N, Li, Na
- 8 Which of the following correctly ranks the ionization energies of O, F, Na, S, and Cs from smallest to largest?
- A Cs, Na, S, O, F
B Cs, S, Na, O, F
C F, O, Na, S, Cs
D F, O, S, Na, Cs
- 9 Write the expected electron configuration for the ground state of iron.
- 10 Fill in the following orbital diagram for a ground state oxygen atom
- 1s 2s 2p 3s 3p
- 11 Fill in the following orbital diagram for a boron atom in an excited state.
- 1s 2s 2p 3s 3p
- 12 True or false? Across a row of the periodic table, smaller atoms generally have smaller ionization energies. If true, provide an example and explain your answer. If false, provide a counterexample and explain your answer.

LOOKING AHEAD

12.1 Characteristics of Chemical Bonds

- A. Types of Chemical Bonds
- B. Electronegativity
- C. Bond Polarity and Dipole Moments

12.2 Characteristics of Ions and Ionic Compounds

- A. Stable Electron Configurations and Charges on Ions
 - Predicting Formulas of Ionic Compounds
- B. Ionic Bonding and Structures of Ionic Compounds
 - Ionic Compounds Containing Polyatomic Ions

12.3 Lewis Structures

- A. Writing Lewis Structures
- B. Lewis Structures of Molecules with Multiple Bonds
 - Some Exceptions to the Octet Rule

12.4 Structures of Molecules

- A. Molecular Structure
- B. The VSEPR Model
- C. Molecules with Double Bonds

The atomium in Brussels, Belgium, is a giant-sized model of a crystal of iron in which the atoms are bound to each other.



IN YOUR LIFE

The world around us is composed almost entirely of compounds and mixtures of compounds. Rocks, coal, soil, petroleum, trees, and human beings are complex mixtures of chemical compounds in which different kinds of atoms are bound together. Most of the pure elements found in the earth's crust also contain many atoms bound together. In a gold nugget, each gold atom is bound to many other gold atoms; and in a diamond, many carbon atoms are bonded very strongly to each other. Substances composed of unbound atoms do exist in nature, but they are very rare. (Examples include the argon atoms in the atmosphere and the helium atoms found in natural gas reserves.)

The manner in which atoms are bound together has a profound effect on the chemical and physical properties of substances. For example, both graphite and diamond are composed solely of carbon atoms. However, graphite is a soft, slippery material used as a lubricant in locks; and diamond is one of the hardest materials known, valuable both as a gemstone and in industrial cutting tools. Why do these materials, both composed solely of carbon atoms, have such different properties? The answer lies in the different ways in which the carbon atoms are bound to each other in these substances.

Diamond, composed of carbon atoms bonded together to produce one of the hardest materials known, makes a beautiful gemstone.

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WHAT DO YOU KNOW?

Prereading Questions

1. When representing a water molecule as H—O—H, what do the lines between the letters symbolize?
2. Write the electron configurations for fluorine, neon, and sodium.
3. Why is the term *sodium chloride molecule* incorrect?

SECTION 12.1

Characteristics of Chemical Bonds

Key Terms

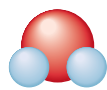
- Bond
- Bond energy
- Ionic bonding
- Ionic compound
- Covalent bonding
- Polar covalent bond
- Electronegativity
- Dipole moment

Objectives

- To learn about ionic and covalent bonds and explain how they are formed
- To learn about the polar covalent bond
- To understand the nature of bonds and their relationship to electronegativity
- To understand bond polarity and how it is related to molecular polarity

Molecular bonding and structure play the central role in determining the course of chemical reactions, many of which are vital to our survival. Most reactions in biological systems are very sensitive to the structures of the participating molecules; in fact, very subtle differences in shape sometimes serve to channel the chemical reaction one way rather than another. Molecules that act as drugs must have exactly the right structure to perform their functions correctly. Structure also plays a central role in our senses of smell and taste. Substances have a particular odor because they fit into the specially shaped receptors in our nasal passages, and taste is also dependent on molecular shape.

To understand the behavior of natural materials, we must understand the nature of chemical bonding and the factors that control the structures of compounds. In this chapter, we will present various classes of compounds that illustrate the different types of bonds. We will then develop models to describe the structure and bonding that characterize the materials found in nature.



A water molecule

A. Types of Chemical Bonds

What is a chemical bond? Although there are several ways to answer this question, we will define a **bond** as a force that holds groups of two or more atoms together and makes them function as a unit. For example, in water the fundamental unit is the H—O—H molecule, which we describe as being held together by the two O—H bonds. We can obtain information about the strength of a bond by measuring the energy required to break the bond, the **bond energy**.

Atoms can interact with one another in several ways to form aggregates. We will consider specific examples to illustrate the various types of chemical bonds.

Ionic Bonding In Chapter 8 we saw that when solid sodium chloride is dissolved in water, the resulting solution conducts electricity, a fact that convinces chemists that sodium chloride is composed of Na^+ and Cl^- ions. Thus, when sodium and chlorine react to form sodium chloride, electrons are transferred from the sodium atoms to the chlorine atoms to form Na^+ and Cl^- ions, which then aggregate to form solid sodium chloride. The resulting solid sodium chloride is a very sturdy material; it has a melting point of approximately $800\text{ }^\circ\text{C}$. See **Figure 12.1**.

Bond (chemical bond)

The force that holds two or more atoms together and makes them function as a unit

Bond energy

The energy required to break a given chemical bond

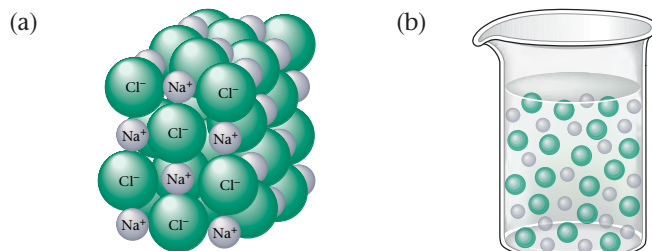
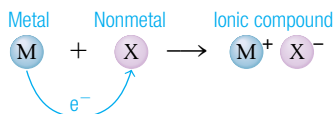


Figure 12.1

(a) Sodium chloride contains ions closely packed together in the solid. (b). When dissolved in water, these ions become free to move around

The strong bonding forces present in sodium chloride result from the attractions among the closely packed, oppositely charged ions. This is an example of **ionic bonding**. Ionic substances are formed when an atom that loses electrons relatively easily reacts with an atom that has a high affinity for electrons. In other words, an **ionic compound** results when a metal reacts with a nonmetal.



We have seen that a bonding force develops when two very different types of atoms react to form oppositely charged ions. But how does a bonding force develop between two identical atoms? Let's explore this situation by considering what happens when two hydrogen atoms are brought close together, as shown in **Figure 12.2**. When hydrogen atoms are close together, the two electrons are simultaneously attracted to both nuclei. Note in Figure 12.2b how the electron probability increases between the two nuclei, indicating that the electrons are shared by the two nuclei.

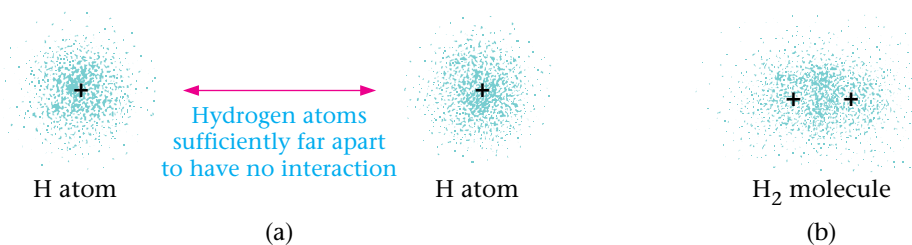


Figure 12.2

The formation of a bond between two hydrogen atoms. (a) Two separate hydrogen atoms. (b) When two hydrogen atoms come close together, the two electrons are attracted simultaneously by both nuclei. This produces the bond. Note the relatively large electron probability between the nuclei, indicating sharing of the electrons.

Covalent bonding The type of bonding we encounter in the hydrogen molecule and in many other molecules where *electrons are shared by nuclei* is called **covalent bonding**. Note that in the H₂ molecule, the electrons reside primarily in the space between the two nuclei, where they are attracted simultaneously by both protons. Although we will not go into detail about it here, the increased attractive forces in this area lead to the formation of the H₂ molecule from the two separated hydrogen atoms. When we say that a bond is formed between the hydrogen atoms, we mean that the H₂ molecule is more stable than two separated hydrogen atoms by a certain quantity of energy (the bond energy).

So far we have considered two extreme types of bonding. In ionic bonding, the participating atoms are so different that one or more electrons are transferred to form oppositely charged ions. The bonding results from the attractions between these ions. In covalent bonding, two identical atoms share electrons equally. The bonding results from the mutual attraction of the two nuclei for the shared electrons. Between these extremes are intermediate cases in which the atoms are not so different that electrons are completely transferred but are different enough so that unequal sharing

Ionic bonding

The attraction between oppositely charged ions

Ionic compound

A compound that results when a metal reacts with a nonmetal to form cations and anions

Covalent bonding

A type of bonding in which atoms share electrons

Polar covalent bond

A covalent bond in which the electrons are not shared equally because one atom attracts the shared electrons more than the other atom

of electrons results, forming what is called a **polar covalent bond**. The hydrogen fluoride (HF) molecule contains this type of bond, which produces the charge distribution



where δ (delta) is used to indicate a partial or fractional charge.

The most logical explanation for the development of *bond polarity* (the partial positive and negative charges on the atoms in such molecules as HF) is that the electrons in the bonds are not shared equally. For example, we can account for the polarity of the HF molecule by assuming that the fluorine atom has a stronger attraction than the hydrogen atom for the shared electrons (see **Figure 12.3**). Because bond polarity has important chemical implications, we find it useful to assign a number that indicates an atom's ability to attract shared electrons.

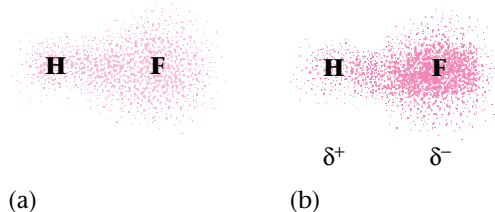


Figure 12.3

Probability representations of the electron sharing in HF. (a) What the probability map would look like if the two electrons in the H—F bond were shared equally. (b) The actual situation, where the shared pair spends more time close to the fluorine atom than to the hydrogen atom. This gives fluorine a slight excess of negative charge and the hydrogen a slight deficit of negative charge (a slight positive charge).

Active Reading Question

Provide an example of an ionic bond, an example of a polar covalent bond, and an example of a covalent bond.

B. Electronegativity

We have seen that when a metal and a nonmetal react, one or more electrons are transferred from the metal to the nonmetal to give ionic bonding. On the other hand, two identical atoms react to form a covalent bond in which electrons are shared equally. When *different* nonmetals react, a bond forms in which electrons are shared *unequally*, giving a polar covalent bond. The unequal sharing of electrons between two atoms is described by a property called **electronegativity**: *the relative ability of an atom in a molecule to attract shared electrons to itself*.

Chemists determine electronegativity values for the elements (see **Figure 12.4**) by measuring the polarities of the bonds between various atoms.

Electronegativity generally

- increases going from left to right across a period
- decreases going down a group for the representative elements.

The range of electronegativity values goes from 4.0 for fluorine to 0.7 for cesium and francium. Remember, the higher the atom's electronegativity value, the closer the shared electrons tend to be to that atom when it forms a bond.

Electronegativity

The tendency of an atom in a molecule to attract shared electrons to itself

Increasing electronegativity

		H 2.1															
Decreasing electronegativity	Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
	Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
	K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
	Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5
	Cs 0.7	Ba 0.9	La-Lu 1.0-1.2	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.9	Bi 1.9	Po 2.0	At 2.2
	Fr 0.7	Ra 0.9	Ac 1.1	Th 1.3	Pa 1.4	U 1.4	Np-No 1.4-1.3										

Key

- < 1.5
- 1.5–1.9
- 2.0–2.9
- 3.0–4.0

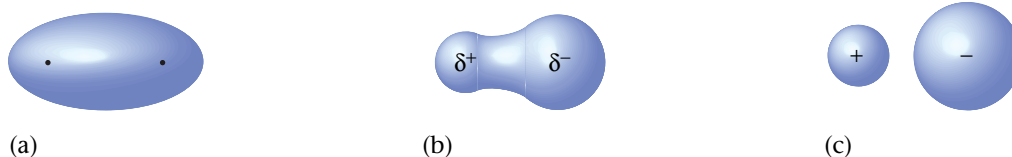
Figure 12.4

Electronegativity values for selected elements. Note that metals have relatively low electronegativity values and that nonmetals have relatively high values.

The polarity of a bond depends on the *difference* between the electronegativity values of the atoms forming the bond.

If the atoms have very similar electronegativities, the electrons are shared almost equally and the bond shows little polarity. If the atoms have very different electronegativity values, a very polar bond is formed. In extreme cases one or more electrons are actually transferred, forming ions and an ionic bond. For example, when an element from Group 1 (electronegativity values of about 0.8) reacts with an element from Group 7 (electronegativity values of about 3), ions are formed and an ionic substance results. In general, if the *difference* between the electronegativities of two elements is about 2.0 or greater, the bond is considered to be ionic.

The relationship between electronegativity and bond type is shown in **Table 12.1**. The various types of bonds are summarized in **Figure 12.5**.

**Figure 12.5**

The three possible types of bonds are (a) a covalent bond formed between identical atoms; (b) a polar covalent bond, with both ionic and covalent components; and (c) an ionic bond, with no electron sharing.

Active Reading Question

In general, how do electronegativity trends compare to trends of atomic size?

CRITICAL THINKING

We use differences in electronegativity to account for certain properties of bonds.

What if all atoms had the same electronegativity values? How would bonding between atoms be affected? What are some differences we would notice?

Table 12.1

The Relationship Between Electronegativity and Bond Type

Electronegativity Difference Between the Bonding Atoms	Bond Type	Covalent Character	Ionic Character
Zero	Covalent		
Intermediate	Polar covalent		
Large	Ionic		

EXAMPLE 12.1

Using Electronegativity to Determine Bond Polarity

Using the electronegativity values given in Figure 12.4, arrange the following bonds in order of increasing polarity: H—H, O—H, Cl—H, S—H, and F—H.

Solution

The polarity of the bond increases as the difference in electronegativity increases. From the electronegativity values in Figure 12.4, the following variation in bond polarity is expected (the electronegativity value appears in parentheses below each element).

Bond	Electronegativity Value	Difference in Electronegativity Values	Bond Type	Polarity
H—H	(2.1) (2.1)	$2.1 - 2.1 = 0$	Covalent	
S—H	(2.5) (2.1)	$2.5 - 2.1 = 0.4$	Polar covalent	
Cl—H	(3.0) (2.1)	$3.0 - 2.1 = 0.9$	Polar covalent	
O—H	(3.5) (2.1)	$3.5 - 2.1 = 1.4$	Polar covalent	
F—H	(4.0) (2.1)	$4.0 - 2.1 = 1.9$	Polar covalent	

Therefore, in order of increasing polarity, we have



Least polar

Most polar

Practice Problem • Exercise 12.1

For each of the following pairs of bonds, choose the bond that is more polar.

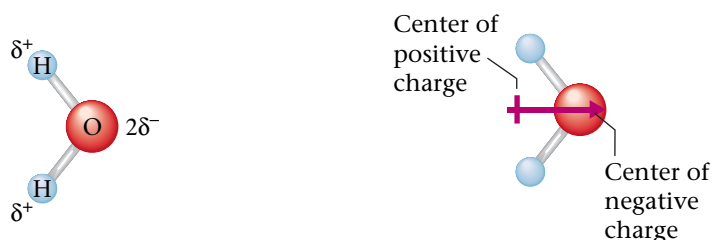
- a. H—P, H—C c. N—O, S—O
 b. O—F, O—I d. N—H, Si—H

C. Bond Polarity and Dipole Moments

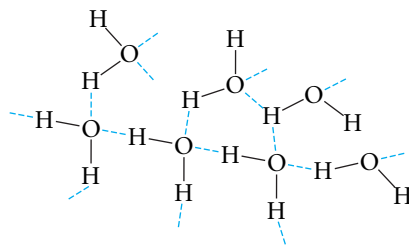
We have seen that hydrogen fluoride has a positive end and a negative end. A molecule such as HF that has a center of positive charge and a center of negative charge is said to have a **dipole moment**. The dipolar character of a molecule is often represented by an arrow. This arrow points toward the negative charge center, and its tail indicates the positive center of charge:



Any diatomic (two-atom) molecule that has a polar bond has a dipole moment. Some polyatomic (more than two atoms) molecules also have dipole moments. For example, because the oxygen atom in the water molecule has a greater electronegativity than the hydrogen atoms, the electrons are not shared equally. This results in a charge distribution that causes the molecule to behave as though it had two centers of charge—one positive and one negative. So the water molecule has a dipole moment.



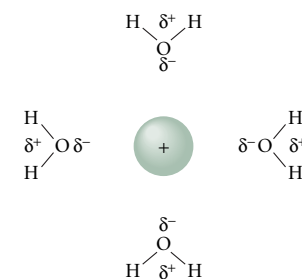
The fact that the water molecule is polar (has a dipole moment) has a profound effect on its properties. In fact, it is not overly dramatic to state that the polarity of the water molecule is crucial to life as we know it on earth. Because water molecules are polar, they can surround and attract both positive and negative ions (see **Figure 12.6**). These attractions allow ionic materials to dissolve in water. Also, the polarity of water molecules causes them to attract each other strongly.



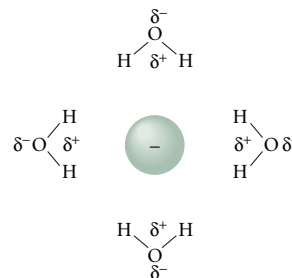
This means that much energy is required to change water from a liquid to a gas (the molecules must be separated from each other to undergo this change of state). Therefore, it is the polarity of the water molecule that causes water to remain a liquid at the temperatures on the earth's surface. If it were nonpolar, water would be a gas and the oceans would be empty.

Dipole moment

A property of a molecule in which the charge distribution can be represented by a center of positive charge and a center of negative charge



(a)



(b)

Figure 12.6

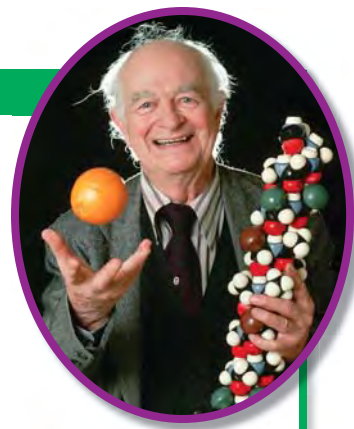
(a) Polar water molecules are strongly attracted to positive ions by their negative ends. (b) They are also strongly attracted to negative ions by their positive ends.

CHEMISTRY EXPLORERS

Linus Pauling • 1901–1994

Linus Pauling is the only person in history to receive two unshared Nobel Prizes. He received the Nobel Prize in Chemistry in 1954, and the Nobel Peace Prize in 1962. As a professor of chemistry at California Institute for Technology, Pauling developed the electronegativity scale of the elements and was awarded the Nobel Prize in Chemistry for his work on the chemical bond. In 1939 he published *The Nature of the Chemical Bond*, which is often cited as the most influential science book of the twentieth century. Pauling was an opponent of nuclear weapons and was awarded the Nobel Peace Prize for his campaign against open-air nuclear testing.

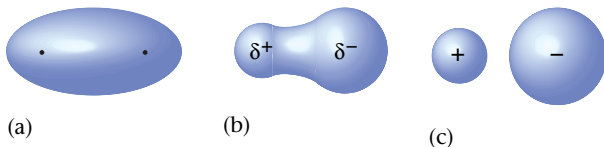
Linus Pauling was a man whose interests varied from chemical bonds to world peace; from electric cars to human health. The British publication *New Scientist* ranked Linus Pauling as one of the twenty greatest scientists of all time, with Albert Einstein being the only other scientist included from the twentieth century. Very good company indeed!



SECTION 12.1

REVIEW QUESTIONS

- 1 What is meant by the term *chemical bond*? What subatomic particles are most important in bonds?
- 2 How are ionic bonds and covalent bonds different?
- 3 How is a polar covalent bond different from a covalent bond?
- 4 Match the compounds HF, NaCl, and O₂ with the following figures.
 - a. H—Cl
 - b. H—H
 - c. H—I
 - d. Br—Br
 - e. CO
- 5 How do electronegativity values help in determining the polarity of a bond?
- 6 For each of the binary molecules below, draw an arrow under the molecules showing its dipole moment. If it has none, write “none.”



Characteristics of Ions and Ionic Compounds

Objectives

- To learn about stable electron configurations
- To learn to predict the formulas of ionic compounds
- To learn about the structures of ionic compounds
- To understand factors governing ionic size

A. Stable Electron Configurations and Charges on Ions

We have seen many times that when a metal and a nonmetal react to form an ionic compound, the metal atom loses one or more electrons to the nonmetal.

Let's Review

Charges on Ions

- Group 1 metals always form 1+ cations.
- Group 2 metals always form 2+ cations.
- Aluminum in Group 3 always forms a 3+ cation.
- Group 7 nonmetals form 1- anions.
- Group 6 elements always form 2- anions.

This is illustrated in **Table 12.2**.

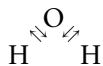
Notice something very interesting about the ions in Table 12.2: all have the electron configuration of neon, a noble gas. That is, sodium loses its one valence electron (the 3s) to form Na^+ , which has a [Ne] electron configuration. Likewise, Mg loses its two valence electrons to form Mg^{2+} , which also has a [Ne] electron configuration. On the other hand, the nonmetal

Table 12.2

The Formation of Ions by Metals and Nonmetals

Group	Ion Formation	Electron Configuration	
		Atom	Ion
1	$\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$	$[\text{Ne}]3s^1$	$[\text{Ne}]$
2	$\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$	$[\text{Ne}]3s^2$	$[\text{Ne}]$
3	$\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-$	$[\text{Ne}]3s^23p^1$	$[\text{Ne}]$
6	$\text{O} + 2\text{e}^- \rightarrow \text{O}^{2-}$	$[\text{He}]2s^22p^4 + 2\text{e}^- \rightarrow [\text{He}]2s^22p^6 = [\text{Ne}]$	
7	$\text{F} + \text{e}^- \rightarrow \text{F}^-$	$[\text{He}]2s^22p^5 + \text{e}^- \rightarrow [\text{He}]2s^22p^6 = [\text{Ne}]$	

In addition, each H shares two electrons with the oxygen atom,



which fills the H 1s orbital, giving it a $1s^2$ or [He] electron configuration. We will have much more to say about covalent bonding later.

Let's Review

Electron Configurations and Bonding

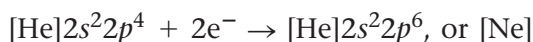
- When a *nonmetal* and a *Group 1, 2, or 3 metal* react to form a binary ionic compound, the ions form in such a way that the valence-electron configuration of the *nonmetal* is *completed* to achieve the configuration of the *next* noble gas, and the valence orbitals of the *metal* are *emptied* to achieve the configuration of the *previous* noble gas. In this way, both ions achieve noble gas electron configurations.
- When *two nonmetals* react to form a covalent bond, they share electrons in a way that completes the valence-electron configurations of both atoms. That is, both nonmetals attain noble gas electron configurations by sharing electrons.

Predicting Formulas of Ionic Compounds

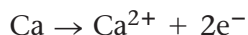
To show how to predict which ions form when a metal reacts with a nonmetal, we will consider the formation of an ionic compound from calcium and oxygen. We can predict the compound that will form by considering the valence electron configurations of the two atoms.



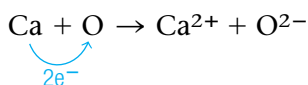
From Figure 12.4 we see that the electronegativity of oxygen (3.5) is much greater than that of calcium (1.0), giving a difference of 2.5. Because of this large difference, electrons are transferred from calcium to oxygen to form an oxygen anion and a calcium cation. How many electrons are transferred? We can base our prediction on the observation that noble gas configurations are the most stable. Note that oxygen needs two electrons to fill its valence orbitals ($2s$ and $2p$) and achieve the configuration of neon ($1s^22s^22p^6$), which is the next noble gas.



By losing two electrons, calcium can achieve the configuration of argon (the previous noble gas).



Two electrons are therefore transferred as follows:



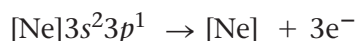
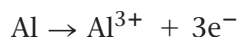
Information

Differences in electronegativities greater than 2.0 lead to ionic bonds.

To predict the formula of the ionic compound, we use the fact that chemical compounds are always electrically neutral—they have the same total quantities of positive and negative charges. In this case, we must have equal numbers of Ca^{2+} and O^{2-} ions, and the empirical formula of the compound is CaO .

Chemical compounds are always electrically neutral.

The same principles can be applied to many other cases. For example, consider the compound formed from aluminum and oxygen. Aluminum has the electron configuration $[\text{Ne}]3s^23p^1$. To achieve the neon configuration, aluminum must lose three electrons, forming the Al^{3+} ion.



MATH



$3 \times (2-)$ balances
 $2 \times (3+)$.

Therefore, the ions will be Al^{3+} and O^{2-} . Because the compound must be electrically neutral, there will be three O^{2-} ions for every two Al^{3+} ions, and the compound has the empirical formula Al_2O_3 .

Table 12.3 shows common elements that form ions with noble gas electron configurations in ionic compounds.

Table 12.3

Common Ions with Noble Gas Configurations in Ionic Compounds

Group 1	Group 2	Group 3	Group 6	Group 7	Electron Configuration
Li^{+}	Be^{2+}				$[\text{He}]$
Na^{+}	Mg^{2+}	Al^{3+}	O^{2-}	F^{-}	$[\text{Ne}]$
K^{+}	Ca^{2+}		S^{2-}	Cl^{-}	$[\text{Ar}]$
Rb^{+}	Sr^{2+}		Se^{2-}	Br^{-}	$[\text{Kr}]$
Cs^{+}	Ba^{2+}		Te^{2-}	I^{-}	$[\text{Xe}]$

Notice that our discussion in this section refers to metals in Groups 1, 2, and 3 (the representative metals). The transition metals exhibit more complicated behavior (they form a variety of ions), which we will not consider in this text.

Active Reading Question

How can we predict that solid aluminum oxide has the formula Al_2O_3 ?

B. Ionic Bonding and Structures of Ionic Compounds

When metals and nonmetals react, the resulting ionic compounds are very stable; large amounts of energy are required to “take them apart.” For example, the melting point of sodium chloride is approximately $800\text{ }^{\circ}\text{C}$. The strong bonding in these ionic compounds results from the attractions between the oppositely charged cations and anions.

Structures of Ionic Compounds We write the formula of an ionic compound such as lithium fluoride simply as LiF , but this is really the empirical, or simplest, formula. The actual solid contains huge and equal numbers of

Li^+ and F^- ions packed together in a way that maximizes the attractions of the oppositely charged ions. A representative part of the lithium fluoride structure is shown in **Figure 12.7a**. In this structure, the larger F^- ions are packed together like hard spheres, and the much smaller Li^+ ions are placed regularly among the F^- ions. The structure shown in **Figure 12.7b** represents only a tiny part of the actual structure, which continues in all three dimensions with the same pattern as that shown.

The structures of virtually all binary ionic compounds can be explained by a model that involves packing the ions as though they were hard spheres. The larger spheres (usually the anions) are packed together, and the small ions occupy the interstices (spaces or holes) among them.

To understand the packing of ions, it helps to realize that *a cation is always smaller than the parent atom, and an anion is always larger than the parent atom*. This makes sense because when a metal loses all of its valence electrons to form a cation, it gets much smaller. On the other hand, in forming an anion, a nonmetal gains enough electrons to achieve the next noble gas electron configuration and so becomes much larger. The relative sizes of the Group 1 and Group 7 atoms and their ions are shown in **Figure 12.8**.



















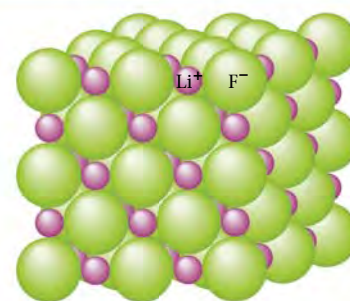
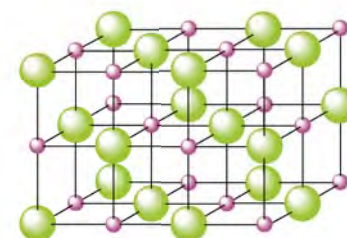
	Atom		Cation		Atom		Anion	
Li 152			Li^+ 60	F 72			F^- 136	
Na 186			Na^+ 95	Cl 99			Cl^- 181	
K 227			K^+ 133	Br 114			Br^- 195	
Rb 248			Rb^+ 148	I 133			I^- 216	
Cs 265			Cs^+ 169					

Figure 12.8

Relative sizes of some ions and their parent atoms. Note that cations are smaller and anions are larger than their parent atoms. The sizes (radii) are given in units of picometers ($1 \text{ pm} = 10^{-12} \text{ m}$).



(a)



(b)

Figure 12.7

The structure of lithium fluoride. (a) This structure represents the ions as packed spheres. (b) This structure shows the positions (centers) of the ions. The spherical ions are packed in a way that maximizes the ionic attractions.

i nformation

When spheres are packed together, they do not fill up all of the space. The spaces (holes) that are left can be occupied by smaller spheres.

Ionic Compounds Containing Polyatomic Ions

So far in this chapter we have discussed only binary ionic compounds, which contain ions derived from single atoms. However, many compounds contain polyatomic ions: charged species composed of several atoms. For example, ammonium nitrate contains the NH_4^+ and NO_3^- ions.

CRITICAL THINKING ?

Ions have different sizes than their parent atoms.

What if ions stayed the same sizes as their parent atoms? How would that change the structures of ionic compounds?



These ions with their opposite charges attract each other in the same way as do the simple ions in binary ionic compounds. However, the *individual* polyatomic ions are held together by covalent bonds, with all of the atoms behaving as a unit. For example, in the ammonium ion, NH_4^+ , there are four N—H covalent bonds. Likewise, the nitrate ion, NO_3^- , contains three covalent N—O bonds. Thus, although ammonium nitrate is an ionic compound because it contains the NH_4^+ and NO_3^- ions, it also contains covalent bonds in the individual polyatomic ions. When ammonium nitrate is dissolved in water, it behaves as a strong electrolyte like the binary ionic compounds sodium chloride and potassium bromide. As we saw in Chapter 8, this occurs because when an ionic solid dissolves, the ions are freed to move independently and can conduct an electric current.

SECTION 12.2 REVIEW QUESTIONS

- How can the periodic table be used to predict how many electrons a certain metal atom loses to form an ion in an ionic compound? Provide two examples.
- How can we predict that oxygen will form an O^{2-} ion and not an O^{3-} ion?
- How do we use electron configurations to predict formulas of ionic compounds?
- Write the electron configurations for the pairs of atoms given below. Use them to predict the formula for an ionic compound formed from these elements.
 - Mg, S
 - K, Cl
 - Cs, F
 - Ba, Br
- Why are cations smaller than their parent atoms? Why are anions larger?
- How are polyatomic anions different from simple anions?
- Choose the elements from Groups 1, 2, 6 and 7 that form ions having the same electron configuration as argon. Write the symbols, with charges for these ions, and their common electron configuration.



Lewis Structures

Objectives

- To learn to write Lewis structures
- To learn to write Lewis structures for molecules with multiple bonds

A. Writing Lewis Structures

Bonding involves only the valence electrons of atoms. Valence electrons are transferred when a metal and a nonmetal react to form an ionic compound. Valence electrons are shared between nonmetals in covalent bonds.

The **Lewis structure** is a representation of a molecule that shows how the valence electrons are arranged among the atoms in the molecule. These representations are named after G. N. Lewis, who conceived the idea while lecturing to a class of general chemistry students in 1902. The rules for writing Lewis structures are based on observations of many molecules from which chemists have learned that the *most important requirement for the formation of a stable compound is that the atoms achieve noble gas electron configurations.*

We have already seen this rule operate in the reaction of metals and nonmetals to form binary ionic compounds. An example is the formation of KBr, where the K^+ ion has the [Ar] electron configuration and the Br^- ion has the [Kr] electron configuration.

In writing Lewis structures, we include only the valence electrons.

Using dots to represent valence electrons, we write the Lewis structure for KBr as follows:



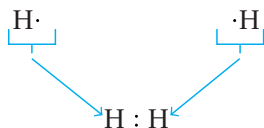
No dots are shown on the K^+ ion because it has lost its only valence electron (the 4s electron). The Br^- ion is shown with eight electrons because it has a filled valence shell.

Active Reading Question

Why do we show only valence electrons when drawing Lewis structures?

Next, we will consider Lewis structures for molecules with covalent bonds involving nonmetals in the first and second periods. The principle of achieving a noble gas electron configuration applies to these elements as follows:

⇒ Hydrogen forms stable molecules where it shares two electrons. That is, it follows a **duet rule**. For example, when two hydrogen atoms, each with one electron, combine to form the H_2 molecule, we have



Key Terms

- Lewis structure
- Duet rule
- Octet rule
- Bonding pair
- Lone pairs (unshared pairs)
- Single bond
- Double bond
- Triple bond
- Resonance

Lewis structure

A representation of a molecule or polyatomic ion showing how valence electrons are arranged among the atoms in the molecule or ion

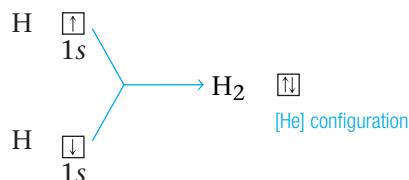
Information

Remember that the electrons in the highest principal energy level of an atom are called the valence electrons.

Information

Lewis structures show only valence electrons.

By sharing electrons, each hydrogen in H_2 has, in effect, two electrons; that is, each hydrogen has a filled valence shell.



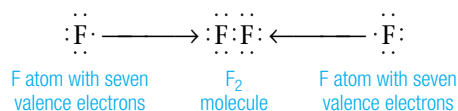
⇒ Helium does not form bonds because its valence orbital is already filled; it is a noble gas. Helium has the electron configuration $1s^2$ and can be represented by the following Lewis structure:



Information

Carbon, nitrogen, oxygen, and fluorine almost always obey the octet rule in stable molecules.

⇒ The second-row nonmetals carbon through fluorine form stable molecules when they are surrounded by enough electrons to fill the valence orbitals—that is, the one $2s$ and the three $2p$ orbitals. Eight electrons are required to fill these orbitals, so these elements typically obey the **octet rule**; they are surrounded by eight electrons. An example is the F_2 molecule, which has the following Lewis structure:



Note that each fluorine atom in F_2 is, in effect, surrounded by eight valence electrons, two of which are shared with the other atom. This is a **bonding pair** of electrons, as we discussed earlier. Each fluorine atom also has three pairs of electrons that are not involved in bonding. These are called **lone pairs** or **unshared pairs**.

⇒ Neon does not form bonds because it already has an octet of valence electrons (it is a noble gas). The Lewis structure is



Note that only the valence electrons ($2s^2 2p^6$) of the neon atom are represented by the Lewis structure. The $1s^2$ electrons are core electrons and are not shown.

Bonding pair

A pair of electrons that are shared between two atoms forming a covalent or polar-covalent bond

Lone pair (unshared pair)

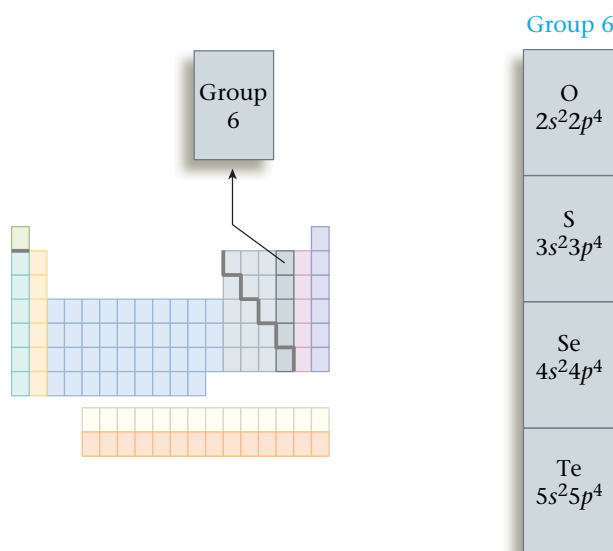
Electron pairs in a Lewis structure that are not involved in bonding



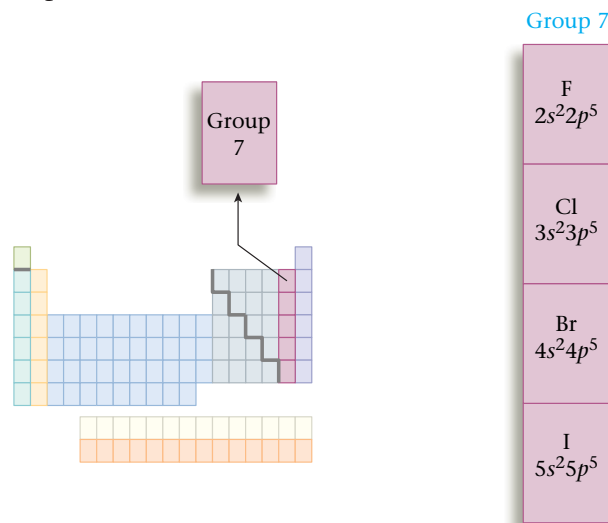
G. N. Lewis in his lab

Next we want to develop some general procedures for writing Lewis structures for molecules. Remember that Lewis structures involve only the valence electrons on atoms, so before we proceed, we will review the relationship of an element's position on the periodic table to the number of valence electrons it has.

Recall that the group number gives the total number of valence electrons. For example, all Group 6 elements have six valence electrons (valence configuration ns^2np^4).



Similarly, all Group 7 elements have seven valence electrons (valence configuration ns^2np^5).



Let's Review

Writing the Lewis structure for a molecule

- We must include all the valence electrons from all atoms. The total number of electrons available is the sum of all the valence electrons from all the atoms in the molecule.
- Atoms that are bonded to each other share one or more pairs of electrons.
- The electrons are arranged so that each atom is surrounded by enough electrons to fill the valence orbitals of that atom. This means two electrons for hydrogen and eight electrons for second-row nonmetals.

The best way to make sure we arrive at the correct Lewis structure for a molecule is to use a systematic approach. We will use the approach summarized by the following rules.

Steps for Writing Lewis Structures

- Step 1** Obtain the sum of the valence electrons from all of the atoms. Do not worry about keeping track of which electrons come from which atoms. It is the *total* number of valence electrons that is important.
- Step 2** Use one pair of electrons to form a bond between each pair of bound atoms. For convenience, a line (instead of a pair of dots) is often used to indicate each pair of bonding electrons.
- Step 3** Arrange the remaining electrons to satisfy the duet rule for hydrogen and the octet rule for each second-row element.

EXAMPLE 12.2

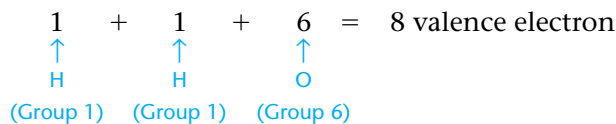
Writing Lewis Structures: Simple Molecules

Write the Lewis structure of the water molecule.

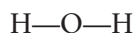
Solution

We will follow the steps listed above.

Step 1 Find the sum of the *valence* electrons for H_2O :

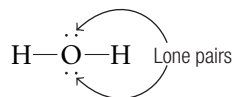


Step 2 Using a pair of electrons per bond, we draw in the two O—H bonds, using a line to indicate each pair of bonding electrons.

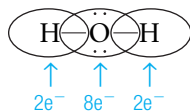


Note that H—O—H represents H : O : H.

Step 3 We arrange the remaining electrons around the atoms to achieve a noble gas electron configuration for each atom. Four electrons have been used in forming the two bonds, so four electrons ($8 - 4$) remain to be distributed. Each hydrogen is satisfied with two electrons (duet rule), but oxygen needs eight electrons to have a noble gas electron configuration. So the remaining four electrons are added to oxygen as two lone pairs. Dots are used to represent the lone pairs.



This is the correct Lewis structure for the water molecule. Each hydrogen shares two electrons, and the oxygen has four electrons and shares four to give a total of eight.



information

The number of valence electrons in an atom is the same as the group number on the periodic table for representative elements.

information

A line is used to represent a shared pair of electrons (bonding electrons) and dots are used to represent unshared pairs.

Resonance

A condition occurring when more than one valid Lewis structure can be written for a particular molecule

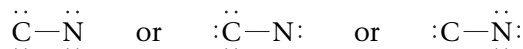
This brings us to a new term, **resonance**. A molecule shows resonance when *more than one Lewis structure can be drawn for the molecule*. In such a case, we call the various Lewis structures *resonance structures*.

Of the three resonance structures for CO_2 shown, the one in the center with two double bonds most closely fits our experimental information about the CO_2 molecule. In this text we will not be concerned about how to choose which resonance structure for a molecule gives the “best” description of that molecule’s properties.

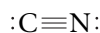
Lewis Structures for Polyatomic Ions Next, let’s consider the Lewis structure of the CN^- (cyanide) ion. Summing the valence electrons, we have



Note that the negative charge means an extra electron must be added. After first drawing a single bond (C—N), we distribute the remaining electrons to achieve a noble gas configuration for each atom. Eight electrons remain to be distributed. We can try various possibilities, such as



These structures are incorrect. To show why none is a valid Lewis structure, count the electrons around the C and N atoms. In the left structure, neither atom satisfies the octet rule. In the center structure, C has eight electrons but N has only four. In the right structure, the opposite is true. Remember that both atoms must simultaneously satisfy the octet rule. Therefore, the correct arrangement is



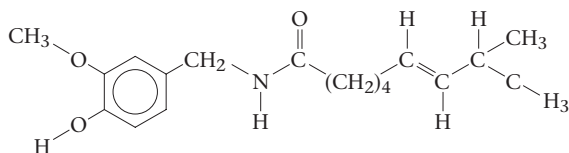
Information



represents



CELEBRITY CHEMICAL Capsaicin



Can you imagine why someone might add pepper to things like paint and fiber-optic cables? It’s certainly not to flavor them so we can eat them! In fact, it’s just the opposite. The idea is to use pepper as an environmentally safe animal deterrent.

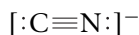
The compound that gives pepper its characteristic “hot” taste is called capsaicin. Burlington Biomedical and Scientific Corporation in Farmingdale, New York, has found a way to manufacture a compound closely related to capsaicin that tastes painfully spicy and intensely bitter. One application for this compound is to add it to paint used for boats to prevent barnacles



The hot taste of peppers is from capsaicin.

from sticking to the surface. Other applications include adding capsaicin to fiber-optic cables to prevent rodents from gnawing on them and using it to coat sutures employed in animal surgery to prevent pets from disturbing a healing wound.

(Satisfy yourself that both carbon and nitrogen have eight electrons.) In this case, we have a triple bond between C and N, in which three electron pairs are shared. Because this is an anion, we indicate the charge outside of square brackets around the Lewis structure.



Active Reading Question

When drawing a Lewis structure for a negatively charged ion, how do you determine the number of valence electrons that are present?

In summary, sometimes we need double or triple bonds to satisfy the octet rule. Writing Lewis structures is a trial-and-error process. Start with single bonds between the bonded atoms, and add multiple bonds as needed.

We will write the Lewis structure for NO_2^- in Example 12.3 to make sure the procedures for writing Lewis structures are clear.

EXAMPLE 12.3

Writing Lewis Structures: Resonance Structures

Write the Lewis structure for the NO_2^- anion.

Solution

Step 1 Find the sum of the valence electrons for NO_2^- .

$$\text{Valence electrons: } \underset{\text{O}}{6} + \underset{\text{N}}{5} + \underset{\text{O}}{6} + \underset{\substack{-1 \\ \text{charge}}}{1} = 18 \text{ electrons}$$

Step 2 Put in single bonds.



Step 3 Satisfy the octet rule. In placing the electrons, we find that there are two Lewis structures that satisfy the octet rule:



Verify that each atom in these structures is surrounded by an octet of electrons. Try some other arrangements to see whether other structures exist in which the eighteen electrons can be used to satisfy the octet rule.

It turns out that these are the only two that work. Note that this is another case in which resonance occurs; there are two valid Lewis structures.

Practice Problem • Exercise 12.3

Ozone is a very important constituent of the atmosphere. At upper levels, it protects us by absorbing high-energy radiation from the sun. Near the earth's surface, it produces harmful air pollution. Write the Lewis structure for ozone, O_3 .

Now let's consider a few more cases in Example 12.4.

EXAMPLE 12.4

Writing Lewis Structures: Summary

Give the Lewis structure for each of the following:

- a. HF e. CF₄
 b. N₂ f. NO⁺
 c. NH₃ g. NO₃⁻
 d. CH₄

Solution

In each case, we apply the three steps for writing Lewis structures. Recall that lines are used to indicate shared electron pairs and that dots are used to indicate nonbonding pairs (lone pairs). The table below summarizes our results.

Molecule or Ion	Total Valence Electrons	Draw Single Bonds	Calculate Number of Electrons Remaining	Use Remaining Electrons to Achieve Noble Gas Configurations	Check	
					Atom	Electrons
HF	1 + 7 = 8	H—F	8 - 2 = 6	H— $\ddot{\text{F}}$:	H F	2 8
N ₂	5 + 5 = 10	N—N	10 - 2 = 8	:N≡N:	N	8
NH ₃	5 + 3(1) = 8	$\begin{array}{c} \text{H}-\text{N}-\text{H} \\ \\ \text{H} \end{array}$	8 - 6 = 2	$\begin{array}{c} \text{H}-\ddot{\text{N}}-\text{H} \\ \\ \text{H} \end{array}$	H N	2 8
CH ₄	4 + 4(1) = 8	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$	8 - 8 = 0	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$	H C	2 8
CF ₄	4 + 4(7) = 32	$\begin{array}{c} \text{F} \\ \\ \text{F}-\text{C}-\text{F} \\ \\ \text{F} \end{array}$	32 - 8 = 24	$\begin{array}{c} :\ddot{\text{F}}: \\ \\ :\ddot{\text{F}}-\text{C}-\ddot{\text{F}}: \\ \\ :\ddot{\text{F}}: \end{array}$	F C	8 8
NO ⁺	5 + 6 - 1 = 10	N—O	10 - 2 = 8	[:N≡O:] ⁺	N O	8 8
NO ₃ ⁻	5 + 3(6) + 1 = 24	$\left[\begin{array}{c} \text{O} \\ \\ \text{N} \\ \\ \text{O} \quad \text{O} \end{array} \right]$	24 - 6 = 18	$\left\{ \begin{array}{l} \left[\begin{array}{c} :\ddot{\text{O}}: \\ \\ \text{O}=\text{N}-\ddot{\text{O}}: \\ \\ \text{O} \end{array} \right]^- \\ \left[\begin{array}{c} :\ddot{\text{O}}: \\ \\ \ddot{\text{O}}-\text{N}=\text{O}: \\ \\ \text{O} \end{array} \right]^- \\ \left[\begin{array}{c} :\ddot{\text{O}}: \\ \\ \ddot{\text{O}}-\text{N}=\text{O}: \\ \\ \text{O} \end{array} \right]^- \end{array} \right.$ NO ₃ ⁻ shows resonance	N O N O N O	8 8 8 8 8 8

Practice Problem • Exercise 12.4

Write the Lewis structures for the following molecules:

- a. NF_3
- b. O_2
- c. CO
- d. PH_3
- e. H_2S
- f. SO_4^{2-}
- g. NH_4^+
- h. ClO_3^-
- i. SO_2

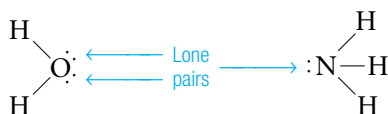
Information

In a molecule of a binary compounds where there is one atom of a given element and several atoms of a second element, the single atom is almost always the central atom.

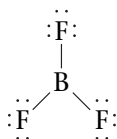
Remember that when writing Lewis structures, you don't have to worry about which electrons come from which atoms in a molecule. It is best to think of a molecule as a new entity that uses all of the available valence electrons from the various atoms to achieve the strongest possible bonds. Think of the valence electrons as belonging to the molecule, rather than to the individual atoms. Simply distribute the valence electrons so that noble gas electron configurations are obtained for each atom, without regard to the origin of each particular electron.

Some Exceptions to the Octet Rule

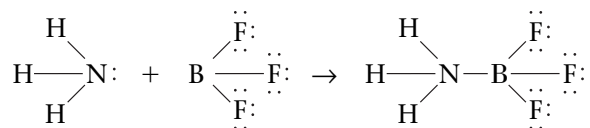
The idea that covalent bonding can be predicted by achieving noble gas electron configurations for all atoms is a simple and very successful idea. The rules we have used for Lewis structures describe correctly the bonding in most molecules. However, with such a simple model, we should expect some exceptions. Boron, for example, tends to form compounds in which the boron atom has fewer than eight electrons around it—that is, it does not have a complete octet. Boron trifluoride, BF_3 , a gas at normal temperatures and pressures, reacts very energetically with molecules such as water and ammonia that have unshared electron pairs (lone pairs).



The violent reactivity of BF_3 with electron-rich molecules arises because the boron atom is electron-deficient. The Lewis structure that seems most consistent with the properties of BF_3 (twenty-four valence electrons) is



Note that in this structure the boron atom has only six electrons around it. The octet rule for boron could be satisfied by drawing a structure with a double bond between the boron and one of the fluorines. However, experiments indicate that each B—F bond is a single bond in accordance with the above Lewis structure. This structure is also consistent with the reactivity of BF_3 with electron-rich molecules. For example, BF_3 reacts vigorously with NH_3 to form H_3NBF_3 .



Note that in the product H_3NBF_3 , which is very stable, boron has an octet of electrons. It is also characteristic of beryllium to form molecules where the beryllium atom is electron-deficient.

Any molecule that contains an odd number of electrons does not conform to our rules for Lewis structures. For example, NO and NO_2 have eleven and seventeen valence electrons, respectively, and conventional Lewis structures cannot be drawn for these cases.

Even though there are exceptions, most molecules can be described by Lewis structures in which all the atoms have noble gas electron configurations, and this is a very useful model for chemists.

Active Reading Question

Why is there no conventional Lewis structure for the molecule NO_2 ?

SECTION 12.3 REVIEW QUESTIONS

- 1 Why is bonding based primarily on an octet rule? Why not a sextet rule?
- 2 What is the difference between a bonding pair of electrons and a lone pair of electrons?
- 3 Why do pairs of atoms share pairs (or multiples of pairs) of electrons? Why not share odd numbers?
- 4 For each molecule below, do the following:
 - i. Give the sum of the valence electrons for all atoms.
 - ii. Draw the Lewis structure.
 - iii. Circle the octet (or duet) for each atom in the structure.
 - iv. Include any resonance structures.
 - a. ClF
 - b. Br_2
 - c. H_2S
 - d. O_3
 - e. CH_4
 - f. C_2H_6
 - g. HCN
 - h. CO



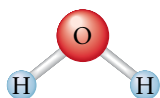
Structures of Molecules

Objectives

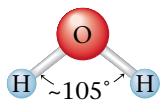
- To understand molecular structure and bond angles
- To learn to predict molecular geometry from the number of electron pairs
- To learn to apply the VSEPR model to molecules with double bonds

A. Molecular Structure

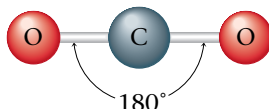
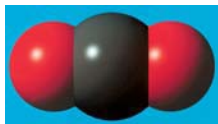
So far in this chapter we have considered the Lewis structures of molecules. These structures represent the arrangement of the *valence electrons* in a molecule. We use the word *structure* in another way when we talk about the **molecular structure** or **geometric structure** of a molecule. These terms refer to the three-dimensional arrangement of the *atoms* in a molecule. For example, the water molecule is known to have the molecular structure



which is often called “bent” or “V-shaped.” To describe the structure more precisely, we often specify the *bond angle*. For the H_2O molecule the bond angle is about 105° .

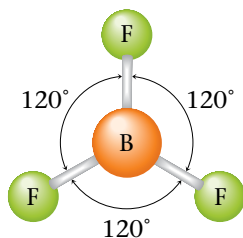


On the other hand, some molecules exhibit a **linear structure** (all atoms in a line). An example is the CO_2 molecule.



Note that a linear molecule has a 180° bond angle.

A third type of molecular structure is illustrated by BF_3 , which is planar or flat (all four atoms in the same plane) with 120° bond angles.



The name usually given to this structure is **trigonal planar**, although triangular might seem to make more sense.

Another type of molecular structure is illustrated by methane, CH_4 . This molecule has the molecular structure called a **tetrahedral structure** or a *tetrahedron*.

Key Terms

- Molecular structure (geometric structure)
- Linear structure
- Trigonal planar structure
- Tetrahedral structure
- Valence shell electron pair repulsion (VSEPR) model
- Trigonal pyramid

Molecular structure (geometric structure)

The three-dimensional arrangement of atoms in a molecule

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

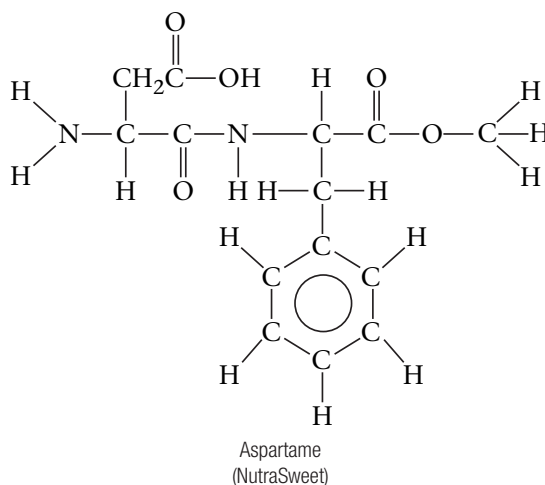
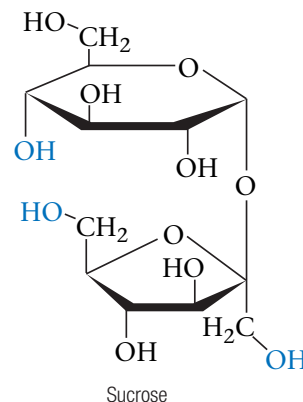
Taste—It's the Structure That Counts

Why do certain substances taste sweet, sour, bitter, or salty? Of course, it has to do with the taste buds on our tongues. But how do these taste buds work? For example, why does sugar taste sweet to us? The answer to this question remains elusive, but it does seem clear that sweet taste depends on how certain molecules fit the "sweet receptors" in our taste buds.

One of the mysteries of the sweet taste sensation is the wide variety of molecules that tastes sweet. The chemical structure of common table sugar (called sucrose) is shown on the right.

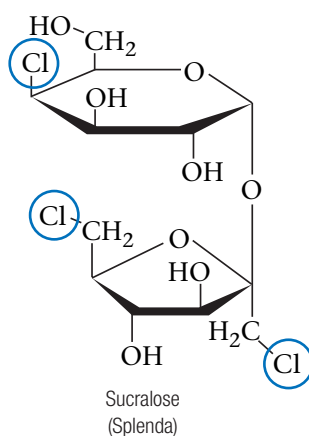


Sugar cubes are pure sucrose.



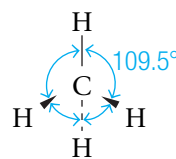
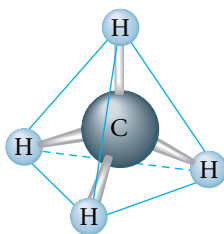
One artificial sweetener, aspartame, is used in many products including candy and diet soda. Some people are allergic to aspartame.

Another widely used artificial sweetener is sucralose which is made by modifying sucrose (compare the structure of sucralose to sucrose shown above). Look at the groups with blue circles in each structure. Sucralose is 600 times as sweet as sucrose and is used to sweeten beverages, baked goods, yogurt, and desserts.



Some brands of yogurt are sweetened with sucralose instead of sugar.





The blue lines shown connecting the H atoms in the center structure above define the four identical triangular faces of the tetrahedron.

In the next section we will discuss these various molecular structures in more detail and learn how to predict the molecular structure of a molecule by looking at the molecule's Lewis structure.

B. The VSEPR Model

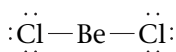
The structures of molecules play a very important role in determining their properties. For example, as we can see in the “Chemistry In Your World” feature, taste is directly related to molecular structure. Structure is particularly important for biological molecules; a slight change in the structure of a large biomolecule can completely destroy its usefulness to a cell and may even change the cell from a normal one to a cancerous one.

Many experimental methods now exist for determining the molecular structure of a molecule—that is, the three-dimensional arrangement of the atoms. These methods must be used when accurate information about the structure is required. However, it is often useful to be able to predict the *approximate* molecular structure of a molecule. Now we will consider a simple model that allows us to do this. The **valence shell electron pair repulsion (VSEPR) model** is useful for predicting the molecular structures of molecules formed from nonmetals. The main idea of this model is that

The structure around a given atom is determined by minimizing repulsions between electron pairs.

This means that the bonding and nonbonding electron pairs (lone pairs) around a given atom are positioned *as far apart as possible*.

Two Pairs of Electrons To see how this model works, we will first consider the molecule BeCl_2 , which has the following Lewis structure (it is an exception to the octet rule):



Note that there are two pairs of electrons around the beryllium atom. What arrangement of these electron pairs allows them to be as far apart as possible to minimize the repulsions? The best arrangement places the pairs on opposite sides of the beryllium atom at 180° from each other.

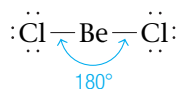


This is the maximum possible separation for two electron pairs. Now that we have determined the optimal arrangement of the electron pairs around the central atom, we can specify the molecular structure of BeCl_2 , that is, the

Valence shell electron pair repulsion (VSEPR) model

A model used to predict molecular geometry. Based on the idea that pairs of electrons surrounding an atom repel each other and that the atoms in a molecule are positioned to minimize this repulsion

positions of the atoms. Because each electron pair on beryllium is shared with a chlorine atom, the molecule has a *linear structure* with a 180° bond angle.

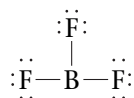


Whenever two pairs of electrons are present around an atom, they should always be placed at an angle of 180° to each other to give a linear arrangement.

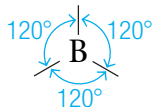
Active Reading Question

What is the main idea of the valence shell electron pair repulsion (VSEPR) model? How does it help us determine the shape of a molecule?

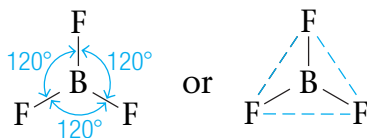
Three Pairs of Electrons Next let's consider BF_3 , which has the following Lewis structure (it is another exception to the octet rule):



Here the boron atom is surrounded by three pairs of electrons. What arrangement minimizes the repulsions among the three pairs of electrons? Here the greatest distance between the electron pairs is achieved by angles of 120° .



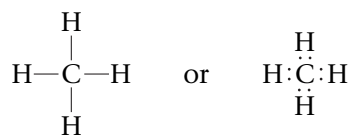
Because each of the electron pairs is shared with a fluorine atom, the molecular structure is



This is a planar (flat) molecule with a triangular arrangement of F atoms, commonly described as a trigonal planar structure.

Whenever three pairs of electrons are present around an atom, they should always be placed at the corners of a triangle (in a plane at an angle of 120° to each other).

Four Pairs of Electrons Next let's consider the methane molecule, which has the Lewis structure



Geometric Balloons**Materials**

- Nine round balloons

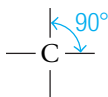
Procedure

1. Obtain nine round balloons.
2. Blow up all of the balloons to approximately the same size and tie each one.
3. Tie two of the balloons together. Then tie three of the remaining balloons together. Tie the final four balloons together.
4. Observe the geometry of each of the balloon figures. What are the angles between each balloon for each balloon figure?

Results/Analysis

1. Relate your findings to VSEPR theory. What does the knot in the center of each cluster represent? What does each balloon represent? Why do these geometries occur naturally?
2. Consider the cluster with three balloons. What is the arrangement of atoms (shape) if one balloon represents a lone pair of electrons? What if there were two lone pairs of electrons?
3. Consider the cluster with four balloons. What is the arrangement of atoms (shape) if one balloon represents a lone pair of electrons? What if there were two lone pairs of electrons? What if there were three lone pairs of electrons?

There are four pairs of electrons around the central carbon atom. What arrangement of these electron pairs best minimizes the repulsions? First we try a square planar arrangement:

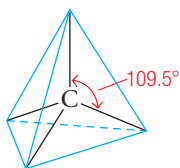


The carbon atom and the electron pairs are all in a plane represented by the surface of the paper, and the angles between the pairs are all 90° .

Is there another arrangement with angles greater than 90° that would put the electron pairs even farther away from each other? The answer is yes. We can get larger angles than 90° by using the following three-dimensional structure, which has angles of approximately 109.5° .



In this drawing the wedge indicates a position above the surface of the paper and the dashed lines indicate positions behind that surface. The solid line indicates a position on the surface of the page. The figure formed by connecting the lines is called a tetrahedron, so we call this arrangement of electron pairs the *tetrahedral arrangement*.



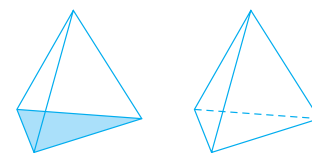
This is the maximum possible separation of four pairs around a given atom.

Whenever four pairs of electrons are present around an atom, they should always be placed at the corners of a tetrahedron (the tetrahedral arrangement).

CRITICAL THINKING

You have seen that molecules with four electron pairs, such as CH_4 , are tetrahedral.

What if a molecule had six electron pairs like SF_6 ? Predict the geometry and bond angles of SF_6 .

information

A tetrahedron has four equal triangular faces.

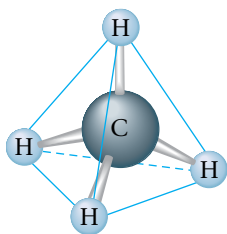


Figure 12.9

The molecular structure of methane. The tetrahedral arrangement of electron pairs produces a tetrahedral arrangement of hydrogen atoms.

Now that we have the arrangement of electron pairs that gives the least repulsion, we can determine the positions of the atoms and thus the molecular structure of CH_4 . In methane each of the four electron pairs is shared between the carbon atom and a hydrogen atom. Thus the hydrogen atoms are placed as shown in **Figure 12.9**, and the molecule has a tetrahedral structure with the carbon atom at the center.

Recall that the main idea of the VSEPR model is to find the arrangement of electron pairs around the central atom that minimizes the repulsions. Then we can determine the *molecular structure* by knowing how the electron pairs are shared with the peripheral atoms. A systematic procedure for using the VSEPR model to predict the structure of a molecule is outlined here.

Steps for Predicting Molecular Structure Using the VSEPR Model

- Step 1** Draw the Lewis structure for the molecule.
- Step 2** Count the electron pairs and arrange them in the way that minimizes repulsions (that is, put the pairs as far apart as possible).
- Step 3** Determine the positions of the atoms from the way the electron pairs are shared.
- Step 4** Determine the name of the molecular structure from the positions of the *atoms*.

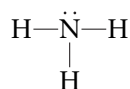
EXAMPLE 12.5

Predicting Molecular Structure Using the VSEPR Model, I

Ammonia, NH_3 , is used as a fertilizer (injected into the soil) and as a household cleaner (in aqueous solution). Predict the structure of ammonia using the VSEPR model.

Solution

Step 1 Draw the Lewis structure.



Step 2 Count the pairs of electrons and arrange them to minimize repulsions. The NH_3 molecule has four pairs of electrons around the N atom: three bonding pairs and one nonbonding pair. From the discussion of the methane molecule, we know that the best arrangement of four electron pairs is the tetrahedral structure shown in Figure 12.10a.

Step 3 Determine the positions of the atoms. The three H atoms share electron pairs as shown in Figure 12.10c.

Step 4 Name the molecular structure. It is very important to recognize that the name of the molecular structure is always based on the *positions of the atoms*. *The placement of the electron pairs determines the structure, but the name is based on the positions of the atoms*. Thus it is incorrect to say that the NH_3 molecule is tetrahedral. It has a tetrahedral arrangement of electron pairs but *not* a tetrahedral arrangement of atoms. The molecular structure of ammonia is a **trigonal pyramid** (one side is different from the other three) rather than a tetrahedron.

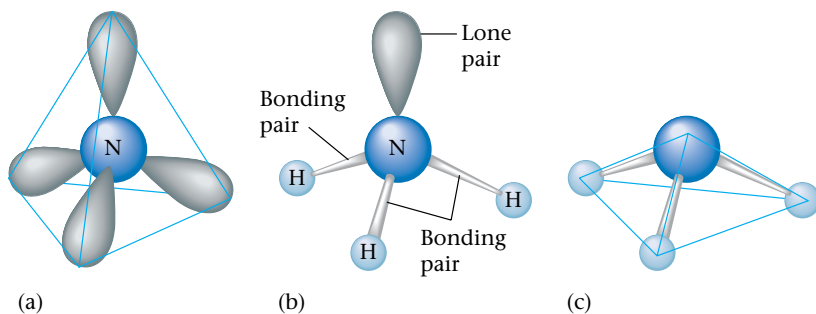


Figure 12.10

(a) The tetrahedral arrangement of electron pairs around the nitrogen atom in the ammonia molecule. (b) Three of the electron pairs around nitrogen are shared with hydrogen atoms as shown, and one is a lone pair. Although the arrangement of *electron pairs* is tetrahedral, as in the methane molecule, the hydrogen atoms in the ammonia molecule occupy only three corners of the tetrahedron. A lone pair occupies the fourth corner (c) The NH_3 molecule has the trigonal pyramid structure (a pyramid with a triangle as a base).

CRITICAL THINKING ?

What if NH_3 did not have a lone pair? How would that change its molecular structure?

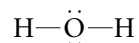
EXAMPLE 12.6

Predicting Molecular Structure Using the VSEPR Model, II

Describe the molecular structure of the water molecule.

Solution

Step 1 The Lewis structure for water is:



Step 2 There are four pairs of electrons: two bonding pairs and two nonbonding pairs. To minimize repulsions, these are best arranged in a tetrahedral structure as shown in **Figure 12.11a**.

Step 3 Although H_2O has a tetrahedral arrangement of *electron pairs*, it is *not a tetrahedral molecule*. The *atoms* in the H_2O molecule form a V-shape, as shown in **Figure 12.11b** and **c**.

Step 4 The molecular structure is called V-shaped or bent.

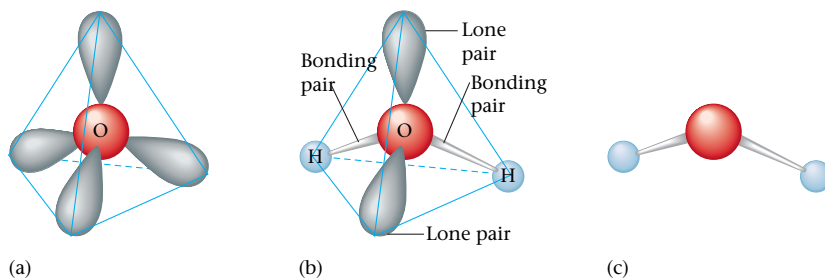


Figure 12.11

(a) The tetrahedral arrangement of the four electron pairs around oxygen in the water molecule. (b) Two of the electron pairs are shared between oxygen and the hydrogen atoms, and two are lone pairs. (c) The V-shaped molecular structure of the water molecule.

Practice Problem • Exercise 12.6

Predict the arrangement of electron pairs around the central atom. Then sketch and name the molecular structure for each of the following molecules or ions.

- | | | |
|-----------------------|-------------------------|---------------------|
| a. NH_4^+ | c. NF_3 | e. ClO_3^- |
| b. SO_4^{2-} | d. H_2S | f. BeF_2 |

Tools for Predicting Molecular Structure Using the VSEPR Model

- Two pairs of electrons on a central atom in a molecule are always placed 180° apart. This is a linear arrangement of pairs.
- Three pairs of electrons on a central atom in a molecule are always placed 120° apart in the same plane as the central atom. This is a trigonal planar (triangular) arrangement of pairs.
- Four pairs of electrons on a central atom in a molecule are always placed 109.5° apart. This is a tetrahedral arrangement of electron pairs.
- When *every pair* of electrons on the central atom is *shared* with another atom, the molecular structure has the same name as the arrangement of electron pairs.

Number of Pairs	Name of Arrangement
------------------------	----------------------------

2	linear
3	trigonal planar
4	tetrahedral

- When one or more of the electron pairs around a central atom are unshared (lone pairs), the name for the molecular structure is *different* from that for the arrangement of electron pairs (see cases 4 and 5 in Table 12.4).

Active Reading Question

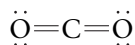
Why must we draw the Lewis structure for a molecule before trying to determine its shape?

The various cases we have considered are summarized in **Table 12.4** on the following page.

C. Molecules with Double Bonds

Up to this point we have applied the VSEPR model only to molecules (and ions) that contain single bonds. Now we will show that this model applies equally well to species with one or more double bonds. We will develop the procedures for dealing with molecules with double bonds by considering examples whose structures are known.

First we will examine the structure of carbon dioxide, a substance that may be contributing to the warming of Earth. The carbon dioxide molecule has the Lewis structure:



Carbon dioxide is known by experiment to be a linear molecule. That is, it has a 180° bond angle.

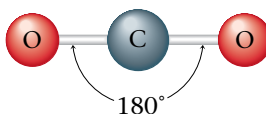
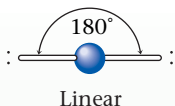
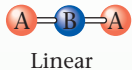
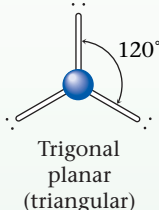
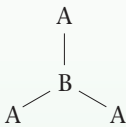
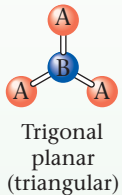
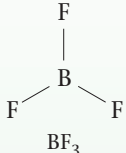
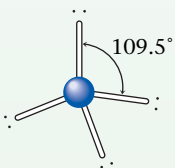
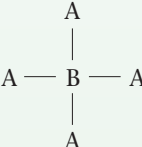
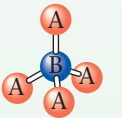
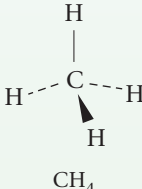
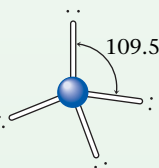
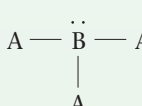
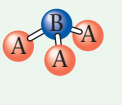
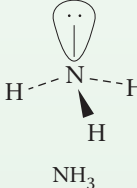
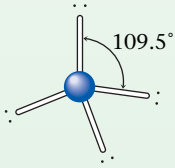
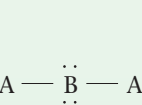
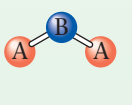
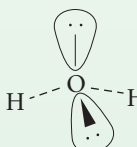
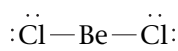


Table 12.4

Arrangements of Electron Pairs and the Resulting Molecular Structures for Two, Three, and Four Electron Pairs

Case	Number of Electron Pairs	Bonds	Electron Pair Arrangement	Partial Lewis Structure	Molecular Structure	Example
1	2	2	 Linear	A — B — A	 Linear	F — Be — F BeF ₂
2	3	3	 Trigonal planar (triangular)		 Trigonal planar (triangular)	 BF ₃
3	4	4	 Tetrahedral		 Tetrahedral	 CH ₄
4	4	3	 Trigonal pyramid		 Trigonal pyramid	 NH ₃
5	4	2	 Tetrahedral		 Bent or V-shaped	 H ₂ O

Recall that two electron pairs around a central atom can minimize their mutual repulsions by taking positions on opposite sides of the atom (at 180° from each other). This causes a molecule like BeCl₂, which has the Lewis structure



to have a linear structure. Now recall that CO₂ has two double bonds and is known to be linear, so the double bonds must be at 180° from each other. Therefore, we conclude that each double bond in this molecule acts *effectively* as one repulsive unit. This conclusion makes sense if we think of a bond in terms of an electron density “cloud” between two atoms. For example, we can picture the single bonds in BeCl₂ as follows:



The minimum repulsion between these two electron density clouds occurs when they are on opposite sides of the Be atom (180° angle between them).

Each double bond in CO_2 involves the sharing of four electrons between the carbon atom and an oxygen atom. Thus we might expect the bonding cloud to be “fatter” than for a single bond:

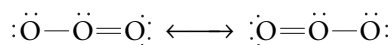


However, the repulsive effects of these two clouds produce the same result as for single bonds; the bonding clouds have minimum repulsions when they are positioned on opposite sides of the carbon. The bond angle is 180° , and so the molecule is linear:

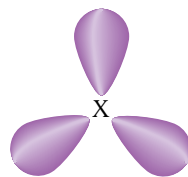


In summary, examination of CO_2 leads us to the conclusion that in using the VSEPR model for molecules with double bonds, each double bond should be treated the same as a single bond. In other words, although a double bond involves four electrons, these electrons are restricted to the space between a given pair of atoms. Therefore, these four electrons do not function as two independent pairs but are “tied together” to form one effective repulsive unit.

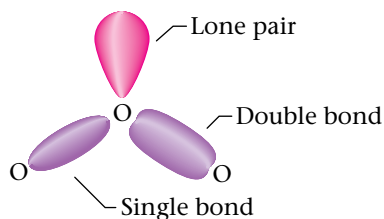
We reach this same conclusion by considering the known structures of other molecules that contain double bonds. For example, consider the ozone molecule, which has eighteen valence electrons and exhibits two resonance structures:



The ozone molecule is known to have a bond angle close to 120° . Recall that 120° angles represent the minimum repulsion for three pairs of electrons.



This indicates that the double bond in the ozone molecule is behaving as one effective repulsive unit:



When using the VSEPR model to predict the molecular geometry of a molecule, a double bond is counted the same as a single electron pair.

Thus CO_2 has two “effective pairs” that lead to its linear structure, whereas O_3 has three “effective pairs” that lead to its bent structure with a 120° bond angle. Therefore, to use the VSEPR model for molecules (or ions) that have double bonds, we use the same tools as those given on p. 430, but we count any double bond the same as a single electron pair. Although we have not shown it here, triple bonds also count as one repulsive unit in applying the VSEPR model.

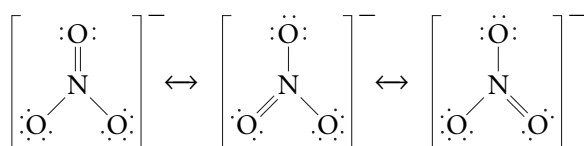
EXAMPLE 12.7

Predicting Molecular Structure Using the VSEPR Model, III

Predict the structure of the nitrate ion.

Solution

Step 1 The Lewis structures for NO_3^- are



Step 2 In each resonance structure there are effectively three pairs of electrons: the two single bonds and the double bond (which counts as one pair). These three “effective pairs” will require a trigonal planar arrangement (120° angles).

Step 3 The atoms are all in a plane, with the nitrogen at the center and the three oxygens at the corners of a triangle (trigonal planar arrangement).

Step 4 The NO_3^- ion has a trigonal planar structure.



SECTION 12.3

REVIEW QUESTIONS

- How does drawing a Lewis structure for a molecule help in determining its molecular shape?
- When is the molecular structure for a molecule the same as the arrangement of electron pairs?
- What causes the name for the molecular structure to be different from the name for the arrangement of electron pairs?
- If double bonds contain four electrons instead of two, why should they be treated the same as single bonds when determining molecular structure?
- Predict the molecular structure for each of the following molecules or ions.
 - SO_3^{2-}
 - HCN
 - Cl_2O
 - PCl_3



Key Terms

12.1 Bond
Bond energy
Ionic bonding
Ionic compound
Covalent bonding
Polar covalent bond
Electronegativity
Dipole moment

12.3 Lewis structure
Duet rule
Octet rule
Bonding pair
Lone pairs (unshared pairs)
Single bond
Double bond
Triple bond
Resonance

12.4 Molecular structure
(geometric structure)
Linear structure
Trigonal planar structure
Tetrahedral structure
Valence shell electron pair
repulsion (VSEPR) model
Trigonal pyramid

Key Ideas

12.1 Characteristics of Chemical Bonds

- Chemical bonds hold groups of atoms together to form molecules and ionic solids.
- Bonds are classified as:
 - Ionic: Formed when one or more electrons are transferred to form positive and negative ions
 - Covalent: Electrons are shared equally between identical atoms
 - ◆ Polar covalent bond: unequal electron sharing between different atoms
- Electronegativity is the relative ability of an atom to attract the electrons shared with another atom in a bond.
- The difference in electronegativity of the atoms forming a bond determines the polarity of that bond.

12.2 Characteristics of Ions and Ionic Compounds

- In stable compounds, atoms tend to achieve the electron configuration of the nearest noble gas atom.
- In ionic compounds:
 - Nonmetals tend to gain electrons to reach the electron configuration of the next noble gas atom.
 - Metals tend to lose electrons to reach the electron configuration of the previous noble gas atom.
- Ions group together to form compounds which are electrically neutral.

12.3 Lewis Structures

- In covalent compounds, nonmetals share electrons so that both atoms achieve noble gas configurations.
- Lewis structures represent the valence electron arrangements of the atoms in a compound.
- The rules for drawing Lewis structures recognize the importance of noble gas electron configurations.
 - Duet rule for hydrogen
 - Octet rule for most other atoms
- Some molecules have more than one valid Lewis structure, called resonance.
- Some molecules violate the octet rule for the component atoms.
 - Examples are BF_3 , NO_2 , and NO .

12.4 Structures of Molecules

- Molecular structure describes how the atoms in a molecule are arranged in space.
- Molecular structure can be predicted by using the valence shell electron pair repulsion (VSEPR) model.

All exercises with **blue numbers** have answers in the back of this book.

12.1 Characteristics of Chemical Bonds

A. Types of Chemical Bonds

1. Define a chemical bond.
2. What types of elements react to form *ionic* compounds? Give an example of the formation of an ionic compound from its elements.
3. What type of bonding requires the complete *transfer* of an electron from one atom to another? What type of bonding involves the sharing (either equally or unequally) of electrons between atoms?
4. Describe the type of chemical bonding that exists between the atoms in the hydrogen molecule, H_2 .
5. Describe the type of chemical bonding that exists between the atoms in the hydrogen fluoride molecule, HF.

B. Electronegativity

6. What do chemists mean by the term *electronegativity*? What does its electronegativity tell us about the atom?
7. What does it mean to say that a bond is *polar*? What are the conditions that give rise to a bond's being polar?
8. For each of the following sets of elements, identify the element expected to be most electronegative and that expected to be least electronegative.

a. K, Sc, Ca b. Br, F, At c. C, O, N

9. On the basis of the electronegativity values given in Figure 12.4, indicate whether each of the following bonds would be expected to be ionic, covalent, or polar covalent.

a. S—S
b. S—O
c. S—H
d. S—K

10. Which of the following molecules contain polar covalent bonds?

a. phosphorus, P_4
b. oxygen, O_2
c. ozone, O_3
d. hydrogen fluoride, HF

11. On the basis of the electronegativity values given in Figure 12.4, indicate which is the more polar bond in each of the following pairs.

a. H—O or H—N
b. H—N or H—F
c. H—O or H—F
d. H—O or H—Cl

12. On the basis of the electronegativity values given in Figure 12.4, indicate which bond of the following pairs has a more ionic character.

a. Na—O or Na—N
b. K—S or K—P
c. Na—Cl or K—Cl
d. Na—Cl or Mg—Cl

C. Bond Polarity and Dipole Moments

13. What is a *dipole moment*? Give four examples of molecules that possess dipole moments, and draw the direction of the dipole as shown in this section.

14. Why is the presence of a dipole moment in the water molecule so important? What are some properties of water that are determined by its polarity?

15. In each of the following diatomic molecules, which end of the molecule is positive relative to the other end?

a. hydrogen fluoride, HF
b. chlorine monofluoride, ClF
c. iodine monochloride, ICl

16. For each of the following bonds, draw a figure indicating the direction of the bond dipole, including which end of the bond is positive and which is negative.

a. P—F
b. P—O
c. P—C
d. P—H

17. For each of the following bonds, draw a figure indicating the direction of the bond dipole, including which end of the bond is positive and which is negative.

a. S—P
b. S—O
c. S—N
d. S—Cl

12.2 Characteristics of Ions and Ionic Compounds

A. Stable Electron Configurations and Charges on Ions

18. Nonmetals form negative ions by (losing/gaining) enough electrons to achieve the electron configuration of the next noble gas.
19. Explain how the atoms in *covalent* molecules achieve configurations similar to those of the noble gases. How does this differ from the situation in ionic compounds?
20. Write the electron configuration for each of the following atoms and for the simple ion that the element most commonly forms. In each case, indicate which noble gas has the same electron configuration as the ion.
- a. lithium, $Z = 3$ d. sulfur, $Z = 16$
b. bromine, $Z = 35$ e. magnesium, $Z = 12$
c. cesium, $Z = 55$
21. What simple ion does each of the following elements most commonly form?
- a. calcium, $Z = 20$
b. nitrogen, $Z = 7$
c. bromine, $Z = 35$
d. magnesium, $Z = 12$
22. On the basis of their electron configurations, predict the formula of the simple binary ionic compound likely to form when the following pairs of elements react with each other.
- a. sodium, Na, and sulfur, S
b. barium, Ba, and selenium, Se
c. magnesium, Mg, and bromine, Br
d. lithium, Li, and nitrogen, N
e. potassium, K, and hydrogen, H
23. Name the noble gas atom that has the same electron configuration as each of the ions in the following compounds.
- a. aluminum sulfide, Al_2S_3
b. magnesium nitride, Mg_3N_2
c. rubidium oxide, Rb_2O
d. cesium iodide, CsI

B. Ionic Bonding and Structures of Ionic Compounds

24. Is the formula we write for an ionic compound the *molecular* formula or the *empirical* formula? Why?

25. Describe in general terms the structure of ionic solids such as NaCl. How are the ions packed in the crystal?
26. Why are cations always smaller than the atoms from which they are formed?
27. Why are anions always larger than the atoms from which they are formed?
28. For each of the following pairs, indicate which species is larger. Explain your reasoning in terms of the electron structure of each species.
- a. Li^+ or F^- c. Ca^{2+} or Ca
b. Na^+ or Cl^- d. Cs^+ or I^-
29. For each of the following pairs, indicate which is larger.
- a. Cl^- or I^- c. Cl^- or Cl
b. Cl^- or Na^+ d. Cl^- or S^{2-}

12.3 Lewis Structures

A. Writing Lewis Structures

30. Why are the *valence* electrons of an atom the only electrons likely to be involved in bonding to other atoms?
31. Explain the “duet” and “octet” rules and how they are used to describe the arrangement of electrons in a molecule.
32. How many electrons are involved when two atoms in a molecule are connected by a “double bond”? Write the Lewis structure of a molecule containing a double bond.
33. Give the *total* number of *valence* electrons in each of the following molecules.
- a. CBr_4 c. C_6H_6
b. NO_2 d. H_2O_2
34. Write a Lewis structure for each of the following simple molecules. Show all bonding valence electron pairs as lines and all nonbonding valence electron pairs as dots.
- a. NH_3 c. NCl_3
b. CCl_4 d. SiBr_4

B. Lewis Structures of Molecules with Multiple Bonds

35. Write a Lewis structure for each of the following simple molecules. Show all bonding valence electron pairs as lines and all nonbonding valence electron pairs as dots.
- a. H_2S c. C_2H_4
b. SiF_4 d. C_3H_8

36. Write a Lewis structure for each of the following simple molecules. Show all bonding valence electron pairs as lines and all nonbonding valence electron pairs as dots. For those molecules that exhibit resonance, draw the various possible resonance forms.

- Cl_2O
- CO_2
- SO_3

37. Write a Lewis structure for each of the following polyatomic ions. Show all bonding valence electron pairs as lines and all nonbonding valence electron pairs as dots. For those ions that exhibit resonance, draw the various possible resonance forms.

- sulfate ion, SO_4^{2-}
- phosphate ion, PO_4^{3-}
- sulfite ion, SO_3^{2-}

38. Write a Lewis structure for each of the following polyatomic ions. Show all bonding valence electron pairs as lines and all nonbonding valence electron pairs as dots. For those ions that exhibit resonance, draw the possible resonance forms.

- hydrogen phosphate ion, HPO_4^{2-}
- dihydrogen phosphate ion, H_2PO_4^-
- phosphate ion, PO_4^{3-}

12.4 Structures of Molecules

A. Molecular Structure

39. What is the geometric structure of the water molecule? How many pairs of valence electrons are there on the oxygen atom in the water molecule? What is the approximate H—O—H bond angle in water?

40. What is the geometric structure of the ammonia molecule? How many pairs of electrons surround the nitrogen atom in NH_3 ? What is the approximate H—N—H bond angle in ammonia?

41. What is the geometric structure of the boron trifluoride molecule, BF_3 ? How many pairs of valence electrons are present on the boron atom in BF_3 ? What are the approximate F—B—F bond angles in BF_3 ?

42. What is the geometric structure of the CH_4 molecule? How many pairs of valence electrons are present on the carbon atom of CH_4 ? Refer to Figure 12.9 and estimate the H—C—H bond angles in CH_4 .

B. The VSEPR Model

43. Why is the geometric structure of a molecule important, especially for biological molecules?

44. What general principles determine the molecular structure (shape) of a molecule?

45. Although the valence electron pairs in ammonia have a tetrahedral arrangement, the overall geometric structure of the ammonia molecule is *not* described as being tetrahedral. Explain.

46. Although both the BF_3 and NF_3 molecules contain the same number of atoms, the BF_3 molecule is flat, whereas the NF_3 molecule is trigonal pyramidal. Explain.

47. For the indicated atom in each of the following molecules or ions, give the number and arrangement of the electron pairs around that atom.

- P in PH_3
- Cl in ClO_4^-
- O in H_2O

48. Using the VSEPR theory, predict the molecular structure of each of the following molecules.

- CCl_4
- H_2S
- GeI_4

49. For each of the following molecules or ions, indicate the bond angle expected between the central atom and any two adjacent hydrogen atoms.

- H_2O
- NH_3
- NH_4^+
- CH_4

C. Molecules with Double Bonds

50. Using the VSEPR theory, predict the molecular structure of each of the following polyatomic ions.

- dihydrogen phosphate ion, H_2PO_4^-
- perchlorate ion, ClO_4^-
- sulfite ion, SO_3^{2-}

51. Predict the geometric structure of the carbonate ion, CO_3^{2-} . What are the bond angles in this molecule?

52. Predict the geometric structure of the acetylene molecule, C_2H_2 . What are the bond angles in this molecule?

Critical Thinking

- 53.** What is *resonance*? Give three examples of molecules or ions that exhibit resonance, and draw Lewis structures for each of the possible resonance forms.
- 54.** In each case, which of the following pairs of bonded elements forms the more polar bond?
- S—F or S—Cl
 - N—O or P—O
 - C—H or Si—H
- 55.** What do we mean by the *bond energy* of a chemical bond?
- 56.** For each of the following pairs of elements, identify which element would be expected to be more electronegative. It should not be necessary to look at a table of actual electronegativity values.
- Be or Ba
 - N or P
 - F or Cl
- 57.** In each of the following molecules, which end of the molecule is negative relative to the other end?
- carbon monoxide, CO
 - iodine monobromide, IBr
 - hydrogen iodide, HI
- 58.** What simple ion does each of the following elements most commonly form?
- sodium
 - iodine
 - potassium
 - calcium
 - sulfur
 - magnesium
 - aluminum
 - nitrogen
- 59.** Using the VSEPR theory, predict the molecular structure of each of the following molecules or ions containing multiple bonds.
- SO₂
 - SO₃
 - HCO₃⁻ (hydrogen is bonded to oxygen)
 - HCN
- 60.** Explain the difference between a covalent bond formed between two atoms of the same element and a covalent bond formed between atoms of two different elements.
- 61.** Rank the following from smallest to greatest radius: K⁺ Ca²⁺ Cl⁻ S²⁻ Ar
Explain.
- 62.** Consider the following molecules: O₃ H₂O SO₂ OCl₂
Which molecules can be described as having a bent shape? Do the bent molecules all have the same geometry? Explain.
- 63.** Consider the molecules CO₂ and OF₂.
- Draw a valid Lewis structure for each molecule.
 - Name the molecular geometry of each molecule.
 - Name the molecular structure (shape) of each molecule.
 - Both molecules have the general formula XY₂. Do they have the same shape? Why or why not?

1 Which of the following atoms is the most electronegative?

- A Na C Cl
B P D I

2 Which of the following is **true** concerning trends on the periodic table?

- A In general, smaller atoms have larger ionization energies and smaller electronegativity values.
B In general, smaller atoms have smaller ionization energies and smaller electronegativity values.
C In general, smaller atoms have smaller ionization energies and larger electronegativity values.
D In general, smaller atoms have larger ionization energies and larger electronegativity values.

3 Diatomic molecules that are formed from two different types of nonmetal atoms generally form _____.

- A ionic bonds
B polar covalent bonds
C nonpolar covalent bonds
D no bonds

4 Consider the carbon monoxide molecule pictured below. Note that the bond may be a single, double, or triple bond. What can best be said about the electrons involved in the bond?



- A There is an equal chance of a bonding electron's being in the area marked by either box.
B A bonding electron is more likely to exist in the area marked by Box A because carbon needs more electrons to get to a noble gas electron configuration.
C A bonding electron is more likely to exist in the area marked by Box B because the oxygen nucleus has a stronger attraction for electrons than does the carbon nucleus.
D Because we never know where the electron is, we cannot say where a bonding electron is more likely to be.

5 Which of the following does not have a noble gas electron configuration?

- A The Na^+ ion
B The Mg^{2+} ion
C The S^{2-} ion
D The Fe^{3+} ion

6 The electron configuration $1s^2 2s^2 2p^6 3s^2 3p^6$ is the correct electron configuration for the most stable form of which ion?

- A The calcium ion
B The magnesium ion
C The fluoride ion
D The oxide ion

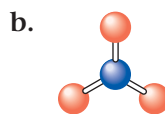
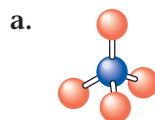
7 What is the expected ground state electron configuration for the most stable ion of bromine?

- A $[\text{Ar}] 4s^2 3d^{10} 4p^5$
B $[\text{Ar}] 4s^2 3d^{10} 4p^6$
C $[\text{Ar}] 4s^2 4d^{10} 4p^6$
D $[\text{Ar}] 4s^2 4d^{10} 4p^5$

8 How many resonance structures can be drawn for the carbonate ion?

- A 1
B 2
C 3
D 4

9 Write the name of each of the following shapes of molecules



LOOKING AHEAD

13.1 Describing the Properties of Gases


- A. Pressure
 - Atmospheric Pressure
 - Units of Pressure
- B. Pressure and Volume: Boyle's Law
 - A Closer Look
- C. Volume and Temperature: Charles's Law
- D. Volume and Moles: Avogadro's Law

13.2 Using Gas Laws to Solve Problems

- A. The Ideal Gas Law
- B. Dalton's Law of Partial Pressures
- C. Gas Stoichiometry
 - Molar Volume

13.3 Using a Model to Describe Gases

- A. Laws and Models: A Review
- B. The Kinetic Molecular Theory of Gases
- C. The Implications of the Kinetic Molecular Theory
 - The Meaning of Temperature
 - The Relationship Between Pressure and Temperature
 - The Relationship Between Volume and Temperature
- D. Real Gases



Hoop of steam being ejected from the Bocca Nuova crater on Mount Etna in Sicily.

IN YOUR LIFE

We are all familiar with gases. In fact, we live immersed in a gaseous “sea”—a mixture of nitrogen [$\text{N}_2(\text{g})$], oxygen [$\text{O}_2(\text{g})$], water vapor [$\text{H}_2\text{O}(\text{g})$], and small amounts of other gases. So it is important from a practical point of view for us to understand the properties of gases.

For example, you know that blowing air into a balloon causes it to expand—the volume of the balloon increases as you put more air into it. On the other hand, when you add more air to an inflated basketball, the ball doesn’t expand, it gets “harder.” In this case the added air increases the pressure inside the ball rather than causing an increase in volume. Although you probably have never tried this, can you guess what happens when an inflated balloon is placed in a freezer? The balloon gets smaller—its volume decreases. (You can easily do this experiment at home.) When we cool a gas, its volume decreases.

Image not available for electronic use. Please refer to the image in the textbook.

The Breitling Orbiter 3, shown over the Swiss Alps, recently completed a nonstop trip around the world.

WHAT DO YOU KNOW?

Prereading Questions

1. How does a gas differ from a solid and a liquid?
2. Have you heard of the term *barometric pressure*? What does it mean?
3. How does a law differ from a theory?
4. What does the temperature of a sample measure?

SECTION 13.1

Describing the Properties of Gases

Key Terms

- Barometer
- Torr (mm Hg)
- Standard atmosphere
- Pascal
- Boyle's law
- Absolute zero
- Charles's law
- Avogadro's law

Objectives

- To learn about atmospheric pressure and how barometers work
- To learn the units of pressure
- To understand how the pressure and volume of a gas are related
- To do calculations involving Boyle's law
- To learn about absolute zero
- To understand how the volume and temperature of a gas are related
- To do calculations involving Charles's law
- To understand how the volume and number of moles of a gas are related
- To do calculations involving Avogadro's law

Gases provide an excellent example of the scientific method (first discussed in Chapter 1). Recall that scientists study matter by making *observations* that are summarized into *laws*. We try to explain the observed behavior by hypothesizing what the atoms and molecules of the substance are doing. This explanation based on the microscopic world is called a *model* or *theory*.

Our experiences show us that when we make a change in a property of a gas, other properties change in a predictable way. In this chapter we will discuss relationships among the characteristics of gases such as pressure, volume, temperature, and amount of gas. These relationships were discovered by making observations as simple as seeing that a balloon expands when you blow into it, or that a sealed balloon will shrink if you put it into the freezer. But we also want to explain these observations. Why do gases behave the way they do? To explain gas behavior we will propose a model called the kinetic molecular theory.

DID YOU KNOW

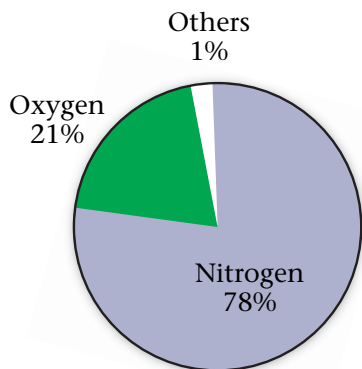
As a gas, water occupies 1300 times as much space as it does as a liquid at 25 °C and atmospheric pressure.

A. Pressure

We know that a gas uniformly fills any container, is easily compressed, and mixes completely with any other gas. One of the most obvious properties of a gas is that it exerts pressure on its surroundings. For example, when you blow up a balloon, the air pushes against the elastic sides of the balloon and keeps it firm.

The gases most familiar to us form the earth's atmosphere. The pressure exerted by the gaseous mixture that we call air can be dramatically demonstrated by the experiment shown in **Figure 13.1**. A small volume of water is placed in a metal can and the water is boiled, which fills the can with steam. The can is then sealed and allowed to cool. Why does the can collapse as it cools? It is the atmospheric pressure that crumples the can. When the can is cooled after being sealed so that no air can flow in, the water vapor (steam) inside the can condenses to a very small volume of liquid water. As a gas, the water vapor filled the can, but when it is condensed to a liquid, the liquid does not come close to filling the can. The H₂O molecules formerly present as a gas are now collected in a much smaller volume of liquid, and there are very few molecules of gas left to exert pressure outward and counteract the air pressure. As a result, the pressure exerted by the gas molecules in the atmosphere smashes the can.

Percent Composition of Dry Air



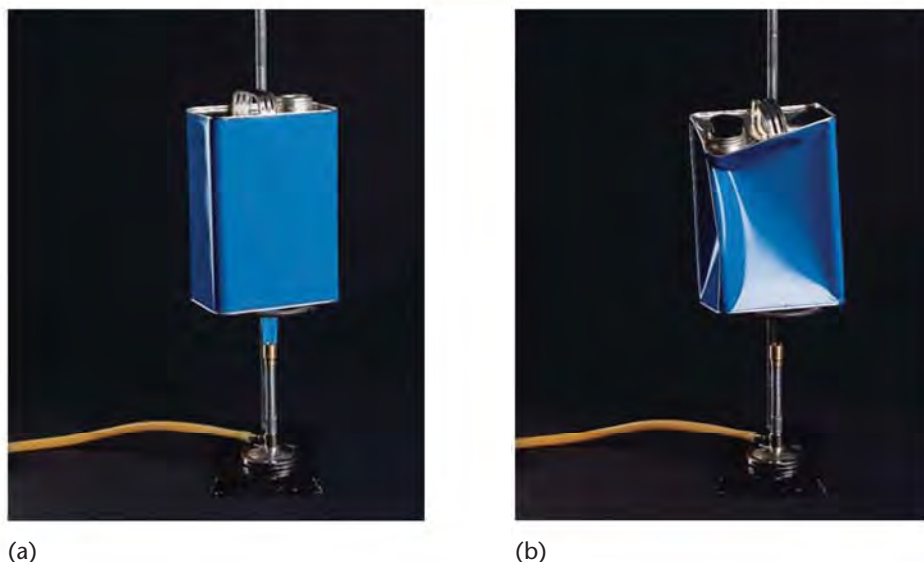
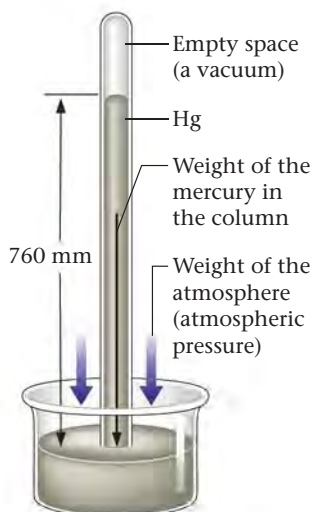


Figure 13.1

The pressure exerted by the gases in the atmosphere can be demonstrated by boiling water in a can (a), and then turning off the heat and sealing the can. As the can cools, the water vapor condenses, lowering the gas pressure inside the can. This causes the can to crumple (b).

Measuring Pressure A device that measures atmospheric pressure, the **barometer**, was invented in 1643 by an Italian scientist named Evangelista Torricelli (1608–1647), who had been a student of the famous astronomer Galileo. Torricelli's barometer is constructed by filling a glass tube with liquid mercury and inverting it in a dish of mercury.



Notice that a large quantity of mercury stays in the tube. In fact, at sea level the height of this column of mercury averages 760 mm. Why does this mercury stay in the tube, seemingly in defiance of gravity? The pressure exerted by the atmospheric gases on the surface of the mercury in the dish keeps the mercury in the tube.

Active Reading Question

In the mercury barometer shown above, what keeps all of the mercury from flowing out of the tube? Why does some of the mercury flow out of the tube?

Barometer

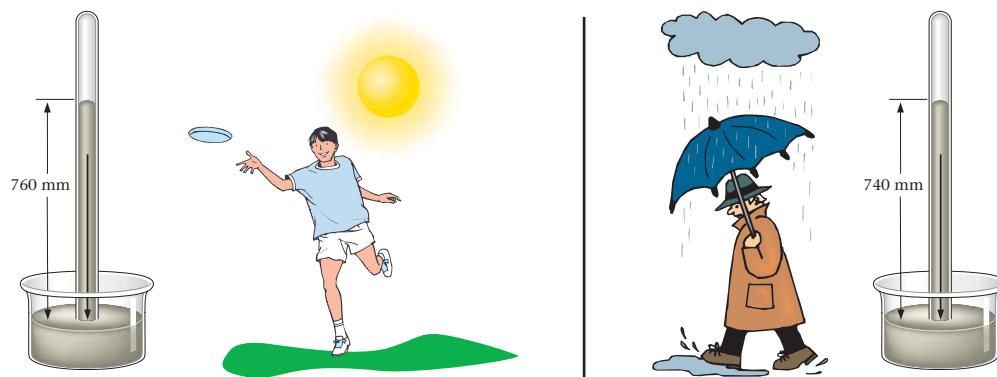
A device that measures atmospheric pressure

DID YOU KNOW

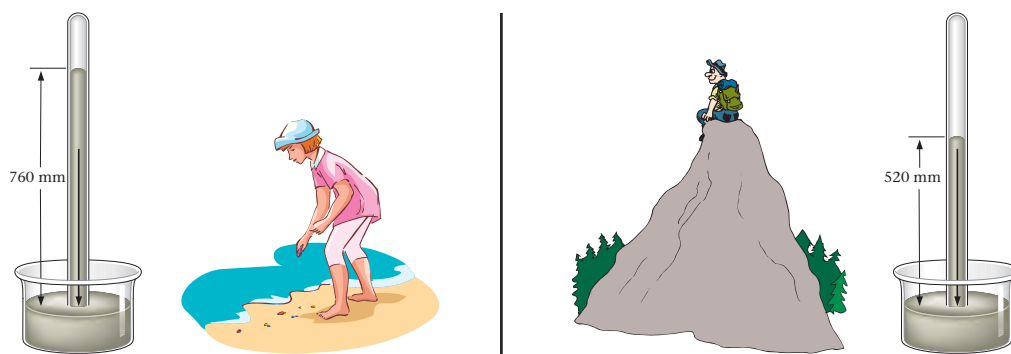
Soon after Torricelli died, a German physicist named Otto von Guericke invented an air pump. In a famous demonstration for the King of Prussia in 1683, Guericke placed two hemispheres together, pumped the air out of the resulting sphere through a valve, and showed that teams of horses could not pull the hemispheres apart. Then, after secretly opening the air valve, Guericke easily separated the hemispheres by hand. The King of Prussia was so impressed that he awarded Guericke a lifetime pension!

Atmospheric Pressure

Atmospheric pressure results from the mass of air being pulled toward the center of the earth by gravity—in other words it results from the weight of the air.



Changing weather conditions cause the atmospheric pressure to vary so that the barometric pressure at sea level is not always 760 mm. A “low” pressure system is often found during stormy weather. A “high” pressure often indicates fair weather.



Atmospheric pressure varies with altitude. In Breckenridge, Colorado (elevation 9600 feet), the atmospheric pressure is about 520 mm because there is less air pushing down on the earth’s surface than at sea level.

Information

Mercury is used to measure pressure because of its high density. The column of water required to measure a given pressure would be 13.6 times as high as a mercury column.

Units of Pressure

Because instruments used for measuring pressure (see **Figure 13.2**) often contain mercury, the most commonly used units for pressure are based on the height of the mercury column (in millimeters) that the gas pressure can support. The unit **mm Hg** (millimeters of mercury) is often called the **torr** in honor of Torricelli. The terms *torr* and *mm Hg* are both used by chemists. A related unit for pressure is the **standard atmosphere** (abbreviated atm).

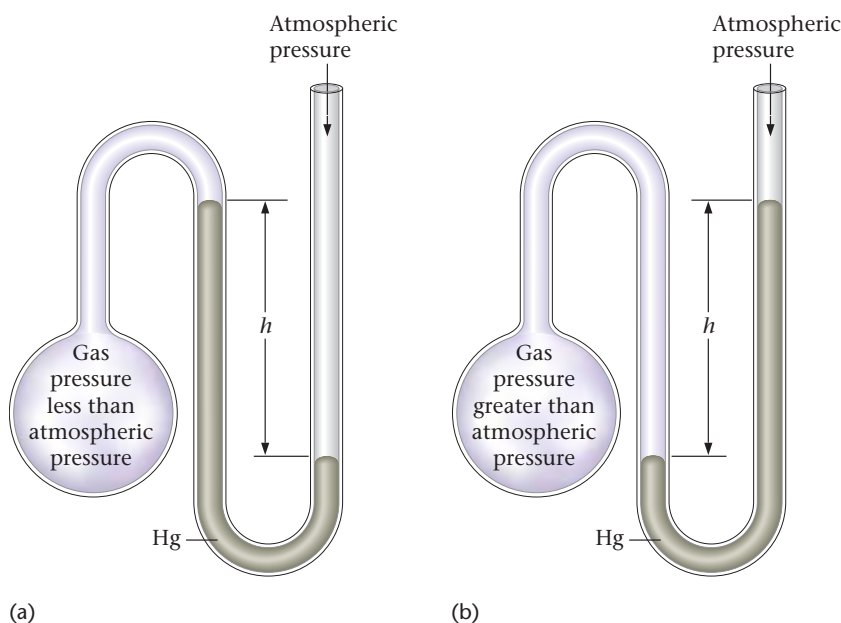
$$1 \text{ standard atmosphere} = 1.000 \text{ atm} = 760.0 \text{ mm Hg} = 760.0 \text{ torr}$$

The SI unit for pressure is the **pascal** (abbreviated Pa).

$$1 \text{ standard atmosphere} = 101,325 \text{ Pa}$$

Thus 1 atmosphere is about 100,000 or 10^5 pascals. Because the pascal is so small we will use it sparingly in this book. A unit of pressure that is employed in the engineering sciences and that we use for measuring tire pressure is pounds per square inch, abbreviated psi.

$$1.000 \text{ atm} = 14.69 \text{ psi}$$

**Figure 13.2**

A device called a manometer is used for measuring the pressure of a gas in a container. The pressure of the gas is equal to h (the difference in mercury levels) in units of torr (equivalent to mm Hg). (a) Gas pressure = atmospheric pressure $- h$. (b) Gas pressure = atmospheric pressure $+ h$.

i nformation

To convert pressure to the units needed, remember

$$\begin{array}{r} 1.000 \text{ atm} \\ \hline 760.0 \text{ mm Hg} \\ 760.0 \text{ torr} \\ 14.69 \text{ psi} \\ 101,325 \text{ Pa} \end{array}$$

Sometimes we need to convert from one unit of pressure to another. We do this by using conversion factors. The process is illustrated in Example 13.1.

EXAMPLE 13.1

Pressure Unit Conversions

The pressure of the air in a tire is measured to be 28 psi. Represent this pressure in atmospheres, torr, and pascals.

Solution

Where do we want to go?

$$28 \frac{\text{lb}}{\text{in.}^2} \rightarrow ? \text{ atm} \rightarrow ? \text{ torr} \rightarrow ? \text{ pascals}$$

What do we know?

- $28 \frac{\text{lb}}{\text{in.}^2}$
- $1.000 \text{ atm} = 14.69 \frac{\text{lb}}{\text{in.}^2}$
- $1.000 \text{ atm} = 760.0 \text{ torr}$
- $1.000 \text{ atm} = 101,325 \text{ Pa}$

How do we get there?

To convert from pounds per square inch to atmospheres, we need the equivalence statement

$$1.000 \text{ atm} = 14.69 \text{ psi}$$

which leads to the conversion factor

$$\frac{1.000 \text{ atm}}{14.69 \text{ psi}}$$

$$28 \text{ psi} \times \frac{1.000 \text{ atm}}{14.69 \text{ psi}} = 1.9 \text{ atm}$$

To convert from atmospheres to torr, we use the equivalence statement

$$1.000 \text{ atm} = 760.0 \text{ torr}$$



Checking the air pressure in a tire

MATH

$$1.9 \times 760.0 = 1444$$

$$1444 \xrightarrow{\text{Rounds to}} 1400$$

$$1400 = 1.4 \times 10^3$$

MATH



$$1.9 \times 101,325.0 \\ = 192,517.5$$

$$192,517.5 \xrightarrow{\text{Rounds to}} 190,000$$

$$190,000 = 1.9 \times 10^5$$

which leads to the conversion factor

$$\frac{760.0 \text{ torr}}{1.000 \text{ atm}}$$

$$1.9 \text{ atm} \times \frac{760.0 \text{ torr}}{1.000 \text{ atm}} = 1.4 \times 10^3 \text{ torr}$$

To change from torr to pascals, we need the equivalence statement

$$1.000 \text{ atm} = 101,325 \text{ Pa}$$

which leads to the conversion factor

$$\frac{101,325 \text{ Pa}}{1.000 \text{ atm}}$$

$$1.9 \text{ atm} \times \frac{101,325 \text{ Pa}}{1.000 \text{ atm}} = 1.9 \times 10^5 \text{ Pa}$$

Does it make sense?

The best way to check a problem like this is to make sure the units on the answer are the units required.

Practice Problem • Exercise 13.1

On a summer day in Breckenridge, Colorado, the atmospheric pressure is 525 mm Hg. What is this air pressure in atmospheres?

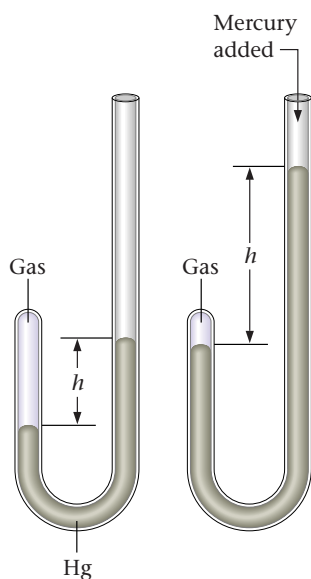


Figure 13.3

A J-tube similar to the one used by Boyle. The pressure on the trapped gas can be changed by adding or withdrawing mercury.

B. Pressure and Volume: Boyle's Law

The first careful experiments on gases were performed by the Irish scientist Robert Boyle (1627–1691). Using a J-shaped tube closed at one end (see **Figure 13.3**), which he reportedly set up in the multistory entryway of his house, Boyle studied the relationship between the pressure of the trapped gas and its volume. Representative values from Boyle's experiments are given in **Table 13.1**. The units given for the volume (cubic inches) and pressure (inches of mercury) are the ones Boyle used. Keep in mind that the metric system was not in use at this time.

Table 13.1

A Sample of Boyle's Observations (moles of gas and temperature both constant)

Experiment	Pressure (in. Hg)	Volume (in. ³)	Pressure × Volume (in. Hg) × (in. ³)	
			Actual	Rounded*
1	29.1	48.0	1396.8	1.40×10^3
2	35.3	40.0	1412.0	1.41×10^3
3	44.2	32.0	1414.4	1.41×10^3
4	58.2	24.0	1396.8	1.40×10^3
5	70.7	20.0	1414.0	1.41×10^3
6	87.2	16.0	1395.2	1.40×10^3
7	117.5	12.0	1410.0	1.41×10^3

*Three significant figures are allowed in the product because both of the numbers that are multiplied together have three significant figures.

The Cartesian Diver**Materials**

- Cartesian diver
- water

Procedure

1. Obtain a Cartesian diver from your teacher.
2. Squeeze the diver. What happens? Make careful observations.

Results/Analysis

1. Explain your observations. Feel free to experiment with the diver. It is a good idea to take the diver apart and experiment with variables (for example, what happens if the bottle is not completely filled with water?). Be sure to reconstruct the diver so that it works again—this effort will help you better understand it.

First let's examine Boyle's observations (Table 13.1) for general trends. Note that as the pressure increases, the volume of the trapped gas decreases. In fact, if you compare the data from experiments 1 and 4, you can see that as the pressure is doubled (from 29.1 to 58.2), the volume of the gas is halved (from 48.0 to 24.0). The same relationship can be seen in experiments 2 and 5 and in experiments 3 and 6 (approximately).

We can see the relationship between the volume of a gas and its pressure more clearly by looking at the product of the values of these two properties ($P \times V$) using Boyle's observations. This product is shown in the last column of Table 13.1. Note that for all the experiments,

$$P \times V = 1.4 \times 10^3 \text{ (in. Hg)} \times \text{in.}^3$$

with only a slight variation due to experimental error. Other similar measurements on gases show the same behavior. This means that the relationship of the pressure and volume of a gas can be expressed as

pressure times volume equals a constant

or in terms of an equation as

$$PV = k$$

which is called **Boyle's law**, where k is a constant at a specific temperature for a given amount of gas. For the data we used from Boyle's experiment, $k = 1.41 \times 10^3 \text{ (in. Hg)} \times \text{in.}^3$

It is often easier to visualize the relationships between two properties if we make a graph. **Figure 13.4** uses the data given in Table 13.1 to show how pressure is related to volume. This graph shows that V decreases as P increases. When this type of relationship exists, we say that volume and pressure are *inversely proportional*; when one increases, the other decreases. Boyle's law is illustrated by the gas samples below.

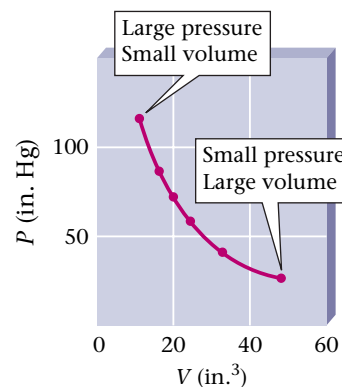
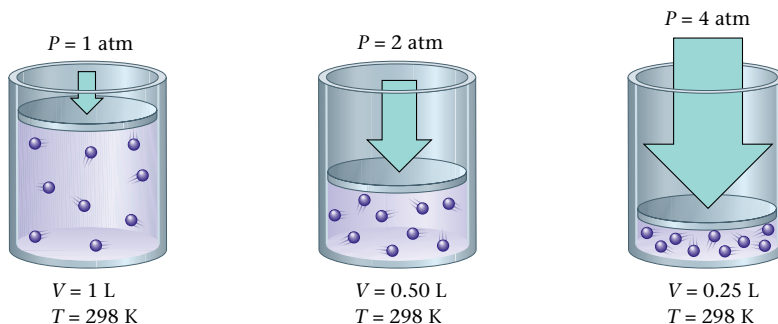


Figure 13.4

A plot of P versus V from Boyle's data in Table 13.1

**Boyle's law**

The pressure of a given sample of a gas is inversely related to the volume of the gas at constant temperature. $PV = k$

Information

For Boyle's law to hold, the amount of gas (moles) must not be changed. The temperature must also be constant.

A Closer Look

Boyle's law means that if we know the volume of a gas at a given pressure, we can predict the new volume if the pressure is changed, *provided that neither the temperature nor the amount of gas is changed*. For example, if we represent the original pressure and volumes as P_1 and V_1 and the final values as P_2 and V_2 , using Boyle's law we can write

$$P_1 V_1 = k$$

and

$$P_2 V_2 = k$$

We can also say

$$P_1 V_1 = k = P_2 V_2$$

$$P_1 V_1 = P_2 V_2 \quad \text{Boyle's law (constant temperature and amount of gas)}$$

We can solve for the final volume (V_2) by dividing both sides of the equation by P_2 .

$$\frac{P_1 V_1}{P_2} = \frac{P_2 V_2}{P_2}$$

Canceling the P_2 terms on the right gives

$$\frac{P_1}{P_2} \times V_1 = V_2$$

or

$$V_2 = V_1 \times \frac{P_1}{P_2}$$

This equation tells us that we can calculate the new gas volume (V_2) by multiplying the original volume (V_1) by the ratio of the original pressure to the final pressure (P_1/P_2).

Active Reading Question

Provide a real world example of Boyle's law.

EXAMPLE 13.2

Calculating Volume Using Boyle's Law

Freon-12 (the common name for the compound CCl_2F_2) was once widely used in refrigeration systems, but has now been replaced by other compounds that do not lead to the breakdown of the protective ozone in the upper atmosphere. Consider a 1.5-L sample of gaseous CCl_2F_2 at a pressure of 56 torr. If the pressure is changed to 150 torr at a constant temperature,

- Will the volume of the gas increase or decrease?
- What will be the new volume of the gas?

Solution

Where do we want to go?

Will the volume of the gas increase or decrease?

$$V_{\text{final}} = ? \text{ L}$$



Neon signs in Hong Kong

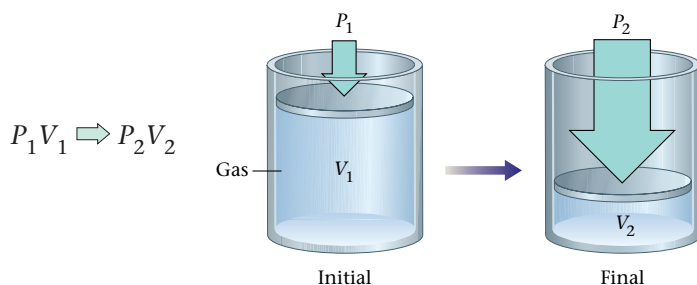
What do we know?

Initial	Final
$P_1 = 56 \text{ torr}$	$P_2 = 150 \text{ torr}$
$V_1 = 1.5 \text{ L}$	$V_2 = ? \text{ L}$

- Temperature is constant
- Boyle's law $P_1V_1 = P_2V_2$

How do we get there?

Drawing a picture is often helpful as we solve a problem. Notice that the pressure is increased from 56 torr to 150 torr, so the volume must decrease:



We can verify this by using Boyle's law in the form

$$V_2 = V_1 \times \frac{P_1}{P_2}$$

Note that V_2 is obtained by “correcting” V_1 using the ratio P_1/P_2 . Because P_1 is less than P_2 , the ratio P_1/P_2 is a fraction that is less than 1. Thus V_2 must be a fraction of (smaller than) V_1 ; the volume decreases.

We calculate V_2 as follows:

$$V_2 = V_1 \times \frac{P_1}{P_2} = 1.5 \text{ L} \times \frac{56 \text{ torr}}{150 \text{ torr}} = 0.56 \text{ L}$$

\uparrow
 \uparrow

V_1
 P_2

Does it make sense?

Since the pressure increases we would expect the volume to decrease. The volume of the gas decreases from 1.5 L to 0.56 L.

Practice Problem • Exercise 13.2

A sample of neon to be used in a neon sign has a volume of 1.51 L at a pressure of 635 torr. Calculate the volume of the gas after it is pumped into the glass tubes of the sign, where it shows a pressure of 785 torr.

information

The fact that the volume decreases in Example 13.2 makes sense because the pressure was increased. *To help catch errors, make it a habit to check whether an answer to a problem makes physical sense.*

EXAMPLE 13.3

Calculating Pressure Using Boyle's Law

In an automobile engine the gaseous fuel–air mixture enters the cylinder and is compressed by a moving piston before it is ignited. In a certain engine the initial cylinder volume is 0.725 L. After the piston moves up, the volume is 0.075 L. The fuel–air mixture initially has a pressure of 1.00 atm. Calculate the pressure of the compressed fuel–air mixture, assuming that both the temperature and the amount of gas remain constant.

Solution

Where do we want to go?

Pressure of the compressed fuel–air mixture = ? atm

What do we know?

Initial	Final
$P_1 = 1.00 \text{ atm}$	$P_2 = ?$
$V_1 = 0.725 \text{ L}$	$V_2 = 0.075 \text{ L}$

- Temperature is constant
- Amount of gas is constant
- Boyle's law $P_1V_1 = P_2V_2$

How do we get there?

We solve Boyle's law in the form $P_1V_1 = P_2V_2$ for P_2 by dividing both sides by V_2 to give the equation

$$P_2 = P_1 \times \frac{V_1}{V_2} = 1.00 \text{ atm} \times \frac{0.725 \cancel{\text{L}}}{0.075 \cancel{\text{L}}} = 9.7 \text{ atm}$$

Does it make sense?

Since the volume gets smaller the pressure must increase. Pressure and volume are inversely related.

MATH



$$\frac{0.725}{0.075} = 9.666 \dots$$

$$9.666 \xrightarrow{\text{Rounds to}} 9.7$$

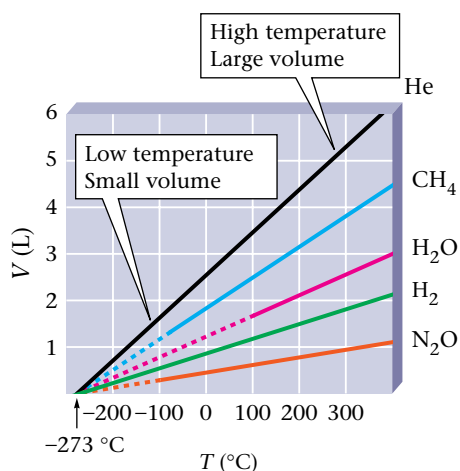


Figure 13.5

Plots of V (L) versus T ($^{\circ}\text{C}$) for several gases. Note that each sample of gas contains a different number of moles to spread out the plots.

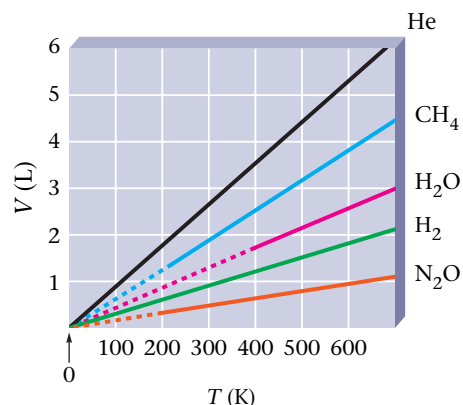
C. Volume and Temperature: Charles's Law

In the century following Boyle's findings, scientists continued to study the properties of gases. The French physicist Jacques Charles (1746–1823), who was the first person to fill a balloon with hydrogen gas and who made the first solo balloon flight, showed that the volume of a given amount of gas (at constant pressure) increases with the temperature of the gas. That is, the volume increases when the temperature increases. A plot of the volume of a given sample of gas (at constant pressure) versus its temperature (in Celsius degrees) gives a straight line. This type of relationship is called *linear*, and this behavior is shown for several gases in **Figure 13.5**.

The solid lines in Figure 13.5 are based on actual measurements of temperature and volume for the gases listed. As we cool the gases they eventually liquefy, so we cannot determine any experimental points below this temperature. However, when we extend each straight line (which is called

extrapolation and is shown here by a dashed line), something very interesting happens. *All* of the lines extrapolate to zero volume at the same temperature: $-273\text{ }^{\circ}\text{C}$. This suggests that $-273\text{ }^{\circ}\text{C}$ is the lowest possible temperature, because a negative volume is physically impossible. In fact, experiments have shown that matter cannot be cooled to temperatures lower than $-273\text{ }^{\circ}\text{C}$. Therefore, this temperature is defined as **absolute zero** on the Kelvin scale.

When the volumes of the gases shown in Figure 13.5 are plotted against temperature on the Kelvin scale rather than the Celsius scale, the graph looks like this.



These plots show that the volume of each gas is *directly proportional to the temperature* (in kelvins) and extrapolates to zero when the temperature is 0 K. Let's illustrate this statement with an example. Suppose we have 1 L of gas at 300 K. When we double the temperature of this gas to 600 K (without changing its pressure), the volume also doubles, to 2 L. Verify this type of behavior by looking carefully at the lines for various gases shown in the graph above.

The direct proportionality between volume and temperature (in kelvins) is represented by the equation known as **Charles's law**:

$$V = bT$$

where T is temperature in kelvins and b is the proportionality constant. Charles's law holds for a given sample of gas at constant pressure. It tells us that (for a given amount of gas at a given pressure) the volume of the gas is directly proportional to the temperature on the Kelvin scale:

$$V = bT \quad \text{or} \quad \frac{V}{T} = b = \text{constant}$$

Notice that in the second form, this equation states that the *ratio* of V to T (in kelvins) must be constant. Thus, when we triple the temperature (in kelvins) of a sample of gas, the volume of the gas triples also.

$$\frac{V}{T} = \frac{3 \times V}{3 \times T} = b = \text{constant}$$

We can also write Charles's law in terms of V_1 and T_1 (the initial conditions) and V_2 and T_2 (the final conditions).

$$\frac{V_1}{T_1} = b \quad \text{and} \quad \frac{V_2}{T_2} = b$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad \text{Charles's law (constant } P \text{ and amount of gas)}$$



The air in a balloon expands when it is heated. This means that some of the air escapes from the balloon, lowering the air density inside and thus making the balloon buoyant.

Absolute zero

$-273\text{ }^{\circ}\text{C}$

Charles's law

The volume of a given sample of a gas is directly related to the temperature of the gas at constant pressure. $V = bT$

CRITICAL THINKING ?

According to Charles's law, doubling the Kelvin temperature of a gas doubles its volume at constant pressure and number of moles.

What if doubling the Celsius temperature of a gas doubled its volume at constant pressure and number of moles? How would the world be different?

DID YOU KNOW

Temperatures such as 0.00000002 K have been obtained in the laboratory, but 0 K has never been reached.

**Active Reading Questions**

1. Provide a real world example of Charles's law.
2. Why must temperature always be in units of Kelvin before using Charles's law?

EXAMPLE 13.4**Calculating Volume Using Charles's Law, I**

A 2.0-L sample of air is collected at 298 K and then cooled to 278 K. The pressure is held constant at 1.0 atm.

- Does the volume increase or decrease?
- Calculate the volume of the air at 278 K.

Solution**Where do we want to go?**

Will the volume increase or decrease?

$$V_{\text{final}} = ? \text{ L}$$

What do we know?

Initial	Final
$T_1 = 298 \text{ K}$	$T_2 = 278 \text{ K}$
$V_1 = 2.0 \text{ L}$	$V_2 = ? \text{ L}$

- Pressure is constant
- Charles's law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

Information

$$\frac{V_1}{T_1} \longrightarrow \frac{V_2}{T_2}$$

Temperature smaller,
volume smaller

How do we get there?

Because the gas is cooled the volume of the gas must decrease (T and V are directly proportional).

To calculate the new volume, V_2 , we will use Charles's law in the form

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

for V_2 . We can do this by multiplying both sides by T_2 and canceling.

$$T_2 \times \frac{V_1}{T_1} = \frac{V_2}{\cancel{T_2}} \times \cancel{T_2} = V_2$$

Thus

$$V_2 = T_2 \times \frac{V_1}{T_1} = 278 \text{ K} \times \frac{2.0 \text{ L}}{298 \text{ K}} = 1.9 \text{ L}$$

Does it make sense?

Since the temperature decreases we would expect the volume to decrease. The volume of the gas decreases from 2.0 L to 1.9 L.

EXAMPLE 13.5**Calculating Volume Using Charles's Law, II**

A sample of gas at 15 °C (at 1 atm) has a volume of 2.58 L. The temperature is then raised to 38 °C (at 1 atm).

- Does the volume of the gas increase or decrease?
- Calculate the new volume.

Solution**Where do we want to go?**

Will the volume increase or decrease?

$$V_{\text{final}} = ? \text{ L}$$

What do we know?

Initial	Final
$T_1 = 15 \text{ }^\circ\text{C}$	$T_2 = 38 \text{ }^\circ\text{C}$
$V_1 = 2.58 \text{ L}$	$V_2 = ? \text{ L}$

- Pressure is constant (1 atm)
- Charles's law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

How do we get there?

Because the gas is heated the volume of the gas must increase (T and V are directly proportional).

The temperatures are given in Celsius. To use Charles's law the temperatures must be in kelvins. To convert from Celsius to Kelvin:

$$15 \text{ }^\circ\text{C} = 15 + 273 = 299 \text{ K}$$

$$38 \text{ }^\circ\text{C} = 38 + 273 = 311 \text{ K}$$

Solving for V_2 gives

$$V_2 = V_1 \times \frac{T_2}{T_1} = 2.58 \text{ L} \left(\frac{311 \text{ K}}{299 \text{ K}} \right) = 2.79 \text{ L}$$

Does it make sense?

Since the temperature increases we would expect the volume to increase. The volume of the gas increases from 2.58 L to 2.79 L.

Practice Problem • Exercise 13.5

A child blows a soap bubble that contains air at 28 °C and has a volume of 23 cm³ at 1 atm. As the bubble rises, it encounters a pocket of cold air (temperature 18 °C). If there is no change in pressure, will the bubble get larger or smaller as the air inside cools to 18 °C? Calculate the new volume of the bubble.

EXAMPLE 13.6

Calculating Temperature Using Charles's Law

In former times, gas volume was used as a way to measure temperature by using devices called gas thermometers. Consider a gas that has a volume of 0.675 L at 35 °C and 1 atm pressure. What is the temperature (in units of °C) of a room where this gas has a volume of 0.535 L at 1 atm pressure?

Solution

Where do we want to go?

Temperature of the room = ? °C

What do we know?

Initial	Final
$T_1 = 35\text{ °C} = 308\text{ K}$	$T_2 = ?\text{ °C}$
$V_1 = 0.675\text{ L}$	$V_2 = 0.535\text{ L}$

- Pressure is constant (1 atm)
- Charles's Law $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

How do we get there?

First we multiply both sides of the Charles's law equation by T_2 .

$$T_2 \times \frac{V_1}{T_1} = \frac{V_2}{\cancel{T_2}} \times \cancel{T_2} = V_2$$

Next we multiply both sides by T_1 .

$$\cancel{T_1} \times T_2 \times \frac{V_1}{\cancel{T_1}} = T_1 \times V_2$$

This gives

$$T_2 \times V_1 = T_1 \times V_2$$

Now we divide both sides by V_1 (or multiply by $\frac{1}{V_1}$),

$$\frac{1}{\cancel{V_1}} \times T_2 \times \cancel{V_1} = \frac{1}{\cancel{V_1}} \times T_1 \times V_2$$

and obtain

$$T_2 = T_1 \times \frac{V_2}{V_1}$$

We have now isolated T_2 on one side of the equation, and we can do the calculation.

$$T_2 = T_1 \times \frac{V_2}{V_1} = (308\text{ K}) \times \frac{0.535\cancel{\text{ L}}}{0.675\cancel{\text{ L}}} = 244\text{ K}$$

To convert from units of K to units of °C, we subtract 273 from the Kelvin temperature.

$$T_{\text{°C}} = T_{\text{K}} - 273 = 244 - 273 = -29\text{ °C}$$

The room is very cold; the temperature is -29 °C .

Does it make sense?

The units are as required and the room is very cold (-29 °C).

D. Volume and Moles: Avogadro's Law

What is the relationship between the volume of a gas and the number of molecules present in the gas sample? Experiments show that when the number of moles of gas is doubled (at constant temperature and pressure), the volume doubles. In other words, the volume of a gas is directly proportional to the number of moles if temperature and pressure remain constant.

Figure 13.6 illustrates this relationship, which can also be represented by the equation

$$V = an \quad \text{or} \quad \frac{V}{n} = a$$

where V is the volume of the gas, n is the number of moles, and a is the proportionality constant. Note that this equation means that the ratio of V to n is constant as long as the temperature and pressure remain constant. Thus, when the number of moles of gas is increased by a factor of 5, the volume also increases by a factor of 5,

$$\frac{V}{n} = \frac{5 \times V}{5 \times n} = a = \text{constant}$$

and so on. In words, this equation means that *for a gas at constant temperature and pressure, the volume is directly proportional to the number of moles of gas*. This relationship is called **Avogadro's law** after the Italian scientist Amadeo Avogadro, who first postulated it in 1811.

For cases where the number of moles of gas is changed from an initial amount to another amount (at constant T and P), we can represent Avogadro's law as

$$\frac{V_1}{n_1} = a = \frac{V_2}{n_2}$$

Initial amount Final amount

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \quad \text{Avogadro's law (constant } T \text{ and } P)$$

Active Reading Question

Provide a real world example of Avogadro's law.

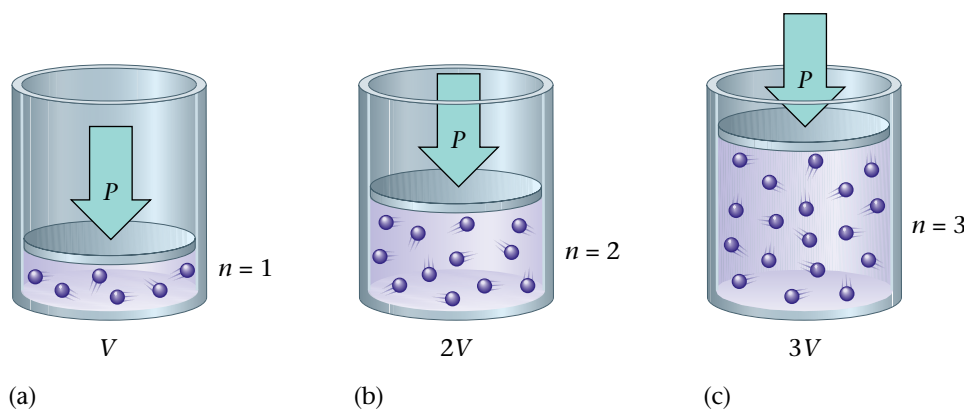


Figure 13.6

The relationship between volume V and number of moles n . As the number of moles is increased from 1 to 2 (a to b), the volume doubles. When the number of moles is tripled (c), the volume is also tripled. The temperature and pressure remain the same in these cases.

**The Candle and the Tumbler****Materials**

- candle
- glass
- water
- test tube

Procedure

1. Light a candle and let some wax drip onto the bottom of a glass that is taller than the candle. Blow out the candle and stick the candle to the glass before the wax solidifies.

2. Add water to the glass until the candle is about half-submerged. Do not get the wick of the candle wet.
3. Light the candle. Take a test tube that is wider than the candle and quickly place it over the candle, submerging the opening of the test tube. What happens? Make careful observations.

Results/Analysis

1. Explain your results.

EXAMPLE 13.7**Using Avogadro's Law in Calculations**

Suppose we have a 12.2-L sample containing 0.50 mol of oxygen gas, O_2 , at a pressure of 1 atm and a temperature of 25 °C. If all of this O_2 is converted to ozone, O_3 , at the same temperature and pressure, what will be the volume of the ozone formed?

Solution**Where do we want to go?**

Volume of the ozone formed = ? L

What do we know?

Initial (oxygen)	Final (ozone)
$n_1 = 0.50 \text{ mol}$	$n_2 = ? \text{ mol}$
$V_1 = 12.2 \text{ L}$	$V_2 = ? \text{ L}$

- Pressure is constant (1 atm)
- Temperature is constant (25 °C)
- $3 O_2(g) \rightarrow 2 O_3(g)$
- Avogadro's law $\frac{V_1}{n_1} = \frac{V_2}{n_2}$

How do we get there?

To do this problem we need to compare the moles of gas originally present to the moles of gas present after the reaction. We know that 0.50 mol of O_2 is present initially. To find out how many moles of O_3 will be present after the reaction, we need to use the balanced equation for the reaction.



We calculate the moles of O_3 produced by using the appropriate mole ratio from the balanced equation.

$$0.50 \text{ mol } O_2 \times \frac{2 \text{ mol } O_3}{3 \text{ mol } O_2} = 0.33 \text{ mol } O_3$$

Avogadro's law states that

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

where V_1 is the volume of n_1 moles of O_2 gas and V_2 is the volume of n_2 moles of O_3 gas. In this case we have

Initial Conditions **Final Conditions**

$$n_1 = 0.50 \text{ mol} \qquad n_2 = 0.33 \text{ mol}$$

$$V_1 = 12.2 \text{ L} \qquad V_2 = ?$$

Solving Avogadro's law for V_2 gives

$$V_2 = V_1 \times \frac{n_2}{n_1} = 12.2 \text{ L} \left(\frac{0.33 \text{ mol}}{0.50 \text{ mol}} \right) = 8.1 \text{ L}$$

Does it make sense?

Note that the volume decreases, as it should, because fewer molecules are present in the gas after O_2 is converted to O_3 .

Practice Problem • Exercise 13.7

Consider two samples of nitrogen gas (composed of N_2 molecules). Sample 1 contains 1.5 mol of N_2 and has a volume of 36.7 L at 25 °C and 1 atm. Sample 2 has a volume of 16.5 L at 25 °C and 1 atm. Calculate the number of moles of N_2 in Sample 2.

MATH

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

$$n_2 \times \frac{V_1}{n_1} = \frac{V_2}{n_2} \times n_2$$

$$V_1 \times \frac{n_2}{n_1} = V_2$$

SECTION 13.1

REVIEW QUESTIONS

- Mercury is a toxic substance. Why is it used instead of water in barometers and manometers?
- What is the SI unit for pressure? What is the unit commonly used in chemistry for pressure? Why aren't they the same?
- Copy and complete the following table.

torr	atm	pascals	mm Hg
459 torr			
		132,874 Pa	
	3.14 atm		
			842 mm Hg

- A 1.04-L sample of gas at 759 mm Hg pressure is expanded until its volume is 2.24 L. What is the pressure in the expanded gas sample (at constant temperature)?

- Compare Boyle's law, Charles's law, and Avogadro's law.
 - What remains constant in each law?
 - What are the variables in each law?
 - What do the graphs for these laws look like?
 - Write each law with V isolated. Using these equations and the graphs from part c, which law(s) show a directly proportional relationship? How can you tell?
- If 525 mL of gas at 25.0 °C is heated to 50.0 °C at constant pressure, calculate the new volume of the sample.
- A 1.50 mol sample of helium at a certain temperature and pressure has a volume of 31.4 L. A second sample of helium at the same temperature and pressure has a volume of 42.4 L. How many moles of helium are in the second sample?



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SECTION 13.2

Using Gas Laws to Solve Problems

Key Terms

- Universal gas constant
- Ideal gas law
- Ideal gas
- Combined gas law
- Partial pressure
- Dalton's law of partial pressures
- Molar volume
- Standard temperature and pressure (STP)

Objectives

- To understand the ideal gas law and use it in calculations
- To understand the relationship between the partial and total pressure of a gas mixture
- To do calculations involving Dalton's law of partial pressures
- To understand the molar volume of an ideal gas
- To learn the definition of STP
- To do stoichiometry calculations using the ideal gas law

A. The Ideal Gas Law

We have considered three laws that describe the behavior of gases as revealed by experimental observations.

$$\text{Boyle's law: } PV = k \quad \text{or} \quad V = \frac{k}{P} \quad (\text{at constant } T \text{ and } n)$$

$$\text{Charles's law:} \quad V = bT \quad (\text{at constant } P \text{ and } n)$$

$$\text{Avogadro's law:} \quad V = an \quad (\text{at constant } T \text{ and } P)$$

These relationships, which show how the volume of a gas depends on pressure, temperature, and number of moles of gas present, can be combined as follows:

$$V = R \left(\frac{Tn}{P} \right)$$

where R is the combined constant and is called the **universal gas constant**. When the pressure is expressed in atmospheres and the volume is in liters, R always has the value $0.08206 \frac{\text{L atm}}{\text{K mol}}$. We can rearrange the above equation by multiplying both sides by P ,

$$P \times V = P \times R \left(\frac{Tn}{P} \right)$$

to obtain the **ideal gas law**.

$$PV = nRT \quad \text{ideal gas law} \quad R = 0.08206 \frac{\text{L atm}}{\text{mol K}}$$

The ideal gas law involves all the important characteristics of a gas:

- pressure (P),
- volume (V),
- number of moles (n),
- temperature (T).

Knowledge of any three of these properties is enough to define completely the condition of the gas, because the fourth property can be determined by using the ideal gas law.

It is important to recognize that the ideal gas law is based on experimental measurements of the properties of gases. A gas that obeys this equation is said to behave *ideally*. That is, this equation defines the behavior of an **ideal gas**. Most gases obey this equation closely at pressures of approximately 1 atm or lower, when the temperature is approximately 0 °C or higher. You should assume ideal gas behavior when working problems involving gases in this text.

Information

Constant n means a constant number of moles of gas.

Ideal gas law

$$PV = nRT$$

Ideal gas

A hypothetical gas that exactly obeys the ideal gas law.

EXAMPLE 13.8**Using the Ideal Gas Law in Calculations**

A sample of hydrogen gas, H_2 , has a volume of 8.56 L at a temperature of 0°C and a pressure of 1.5 atm. Calculate the number of moles of H_2 present in this gas sample. Assume that the gas behaves ideally.

Solution**Where do we want to go?**

Moles of $\text{H}_2 = ? \text{ mol}$

What do we know?

- $P = 1.5 \text{ atm}$
- $T = 0^\circ\text{C} = 273 \text{ K}$
- $V = 8.56 \text{ L}$
- Ideal gas law $PV = nRT$

How do we get there?

We can calculate the number of moles of gas present by using the ideal gas law, $PV = nRT$. We solve for n by dividing both sides by RT :

$$\frac{PV}{RT} = \frac{nRT}{RT}$$

to give

$$\frac{PV}{RT} = n$$

Thus

$$n = \frac{PV}{RT} = \frac{(1.5 \text{ atm})(8.56 \text{ L})}{\left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}}\right)(273 \text{ K})} = 0.57 \text{ mol}$$

Practice Problem • Exercise 13.8

A weather balloon contains $1.10 \times 10^5 \text{ mol}$ of helium and has a volume of $2.70 \times 10^6 \text{ L}$ at 1.00 atm pressure. Calculate the temperature of the helium in the balloon in kelvins and in Celsius degrees.

EXAMPLE 13.9

Ideal Gas Law Calculations Involving Conversion of Units

What volume is occupied by 0.250 mol of carbon dioxide gas at 25 °C and 371 torr?

Solution

Where do we want to go?

Volume of CO₂ = ? L

What do we know?

- $P = 371$ torr
- $T = 25$ °C = 298 K
- $n = 0.250$ mol CO₂
- Ideal gas law $PV = nRT$

How do we get there?

We can use the ideal gas law to calculate the volume, but we must first convert pressure to atmospheres and temperature to the Kelvin scale.

$$P = 371 \text{ torr} = 371 \text{ torr} \times \frac{1.000 \text{ atm}}{760.0 \text{ torr}} = 0.488 \text{ atm}$$

$$T = 25 \text{ °C} = 25 + 273 = 298 \text{ K}$$

MATH

$$\begin{aligned} PV &= nRT \\ \frac{PV}{P} &= \frac{nRT}{P} \\ V &= \frac{nRT}{P} \end{aligned}$$

We solve for V by dividing both sides of the ideal gas law ($PV = nRT$) by P .

$$V = \frac{nRT}{P} = \frac{(0.250 \text{ mol}) \left(0.08206 \frac{\text{L atm}}{\text{K mol}}\right) (298 \text{ K})}{0.488 \text{ atm}} = 12.5 \text{ L}$$

The volume of the sample of CO₂ is 12.5 L.

Practice Problem • Exercise 13.9

Radon, a radioactive gas formed naturally in the soil, can cause lung cancer. It can pose a hazard to humans by seeping into houses, and there is concern about this problem in many areas. A 1.5-mol sample of radon gas has a volume of 21.0 L at 33 °C. What is the pressure of the gas?

Note that R has units of $\frac{\text{L atm}}{\text{K mol}}$. Accordingly, whenever we use the ideal gas law, we must express the volume in units of liters, the temperature in kelvins, and the pressure in atmospheres. When we are given data in other units, we must first convert to the appropriate units.

The ideal gas law can also be used to calculate the changes that will occur when the conditions of the gas are changed as illustrated in Example 13.10.

EXAMPLE 13.10**Using the Ideal Gas Law Under Changing Conditions**

Suppose we have a 0.240-mol sample of ammonia gas at 25 °C with a volume of 3.5 L at a pressure of 1.68 atm. The gas is compressed to a volume of 1.35 L at 25 °C. Use the ideal gas law to calculate the final pressure.

Solution**Where do we want to go?**

pressure of $\text{NH}_3 = ?$ atm

What do we know?

Initial	Final
$V_1 = 3.5 \text{ L}$	$V_2 = 1.35 \text{ L}$
$P_1 = 1.68 \text{ atm}$	$P_2 = ?$
$T_1 = 25 \text{ }^\circ\text{C} = 298 \text{ K}$	$T_2 = 25 \text{ }^\circ\text{C} = 298 \text{ K}$
$n_1 = 0.240 \text{ mol}$	$n_2 = 0.240 \text{ mol}$

- Ideal gas law $PV = nRT$

How do we get there?

Note that both n and T remain constant—only P and V change. Thus we could simply use Boyle's law ($P_1V_1 = P_2V_2$) to solve for P_2 . However, we will use the ideal gas law to solve this problem to introduce the idea that one equation—the ideal gas equation—can be used to solve almost any gas problem.

The key idea here is that in using the ideal gas law to describe a change in conditions for a gas, we always *solve the ideal gas equation in such a way that the variables that change are on one side of the equal sign and the constant terms are on the other side*. That is, we start with the ideal gas equation in the conventional form ($PV = nRT$) and rearrange it so that all the terms that change are moved to one side and all the terms that do not change are moved to the other side. In this case the pressure and volume change and the temperature and number of moles remain constant (as does R , by definition). So we write the ideal gas law as

$$\begin{array}{ccc} PV & = & nRT \\ \text{Change} & & \text{Remain constant} \end{array}$$

Because n , R , and T remain the same in this case, we can write $P_1V_1 = nRT$ and $P_2V_2 = nRT$. Combining these gives

$$P_1V_1 = nRT = P_2V_2 \text{ or } P_1V_1 = P_2V_2$$

and

$$P_2 = P_1 \times \frac{V_1}{V_2} = (1.68 \text{ atm}) \left(\frac{3.5 \cancel{\text{L}}}{1.35 \cancel{\text{L}}} \right) = 4.4 \text{ atm}$$

Practice Problem • Exercise 13.10

A sample of methane gas that has a volume of 3.8 L at 5 °C is heated to 86 °C at constant pressure. Calculate its new volume.

Information

Does this answer make sense? The volume was decreased (at constant temperature and constant number of moles), which means that the pressure should increase, as the calculation indicates.

Note that in solving Example 13.10, we actually obtained Boyle's law ($P_1V_1 = P_2V_2$) from the ideal gas equation. You might well ask, "Why go to all this trouble?" The idea is to learn to use the ideal gas equation to solve all types of gas law problems. This way you will never have to ask yourself, "Is this a Boyle's law problem or a Charles's law problem?"

We continue to practice using the ideal gas law in Example 13.11. Remember, the key idea is to rearrange the equation so that the quantities that change are moved to one side of the equation and those that remain constant are moved to the other.

EXAMPLE 13.11

Calculating Volume Changes Using the Ideal Gas Law

A sample of diborane gas, B_2H_6 , a substance that bursts into flames when exposed to air, has a pressure of 0.454 atm at a temperature of $-15^\circ C$ and a volume of 3.48 L. If conditions are changed so that the temperature is $36^\circ C$ and the pressure is 0.616 atm, what will be the new volume of the sample?

Solution

Where do we want to go?

Volume = ? L

What do we know?

Initial	Final
$V_1 = 3.48 \text{ L}$	$V_2 = ? \text{ L}$
$P_1 = 0.454 \text{ atm}$	$P_2 = 0.616 \text{ atm}$
$T_1 = -15^\circ C = 258 \text{ K}$	$T_2 = 36^\circ C = 309 \text{ K}$

- $n = \text{constant}$ (no gas is added or removed)
- Ideal gas law $PV = nRT$

How do we get there?

We rearrange the ideal gas equation ($PV = nRT$) by dividing both sides by T ,

$$\begin{aligned} PV &= nRT \\ \frac{PV}{T} &= n\frac{RT}{T} \\ \frac{PV}{T} &= nR \end{aligned}$$

$$\frac{PV}{T} = nR$$

Change Constant

which leads to the equation

$$\frac{P_1V_1}{T_1} = nR = \frac{P_2V_2}{T_2}$$

or

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

We can now solve for V_2 by dividing both sides by P_2 and multiplying both sides by T_2 .

$$\begin{aligned} \frac{1}{P_2} \times \frac{P_1V_1}{T_1} &= \frac{\cancel{P_2}V_2}{T_2} \times \frac{1}{\cancel{P_2}} = \frac{V_2}{T_2} \\ T_2 \times \frac{P_1V_1}{\cancel{P_2}T_1} &= \frac{V_2}{\cancel{T_2}} \times \cancel{T_2} = V_2 \end{aligned}$$

Information

Always convert the temperature to the Kelvin scale and the pressure to atmospheres when applying the ideal gas law.

MATH



That is,

$$\frac{T_2 P_1 V_1}{P_2 T_1} = V_2$$

It is sometimes convenient to think in terms of the ratios of the initial temperature and pressure and the final temperature and pressure. That is,

$$V_2 = \frac{T_2 P_1 V_1}{T_1 P_2} = V_1 \times \frac{T_2}{T_1} \times \frac{P_1}{P_2}$$

Substituting the information given yields

$$V_2 = \frac{309 \text{ K}}{258 \text{ K}} \times \frac{0.454 \text{ atm}}{0.616 \text{ atm}} \times 3.48 \text{ L} = 3.07 \text{ L}$$

Practice Problem • Exercise 13.11

A sample of argon gas with a volume of 11.0 L at a temperature of 13 °C and a pressure of 0.747 atm is heated to 56 °C and a pressure of 1.18 atm. Calculate the final volume.

CHEMISTRY IN YOUR WORLD

Consumer Connection

Snacks Need Chemistry, Too!

Have you ever wondered what makes popcorn pop? The popping is linked with the properties of gases. What happens when a gas is heated? Charles's law tells us that if the pressure is held constant, the volume of the gas must increase as the temperature is increased. But what happens if the gas being heated is trapped at a constant volume? We can see what happens by rearranging the ideal gas law ($PV = nRT$) as follows:

$$P = \left(\frac{nR}{V}\right)T$$

When n , R , and V are held constant, the pressure of a gas is directly proportional to the temperature. Thus, as the temperature of the trapped gas increases, its pressure also increases. This is exactly what happens inside a kernel of popcorn as it is heated. The moisture inside the kernel vaporized by the heat produces increasing pressure. The pressure finally becomes so great that the kernel breaks open, allowing the starch inside to expand to about 40 times its original size.

What's special about popcorn? Why does it pop while "regular" corn doesn't? William da Silva, a biologist at the University of Campinas in Brazil, has traced the "popability" of popcorn to its outer casing, called the pericarp. The molecules in the pericarp of popcorn, which are packed in a much more orderly way than in regular corn, transfer heat unusually quickly, producing a very fast pressure jump that pops the kernel. In addition, because the pericarp of popcorn is much thicker and stronger than that of regular corn, it can withstand more pressure, leading to a more explosive pop when the moment finally comes.

Popcorn popping



Image not available for electronic use. Please refer to the image in the textbook.

The equation obtained in Example 13.11,

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

is often called the **combined gas law** equation. It holds when the amount of gas (moles) is held constant. While it may be convenient to remember this equation, it is not necessary because you can always use the ideal gas equation.

Divers use a mixture of oxygen and helium in their breathing tanks when diving to depths greater than 150 feet.

B. Dalton's Law of Partial Pressures

Many important gases contain a mixture of components. One notable example is air. Scuba divers who are going deeper than 150 feet use another important mixture, helium and oxygen. Normal air is not used because the nitrogen present dissolves in the blood in large quantities as a result of the high pressures experienced by the diver under several hundred feet of water. When the diver returns too quickly to the surface, the nitrogen bubbles out of the blood just as soda fizzes when it's opened, and the diver gets "the bends"—a very painful and potentially fatal condition. Because helium gas is only sparingly soluble in blood, it does not cause this problem.

Studies of gaseous mixtures show that each component behaves independently of the others. In other words, a given amount of oxygen exerts the same pressure in a 1.0-L vessel whether it is alone or in the presence of nitrogen (as in the air) or helium.

Among the first scientists to study mixtures of gases was John Dalton. In 1803 Dalton summarized his observations in this statement:

Top Ten Components of Air (dry air at sea level)

Component	Volume Percent
N ₂	78
O ₂	21
Ar	0.93
CO ₂	0.01–0.10
Ne	0.0018
He	5.2 × 10 ⁻⁴
CH ₄	2.0 × 10 ⁻⁴
Kr	1.1 × 10 ⁻⁴
H ₂ , N ₂ O	5.0 × 10 ⁻⁴

MATH

$$PV = nRT$$

$$\frac{PV}{V} = \frac{nRT}{V}$$

$$P = \frac{nRT}{V}$$

Partial pressure

The independent pressure exerted by a gas in a mixture of gases.

Dalton's law of partial pressures

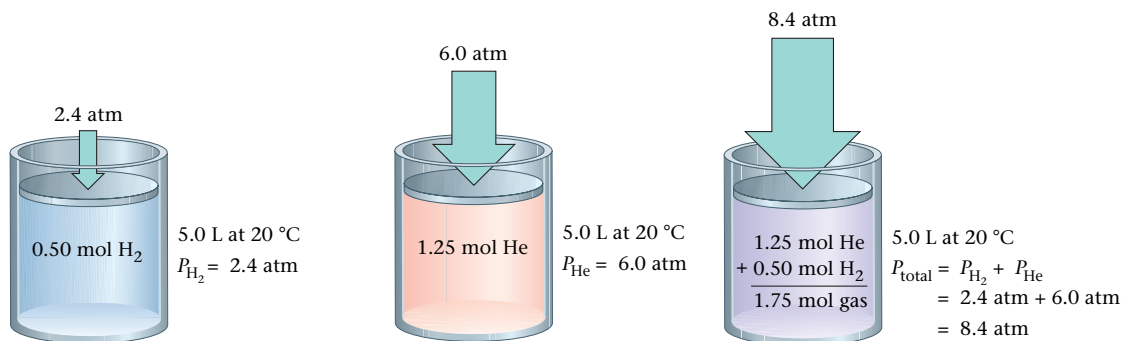
For a mixture of gases in a container, the total pressure exerted is the sum of the partial pressures of each of the gases.

Dalton's law of partial pressures For a mixture of gases in a container, the total pressure exerted is the sum of the partial pressures of the gases present.

The **partial pressure** of a gas is the pressure that the gas would exert if it were alone in the container. **Dalton's law of partial pressures** can be expressed as follows for a mixture containing three gases:

$$P_{\text{total}} = P_1 + P_2 + P_3$$

where the subscripts refer to the individual gases (gas 1, gas 2, and gas 3). The pressures P_1 , P_2 , and P_3 are the partial pressures; that is, each gas is responsible for only part of the total pressure.



Assuming that each gas behaves ideally, we can calculate the partial pressure of each gas from the ideal gas law:

$$P_1 = \frac{n_1RT}{V} \quad P_2 = \frac{n_2RT}{V} \quad P_3 = \frac{n_3RT}{V}$$

The total pressure of the mixture, P_{total} , can be represented as

$$\begin{aligned} P_{\text{total}} &= P_1 + P_2 + P_3 = \frac{n_1RT}{V} + \frac{n_2RT}{V} + \frac{n_3RT}{V} \\ &= n_1\left(\frac{RT}{V}\right) + n_2\left(\frac{RT}{V}\right) + n_3\left(\frac{RT}{V}\right) \\ &= (n_1 + n_2 + n_3)\left(\frac{RT}{V}\right) \\ &= n_{\text{total}}\left(\frac{RT}{V}\right) \end{aligned}$$

where n_{total} is the sum of the numbers of moles of the gases in the mixture. Thus, for a mixture of ideal gases, it is the *total number of moles of particles* that is important, not the *identity* of the individual gas particles. This idea is illustrated in **Figure 13.7**.

Active Reading Question

What is meant by the partial pressure of a gas?

The fact that the pressure exerted by an ideal gas is affected by the *number* of gas particles and is independent of the *nature* of the gas particles tells us two crucial things about ideal gases:

- The volume of the individual gas particle (atom or molecule) must not be very important.
- The forces among the particles must not be very important.

If these factors were important, the pressure of the gas would depend on the nature of the individual particles. For example, an argon atom is much larger than a helium atom. Yet 1.75 mol of argon gas in a 5.0-L container at 20 °C exerts the same pressure as 1.75 mol of helium gas in a 5.0-L container at 20 °C.

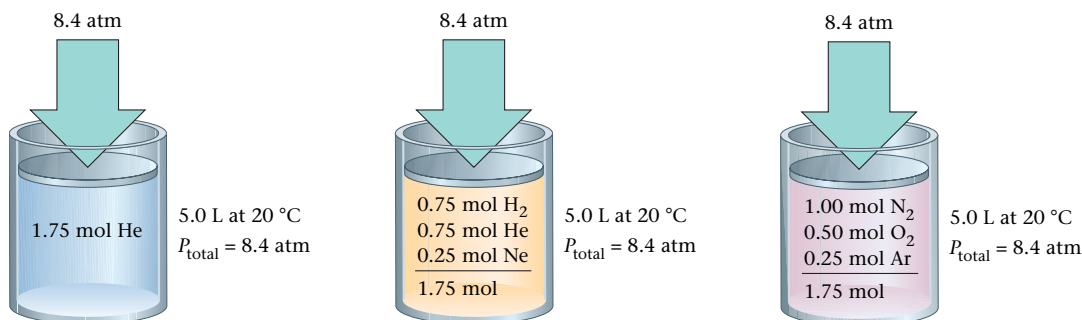


Figure 13.7

The total pressure of a mixture of gases depends on the number of moles of gas particles (atoms or molecules) present, not on the identities of the particles. Note that these three samples show the same total pressure because each contains 1.75 mol of gas. The detailed nature of the mixture is unimportant.

The same idea applies to the forces among the particles. Although the forces among gas particles depend on the nature of the particles, this seems to have little influence on the behavior of an ideal gas. We will see that these observations strongly influence the model that we will construct to explain ideal gas behavior.

EXAMPLE 13.12

Using Dalton's Law of Partial Pressures, I

Mixtures of helium and oxygen are used in the "air" tanks of underwater divers for deep dives. For a particular dive, 12 L of O₂ at 25 °C and 1.0 atm and 46 L of He at 25 °C and 1.0 atm were pumped into a 5.0-L tank. Calculate the partial pressure of each gas and the total pressure in the tank at 25 °C.

Solution

Where do we want to go?

$$P_{\text{O}_2} = ? \text{ atm} \quad P_{\text{He}} = ? \text{ atm} \quad P_{\text{total}} = ? \text{ atm}$$

MATH

$$\begin{aligned} PV &= nRT \\ \frac{PV}{RT} &= \frac{nRT}{RT} \\ \frac{PV}{RT} &= n \end{aligned}$$

What do we know?

oxygen	helium
$V_{\text{O}_2} = 12 \text{ L}$	$V_{\text{He}} = 46 \text{ L}$
$P_{\text{O}_2} = ? \text{ atm}$	$P_{\text{He}} = ? \text{ atm}$
$T_{\text{O}_2} = 25 \text{ }^\circ\text{C} = 298 \text{ K}$	$T_{\text{He}} = 25 \text{ }^\circ\text{C} = 298 \text{ K}$

- Volume of tank = 5.0 L
- Ideal gas law $PV = nRT$
- $P_{\text{total}} = P_{\text{O}_2} + P_{\text{He}}$

How do we get there?

Because the partial pressure of each gas depends on the moles of that gas present, we must first calculate the number of moles of each gas by using the ideal gas law in the form

$$n = \frac{PV}{RT}$$

$$\text{Moles of O}_2 = n_{\text{O}_2} = \frac{(1.0 \text{ atm})(12 \text{ L})}{(0.08206 \text{ L}\cdot\text{atm}/\text{K}\cdot\text{mol})(298 \text{ K})} = 0.49 \text{ mol}$$

$$\text{Moles of He} = n_{\text{He}} = \frac{(1.0 \text{ atm})(46 \text{ L})}{(0.08206 \text{ L}\cdot\text{atm}/\text{K}\cdot\text{mol})(298 \text{ K})} = 1.9 \text{ mol}$$

The tank containing the mixture has a volume of 5.0 L, and the temperature is 25 °C (298 K). We can use these data and the ideal gas law to calculate the partial pressure of each gas.

$$P = \frac{nRT}{V}$$

$$P_{\text{O}_2} = \frac{(0.49 \text{ mol})(0.08206 \text{ L}\cdot\text{atm}/\text{K}\cdot\text{mol})(298 \text{ K})}{5.0 \text{ L}} = 2.4 \text{ atm}$$

$$P_{\text{He}} = \frac{(1.9 \text{ mol})(0.08206 \text{ L}\cdot\text{atm}/\text{K}\cdot\text{mol})(298 \text{ K})}{5.0 \text{ L}} = 9.3 \text{ atm}$$

The total pressure is the sum of the partial pressures.

$$P_{\text{total}} = P_{\text{O}_2} + P_{\text{He}} = 2.4 \text{ atm} + 9.3 \text{ atm} = 11.7 \text{ atm}$$

Practice Problem • Exercise 13.12

A 2.0-L flask contains a mixture of nitrogen gas and oxygen gas at 25 °C. The total pressure of the gaseous mixture is 0.91 atm, and the mixture is known to contain 0.050 mol N₂. Calculate the partial pressure of oxygen and the moles of oxygen present.

A mixture of gases always occurs when a gas is collected by displacement of water. For example, **Figure 13.8** shows the collection of the oxygen gas that was produced by the decomposition of solid potassium chlorate. The gas is collected by bubbling it into a bottle initially filled with water. Thus the gas in the bottle is really a mixture of water vapor and oxygen. (Water vapor is present because molecules of water escape from the surface of the liquid and collect as a gas in the space above the liquid.) Therefore, the total pressure exerted by this mixture is the sum of the partial pressure of the gas being collected and the partial pressure of the water vapor. The partial pressure of the water vapor is called the vapor pressure of water. Because water molecules are more likely to escape from hot water than from cold water, the *vapor pressure* of water increases with temperature. This is shown by the values of water vapor pressure at various temperatures shown in **Table 13.2**.

Table 13.2

The Vapor Pressure of Water as a Function of Temperature

T (°C)	P (torr)
0.0	4.579
10.0	9.209
20.0	17.535
25.0	23.756
30.0	31.824
40.0	55.324
60.0	149.4
70.0	233.7
90.0	525.8

Active Reading Question

Why must we worry about the partial pressure of water vapor when collecting a gas by the displacement of water?

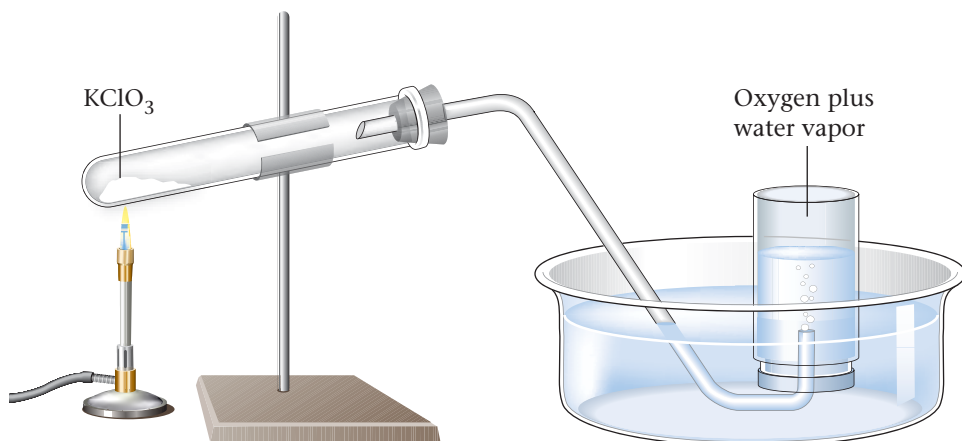


Figure 13.8

The production of oxygen by thermal decomposition of KClO₃

EXAMPLE 13.13

Using Dalton's Law of Partial Pressures, II

A sample of solid potassium chlorate, KClO_3 , was heated in a test tube (see Figure 13.8) and decomposed according to the reaction



The oxygen produced was collected by displacement of water at 22°C . The resulting mixture of O_2 and H_2O vapor had a total pressure of 754 torr and a volume of 0.650 L. Calculate the partial pressure of O_2 in the gas collected and the number of moles of O_2 present. The vapor pressure of water at 22°C is 21 torr.

Solution

Where do we want to go?

$$P_{\text{O}_2} = ? \text{ atm}$$

$$n_{\text{O}_2} = ? \text{ mol}$$

What do we know?

oxygen	water vapor
$n_{\text{O}_2} = ? \text{ mol}$	
$P_{\text{O}_2} = ? \text{ atm}$	$P_{\text{H}_2\text{O}} = 21 \text{ torr}$
$T_{\text{O}_2} = 22^\circ\text{C} = 295 \text{ K}$	$T_{\text{H}_2\text{O}} = 22^\circ\text{C} = 295 \text{ K}$

- $P_{\text{total}} = 754 \text{ torr} = P_{\text{O}_2} + P_{\text{H}_2\text{O}}$
- Ideal Gas Law $PV = nRT$
- $V_{\text{total}} = 0.650 \text{ L} = V_{\text{O}_2} + V_{\text{H}_2\text{O}}$

How do we get there?

We can find the partial pressure of O_2 from Dalton's law of partial pressures:

$$P_{\text{total}} = P_{\text{O}_2} + P_{\text{H}_2\text{O}} = P_{\text{O}_2} + 21 \text{ torr} = 754 \text{ torr}$$

or

$$P_{\text{O}_2} + 21 \text{ torr} = 754 \text{ torr}$$

We can solve for P_{O_2} by subtracting 21 torr from both sides of the equation.

$$P_{\text{O}_2} = 754 \text{ torr} - 21 \text{ torr} = 733 \text{ torr}$$

Next we solve the ideal gas law for the number of moles of O_2 .

$$n_{\text{O}_2} = \frac{P_{\text{O}_2} V}{RT}$$

In this case, $P_{\text{O}_2} = 733 \text{ torr}$. We change the pressure to atmospheres as follows:

$$\frac{733 \text{ torr}}{760 \text{ torr/atm}} = 0.964 \text{ atm}$$

Then,

$$V = 0.650 \text{ L}$$

MATH

$$\begin{aligned} PV &= nRT \\ \frac{PV}{RT} &= \frac{nRT}{RT} \\ \frac{PV}{RT} &= n \end{aligned}$$

$$T = 22\text{ }^{\circ}\text{C} = 22 + 273 = 295\text{ K}$$

$$R = 0.08206\text{ L atm/K mol}$$

so

$$n_{\text{O}_2} = \frac{(0.964\text{ atm})(0.650\text{ L})}{(0.08206\text{ L atm/K mol})(295\text{ K})} = 2.59 \times 10^{-2}\text{ mol}$$

Practice Problem • Exercise 13.13

Consider a sample of hydrogen gas collected over water at 25 °C where the vapor pressure of water is 24 torr. The volume occupied by the gaseous mixture is 0.500 L, and the total pressure is 0.950 atm. Calculate the partial pressure of H₂ and the number of moles of H₂ present.

C. Gas Stoichiometry

We have seen in this chapter just how useful the ideal gas equation is. For example, if we know the pressure, volume, and temperature for a given sample of gas, we can calculate the number of moles present: $n = PV/RT$. This fact makes it possible to do stoichiometric calculations for reactions involving gases.

EXAMPLE 13.14

Gas Stoichiometry: Calculating Volume

Calculate the volume of oxygen gas produced at 1.00 atm and 25 °C by the complete decomposition of 10.5 g of potassium chlorate. The balanced equation for the reaction is



Solution

Where do we want to go?

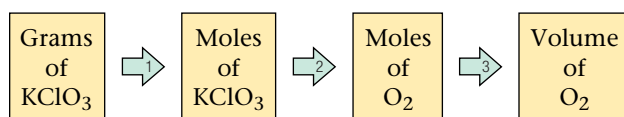
Volume of O₂ produced = ? L

What do we know?

- $2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g)$
- $P = 1.00\text{ atm}$
- $T = 25\text{ }^{\circ}\text{C} = 298\text{ K}$
- 10.5 g KClO₃
- Ideal gas law $PV = nRT$

How do we get there?

We'll summarize the steps required to do this problem in the following schematic:



MATH



$$\frac{10.5}{122.6} = 0.085644$$

0.085644 Rounds to 0.0856

$$0.0856 = 8.56 \times 10^{-2}$$

Step 1 To find the moles of KClO_3 in 10.5 g, we use the molar mass of KClO_3 (122.6 g).

$$10.5 \text{ g } \text{KClO}_3 \times \frac{1 \text{ mol } \text{KClO}_3}{122.6 \text{ g } \text{KClO}_3} = 8.56 \times 10^{-2} \text{ mol } \text{KClO}_3$$

Step 2 To find the moles of O_2 produced, we use the mole ratio of O_2 to KClO_3 derived from the balanced equation.

$$8.56 \times 10^{-2} \text{ mol } \text{KClO}_3 \times \frac{3 \text{ mol } \text{O}_2}{2 \text{ mol } \text{KClO}_3} = 1.28 \times 10^{-1} \text{ mol } \text{O}_2$$

Step 3 To find the volume of oxygen produced, we use the ideal gas law $PV = nRT$, where

$$P = 1.00 \text{ atm}$$

$$V = ?$$

$$n = 1.28 \times 10^{-1} \text{ mol, the moles of } \text{O}_2 \text{ we calculated}$$

$$R = 0.08206 \text{ L atm/K mol}$$

$$T = 25 \text{ }^\circ\text{C} = 25 + 273 = 298 \text{ K}$$

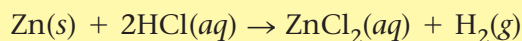
Solving the ideal gas law for V gives

$$V = \frac{nRT}{P} = \frac{(1.28 \times 10^{-1} \text{ mol})(0.08206 \frac{\text{L atm}}{\text{K mol}})(298 \text{ K})}{1.00 \text{ atm}} = 3.13 \text{ L}$$

Thus 3.13 L of O_2 will be produced.

Practice Problem • Exercise 13.14

Calculate the volume of hydrogen produced at 1.50 atm and 19 °C by the reaction of 26.5 g of zinc with excess hydrochloric acid according to the balanced equation



Molar Volume

It is useful to define the volume occupied by 1 mol of a gas under certain specified conditions. For 1 mol of an ideal gas at 0 °C (273 K) and 1 atm, the volume of the gas given by the ideal gas law is

$$V = \frac{nRT}{P} = \frac{(1.00 \text{ mol})(0.08206 \text{ L atm/K mol})(273 \text{ K})}{1.00 \text{ atm}} = 22.4 \text{ L}$$

This volume of 22.4 L is called the **molar volume** of an ideal gas.

The conditions 0 °C and 1 atm are called **standard temperature and pressure (STP)**. Properties of gases are often given under these conditions.

The molar volume of an ideal gas is 22.4 L at STP.
22.4 L contains 1 mol of an ideal gas at STP.

Molar volume

The volume of 1 mole of an ideal gas is equal to 22.4 L at 0 °C and 1 atm.

Standard temperature and pressure (STP)

0 °C and 1 atm

EXAMPLE 13.15**Gas Stoichiometry: Calculations Involving Gases at STP**

A sample of nitrogen gas has a volume of 1.75 L at STP. How many moles of N_2 are present?

Solution**Where do we want to go?**

Amount of N_2 produced = ? mol

What do we know?

- $V = 1.75$ L at STP
- 1.00 mol = 22.4 L (STP)

How do we get there?

We could solve this problem by using the ideal gas equation, but we can take a shortcut by using the molar volume of an ideal gas at STP. Because 1 mol of an ideal gas at STP has a volume of 22.4 L, a 1.75-L sample of N_2 at STP contains considerably less than 1 mol. We can find how many moles by using the equivalence statement

$$1.000 \text{ mol} = 22.4 \text{ L (STP)}$$

which leads to the conversion factor we need:

$$1.75 \cancel{\text{ L } N_2} \times \frac{1.000 \text{ mol } N_2}{22.4 \cancel{\text{ L } N_2}} = 7.81 \times 10^{-2} \text{ mol } N_2$$

Practice Problem • Exercise 13.15

Ammonia is commonly used as a fertilizer to provide a source of nitrogen for plants. A sample of $NH_3(g)$ occupies 5.00 L at 25 °C and 15.0 atm. What volume will this sample occupy at STP?

CHEMISTRY IN YOUR WORLD**Consumer Connection****The Chemistry of Air Bags**

Most experts agree that air bags represent a very important advance in automobile safety. These bags, which are stored in the auto's steering wheel or dash, are designed to inflate rapidly (within about 40 ms) in the event of a crash, cushioning the front seat occupants against impact. The bags then deflate immediately to allow vision and movement after the crash. Air bags are activated when a severe deceleration (an impact) causes a steel ball to compress a spring and electrically ignite a detonator cap, which, in turn, causes sodium azide (NaN_3) to decompose explosively, forming sodium and nitrogen gas:



This system works very well and requires a relatively small amount of sodium azide (100 g yields 56 L $N_2(g)$ at 25 °C and 1.0 atm).

When a vehicle containing air bags reaches the end of its useful life, the sodium azide present in the activators must be given proper disposal. Sodium azide, besides being explosive is toxic.

The air bag is an application of chemistry that has saved thousands of lives.



Inflated air bags

EXAMPLE 13.16

Gas Stoichiometry: Reactions Involving Gases at STP

Quicklime, CaO, is produced by heating calcium carbonate, CaCO₃. Calculate the volume of CO₂ produced at STP from the decomposition of 152 g of CaCO₃ according to the reaction



Solution

Where do we want to go?

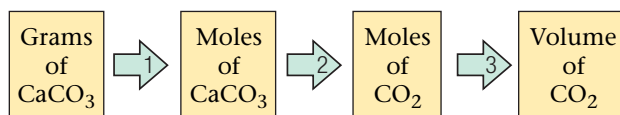
Volume of CO₂ produced = ? L

What do we know?

- CaCO₃(s) + CaO(s) → CO₂(g)
- STP conditions
- 1.00 mol = 22.4 L (STP)
- 152 g CaCO₃

How do we get there?

The strategy for solving this problem is summarized by the following schematic:



Step 1 Using the molar mass of CaCO₃ (100.1 g), we calculate the number of moles of CaCO₃.

$$152 \text{ g CaCO}_3 \times \frac{1 \text{ mol CaCO}_3}{100.1 \text{ g CaCO}_3} = 1.52 \text{ mol CaCO}_3$$

Step 2 Each mole of CaCO₃ produces 1 mol of CO₂, so 1.52 mol of CO₂ will be formed.

Step 3 We can convert the moles of CO₂ to volume by using the molar volume of an ideal gas, because the conditions are STP.

$$1.52 \text{ mol CO}_2 \times \frac{22.4 \text{ L CO}_2}{1 \text{ mol CO}_2} = 34.1 \text{ L CO}_2$$

Thus the decomposition of 152 g of CaCO₃ produces 34.1 L of CO₂ at STP.

Information

Remember that the molar volume of an ideal gas is 22.4 L at STP.

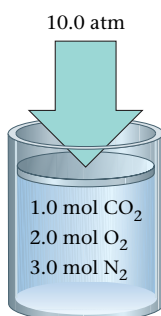
Note that the final step in Example 13.16 involves calculating the volume of gas from the number of moles. Because the conditions were specified as STP, we were able to use the molar volume of a gas at STP. If the conditions of a problem are different from STP, we must use the ideal gas law to compute the volume.



SECTION 13.2

REVIEW QUESTIONS

- When solving gas problems the ideal gas law can be used even when some of the properties are constant. Explain how to do this.
- At what temperature will 6.21 g of oxygen gas exert a pressure of 5.00 atm in a 10.0-L container?
- The pressure of a gas is affected by the number of particles and not affected by the kind of gas in the container. What does this fact tell us about ideal gases?
- A 1.50 L sample of neon gas at 1.10 atm and 25 °C is heated to 45 °C. The neon gas is then subjected to a pressure of 1.50 atm. Determine the new volume of the neon gas.
- Calculate the partial pressure (in torr) of each gas in the mixture.
- In Figure 13.8 oxygen gas is being collected over water. Assume this experiment is being done at 25 °C. Write an equation for it, and find the pressure of the oxygen gas (assume the pressure in the bottle is P_T).
- Magnesium metal reacts with hydrochloric acid to produce hydrogen gas and a solution of magnesium chloride. Write and balance the chemical equation for this reaction, and determine the volume of the hydrogen gas generated (at 1.00 atm, 25 °C) by reacting 10.0 g Mg with an excess of hydrochloric acid.
- When water is subjected to an electric current, it decomposes to hydrogen and oxygen gas. If 1.00 g of water is decomposed at 1.00 atm and 25 °C, what volume of oxygen gas is collected?



Using a Model to Describe Gases

Key Term

- Kinetic molecular theory

Objectives

- To understand the relationship between laws and models (theories)
- To understand the postulates of the kinetic molecular theory
- To understand temperature
- To learn how the kinetic molecular theory explains the gas laws
- To describe the properties of real gases

A. Laws and Models: A Review

In this chapter we have considered several properties of gases and have seen how the relationships among these properties can be expressed by various laws written in the form of mathematical equations. The most useful of these is the ideal gas equation, which relates all the important gas properties. However, under certain conditions gases do not obey the ideal gas equation. For example, at high pressures and/or low temperatures, the properties of gases deviate significantly from the predictions of the ideal gas equation. On the other hand, as the pressure is lowered and/or the temperature is increased, almost all gases show close agreement with the ideal gas equation. This means that an ideal gas is really a hypothetical substance. At low pressures and/or high temperatures, real gases *approach* the behavior expected for an ideal gas.

Building a Model for Gases At this point we want to build a model, a theory, to explain *why* a gas behaves as it does. We want to answer the question, *What are the characteristics of the individual gas particles that cause a gas to behave as it does?*

Let's Review

Scientific Method

- A law is a generalization of observed behavior.
- Laws are useful → We can predict behavior of similar systems.

However, laws do not tell us *why* nature behaves the way it does. Scientists try to answer this question by constructing theories (building models). The models in chemistry are speculations about how individual atoms or molecules (microscopic particles) cause the behavior of macroscopic systems (collections of atoms and molecules in large enough numbers so that we can observe them).

A model is considered successful if it explains known behavior and predicts correctly the results of future experiments.

A model can never be proved absolutely true. By its very nature any model is an approximation and is destined to be modified.

Models range from the simple, to predict approximate behavior, to the extraordinarily complex, to account precisely for observed behavior. In this text, we use relatively simple models that fit most experimental results.

Active Reading Question

Explain the difference between a law and a theory.

B. The Kinetic Molecular Theory of Gases

A relatively simple model that attempts to explain the behavior of an ideal gas is the **kinetic molecular theory**. This model is based on speculations about the behavior of the individual particles (atoms or molecules) in a gas. The assumptions of the kinetic molecular theory (KMT) can be stated as follows:

Assumptions of the Kinetic Molecular Theory of Gases

1. Gases consist of tiny particles (atoms or molecules).
2. These particles are so small, compared with the distances between them, that the volume (size) of the individual particles can be assumed to be negligible (zero).
3. The particles are in constant random motion, colliding with the walls of the container. These collisions with the walls cause the pressure exerted by the gas.
4. The particles are assumed not to attract or to repel each other.
5. The average kinetic energy of the gas particles is directly proportional to the Kelvin temperature of the gas.

The kinetic energy referred to in assumption 5 is the energy associated with the motion of a particle.

Kinetic energy (KE) is given by $KE = \frac{1}{2}mv^2$, where m is the mass of the particle and v is the velocity (speed) of the particle.

The greater the mass or velocity of a particle, the greater its kinetic energy. This means that if a gas is heated to higher temperatures, the average speed of the particles increases; therefore, their kinetic energy increases.

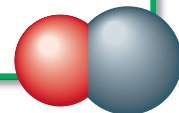
CELEBRITY CHEMICAL

Carbon Monoxide (CO)

Carbon monoxide is a colorless, odorless gas originally discovered in the United States by Joseph Priestley in 1799. It is formed in the combustion of fossil fuels (carbon-based molecules formed by the decay of ancient life forms) when the oxygen supply is limited.

Carbon monoxide is quite toxic, because it binds to hemoglobin 200 times more strongly than oxygen does. Thus, when carbon monoxide is present in inhaled air, it preferentially binds to the hemoglobin and excludes the oxygen needed

for human survival. This gas is especially dangerous because it has no taste or odor. The most common type of carbon monoxide poisonings occur in homes where the flow of oxygen to the furnace has become restricted, often due to squirrels or birds building nests in intake air flues. Today carbon monoxide detectors are available that automatically sound an alarm when the carbon monoxide levels become unsafe.



CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Signs of Pollution

About 6% of the cars on the road produce 50% of auto-based air pollution. However, finding these offenders can be difficult. Scientists at the University of Denver in Colorado have developed a system that can measure the pollution produced by individual cars as they pass a roadside detector. This information is then transmitted to a billboard display. When the emissions are within acceptable limits, a smiling car is displayed. When the emissions are too high, a frowning car is displayed with the message that the problem is “costing you money.” This “smart sign” operates reliably even with traffic flows exceeding 1,000 vehicles per hour. The system has given

millions of readings at a cost of about 2 cents each, approximately 2% of the passing vehicles see the “frowning car.”



Although real gases do not conform exactly to the five assumptions listed, we will see next that these assumptions do indeed explain *ideal* gas behavior—behavior shown by real gases at high temperatures and/or low pressures.

C. The Implications of the Kinetic Molecular Theory

Now we will show how the kinetic molecular theory explains some of the observed properties of gases.

The Meaning of Temperature

The temperature of a substance reflects the vigor of the motions of the components that make up the substance. Thus the temperature of a gas reflects how rapidly, on average, its individual particles are moving. At high temperatures the particles move very fast and hit the walls of the container frequently, and at low temperatures the particles' motions are slower and they collide with the walls of the container much less often. As assumption 5 of the KMT states, the Kelvin temperature of a gas is directly proportional to the *average kinetic energy* of the gas particles.

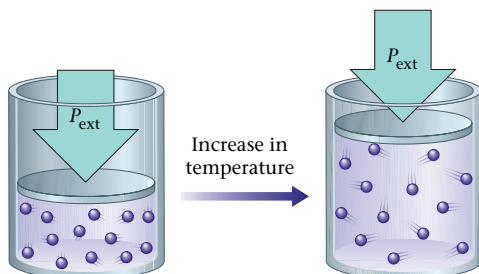
The Relationship Between Pressure and Temperature

To see how the meaning of temperature given above helps to explain gas behavior, picture a gas in a rigid container. As the gas is heated to a higher temperature, the particles move faster, hitting the walls more often. And, of course, the impacts become more forceful as the particles move faster. If the pressure is due to collisions with the walls, the gas pressure should increase as temperature is increased.

Is this what we observe when we measure the pressure of a gas as it is heated? Yes. A given sample of gas in a rigid container (if the volume is not changed) shows an increase in pressure as its temperature is increased.

The Relationship Between Volume and Temperature

Now picture the gas in a container with a movable piston. The gas pressure P_{gas} is balanced by an external pressure P_{ext} . What happens when we heat the gas to a higher temperature?



As the temperature increases, the particles move faster, causing the gas pressure to increase. As soon as the gas pressure P_{gas} becomes greater than P_{ext} (the pressure holding the piston), the piston moves up until $P_{\text{gas}} = P_{\text{ext}}$. Therefore, the KMT predicts that the volume of the gas will increase as we raise its temperature at a constant pressure. This agrees with experimental observations (as summarized by Charles's law).

EXAMPLE 13.17

Using the Kinetic Molecular Theory to Explain Gas Law Observations

Use the KMT to predict what will happen to the pressure of a gas when its volume is decreased (n and T constant). Does this prediction agree with the experimental observations?

Solution

When we decrease the gas's volume (make the container smaller), the particles hit the walls more often because they do not have to travel so far between the walls. This would suggest an increase in pressure. This prediction on the basis of the model is in agreement with experimental observations of gas behavior (as summarized by Boyle's law).

D. Real Gases

CRITICAL THINKING ?

You have learned that no gas behaves ideally.

What if all gases behaved ideally under all conditions? How would the world be different?

So far in our discussion of gases, we have assumed that we are dealing with an ideal gas—one that exactly obeys the equation $PV = nRT$. Of course, there is no such thing as an ideal gas. An ideal gas is a hypothetical substance consisting of particles with zero volumes and no attractions for one another. All is not lost, however, because real gases behave very much like ideal gases under many conditions. For example, in a sample of helium gas at 25 °C and 1 atm pressure, the He atoms are so far apart that the tiny volume of each atom has no importance. Also, because the He atoms are moving so rapidly and have very little attraction for one another, the ideal gas assumption that there are no attractions is virtually true. Thus a sample of helium at 25 °C and 1 atm pressure very closely obeys the ideal gas law.

As real gases are compressed into smaller and smaller volumes (see **Figure 13.9**), the particles of the gas begin to occupy a significant fraction of the available volume. That is, in a very small container the space taken up by the particles becomes important. Also, as the volume of the container gets smaller, the particles move much closer together and are more likely to attract one another. Thus, in a highly compressed state (small V , high P), the facts that real gas molecules take up space and have attractions for one another become important. Under these conditions a real gas does not obey the equation $PV = nRT$ very well. In other words, under conditions of high pressure (small volume), real gases act differently from the ideal gas behavior.

Because we typically deal with gases that have pressures near 1 atm, however, we can safely assume ideal gas behavior in our calculations.

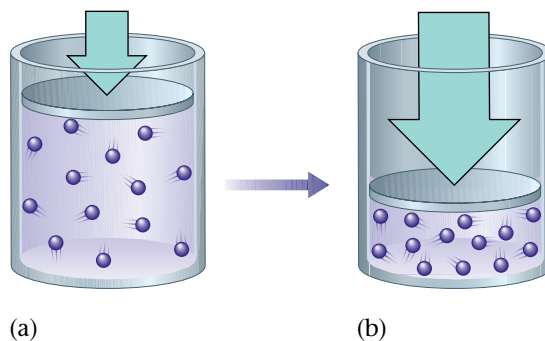


Figure 13.9
A gas sample is compressed (the volume is decreased).

Active Reading Question

What are the two main assumptions of the kinetic molecular theory that are true about ideal gases but not true about real gases?

SECTION 13.3 REVIEW QUESTIONS

- 1 Explain the difference between a *law* and a *theory*.
- 2 When do real gases behave as ideal gases?
- 3 What are the main ideas of the kinetic molecular theory of gases?
- 4 Use the kinetic molecular theory to draw a molecular-level picture of a gas. Briefly describe the motion of the gas particles.
- 5 Use the kinetic molecular theory to explain Avogadro's law.



Key Terms

13.1 Barometer

Torr (mm Hg)
 Standard atmosphere
 Pascal
 Boyle's law
 Absolute zero
 Charles's law
 Avogadro's law

13.2 Universal gas constant

Ideal gas law
 Ideal gas
 Combined gas law
 Partial pressure
 Dalton's law of partial pressures
 Molar volume
 Standard temperature and pressure (STP)

13.3 Kinetic molecular theory

Key Ideas

13.1 Describing the Properties of Gases

- The common units for pressure are mm Hg (torr), atmosphere (atm), and pascal (Pa). The SI unit is the pascal.
- Boyle's law states that the volume of a given amount of gas at constant temperature varies inversely to its pressure. $PV = k$
- Charles's law states that the volume of a given amount of an ideal gas at constant pressure varies directly with its temperature (in kelvins). $V = bT$
 - At absolute zero (0 K; $-273\text{ }^{\circ}\text{C}$) the volume of an ideal gas extrapolates to zero.
- Avogadro's law states for an ideal gas at constant temperature and pressure, the volume varies directly with the number of moles of gas (n). $V = an$

13.2 Using Gas Laws to Solve Problems

- The ideal gas law describes the relationship among P , V , n , and T for an ideal gas. $PV = nRT$

$$R = \left(0.08206 \frac{\text{L atm}}{\text{mol K}}\right)$$

A gas that obeys this law exactly is called an ideal gas.

- From the ideal gas law we can obtain the combined gas law which applies when n is constant.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

- Dalton's law of partial pressures states the total pressure of a mixture of gases is equal to the sum of the individual (partial) pressures of the gases. $P_{\text{total}} = P_1 + P_2 + P_3 + \dots$
- Standard temperature and pressure (STP) is defined as $P = 1\text{ atm}$ and $T = 273\text{ K}$ ($0\text{ }^{\circ}\text{C}$).
- The volume of one mole of an ideal gas (the molar volume) is 22.4 L at STP.

13.3 Using a Model to Describe Gases

- The kinetic molecular theory (KMT) is a model based on the properties of individual gas components that explains the relationship of P , V , T , and n for an ideal gas.
- A law is a summary of experimental observation.
- A model (theory) is an attempt to explain observed behavior.
- The temperature of an ideal gas reflects the average kinetic energy of the gas particles.
- The pressure of a gas increases as its temperature increases because the gas particles speed up.
- The volume of a gas must increase because the gas particles speed up as a gas is heated to a higher temperature.

All exercises with **blue numbers** have answers in the back of this book.

13.1 Describing the Properties of Gases

A. Pressure

- How are the three states of matter similar, and how do they differ?
- What is meant by “the pressure of the atmosphere”? What causes this pressure?
- Describe a simple mercury barometer. How is such a barometer used to measure the pressure of the atmosphere?
- Convert the following pressures into *atmospheres*.
 - 105.2 kPa
 - 75.2 cm Hg
 - 752 mm Hg
 - 767 torr
- Convert the following pressures into units of *mm Hg*.
 - 0.9975 atm
 - 225,400 Pa
 - 99.7 kPa
 - 1.078 atm
- Convert the following pressures into pascals.
 - 774 torr
 - 0.965 atm
 - 112.5 kPa
 - 801 mm Hg

B. Pressure and Volume: Boyle’s Law

- For each of the following sets of pressure and volume data, calculate the missing quantity. Assume that the temperature and the amount of gas remain constant.
 - $V = 53.2 \text{ mL}$ at 785 mm Hg;
 $V = ?$ at 700. mm Hg
 - $V = 2.25 \text{ L}$ at 1.67 atm;
 $V = 2.00 \text{ L}$ at ? atm
 - $V = 5.62 \text{ L}$ at 695 mm Hg;
 $V = ?$ at 1.51 atm

- For each of the following sets of pressure and volume data, calculate the missing quantity. Assume that the temperature and the amount of gas remain the same.
 - $V = 291 \text{ mL}$ at 1.07 atm;
 $V = ?$ at 2.14 atm
 - $V = 1.25 \text{ L}$ at 755 mm Hg;
 $V = ?$ at 3.51 atm
 - $V = 2.71 \text{ L}$ at 101.4 kPa;
 $V = 3.00 \text{ L}$ at ? mm Hg
- If the pressure exerted on the gas in a weather balloon decreases from 1.01 atm to 0.562 atm as it rises, by what factor will the volume of the gas in the balloon increase as it rises?
- What pressure (in atmospheres) is required to compress 1.00 L of gas at 760. mm Hg pressure to a volume of 50.0 mL?

C. Volume and Temperature: Charles’s Law

- What is meant by the *absolute zero* of temperature?
- How can Charles’s law be used to determine absolute zero?
- If a 45.0-mL sample of gas at 26.5 °C is heated to 55.2 °C, what is the new volume of the gas sample (at constant pressure)?
- For each of the following sets of volume and temperature data, calculate the missing quantity. Assume that the pressure and the mass of gas remain constant.
 - $V = 25.0 \text{ L}$ at 0 °C; $V = 50.0 \text{ L}$ at ? °C
 - $V = 247 \text{ mL}$ at 25 °C;
 $V = 255 \text{ mL}$ at ? °C
 - $V = 1.00 \text{ mL}$ at 2272 °C; $V = ?$ at 25 °C
- For each of the following sets of volume and temperature data, calculate the missing quantity. Assume that the pressure and the mass of gas remain constant.
 - $V = 2.01 \times 10^2 \text{ L}$ at 1150 °C;
 $V = 5.00 \text{ L}$ at ? °C
 - $V = 44.2 \text{ mL}$ at 298 K; $V = ?$ at 0 K
 - $V = 44.2 \text{ mL}$ at 298 K; $V = ?$ at 0 °C
- If 5.00 L of an ideal gas is cooled from 24 °C to -272 °C, what will the volume of the gas become?

- 17.** A sample of neon gas occupies 266 mL at 25.2 °C. At what temperature would the volume of this sample of neon be reduced to half its initial size (at constant pressure)? At what temperature would the volume of this sample of neon be doubled (at constant pressure)?
- 18.** The label on an aerosol spray can contains a warning that the can should not be heated to over 130 °F because of the danger of explosion due to the pressure increase as it is heated. Calculate the potential volume of the gas contained in a 500.-mL aerosol can when it is heated from 25 °C to 54 °C (approximately 130 °F), assuming a constant pressure.

D. Volume and Moles: Avogadro's Law

- 19.** If 0.214 mol of argon gas occupies a volume of 652 mL at a particular temperature and pressure, what volume would 0.375 mol of argon occupy under the same conditions?
- 20.** If 46.2 g of oxygen gas occupies a volume of 100. L at a particular temperature and pressure, what volume will 5.00 g of oxygen gas occupy under the same conditions?

13.2 Using Gas Laws to Solve Problems

A. The Ideal Gas Law

- 21.** Under what conditions do *real* gases behave most ideally?
- 22.** Show how Boyle's law can be derived from the ideal gas law.
- 23.** Given each of the following sets of values for an ideal gas, calculate the unknown quantity.
- $P = 782 \text{ mm Hg}$; $V = ?$; $n = 0.210 \text{ mol}$;
 $T = 27 \text{ }^\circ\text{C}$
 - $P = ? \text{ mm Hg}$; $V = 644 \text{ mL}$;
 $n = 0.0921 \text{ mol}$; $T = 303 \text{ K}$
 - $P = 745 \text{ mm Hg}$; $V = 11.2 \text{ L}$;
 $n = 0.401 \text{ mol}$; $T = ? \text{ K}$
- 24.** Given each of the following sets of values for an ideal gas, calculate the unknown quantity.
- $P = 1.01 \text{ atm}$; $V = ?$; $n = 0.00831 \text{ mol}$;
 $T = 25 \text{ }^\circ\text{C}$
 - $P = ? \text{ atm}$; $V = 602 \text{ mL}$;
 $n = 8.01 \times 10^{-3} \text{ mol}$; $T = 310. \text{ K}$
 - $P = 0.998 \text{ atm}$; $V = 629 \text{ mL}$; $n = ? \text{ mol}$;
 $T = 35 \text{ }^\circ\text{C}$

- 25.** What volume does 4.24 g of nitrogen gas occupy at 58.2 °C and 2.04 atm?
- 26.** Suppose two 200.0-L tanks are to be filled separately with the gases helium and hydrogen. What mass of each gas is needed to produce a pressure of 135 atm in its respective tank at 24 °C?
- 27.** At what temperature does 16.3 g of nitrogen gas have a pressure of 1.25 atm in a 25.0-L tank?
- 28.** Determine the pressure in a 125-L tank containing 56.2 kg of oxygen gas at 21 °C.
- 29.** What will be the new volume if 125 mL of He gas at 100. °C and 0.981 atm is cooled to 25 °C and the pressure is increased to 1.15 atm?
- 30.** What will the volume of the sample become if 459 mL of an ideal gas at 27 °C and 1.05 atm is cooled to 15 °C and 0.997 atm?

B. Dalton's Law of Partial Pressures

- 31.** Explain why the measured properties of a mixture of gases depend only on the total number of moles of particles, not on the identities of the individual gas particles. How is this observation summarized as a law?
- 32.** We often collect small samples of gases in the laboratory by bubbling the gas into a bottle or flask containing water. Explain why the gas becomes saturated with water vapor and how we must take the presence of water vapor into account when calculating the properties of the gas sample.
- 33.** If 4.0 g of $\text{O}_2(\text{g})$ and 4.0 g of $\text{He}(\text{g})$ are placed in a 5.0-L vessel at 65 °C, what will be the partial pressure of each gas and the total pressure in the vessel?
- 34.** A tank contains a mixture of 3.0 mol N_2 , 2.0 mol O_2 , and 1.0 mol CO_2 at 25 °C and a total pressure of 10.0 atm. Calculate the partial pressure (in torr) of each gas in the mixture.
- 35.** A sample of oxygen gas is saturated with water vapor at 27 °C. The total pressure of the mixture is 772 torr, and the vapor pressure of water is 26.7 torr at 27 °C. What is the partial pressure of the oxygen gas?

36. A 500.-mL sample of O_2 gas at $24^\circ C$ was prepared by decomposing a 3.0% aqueous solution of hydrogen peroxide, H_2O_2 , in the presence of a small amount of manganese catalyst by the reaction



The oxygen thus prepared was collected by displacement of water. The total pressure of gas collected was 755 mm Hg. What is the partial pressure of O_2 in the mixture? How many moles of O_2 are in the mixture? (The vapor pressure of water at $24^\circ C$ is 23 mm Hg.)

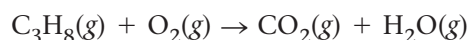
C. Gas Stoichiometry

37. When calcium carbonate is heated strongly, carbon dioxide gas is released



What volume of $CO_2(g)$, measured at STP, is produced if 15.2 g of $CaCO_3(s)$ is heated?

38. Consider the following *unbalanced* chemical equation for the combustion of propane.



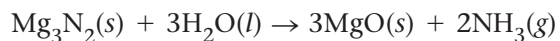
What volume of oxygen gas at $25^\circ C$ and 1.04 atm is needed for the complete combustion of 5.53 g of propane?

39. Ammonia and gaseous hydrogen chloride combine to form ammonium chloride.



If 4.21 L of $NH_3(g)$ at $27^\circ C$ and 1.02 atm is combined with 5.35 L of $HCl(g)$ at $26^\circ C$ and 0.998 atm, what mass of $NH_4Cl(s)$ will be produced? Which gas is the limiting reactant? Which gas is present in excess?

40. If water is added to magnesium nitride, ammonia gas is produced when the mixture is heated.

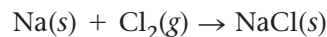


If 10.3 g of magnesium nitride is treated with water, what volume of ammonia gas would be collected at $24^\circ C$ and 752 mm Hg?

41. What volume does a mixture of 14.2 g of He and 21.6 g of H_2 occupy at $28^\circ C$ and 0.985 atm?
42. A sample of hydrogen gas has a volume of 145 mL when measured at $44^\circ C$ and 1.47 atm. What volume would the hydrogen sample occupy at STP?

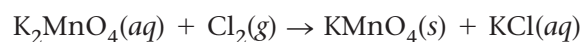
43. A mixture contains 5.00 g *each* of O_2 , N_2 , CO_2 , and Ne gas. Calculate the volume of this mixture at STP. Calculate the partial pressure of each gas in the mixture at STP.

44. Given the following *unbalanced* chemical equation for the combination reaction of sodium metal and chlorine gas



what volume of chlorine gas, measured at STP, is necessary for the complete reaction of 4.81 g of sodium metal?

45. Potassium permanganate, $KMnO_4$, is produced commercially by oxidizing aqueous potassium manganate, K_2MnO_4 , with chlorine gas. The *unbalanced* chemical equation is



What volume of $Cl_2(g)$, measured at STP, is needed to produce 10.0 g of $KMnO_4$?

13.3 Using a Model to Describe Gases

A. Laws and Models: A Review

46. What is a scientific *law*? What is a *theory*? How do these concepts differ? Does a law explain a theory, or does a theory attempt to explain a law?
47. When is a scientific theory considered to be successful? Are all theories successful? Will a theory that has been successful in the past necessarily be successful in the future?

B. The Kinetic Molecular Theory of Gases

48. What do we assume about the volume of the actual molecules themselves in a sample of gas, compared to the bulk volume of the gas overall? Why?
49. How do chemists explain on a molecular basis the fact that gases in containers exert pressure on the walls of the container?
50. Temperature is a measure of the average _____ of the molecules in a sample of gas.

C. The Implications of the Kinetic Molecular Theory

51. How is the phenomenon of temperature explained on the basis of the kinetic molecular theory? What microscopic property of gas molecules is reflected in the temperature measured?
52. Explain, in terms of the kinetic molecular theory, how an increase in the temperature of a gas confined to a rigid container causes an increase in the pressure of the gas.

Critical Thinking

55. A helium tank contains 25.2 L of helium at 8.40 atm pressure. Determine how many 1.50-L balloons at 755 mm Hg can be inflated with the gas in the tank, assuming that the tank will also have to contain He at 755 mm Hg after the balloons are filled (that is, it is not possible to empty the tank completely). The temperature is 25 °C in all cases.
56. As weather balloons rise from the earth's surface, the pressure of the atmosphere becomes less, tending to cause the volume of the balloons to expand. However, the temperature is much lower in the upper atmosphere than at sea level. Would this temperature effect tend to make such a balloon expand or contract? Weather balloons do, in fact, expand as they rise. What does this tell you?
57. Sulfur trioxide, SO_3 , is produced in enormous quantities each year for use in the synthesis of sulfuric acid.
- $$\text{S}(s) + \text{O}_2(g) \rightarrow \text{SO}_2(g)$$
- $$2\text{SO}_2(g) + \text{O}_2(g) \rightarrow 2\text{SO}_3(g)$$
- What volume of $\text{O}_2(g)$ at 350. °C and a pressure of 5.25 atm is needed to completely convert 5.00 g of sulfur to sulfur trioxide?
58. A particular balloon is designed by its manufacturer to be inflated to a volume of no more than 2.5 L. If the balloon is filled with 2.0 L of helium at sea level, is released, and rises to an altitude at which the atmospheric pressure is only 500. mm Hg, will the balloon burst?
59. An expandable vessel contains 729 mL of gas at 22 °C. What volume will the gas sample in the vessel have if it is placed in a boiling water bath (100. °C)?

D. Real Gases

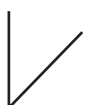



53. What does it mean for a gas to behave nonideally? How do we know that real gases behave nonideally?
54. Why does decreasing the volume of a gas cause nonideal behavior?
60. A weather balloon is filled with 1.0 L of helium at 23 °C and 1.0 atm. What volume does the balloon have when it has risen to a point in the atmosphere where the pressure is 220 torr and the temperature is -31 °C?
61. Consider the following *unbalanced* chemical equation:
- $$\text{Cu}_2\text{S}(s) + \text{O}_2(g) \rightarrow \text{Cu}_2\text{O}(s) + \text{SO}_2(g)$$
- What volume of oxygen gas, measured at 27.5 °C and 0.998 atm, is required to react with 25 g of copper(I) sulfide? What volume of sulfur dioxide gas is produced under the same conditions?
62. A mixture contains 5.0 g of He, 1.0 g of Ar, and 3.5 g of Ne. Calculate the volume of this mixture at STP. Calculate the partial pressure of each gas in the mixture at STP.
63. In an experiment 350.0 mL of hydrogen gas was collected over water at 27 °C and 72 cm Hg. Then the absolute temperature doubled, the total pressure changed to 1.00 atm, and one-third of the gas leaked out of the container. What would the final volume be?
64. A balloon will burst at a volume of 2.00 L. If it is partially filled at a temperature of 20 °C and a pressure of 775 mm Hg to a volume of 0.750 L, what is the temperature at which it will burst?
65. One method for estimating the temperature at the center of the sun is based on the ideal gas law. If the center of the sun is assumed to be a mass of gases with average molar mass of 2.00 g and if the density and pressure are 1.4 g/cm³ and 1.3×10^9 atm, respectively, calculate the temperature.

66. A 0.0712-g sample of X_4H_{10} has a volume of 30.0 cm^3 at 801 mm Hg and $20.0 \text{ }^\circ\text{C}$. What is the element X?
67. A certain German sports car engine has a cylinder volume of 618 cm^3 . The cylinder is full of air at $75 \text{ }^\circ\text{C}$ and 802 mm Hg. How many moles of gas are in the cylinder? If air is 20.0% oxygen by volume, how many molecules of oxygen are present?
68. A certain compound contains 48.6% carbon, 8.18% hydrogen, and 43.2% oxygen by mass. The gas at $150.0 \text{ }^\circ\text{C}$ and 1.00 atm has a density of 2.13 g/L. Suggest two possible Lewis structures for the compound.
69. You buy a deflated balloon that can expand to a maximum volume of 3.00 L. You decide to fill the balloon with hydrogen gas generated by the following reaction:
- $$\text{Zn}(s) + 2\text{HCl}(aq) \rightarrow \text{ZnCl}_2(aq) + \text{H}_2(g)$$
- If you react an excess of zinc with 100.0 mL of 2.00 M HCl and fill the balloon with all of the hydrogen generated (at 1.00 atm and $25 \text{ }^\circ\text{C}$), will the balloon pop? Fully support your answer.

70. Sometimes a law can be written as a statement and a mathematical equation. For example, Boyle's law can be stated as follows: the pressure of a gas is inversely proportional to its volume at constant temperature and number of moles of gas. Boyle's law can also be stated mathematically as $P_1V_1 = P_2V_2$.

Avogadro's law states that the volume of a gas is directly related to the number of moles of gas at constant temperature and pressure.

- Starting with $PV = nRT$, derive Avogadro's law. Hint: See Example 13.10 on page 461 for a derivation of Boyle's law. Use the same ideas to derive Avogadro's law.
- Explain Avogadro's law using the ideas of the kinetic molecular theory. Use complete sentences.
- Use Avogadro's law (as derived in part a) to solve the following problem. Consider two containers, one with 10.0 g of He gas and one with 20.0 g of Ar gas, both at the same pressure and temperature. Which container has the greater volume, and by what factor? Show all work.

- 1 You are holding two helium balloons, a large balloon and a small balloon. How do the pressures of helium compare?
- A The pressure in the large balloon is greater than the pressure in the small balloon. This is best explained by the fact that there are more moles of gas in the large balloon, thus greater pressure.
 - B The pressure in the small balloon is greater than the pressure in the large balloon. This is best explained by the fact that as the volume of a container decreases, the pressure of the gas increases.
 - C The pressure in the large balloon is greater than the pressure in the small balloon. The fact that there is greater pressure in the balloon is why the volume is larger; as the particles push more on the inside, the volume increases.
 - D The pressures are essentially the same and about equal to atmospheric pressure.
- 2 Determine the pressure exerted by 1.80 mol of gas in a 2.92-L container at 32.0 °C.
- A 1.62 atm
 - B 8.57 atm
 - C 15.4 atm
 - D 22.4 atm
- 3 Consider a gas at 1.00 atm in a 5.00-L container at 20.0 °C. What pressure does the gas exert when transferred to a volume of 2.00 L at 43 °C?
- A 0.812 atm
 - B 2.70 atm
 - C 3.14 atm
 - D 5.38 atm
- 4 Consider a mixture of equal masses of helium and hydrogen gases in a steel container. Which gas exerts the greater partial pressure, and by what factor?
- A The helium gas exerts about twice as much pressure as the hydrogen gas.
 - B The helium gas exerts about four times as much pressure as the hydrogen gas.
 - C The hydrogen gas exerts about twice as much pressure as the helium gas.
 - D The helium gas exerts about the same pressure as the hydrogen gas.
- 5 You are holding three balloons each containing 10.0 g of a different gas. The balloon containing which gas is the largest balloon?
- A H₂
 - B He
 - C O₂
 - D Each of the balloons is the same size.
- 6 A 4.37 gram sample of a certain diatomic gas occupies a volume of 3.00 L at 1.00 atm and a temperature of 45 °C. Identify this gas.
- A F₂
 - B N₂
 - C H₂
 - D O₂
- 7 The choices below are all assumptions of the kinetic molecular theory. All of these assumptions are accepted for ideal gases. Which of these does **not** apply for **real** gases?
- A Gases consist of particles in constant motion.
 - B Gas particles have no attraction for one another.
 - C Temperature is a measure of average kinetic energy.
 - D Pressure is due to collisions of gas particles with the walls of the container.
- Use the graphs below to answer questions 8–10.
- A  B  C  D 
- 8 Which of the above graphs best represents the relationship between the pressure and temperature (measured in kelvins) of one mole of an ideal gas?
- 9 Which of the above graphs best represents the relationship between the pressure and volume of one mole of an ideal gas?
- 10 Which of the above graphs best represents the relationship between the volume and temperature (measured in kelvins) of one mole of an ideal gas?

LOOKING AHEAD

14.1 Intermolecular Forces and Phase Changes

- A. Intermolecular Forces
 - Hydrogen Bonding
 - London Dispersion Forces
- B. Water and Its Phase Changes
- C. Energy Requirements for the Changes of State

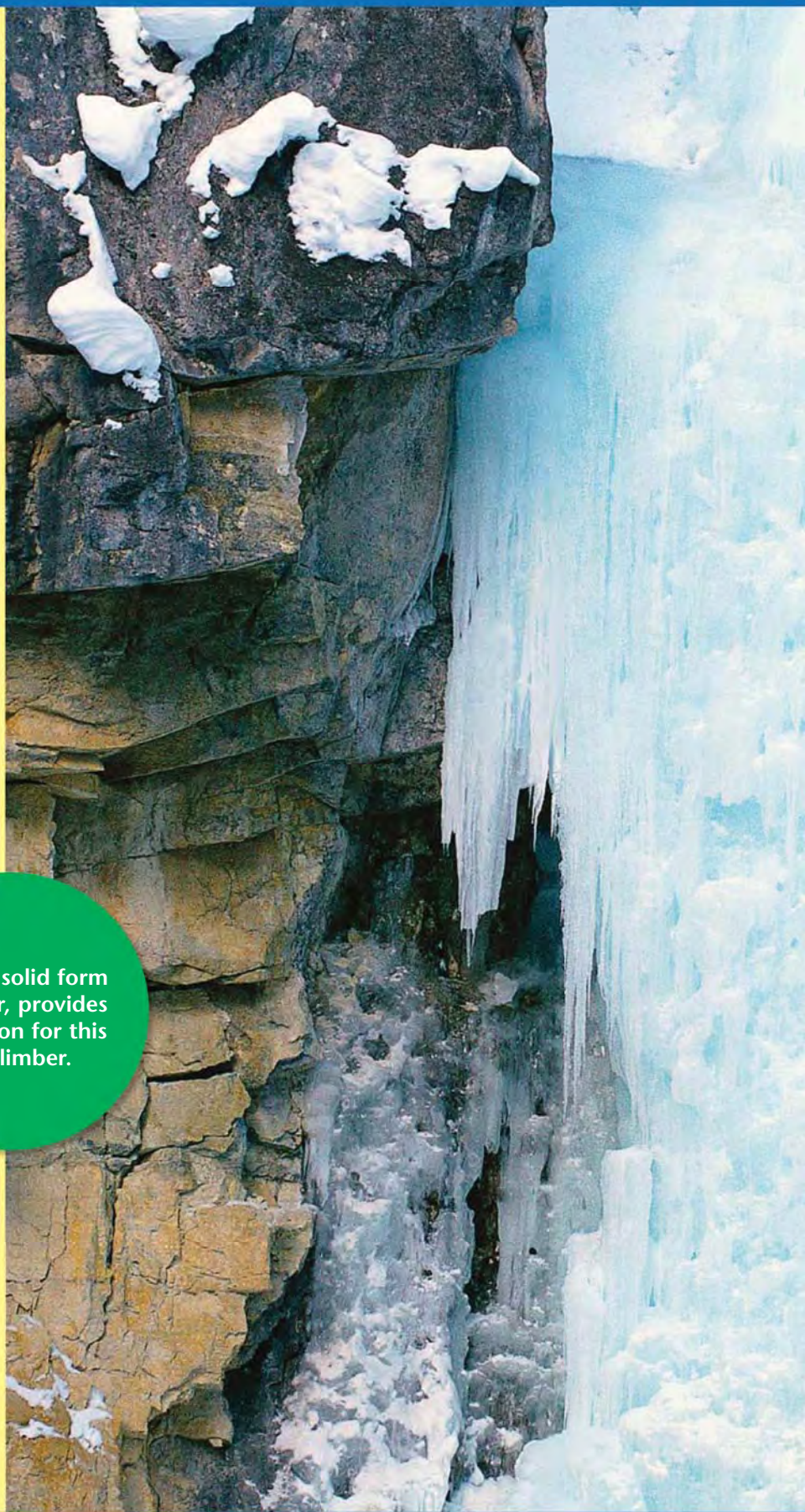
14.2 Vapor Pressure and Boiling Point

- A. Evaporation and Vapor Pressure
 - Vapor Pressure
- B. Boiling Point and Vapor Pressure

14.3 Properties of Solids

- A. The Solid State: Types of Solids
- B. Bonding in Solids
 - Ionic Solids
 - Molecular Solids
 - Atomic Solids
 - Bonding in Metals

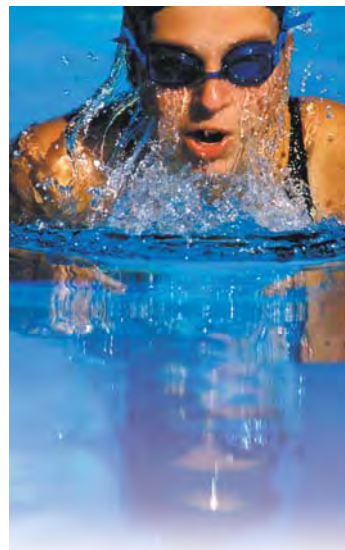
Ice, the solid form of water, provides recreation for this ice climber.



IN YOUR LIFE

You have only to think about water to appreciate how different the three states of matter are.

Flying, swimming, and ice skating are all done in contact with water in its various states. We swim in liquid water and skate on water in its solid form (ice). Airplanes fly in an atmosphere containing water in the gaseous state (water vapor). To allow these various activities, the arrangements of the water molecules must be significantly different in their gas, liquid, and solid forms.



A swimmer in water, one of the earth's most important substances

WHAT DO YOU KNOW?**Prereading Questions**

1. At a molecular level, how do gases, liquids, and solids compare?
2. What is meant by the boiling point of a liquid?
3. What does it mean for a liquid to evaporate?



SECTION 14.1

Intermolecular Forces and Phase Changes

Key Terms

- Intermolecular forces
- Intramolecular forces
- Dipole–dipole attraction
- Hydrogen bonding
- London dispersion forces
- Normal boiling point
- Heating/cooling curve
- Normal freezing point
- Molar heat of fusion
- Molar heat of vaporization

Objectives

- To learn about dipole–dipole, hydrogen bonding, and London dispersion forces
- To understand the effect of intermolecular forces on the properties of liquids
- To learn some of the important features of water
- To learn about interactions among water molecules
- To understand and use heat of fusion and heat of vaporization

In Chapter 13 we saw that the particles of a gas are far apart, are in rapid random motion, and have little effect on each other. Solids are obviously very different from gases.

Gases

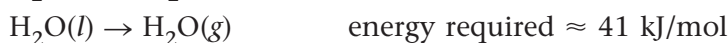
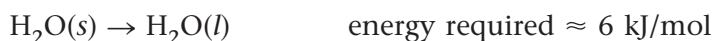
- Low density
- Highly compressible
- Fill container

Solids

- High density
- Slightly compressible
- Rigid (keeps its shape)

These properties indicate that the components of a solid are close together and exert large attractive forces on each other.

The properties of liquids lie somewhere between those of solids and of gases—but not midway between, as can be seen from some of the properties of the three states of water. For example, it takes about seven times more energy to change liquid water to steam (a gas) at 100 °C than to melt ice to form liquid water at 0 °C.



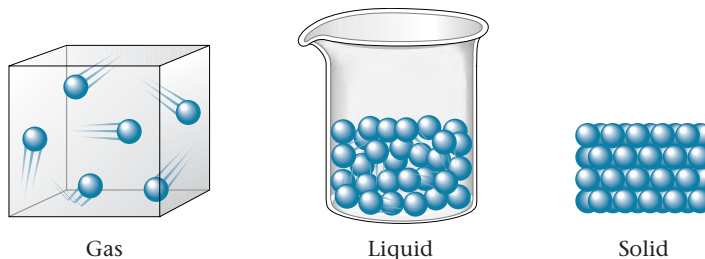
These values indicate that going from the liquid to the gaseous state involves a much greater change than going from the solid to the liquid. Therefore, we can conclude that the solid and liquid states are more similar than the liquid and gaseous states. This is also demonstrated by the densities of the three states of water (see **Table 14.1**). Note that water in its gaseous state is about 2000 times less dense than in the solid and liquid states and that the latter two states have very similar densities.

We find in general that the liquid and solid states show many similarities and are strikingly different from the gaseous state.

Table 14.1

Densities of the Three States of Water

State	Density (g/cm ³)
solid (0 °C, 1 atm)	0.9168
liquid (25 °C, 1 atm)	0.9971
gas (100 °C, 1 atm)	5.88×10^{-4}



The best way to picture the solid state is in terms of closely packed, highly ordered particles in contrast to the widely spaced, randomly arranged particles of a gas. The liquid state lies in between, but its properties indicate that it much more closely resembles the solid than the gaseous state. It is useful to picture a liquid in terms of particles that are generally quite close together,

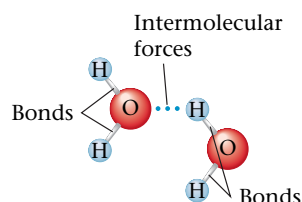
but with a more disordered arrangement than in the solid state and with some empty spaces. For most substances, the solid state has a higher density than the liquid. However, water is an exception to this rule. Ice has an unusual amount of empty space and so is less dense than liquid water, as indicated in Table 14.1.

In this chapter we will explore the important properties of liquids and solids. We will illustrate many of these properties by considering one of the earth's most important substances: water.

A. Intermolecular Forces

Most substances consisting of small molecules are gases at normal temperatures and pressures. Examples are oxygen gas (contains O_2), nitrogen gas (contains N_2), methane gas (contains CH_4), and carbon dioxide gas (contains CO_2). A notable exception to this rule is water. Given the small size of its molecules we might expect water to be a gas at normal temperatures and pressures. Think about how different the world would be if water were a gas at 25 °C and 1 atm pressure. The oceans would be empty chasms, the Mississippi River would be a wide, dry ditch, and it would never rain! Life on earth as we know it would be impossible if water were a gas under normal conditions.

So why is water a liquid rather than a gas at normal temperatures and pressures? The answer has to do with something called **intermolecular forces**—forces that occur between the molecules.



Notice that molecules are held together by bonds, sometimes called **intramolecular forces**, that occur inside the molecules. Intermolecular forces exist *between* molecules. We have seen that covalent bonding forces within molecules arise from the sharing of electrons, but how do intermolecular forces arise? Actually several types of intermolecular forces exist. To illustrate one type, we will consider the forces that exist among water molecules.

Active Reading Question

What type of force is represented by the lines in a Lewis structure?
What type of force is responsible for holding molecules together in a solid or a liquid?

As we saw in Chapter 12, water is a polar molecule—it has a dipole moment. When molecules with dipole moments are put together, they orient themselves to take advantage of their charge distributions. Molecules with dipole moments can attract each other by lining up so that the positive and negative ends are close to each other, as shown here.



Nitrogen, which forms a liquid at 77 K, is being poured causing moisture in the air to freeze to produce a fog.

Intermolecular forces

Attractive forces that occur between molecules

Intramolecular forces

Attractive forces that occur between atoms in a molecule; chemical bonds

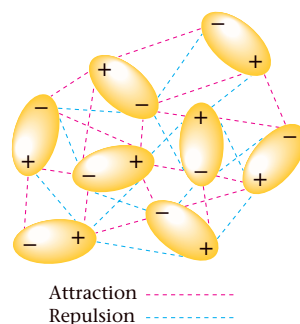
Dipole–dipole attraction

The attractive force between the positively charged end of one polar molecule with the negatively charged end of another polar molecule

Hydrogen bonding

Special name for unusually strong dipole–dipole attractions that occur among molecules in which hydrogen is bonded to a highly electronegative atom (such as nitrogen, oxygen, or fluorine)

This is called a **dipole–dipole attraction**. In the liquid, the dipoles find the best compromise between attraction and repulsion, as shown here.



Dipole–dipole forces are typically only about 1% as strong as covalent or ionic bonds, and they become weaker as the distance between the dipoles increases. In the gas phase, where the molecules are usually very far apart, these forces are relatively unimportant.

Hydrogen Bonding

Particularly strong dipole–dipole forces occur between molecules in which hydrogen is bound to a highly electronegative atom, such as nitrogen, oxygen, or fluorine. Two factors account for the strengths of these interactions: the great polarity of the bond and the close approach of the dipoles, which is made possible by the very small size of the hydrogen atom. Because dipole–dipole attractions of this type are so unusually strong, they are given a special name—**hydrogen bonding**.

CRITICAL THINKING ?

You have learned the difference between intermolecular forces and intramolecular forces.

What if intermolecular forces were stronger than intramolecular forces? What differences could you observe in the world?

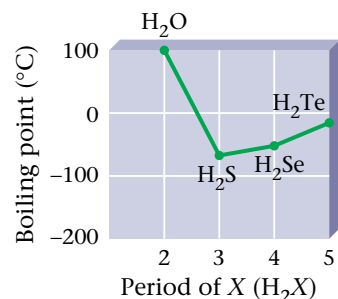
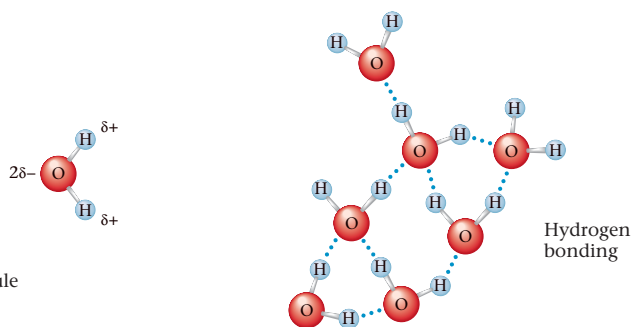


Figure 14.1

The boiling points of the covalent hydrides of elements in Group 6

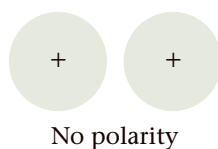
Hydrogen bonding has a very important effect on various physical properties. For example, the boiling points for the covalent compounds of hydrogen with the elements in Group 6 are given in **Figure 14.1**. Note that the boiling point of water is much higher than would be expected from the trend shown by the other members of the series. Why? Because the especially large electronegativity value of the oxygen atom compared with that of other group members causes the O–H bonds to be much more polar than the S–H, Se–H, or Te–H bonds. This leads to very strong hydrogen-bonding forces among the water molecules. An unusually large quantity of energy is required to overcome these interactions and separate the molecules to produce the gaseous state. That is, water molecules tend to remain together in the liquid state even at relatively high temperatures, hence the very high boiling point of water.

Active Reading Question

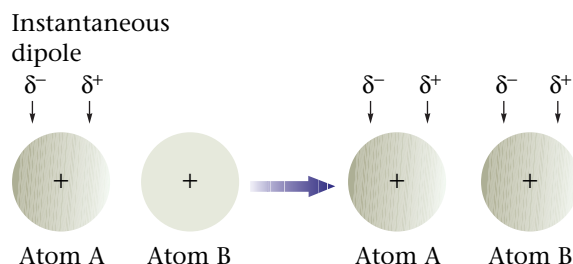
What observation shows us that hydrogen bonding is quite strong?

London Dispersion Forces

Even molecules without dipole moments must exert forces on each other. We know this because all substances—even the noble gases—exist in the liquid and solid states at very low temperatures. There must be forces to hold the atoms or molecules as close together as they are in these condensed states. The forces that exist among noble gas atoms and nonpolar molecules are called **London dispersion forces**. To understand the origin of these forces, consider a pair of noble gas atoms. Although we usually assume that the electrons of an atom are uniformly distributed about the nucleus,



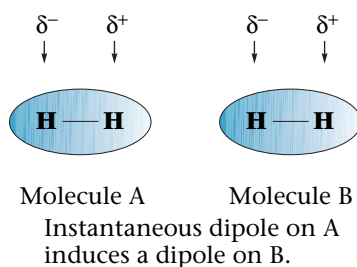
this is apparently not true at every instant. Atoms can develop a temporary dipolar arrangement of charge as the electrons move around the nucleus. This *instantaneous dipole* can then *induce* a similar dipole in a neighboring atom.



The interatomic attraction thus formed is both weak and short-lived, but it can be very significant for large atoms and large molecules, as we will see.

The motions of the atoms must be greatly slowed down before the weak London dispersion forces can lock the atoms into place to produce a solid. This explains, for instance, why the noble gas elements have such low freezing points (see **Table 14.2**).

Nonpolar molecules such as H_2 , N_2 , and I_2 , none of which has a permanent dipole moment, also attract each other by London dispersion forces.



London forces become more significant as the sizes of atoms or molecules increase. Larger size means there are more electrons available to form the dipoles.

Active Reading Question

Rank intermolecular forces from weakest to strongest (given molecules of equal size).

London dispersion forces

Relatively weak intermolecular forces resulting from a temporarily uneven distribution of electrons that induces a dipole in a neighbor

Table 14.2

The Freezing Points of the Group 8 Elements

Element	Freezing Point ($^{\circ}\text{C}$)
helium*	-272.0 (25 atm)
neon	-248.6
argon	-189.4
krypton	-157.3
xenon	-111.9

*Helium will not freeze unless the pressure is increased above 1 atm.



Drinking water is important when exercising.

B. Water and Its Phase Changes

In the world around us we see many solids (soil, rocks, trees, concrete, and so on), and we are immersed in the gases of the atmosphere. But the liquid we most commonly see is water; it is virtually everywhere, covering about 70% of the earth's surface. Approximately 97% of the earth's water is found in the oceans, which are actually mixtures of water and huge quantities of dissolved salts.

Water is one of the most important substances on earth. It is crucial for sustaining the reactions within our bodies that keep us alive, but it also affects our lives in many indirect ways. The oceans help moderate the earth's temperature. Water cools automobile engines and nuclear power plants. Water provides a means of transportation on the earth's surface and acts as a medium for the growth of many of the creatures we use as food, and much more.

Pure water is a colorless, tasteless substance that at 1 atm pressure freezes to form a solid at 0 °C and vaporizes completely to form a gas at 100 °C. This means that (at 1 atm pressure) the liquid range of water occurs between the temperatures 0 °C and 100 °C.

DID YOU KNOW

The water we drink often has a taste because of the substances dissolved in it.

Heating Water What happens when we heat liquid water? First the temperature of the water rises. Just as with gas molecules, the motions of the water molecules increase as it is heated. Eventually the temperature of the water reaches 100 °C; now bubbles develop in the interior of the liquid, float to the surface, and burst—the boiling point has been reached. An interesting thing happens at the boiling point: even though heating continues, the temperature stays at 100 °C until all the water has changed to vapor. Only when all of the water has changed to the gaseous state does the temperature begin to rise again. (We are now heating the vapor.) At 1 atm pressure, liquid water always changes to gaseous water at 100 °C, the **normal boiling point** for water.

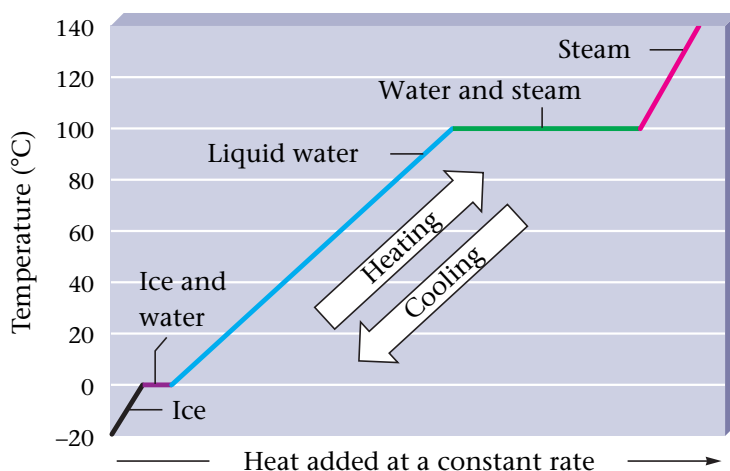
Normal boiling point

The boiling temperature of a liquid under one atmosphere of pressure. The temperature at which the vapor pressure of a liquid is exactly one atmosphere.

The experiment just described is represented in **Figure 14.2**, which is called the **heating/cooling curve** for water. Going from left to right on this graph means energy is being added (heating). Going from right to left on the graph means that energy is being removed (cooling).

Figure 14.2

The heating/cooling curve for water heated or cooled at a constant rate. The plateau at the boiling point is longer than the plateau at the melting point, because it takes almost seven times as much energy (and thus seven times the heating time) to vaporize liquid water as to melt ice.



Cooling Water When liquid water is cooled, the temperature decreases until it reaches 0 °C, where the liquid begins to freeze (see Figure 14.2). The temperature remains at 0 °C until all the liquid water has changed to ice and then begins to drop again as cooling continues. At 1 atm pressure, water freezes (or, in the opposite process, ice melts) at 0 °C. This is called the **normal freezing point** of water. Liquid and solid water can coexist indefinitely if the temperature is held at 0 °C. However, at temperatures below 0 °C liquid water freezes, while at temperatures above 0 °C ice melts.

Interestingly, water expands when it freezes. That is, one gram of ice at 0 °C has a greater volume than one gram of liquid water at 0 °C. This has very important practical implications. For instance, water in a confined space can break its container when it freezes and expands. This accounts for the bursting of water pipes and engine blocks that are left unprotected in freezing weather.

The expansion of water when it freezes also explains why ice cubes float. Recall that density is defined as mass/volume. When one gram of liquid water freezes, its volume becomes greater (it expands). Therefore, the *density* of one gram of ice is less than the density of one gram of water, because in the case of ice we divide by a slightly larger volume. For example, at 0 °C the density of liquid water is

$$\frac{1.00 \text{ g}}{1.00 \text{ mL}} = 1.00 \text{ g/mL}$$

and the density of ice is

$$\frac{1.00 \text{ g}}{1.09 \text{ mL}} = 0.917 \text{ g/mL}$$

The lower density of ice also means that ice floats on the surface of lakes as they freeze, providing a layer of insulation that helps prevent lakes and rivers from freezing solid in the winter. This means that aquatic life continues to have liquid water available through the winter.

C. Energy Requirements for the Changes of State

It is important to recognize that changes of state from solid to liquid and from liquid to gas are *physical* changes. No *chemical* bonds are broken in these processes. Ice, water, and steam all contain H₂O molecules. When water is boiled to form steam, water molecules are separated from each other (see **Figure 14.3**) but the individual molecules remain intact.

It takes energy to melt ice and to vaporize water, because intermolecular forces between water molecules must be overcome. In ice the molecules are virtually locked in place, although they can vibrate about their positions. When energy is added, the vibrational motions increase, and the molecules eventually achieve the greater movement and disorder characteristic of liquid water. The ice has melted. As still more energy is added, the gaseous state is eventually reached, in which the individual molecules are far apart

Normal freezing point

The freezing temperature of a liquid under one atmosphere of pressure

Information

Remember that temperature is a measure of the random motions (average kinetic energy) of the particles in a substance.

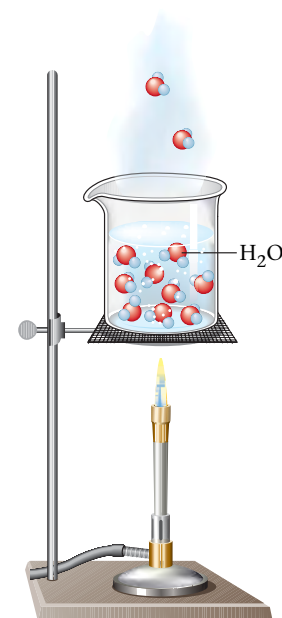


Figure 14.3

Both liquid water and gaseous water contain H₂O molecules. In liquid water the H₂O molecules are close together, whereas in the gaseous state the molecules are widely separated. The bubbles contain gaseous water.

and interact relatively little. However, the gas still consists of water molecules. It would take *much* more energy to overcome the covalent bonds and decompose the water molecules into their component atoms.

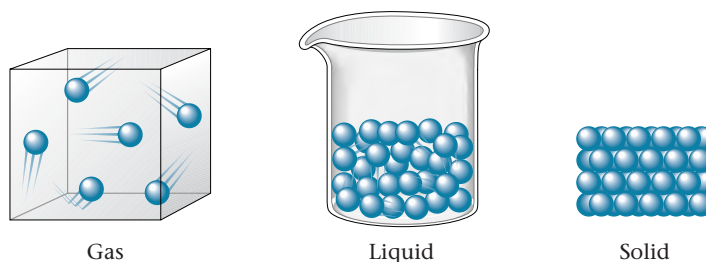
The energy required to melt 1 mol of a substance is called the **molar heat of fusion**. For ice, the molar heat of fusion is 6.02 kJ/mol. The energy required to change 1 mol of liquid to its vapor is called the **molar heat of vaporization**. For water, the molar heat of vaporization is 40.6 kJ/mol at 100 °C. Notice in Figure 14.2 that the plateau that corresponds to the vaporization of water is much longer than that for the melting of ice. This occurs because it takes much more energy (almost seven times as much) to vaporize a mole of water than to melt a mole of ice. This is consistent with our models of solids, liquids, and gases.

Molar heat of fusion

The energy required to melt one mole of a solid

Molar heat of vaporization

The energy required to vaporize one mole of a liquid

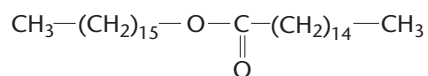


CHEMISTRY IN YOUR WORLD

Connection to Biology

Whales Need Changes of State

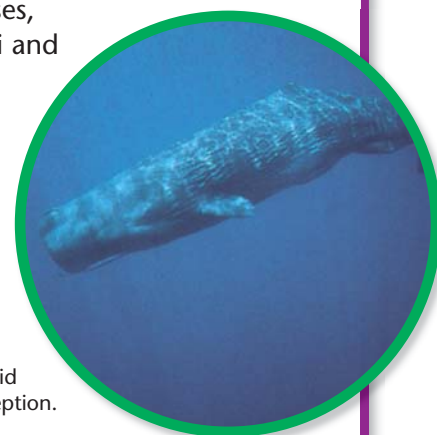
Sperm whales are prodigious divers. They commonly dive a mile or more into the ocean, hovering at that depth in search of schools of squid or fish. To remain motionless at a given depth, the whale must have the same density as the surrounding water. Because the density of seawater increases with depth, the sperm whale has a system that automatically increases its density as it dives. This system involves the spermaceti organ found in the whale's head. Spermaceti is a waxy substance with the formula



which is a liquid above 30 °C. At the ocean surface the spermaceti in the whale's head is a liquid, warmed by the flow of blood through the spermaceti organ. When the whale dives, this blood flow decreases and the colder water causes the spermaceti to begin freezing. Because solid

spermaceti is more dense than the liquid state, the sperm whale's density increases as it dives, matching the increase in the water's density.* When the whale wants to resurface, blood flow through the spermaceti organ increases, remelting the spermaceti and making the whale more buoyant. So the sperm whale's sophisticated density-regulating mechanism is based on a simple change of state.

*For most substances, the solid state is more dense than the liquid state. Water is an important exception.



A sperm whale

In liquids, the particles (molecules) are relatively close together, so most of the intermolecular forces are still present. However, when the molecules go from the liquid to the gaseous state, they must be moved far apart. To separate the molecules enough to form a gas, virtually all of the intermolecular forces must be overcome, and this requires large quantities of energy.

Active Reading Question

Which is greater, the molar heat of fusion of water or the molar heat of vaporization of water? Why?

EXAMPLE 14.1

Calculating Energy Changes: Solid to Liquid

Calculate the energy required to melt 8.5 g of ice at 0 °C. The molar heat of fusion for ice is 6.02 kJ/mol.

Solution

Where do we want to go?

Energy to melt 8.5 g ice at 0 °C = ? kJ

What do we know?

- 8.5 g ice at 0 °C
- molar heat of fusion for ice = 6.02 kJ/mol
- molar mass H₂O = 18.02 g/mol

How do we get there?

- Convert the mass of water to moles since the heat of fusion for water is per mole of water.

$$8.5 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18 \text{ g H}_2\text{O}} = 0.47 \text{ mol H}_2\text{O}$$

- The molar heat of fusion represents the equivalence statement
6.02 kJ required for 1 mol H₂O

This leads to the conversion factor used below to calculate the energy required to melt 8.5 g ice:

$$0.47 \text{ mol H}_2\text{O} \times \frac{6.02 \text{ kJ}}{\text{mol H}_2\text{O}} = 2.8 \text{ kJ}$$

Does it make sense?

Since our sample is about $\frac{1}{2}$ mol the amount of energy should be about half the molar heat of fusion (6.02 kJ). Our answer is 2.8 kJ which is about half the molar heat of fusion.

EXAMPLE 14.2

Calculating Energy Changes: Liquid to Gas

Calculate the energy (in kJ) required to heat 25 g of liquid water from 25 °C to 100. °C and change it to steam at 100. °C. The specific heat capacity of liquid water is 4.18 J/g °C, and the molar heat of vaporization of water is 40.6 kJ/mol.

Top Ten Lowest Melting Points

Element	Melting Point (°C)
Helium	-270.
Hydrogen	-259
Neon	-249
Fluorine	-220.
Oxygen	-218
Nitrogen	-210.
Argon	-189
Krypton	-157
Xenon	-112
Chlorine	-101

Solution

Where do we want to go?

Energy required = ? kJ

What do we know?

- 25 g water to 25 g steam
- $T_{\text{initial}} = 25\text{ °C}$
- $T_{\text{final}} = 100.\text{ °C}$
- Specific heat capacity for water (solid) = 4.184 J/g °C
- molar heat of vaporization for water = 40.6 kJ/mol
- molar mass $\text{H}_2\text{O} = 18.02\text{ g/mol}$
- Energy to heat a liquid = $Q = s \times m \times \Delta T$

How do we get there?

We need to split the problem into two parts:

- heating the water to its boiling point
- converting the water to steam at the boiling point

Part 1 Heating to Boiling

We begin by heating the water from 25 °C to 100. °C.

$$\Delta T = 100.\text{ °C} - 25\text{ °C} = 75\text{ °C}$$

To determine the energy to heat the water we use the equation

Energy required (Q)	=	Specific heat capacity (s)	×	Mass of water (m)	×	Temperature change (ΔT)
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which we can represent by the equation

$$Q = s \times m \times \Delta T$$

Thus

$$\begin{aligned} Q &= 4.184 \frac{\text{J}}{\text{g}\cdot\text{°C}} \times 25\text{ g} \times 75\text{ °C} = 7.8 \times 10^3\text{ J} \\ &= 7.8 \times 10^3\text{ J} \times \frac{1\text{ kJ}}{1000\text{ J}} = 7.8\text{ kJ} \end{aligned}$$

Energy required to heat 25 g of water from 25 °C to 100. °C Specific heat capacity Mass of water Temperature change

MATH



When possible, try to break a large problem into a series of smaller problems.

Part 2 Vaporization

Now we must use the molar heat of vaporization to calculate the energy required to vaporize the 25 g of water at 100. °C. The heat of vaporization is given per mole rather than per gram, so we must first convert the 25 g of water to moles.

$$25 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} = 1.4 \text{ mol H}_2\text{O}$$

We can now calculate the energy required to vaporize the water.

$$\frac{40.6 \text{ kJ}}{\text{mol H}_2\text{O}} \times 1.4 \text{ mol H}_2\text{O} = 57 \text{ kJ}$$

Molar heat of vaporization Moles of water

The total energy is the sum of the two parts.

$$7.8 \text{ kJ} + 57 \text{ kJ} = 65 \text{ kJ}$$

Heat from 25 °C to 100. °C Change to vapor

Practice Problem • Exercise 14.2

Calculate the total energy required to melt 15 g of ice at 0 °C, heat the water to 100. °C, and vaporize it to steam at 100. °C.

Hint: Break the process into three parts and then find the sum.



SECTION 14.1

REVIEW QUESTIONS

- 1 What is the difference between intermolecular forces and intramolecular forces?
- 2 Is a hydrogen bond a true chemical bond? Explain.
- 3 How can molecules without dipoles condense to form liquids or solids?
- 4 Draw a cooling curve for a sample of steam (120 °C) which condenses to liquid water and cools to eventually form ice (0 °C).
- 5 Why are changes of state considered to be physical changes and not chemical changes?
- 6 Why doesn't the temperature of a substance (such as ice) change during melting? What is happening to the energy being added to the system?
- 7 Calculate the energy required to change 1.00 mol of ice at -10. °C to water at 15 °C. The heat capacity of ice is 2.03 J/g °C, the heat capacity of water is 4.184 J/g °C, and the molar heat of fusion for ice is 6.02 kJ/mol.



RESEARCH LINKS
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SECTION 14.2

Vapor Pressure and Boiling Point

Key Terms

- Vaporization (evaporation)
- Condensation
- Vapor pressure

Objectives

- To understand the relationship among vaporization, condensation, and vapor pressure
- To relate the boiling point of water to its vapor pressure

A. Evaporation and Vapor Pressure

We all know that a liquid can evaporate from an open container. This is clear evidence that the molecules of a liquid can escape the liquid's surface and form a gas. This process, which is called **vaporization** or **evaporation**, requires energy to overcome the relatively strong intermolecular forces in the liquid.

To help you understand the energy changes that accompany evaporation, consider the microscopic view of a liquid shown in **Figure 14.4**.

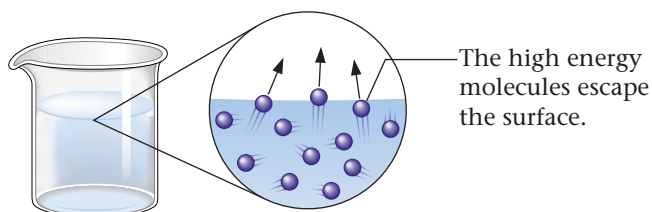


Figure 14.4

The microscopic view of a liquid near its surface

Vaporization (evaporation)

The process in which a liquid is converted to a gas

Remember that the temperature of a substance reflects the *average* kinetic energy of the components of that substance. Not all components have this average value of the kinetic energy, however. In fact, some components have a relatively small kinetic energy (are moving relatively slowly), whereas other components may have a kinetic energy much higher than the average (are moving much faster than the “average component”). To escape into the vapor phase a given component must have sufficient speed to overcome the intermolecular forces of the liquid. Thus only the fastest-moving components can escape the surface of the liquid.

Now consider what will happen if the evaporation process is taking place from an insulated container (no energy can flow in or out as heat). What will happen to the temperature of the liquid as evaporation occurs? Because the higher-energy components escape as evaporation occurs, the *average* energy of the remaining components will decrease. As the average kinetic energy of the components of the liquid drops, the temperature decreases. Thus evaporation is a cooling process—it is endothermic.

Active Reading Question

Why is evaporation a cooling process?

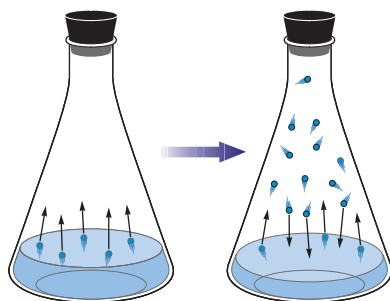
On the other hand, if a liquid evaporates from a container that is not insulated, energy will flow in (heat) as evaporation occurs to keep the temperature constant. Again, this process shows that evaporation is endothermic.

The fact that vaporization requires energy has great practical significance; in fact, one of the most important roles that water plays in our world is to act as a coolant. Because of the strong hydrogen bonding among its molecules in the liquid state, water has an unusually large heat of vaporization

(41 kJ/mol). A significant portion of the sun's energy is spent evaporating water from the oceans, lakes, and rivers rather than warming the earth. The vaporization of water is also crucial to our body's temperature-control system, which relies on the evaporation of perspiration.

Vapor Pressure

When we place a given amount of liquid in a container and then close it, we observe that the amount of liquid at first decreases slightly but eventually becomes constant. The decrease occurs because there is a transfer of molecules from the liquid to the vapor phase.



However, as the number of vapor molecules increases, it becomes more and more likely that some of them will return to the liquid. The process by which vapor molecules form a liquid is called **condensation**. Eventually, the same number of molecules are leaving the liquid as are returning to it: the rate of condensation equals the rate of evaporation. *At this point no further change occurs in the amounts of liquid or vapor, because the two opposite processes exactly balance each other; the system is at equilibrium.* Note that this system is highly *dynamic* on the molecular level—molecules are constantly escaping from and entering the liquid. However, there is no *net* change because the two opposite processes *balance* each other.

As a real life example, consider two island cities connected by a bridge. Suppose the traffic flow on the bridge is the same in both directions. There is motion—we can see the cars traveling across the bridge—but the number



Sweat cools the body by evaporation

Condensation

The process in which a vapor is converted to a liquid

Information

A system at equilibrium is dynamic on the molecular level, but shows no visible changes.

CHEMISTRY EXPLORERS

Dorothy Hodgkin • 1910–1994

Dorothy Hodgkin was born in Cairo, Egypt. She became interested in chemistry and crystals at about the age of 10, and on her sixteenth birthday received a book by the Nobel Prize-winning physicist William Bragg. The book discussed how to use X rays to analyze crystals, and from that point on, her career path was set.

Dr. Hodgkin used X-ray analysis to determine the structures of proteins. She is most famous, however, for three very important discoveries. In 1945 she determined the structure of penicillin, and this knowledge helped manufacturers create penicillin. In 1954 she determined the structure of

vitamin B₁₂, which led to her winning the Nobel Prize in Chemistry in 1964. While both of these are important and useful discoveries, Dr. Hodgkin considered her greatest scientific achievement to be the discovery of insulin in 1969. Insulin is now used in the treatment of diabetes.

Dr. Hodgkin's great achievements helped change the way we think about molecules that are necessary for life. She was one of the first people to use molecular structure to explain biological function.



of cars in each city is not changing because an equal number enter and leave each one. The result is no *net* change in the number of autos in each city: an equilibrium exists.

The pressure of the vapor present at equilibrium with its liquid is called the *equilibrium vapor pressure* or, more commonly, the **vapor pressure** of the liquid. A simple barometer can be used to measure the vapor pressure of a liquid, as shown in **Figure 14.5**. Because mercury is so dense, any common liquid injected at the bottom of the column of mercury floats to the top, where it produces a vapor, and the pressure of this vapor pushes some mercury out of the tube. When the system reaches equilibrium, the vapor pressure can be determined from the change in the height of the mercury column.

In effect, we are using the space above the mercury in the tube as a closed container for each liquid. However, in this case as the liquid vaporizes, the vapor formed creates a pressure that pushes some mercury out of the tube and lowers the mercury level. The mercury level stops changing when the excess liquid floating on the mercury comes to equilibrium with the vapor. The change in the mercury level (in millimeters) from its initial position (before the liquid was injected) to its final position is equal to the vapor pressure of the liquid.

The vapor pressures of liquids vary widely (see Figure 14.5). Liquids with high vapor pressures are said to be *volatile*—they evaporate rapidly.

The vapor pressure of a liquid at a given temperature is determined by the *intermolecular forces* that act among the molecules. Liquids in which the intermolecular forces are large have relatively low vapor pressures, because such molecules need high energies to escape to the vapor phase. For example, although water is a much smaller molecule than diethyl ether, $\text{C}_2\text{H}_5\text{—O—C}_2\text{H}_5$, the strong hydrogen-bonding forces in water cause its vapor pressure to be much lower than that of ether (see Figure 14.5).

Vapor pressure

The pressure exerted by a vapor in equilibrium with its liquid phase at a certain temperature

Information

Vapor, not gas, is the term we customarily use for the gaseous state of a substance that exists naturally as a solid or liquid at 25 °C and 1 atm.

Active Reading Question

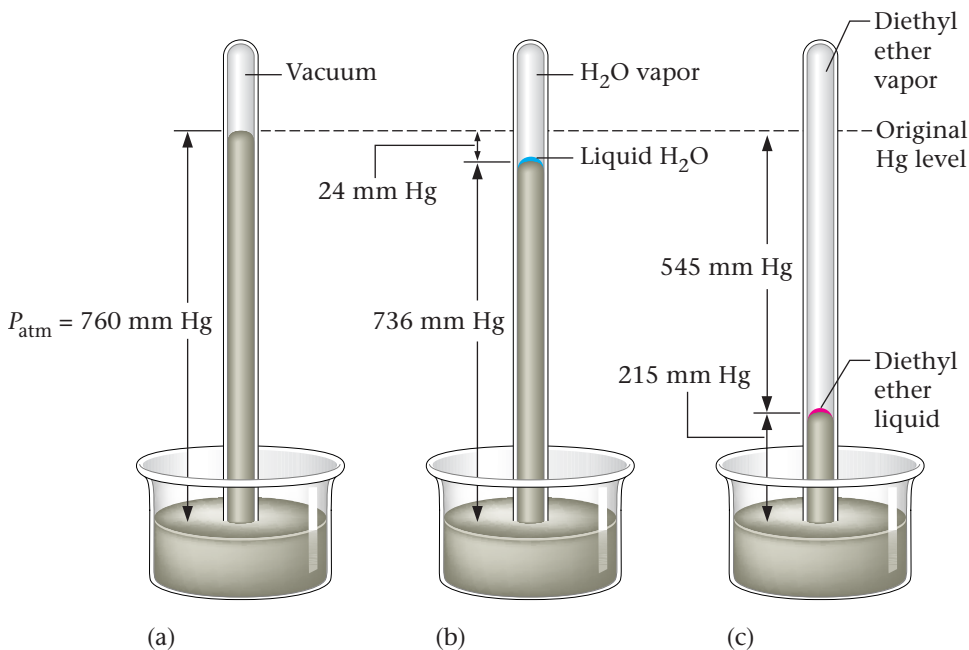
Why do substances have different vapor pressures?

Figure 14.5

(a) It is easy to measure the vapor pressure of a liquid by using a simple barometer of the type shown here.

(b) The water vapor pushed the mercury level down 24 mm (760–736), so the vapor pressure of water is 24 mm Hg at this temperature.

(c) Diethyl ether is much more volatile than water and thus shows a higher vapor pressure. In this case, the mercury level has been pushed down 545 mm (760–215), so the vapor pressure of diethyl ether is 545 mm Hg at this temperature.



EXAMPLE 14.3**Using Knowledge of Intermolecular Forces to Predict Vapor Pressure**




Predict which substance in each of the following pairs will show the largest vapor pressure at a given temperature.

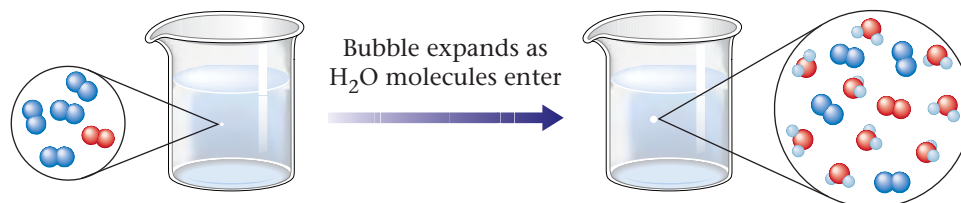
- $\text{H}_2\text{O}(l)$, $\text{CH}_3\text{OH}(l)$
- $\text{CH}_3\text{OH}(l)$, $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}(l)$

Solution

- Water contains two polar O—H bonds; methanol (CH_3OH) has only one. Therefore, the hydrogen bonding among H_2O molecules is expected to be much stronger than that among CH_3OH molecules. This gives water a lower vapor pressure than methanol.
- Each of these molecules has one polar O—H bond. However, because $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$ is a much larger molecule than CH_3OH , it has much greater London forces and thus is less likely to escape from its liquid. Thus $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}(l)$ has a lower vapor pressure than $\text{CH}_3\text{OH}(l)$.

B. Boiling Point and Vapor Pressure

Look at the photo of water boiling. When water boils, bubbles form in the interior of the liquid, and then rise to the surface where they “pop.” What causes these bubbles to form at the boiling point and not at a lower temperature? An air bubble acts as a “birth place” for a bubble growing in the interior of water as high-energy H_2O () molecules enter.  represent N_2 molecules;  represent O_2 molecules.



Assume we start with a tiny air bubble in the water. This bubble expands when high-energy water molecules enter the bubble and produce enough internal pressure to push back the water surrounding the bubble. The formation of a bubble can happen only when the average kinetic energy of the water molecules is great enough to produce a pressure inside the bubble that can oppose the atmospheric pressure pushing down on the surface of the water (see **Figure 14.6**). That is, the vapor pressure of the water must be equal to atmospheric pressure before boiling can occur. This explains why the boiling of water occurs at $100\text{ }^\circ\text{C}$ at an atmospheric pressure of 1 atm and not at a lower temperature. Only when the water temperature reaches $100\text{ }^\circ\text{C}$ are bubbles able to form in the interior of the liquid because the water molecules are energetic enough to sustain a pressure of 1 atm inside the bubbles. At temperatures below $100\text{ }^\circ\text{C}$

**information**

As you heat water, notice that tiny bubbles form well below the boiling point. These bubbles result from dissolved air being expelled from the water.

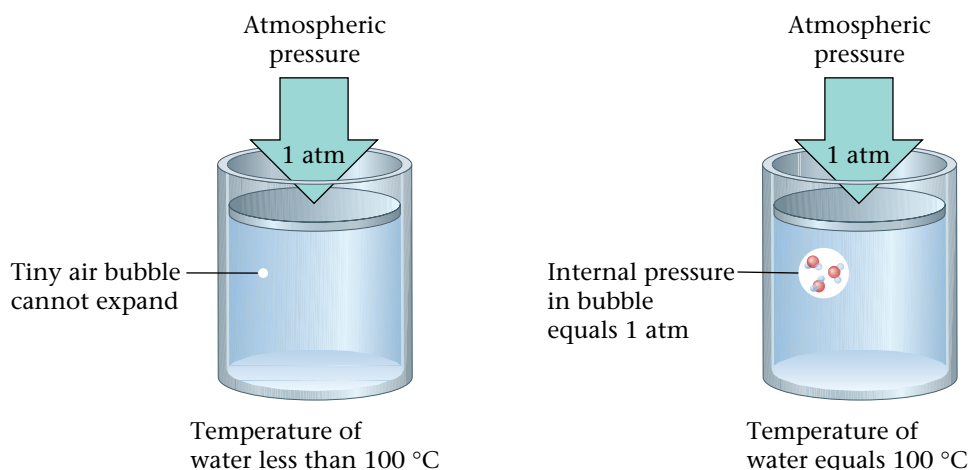


Figure 14.6

The formation of the bubble in the interior of water is opposed by atmospheric pressure.

CRITICAL THINKING ?

You have seen that the water molecule has a bent shape and is therefore a polar molecule. This characteristic accounts for many of water's interesting properties.

What if the water molecule were linear? How would it affect the properties of water? How would life be different?

the external pressure (1 atm) pushing on the surface of the water prevents bubble formation. So water boils at 100 °C (1 atm atmospheric pressure) because its vapor pressure is 1 atm at 100 °C.

Given this information, why do you suppose that water boils at 90 °C in Leadville, Colorado (elevation 10,000 ft)? Because Leadville is located at a high elevation where the atmospheric pressure is much less than 1 atm, the pressure pushing on the surface of water there is much less than at sea level. Thus you do not have to heat water to as high a temperature in Leadville to enable bubbles to form in the interior of the water. As a consequence, water can boil at a lower temperature in Leadville than in Los Angeles.

We have seen that the boiling point of water changes with atmospheric pressure. **Table 14.3** shows the boiling point of water at various locations around the world.

Active Reading Question

It takes longer to boil a potato in Leadville, Colorado, than in Madison, Wisconsin. Why is this?

Table 14.3

Boiling Point of Water at Various Locations

Location	Feet Above Sea Level	P_{atm} (atm)	Boiling Point (°C)
Top of Mt. Everest, Tibet	29,028	0.32	70
Top of Mt. McKinley, Alaska	20,320	0.45	79
Top of Mt. Whitney, California	14,494	0.57	85
Top of Mt. Washington, New Hampshire	6,293	0.78	93
Boulder, Colorado	5430	0.80	94
Madison, Wisconsin	900	0.96	99
New York City, New York	10	1.00	100
Death Valley, California	-282	1.01	100.3

Hot Molecules**Materials**

- molecular model kit

Procedure

1. Obtain a molecular model kit from your teacher.
2. Build models for each pair of molecules.
 - a. H_2O , CH_4
 - b. C_2H_6 , C_6H_{14}
 - c. C_8H_{18} , C_8H_{18} (make one molecule with all the carbons connected in a line and one other molecule)
 - d. $\text{C}_2\text{H}_6\text{O}$, $\text{C}_2\text{H}_6\text{O}$ (make two different molecules)

Results/Analysis

1. Use your models to decide which of the molecules in each pair has the highest boiling point. Explain your reasoning.
2. Use your models to decide which of the molecules in each pair has the highest vapor pressure. Explain your reasoning.

SECTION 14.2**REVIEW QUESTIONS**

- 1 Use your knowledge of evaporation to explain why perspiration cools you on a warm, dry day.
- 2 Would you expect your hand to feel cooler if you dipped it into water and removed it to the air, or if you dipped your hand into methyl alcohol and removed it to the air? Explain. Note: Do not dip your hand in methyl alcohol.
- 3 Explain how a barometer can be used to measure the vapor pressure of a liquid.
- 4 Water and methane (CH_4) have very similar molar masses. However, water is a liquid under normal conditions, and methane is a gas. Use intermolecular forces to explain this.
- 5 Why does water boil at $100\text{ }^\circ\text{C}$ at sea level?
- 6 What is the difference between boiling and evaporation?



RESEARCH LINKS
CLASSZONE.COM

Properties of Solids

Key Terms

- Crystalline solids
- Ionic solids
- Molecular solids
- Atomic solids
- Electron sea model
- Alloy
- Substitutional alloy
- Interstitial alloy

Objectives

- To learn about the types of crystalline solids
- To understand the interparticle forces in crystalline solids
- To learn how the bonding in metals determines metallic properties

A. The Solid State: Types of Solids

Solids play a very important role in our lives. The concrete we drive on, the trees that shade us, the windows we look through, the paper that holds this print, the diamond in an engagement ring, and the plastic lenses in eyeglasses are all important solids. Most solids, such as wood, paper, and glass, contain mixtures of various components. However, some natural solids, such as diamonds and table salt, are nearly pure substances.

Many substances form **crystalline solids**—those with a regular arrangement of their components. This is illustrated by the partial structure of sodium chloride shown in **Figure 14.7**. The highly ordered arrangement of the components in a crystalline solid produces beautiful, regularly shaped crystals such as those shown in **Figure 14.8**.

There are many different types of crystalline solids. For example, both sugar and salt have beautiful crystals that we can easily see. However, although both dissolve readily in water, the properties of the resulting solutions are quite different. The salt solution readily conducts an electric current; the sugar solution does not. This behavior arises from the different natures of the components in these two solids. Common salt, NaCl, is an ionic solid that contains Na^+ and Cl^- ions. When solid sodium chloride dissolves in water, sodium ions and chloride ions are distributed throughout the resulting solution. These ions are free to move through the solution to conduct an electric current. Table sugar (sucrose), on the other hand, is composed of neutral molecules that are dispersed throughout the water when the solid dissolves. No ions are present, and the resulting solution does not conduct electricity. These examples illustrate two important types of crystalline solids: **ionic solids**, represented by sodium chloride; and **molecular solids**, represented by sucrose.

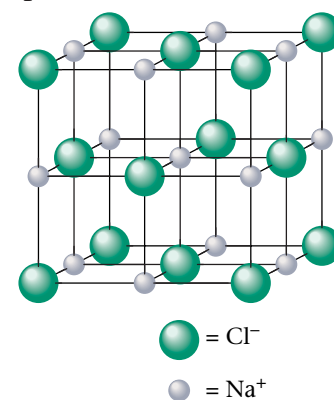


Figure 14.7

The regular arrangement of sodium and chloride ions in sodium chloride, a crystalline solid

Crystalline solid

A solid characterized by the regular arrangement of its components

Ionic solid

A solid that contains cations and anions

Molecular solid

A solid composed of molecules



Quartz, SiO_2



Rock salt, NaCl



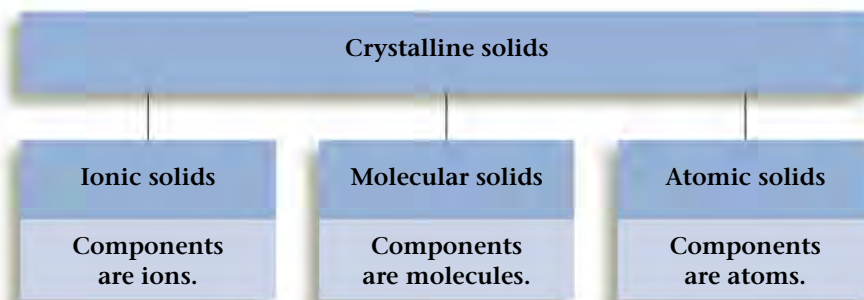
Red beryl, $\text{Be}_3\text{Al}_2\text{Si}_6\text{O}_{18}$

Figure 14.8

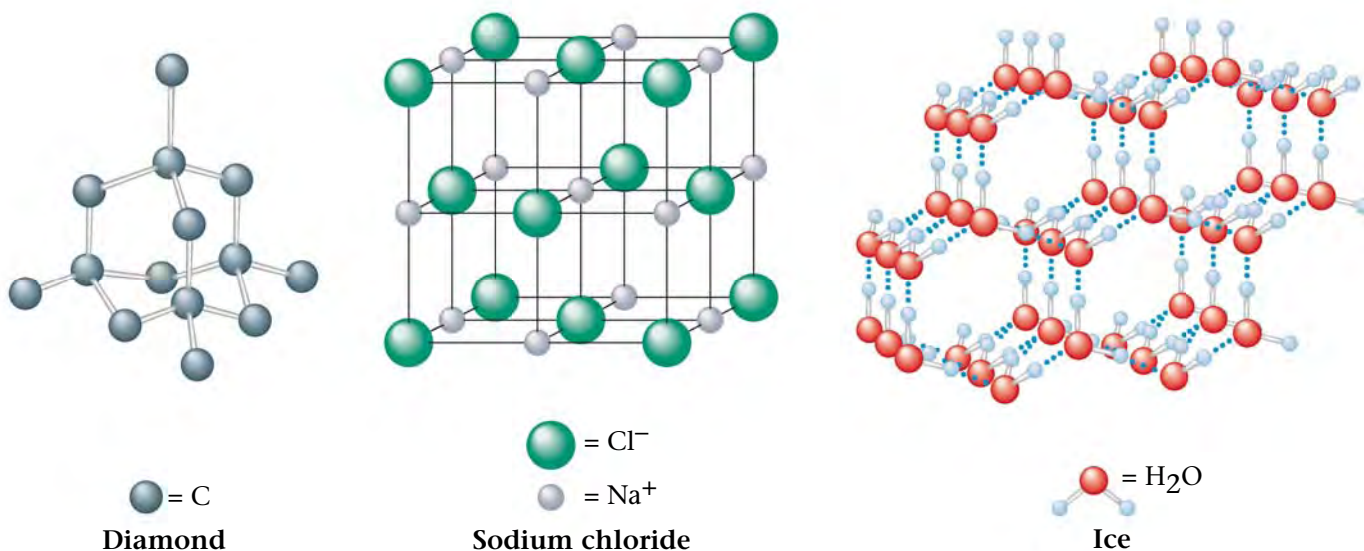
Several crystalline solids

A third type of crystalline solid is represented by elements such as graphite and diamond (both pure carbon), boron, silicon, and all metals. These substances, which contain atoms covalently bonded to each other, are called **atomic solids**.

We have seen that crystalline solids can be grouped conveniently into three classes.



Notice that the names of the three classes come from the components of the solid. An ionic solid contains ions, a molecular solid contains molecules, and an atomic solid contains atoms. Examples of the three types of solids are shown below.



The properties of a solid are determined primarily by the nature of the forces that hold the solid together. For example, although argon, copper, and diamond are all atomic solids (their components are atoms), they have strikingly different properties. Argon has a very low melting point ($-189\text{ }^{\circ}\text{C}$), whereas diamond and copper melt at high temperatures (about $3500\text{ }^{\circ}\text{C}$ and $1083\text{ }^{\circ}\text{C}$, respectively). Copper is an excellent conductor of electricity (it is widely used for electrical wires), whereas both argon and diamond are insulators. The shape of copper can easily be changed; it is both malleable (will form thin sheets) and ductile (can be pulled into a wire). Diamond, on the other hand, is the hardest natural substance known. The marked differences in properties among these three atomic solids are due to differences in bonding. We will explore the bonding in solids next.

Copper can be formed into many different kinds of objects.



Atomic solid

A solid that contains atoms at the lattice points

Information

The internal forces in a solid determine many of the properties of the solid.

B. Bonding in Solids

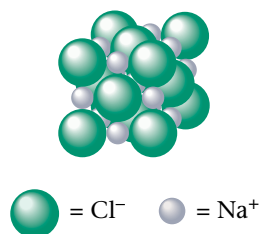
We have seen that crystalline solids can be divided into three classes, depending on the fundamental particle or unit of the solid. Ionic solids consist of oppositely charged ions packed together, molecular solids contain molecules, and atomic solids have atoms as their fundamental particles. Examples of the various types of solids are given in **Table 14.4**.

Active Reading Question

Give an example of an ionic solid, a molecular solid, and an atomic solid.

Ionic Solids

Ionic solids are stable substances with high melting points that are held together by the strong forces that exist between oppositely charged ions. The structures of ionic solids can be visualized best by thinking of the ions as spheres packed together as efficiently as possible. For example, in NaCl the larger Cl^- ions are packed together much like one would pack balls in a box. The smaller Na^+ ions occupy the small spaces (“holes”) left among the spherical Cl^- ions.



Molecular Solids

In a molecular solid the fundamental particle is a molecule. Examples of molecular solids include ice (contains H_2O molecules), dry ice (contains CO_2 molecules), sulfur (contains S_8 molecules), and white phosphorus (contains tetrahedral P_4 molecules).

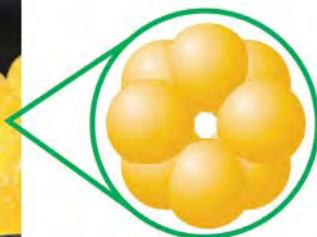
Table 14.4

Examples of the Various Types of Solids

Type of Solid	Examples	Fundamental Unit(s)
ionic	sodium chloride, $\text{NaCl}(s)$	Na^+ , Cl^- ions
ionic	ammonium nitrate, $\text{NH}_4\text{NO}_3(s)$	NH_4^+ , NO_3^- ions
molecular	dry ice, $\text{CO}_2(s)$	CO_2 molecules
molecular	ice, $\text{H}_2\text{O}(s)$	H_2O molecules
atomic	diamond, $\text{C}(s)$	C atoms
atomic	iron, $\text{Fe}(s)$	Fe atoms
atomic	argon, $\text{Ar}(s)$	Ar atoms



Sulfur crystals

 S_8 

White phosphorus

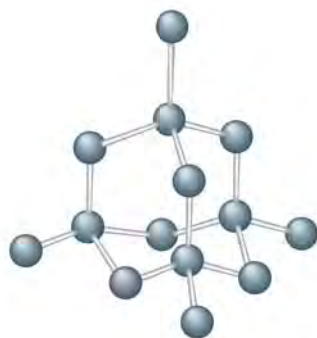
 P_4

Molecular solids tend to melt at relatively low temperatures because the intermolecular forces that exist among the molecules are relatively weak. If the molecule has a dipole moment, dipole–dipole forces hold the solid together. In solids with nonpolar molecules, London dispersion forces hold the solid together.

Atomic Solids

The properties of atomic solids vary greatly because of the different ways in which the fundamental particles, the atoms, can interact with each other. For example, the solids of the Group 8 elements have very low melting (freezing) points (see Table 14.2), because these atoms, having filled valence orbitals, cannot form covalent bonds with each other. So the forces in these solids are the relatively weak London dispersion forces.

On the other hand, diamond, a form of solid carbon, is one of the hardest substances known and has an extremely high melting point (about 3500 °C). The incredible hardness of diamond arises from the very strong covalent carbon–carbon bonds in the crystal, which lead to a giant molecule. In fact, the entire crystal can be viewed as one huge molecule. A small part of the diamond structure is represented below.

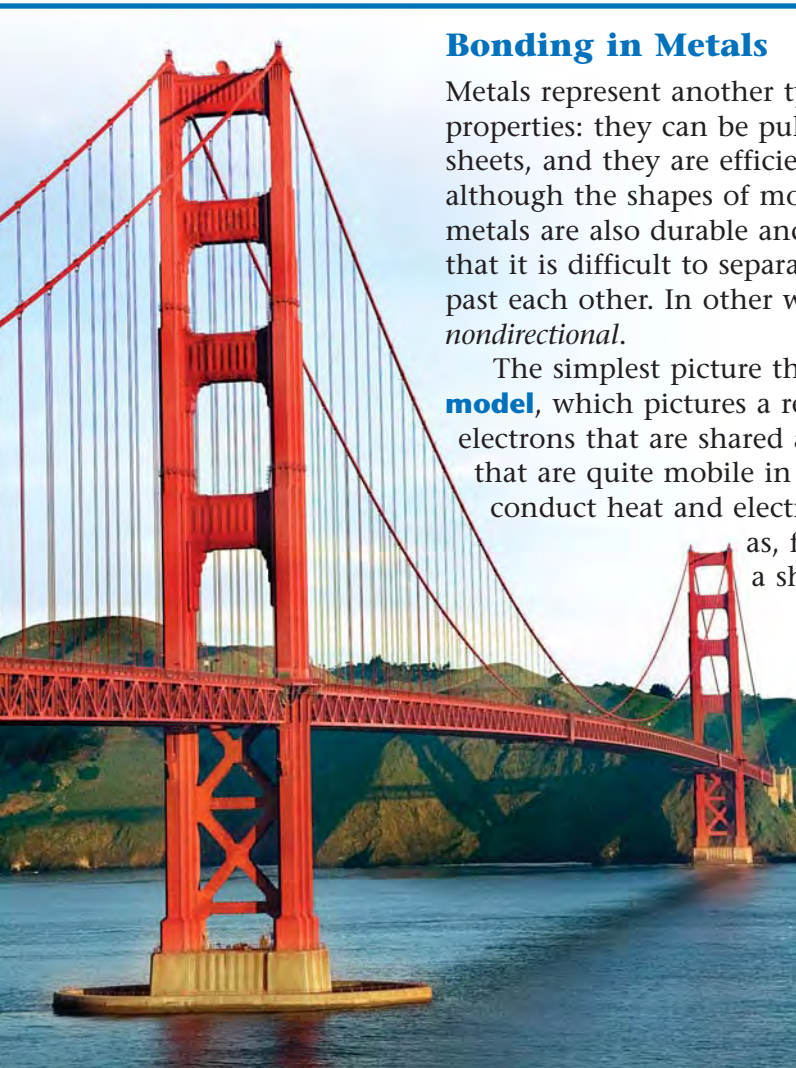


In diamond each carbon atom is bound covalently to four other carbon atoms to produce a very stable solid. Several other elements also form solids whereby the atoms join together covalently to form giant molecules. Silicon and boron are examples.

At this point you might be asking yourself, “Why aren’t solids such as a crystal of diamond, which is a ‘giant molecule,’ classified as molecular solids?” The answer is that, by convention, a solid is classified as a molecular solid only if (like ice, dry ice, sulfur, and phosphorus) it contains small molecules. Substances like diamond that contain giant molecules are called network solids.



Diamonds are prized for use as jewelry because they reflect light and take millions of years to form in nature.



Bonding in Metals

Metals represent another type of atomic solid. Metals have familiar physical properties: they can be pulled into wires, they can be hammered into sheets, and they are efficient conductors of heat and electricity. However, although the shapes of most pure metals can be changed relatively easily, metals are also durable and have high melting points. These facts indicate that it is difficult to separate metal atoms but relatively easy to slide them past each other. In other words, the bonding in most metals is *strong* but *nondirectional*.

The simplest picture that explains these observations is the **electron sea model**, which pictures a regular array of metal atoms in a “sea” of valence electrons that are shared among the atoms in a nondirectional way and that are quite mobile in the metal crystal. The mobile electrons can conduct heat and electricity, and the atoms can be moved rather easily, as, for example, when the metal is hammered into a sheet or pulled into a wire.

Active Reading Question

What properties of a metal does the electron sea model explain?

Metals can be drawn into wires which can be used to construct cables like these on the Golden Gate bridge.

CELEBRITY CHEMICAL

Titanium (Ti)

Titanium is a wonderful structural material. It is 43% less dense than steel but has 30% greater yield strength than steel when alloyed with metals such as aluminum and tin. Titanium also has excellent resistance to fatigue and corrosion. Titanium is especially well suited for making bicycle frames. After their first ride on a bicycle with a titanium frame, most experienced cyclists find themselves searching for the right words to describe the “magic” of titanium. In fact, the magic of titanium results from its combination of toughness, stretchability, and resilience. It resists deforming under pedaling loads but does not transmit road shocks to the rider nearly as much as steel.

Titanium, a lustrous metal with a melting point of 1667 °C, is quite abundant in the earth’s crust,

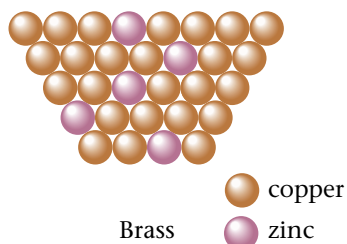
ranking ninth among the elements. Unfortunately, it is difficult to obtain from its ores in pure form. Titanium becomes very brittle when trace impurities such as C, N, and O are present, and it must be processed with great care. Its stretchability also makes it difficult to machine on a lathe to make specific parts. The resulting bicycle makes all these troubles worthwhile, however.

Titanium is also used to make artificial hip joints.



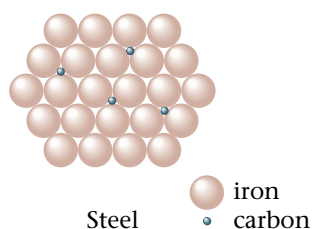
Because of the nature of the metallic crystal, other elements can be introduced relatively easily to produce substances called alloys. An **alloy** is best defined as *a substance that contains a mixture of elements and has metallic properties*. There are two common types of alloys.

In a **substitutional alloy** some of the host metal atoms are *replaced* by other metal atoms of similar sizes. For example, in brass approximately one-third of the atoms in the host copper metal have been replaced by zinc atoms.



Sterling silver (93% silver and 7% copper) and pewter (85% tin, 7% copper, 6% bismuth, and 2% antimony) are other examples of substitutional alloys.

An **interstitial alloy** is formed when some of the interstices (holes) among the closely packed metal atoms are occupied by atoms much smaller than the host atoms.



Steel, the best-known interstitial alloy, contains carbon atoms in the “holes” of an iron crystal. The presence of interstitial atoms changes the properties of the host metal. Pure iron is relatively soft, ductile, and malleable because of the absence of strong directional bonding. The spherical metal atoms can be moved rather easily with respect to each other. However, when carbon, which forms strong directional bonds, is introduced into an iron crystal, the presence of the directional carbon–iron bonds makes the resulting alloy harder, stronger, and less ductile than pure iron. The amount of carbon directly affects the properties of steel. *Mild steels* (containing less than 0.2% carbon) are still ductile and malleable and are used for nails, cables, and chains. *Medium steels* (containing 0.2–0.6% carbon) are harder than mild steels and are used in rails and structural steel beams. *High-carbon steels* (containing 0.6–1.5% carbon) are tough and hard and are used for springs, tools, and cutlery.

Many types of steel also contain elements in addition to iron and carbon. Such steels are often called *alloy steels* and can be viewed as being mixed interstitial (carbon) and substitutional (other metals) alloys. An example is stainless steel, which has chromium and nickel atoms substituted for some of the iron atoms. The addition of these metals greatly increases the steel’s resistance to corrosion.

Alloy

A substance that contains a mixture of elements and has metallic properties



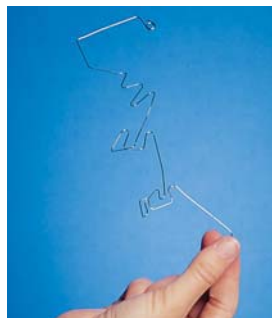
A steel sculpture in Chicago

Metal with a Memory



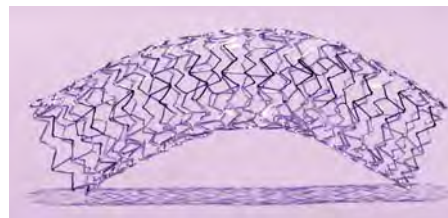
An upset mother walks into the optical shop carrying her mangled pair of \$400 glasses. Her child had gotten into her purse, found her glasses, and twisted them into a pretzel. She hands them to the optometrist with little hope that they can be fixed. The optometrist says not to worry and drops the glasses into a dish of warm water where the glasses spring back into their original shape. The optometrist hands the restored glasses to the woman and says there is no charge for repairing them.

How can the frames “remember” their original shape when placed in warm water? The answer is a nickel-titanium alloy called Nitinol that was developed in the late 1950s and early 1960s at the Naval Ordnance Laboratory in White Oak, Maryland, by William J. Buehler. (The name Nitinol comes from *Nickel Titanium Naval Ordnance Laboratory*.) Nitinol has an amazing ability to remember a shape originally impressed in it.



Nitinol has many medical applications, including hooks used by orthopedic surgeons to attach ligaments and tendons to bones and stents that are inserted in partially blocked arteries to keep them open.

Before insertion, the stents are crushed down so they can be fed into the artery through a thin tube. As the crushed stent heats up in the blood it returns to its original shape.



Nitinol is also used for the wires in braces to straighten teeth. The wires are formed into the correct shape and then inserted into the patient’s mouth. As they warm up the wire tries to return to its original shape putting constant pressure on the teeth out of position. This constant pressure speeds up the moving and straightening process!

EXAMPLE 14.4**Identifying Types of Crystalline Solids**

Name the type of crystalline solid formed by each of the following substances:

- ammonia
- iron
- cesium fluoride
- argon
- sulfur

Solution

- Solid ammonia contains NH_3 molecules, so it is a molecular solid.
- Solid iron contains iron atoms as the fundamental particles. It is an atomic solid.
- Solid cesium fluoride contains the Cs^+ and F^- ions. It is an ionic solid.
- Solid argon contains argon atoms, which cannot form covalent bonds to each other. It is an atomic solid.
- Sulfur contains S_8 molecules, so it is a molecular solid.

 **Practice Problem • Exercise 14.4**

Name the type of crystalline solid formed by each of the following substances:

- sulfur trioxide
- barium oxide
- gold

**SECTION 14.3****REVIEW QUESTIONS**

- Salt and sugar are both crystalline solids. How are they different?
- What are the forces holding each type of crystalline solid together? Predict which type would have the lowest boiling point. Explain your choice.
- What is the difference between a substitutional alloy and an interstitial alloy?
- How are metals different from other atomic solids such as diamond?
- Name the type of crystalline solid formed by each of the following substances.
 - ice
 - copper
 - calcium chloride

**RESEARCH LINKS**
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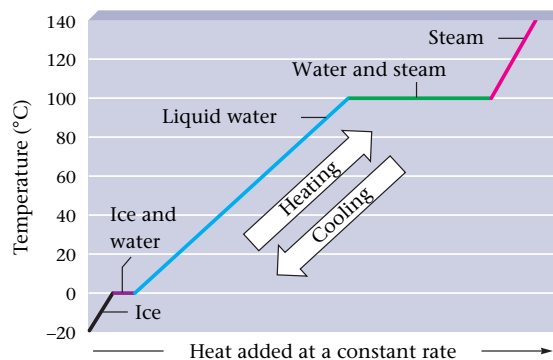
Key Terms

14.1 Intermolecular forces
 Intramolecular forces
 Dipole–dipole attraction
 Hydrogen bonding
 London dispersion forces
 Normal boiling point
 Heating/cooling curve
 Normal freezing point
 Molar heat of fusion
 Molar heat of vaporization

Key Ideas

14.1 Intermolecular Forces and Phase Changes

- Intermolecular forces are the forces that occur among the molecules in a substance.
- The classes of intermolecular forces are
 - Dipole–dipole: due to the attractions among molecules that have dipole moments
 - Hydrogen bonding: particularly strong dipole–dipole forces that occur when hydrogen is bound to a highly electronegative atom (such as O, N, F)
 - London dispersion forces: forces that occur among nonpolar molecules (or atoms) when accidental dipoles develop in the electron “cloud”
- Phase changes can occur when energy enters or leaves a compound.
 - As energy enters a system a substance can change from a solid to a liquid to a gas.



- The molar heat of fusion is the energy required to melt 1 mol of a solid substance.
- The normal melting point is the temperature at which a solid melts at a pressure of 1 atm.
- The normal boiling point is the temperature at which a liquid boils at a pressure of 1 atm.
- A substance with relatively large intermolecular forces requires relatively large amounts of energy to produce phase changes.
 - ◆ Water is an example of a substance with large intermolecular forces (hydrogen bonding).

14.2 Vaporization (evaporation)

Condensation
 Vapor pressure

14.2 Vapor Pressure and Boiling Point

- The equilibrium pressure of vapor over a liquid in a closed container is called the vapor pressure.
- The vapor pressure of a liquid is a balance between condensation and evaporation.
- Vapor pressure is relatively low for a substance (such as water) with large intermolecular forces.
- The boiling point of a liquid occurs at a temperature at which the vapor pressure of the liquid equals the atmospheric pressure.

14.3 Crystalline solids

Ionic solids

Molecular solids

Atomic solids

Electron sea model

Alloy

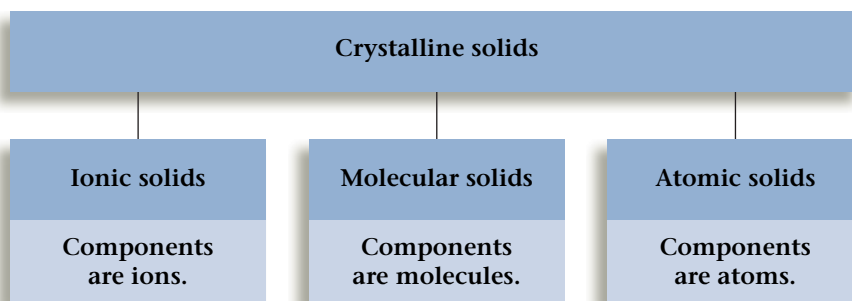
Substitutional alloy

Interstitial alloy

14.3 Properties of Solids

■ Types of solids

- Ionic solid—the components are ions.
- Molecular solid—the components are molecules.
- Atomic solids—the components are atoms.



■ Bonding in solids

- Ionic solids are bound together by the attractions among the oppositely charged ions.
- Molecular solids are bound together by the intermolecular forces among the molecules.
- Atomic solids
 - ◆ The solids produced from noble gases have weak London dispersion forces.
 - ◆ Many atomic solids form “giant molecules” that are held together by covalent bonds
- Metals are held together by nondirectional covalent bonds (called the electron sea model) among the closely packed atoms.
 - Metals form alloys of two types.
 - ◆ Substitutional: different atoms are substituted for the host metal atoms.
 - ◆ Interstitial: small atoms are introduced into the “holes” in the metallic structure.

All exercises with **blue numbers** have answers in the back of this book.

14.1 Intermolecular Forces and Phase Changes

A. Intermolecular Forces

1. What is a dipole–dipole attraction? Give three examples of liquid substances in which you would expect dipole–dipole attractions to be large.
2. How is the strength of dipole–dipole interactions related to the *distance* between polar molecules? Are dipole–dipole forces short-range or long-range forces?
3. What is meant by *hydrogen bonding*? Give three examples of substances that would be expected to exhibit hydrogen bonding in the liquid state.
4. The normal boiling point of water is unusually high, compared to the boiling points of H_2S , H_2Se , and H_2Te . Explain this observation in terms of the *hydrogen bonding* that exists in water, but that does not exist in the other compounds.
5. Why are the dipole–dipole interactions between polar molecules *not* important in the vapor phase?
6. Although the noble gas elements are monatomic and could not give rise to dipole–dipole forces or hydrogen bonding, these elements still can be liquefied and solidified. Explain.
7. Discuss the types of intermolecular forces acting in the liquid state of each of the following substances.

a. Kr	c. NF_3
b. S_8	d. H_2O
8. The boiling points of the noble gas elements are listed below. Comment on the trend in the boiling points. Why do the boiling points vary in this manner?

He	–272 °C	Kr	–152.3 °C
Ne	–245.9 °C	Xe	–107.1 °C
Ar	–185.7 °C	Rn	–61.8 °C
9. When 50 mL of liquid water at 25 °C is added to 50 mL of ethanol (ethyl alcohol), also at 25 °C, the combined volume of the mixture is considerably *less* than 100 mL. Give a possible explanation.

B. Water and Its Phase Changes

10. What are some important physical properties of water? How do these properties of water help moderate the earth's environment?

11. Describe some uses, both in nature and in industry, of water as a *cooling agent*.
12. We all have observed that ice *floats* on liquid water. Why? What is unusual about this?
13. Discuss some implications of the fact that, unlike most substances, water *expands* in volume when it freezes.
14. Describe, on both a microscopic and a macroscopic basis, what happens to a sample of water as it is heated from room temperature to 50 °C above its normal boiling point.
15. Figure 14.2 presents the *cooling curve* for water. Discuss the meaning of the different portions of this curve (for example, explain what each flat section and each sloping section represents).

C. Energy Requirements for the Changes of State

16. Describe in detail the microscopic processes that take place when a solid melts.
17. Describe in detail the microscopic processes that take place when a liquid boils.
18. Explain the difference between *intramolecular* and *intermolecular* forces.
19. Which type of forces (*intramolecular* or *intermolecular*) must be overcome to melt a solid or vaporize a liquid?
20. Discuss the similarities and differences between the arrangements of molecules and the forces between molecules in liquid water versus steam, and in liquid water versus ice.
21. The following data have been collected for substance X. Construct a heating curve for substance X. (The drawing does not need to be absolutely to scale, but it should clearly show relative differences.)

normal melting point	–15 °C
molar heat of fusion	2.5 kJ/mol
normal boiling point	134 °C
molar heat of vaporization	55.3 kJ/mol

22. The molar heats of fusion and vaporization for water are 6.02 kJ/mol and 40.6 kJ/mol, respectively, and the specific heat capacity of liquid water is 4.18 J/g °C. What quantity of heat energy is required to melt 25.0 g of ice at 0 °C? What quantity of heat is required to vaporize 37.5 g of liquid water at 100. °C? What quantity of heat is required to warm 55.2 g of liquid water from 0 °C to 100. °C?

23. Given that the specific heat capacities of ice and steam are $2.06 \text{ J/g } ^\circ\text{C}$ and $2.03 \text{ J/g } ^\circ\text{C}$, respectively, and considering the information about water given in exercise 22, calculate the total quantity of heat evolved when 10.0 g of steam at $200. ^\circ\text{C}$ is condensed, cooled, and frozen to ice at $-50. ^\circ\text{C}$.

14.2 Vapor Pressure and Boiling Point

A. Evaporation and Vapor Pressure

24. Describe, on a microscopic basis, the processes of *evaporation* and *condensation*. Which process requires the input of energy?
25. What is *vapor pressure*? On a microscopic basis, how does a vapor pressure develop in a closed flask containing a small amount of liquid? What processes are going on in the flask?
26. What do we mean by a *dynamic equilibrium*? Describe how the development of a vapor pressure above a liquid represents such an equilibrium.
27. Describe an experimental method that could be used to determine the vapor pressure of a volatile liquid.
28. Which substance in each pair would be expected to be more volatile at a particular temperature? Explain your reasoning.
- $\text{H}_2\text{O}(l)$ or $\text{H}_2\text{S}(l)$
 - $\text{H}_2\text{O}(l)$ or $\text{CH}_3\text{OH}(l)$
 - $\text{CH}_3\text{OH}(l)$ or $\text{CH}_3\text{CH}_2\text{OH}(l)$
29. Although water and ammonia differ in molar mass by only one unit, the boiling point of water is over $100 ^\circ\text{C}$ higher than that of ammonia. What forces in liquid water that do *not* exist in liquid ammonia could account for this observation?
30. Two molecules that contain the same number of each kind of atom but that have different molecular structures are said to be *isomers* of each other. For example, both ethyl alcohol and dimethyl ether (shown below) have the formula $\text{C}_2\text{H}_6\text{O}$ and are isomers. Based on considerations of intermolecular forces, which substance would you expect to be more volatile? Which would you expect to have the higher boiling point? Explain.



B. Boiling Point and Vapor Pressure

31. Potatoes and other vegetables often require longer cooking times at higher elevations. Explain why it takes longer to cook something at higher elevations.
32. Why is the boiling temperature of water less than $100 ^\circ\text{C}$ at high altitudes?

14.3 Properties of Solids

A. The Solid State: Types of Solids

33. What are crystalline solids? What kind of microscopic structure do such solids have? How is this microscopic structure reflected in the macroscopic appearance of such solids?
34. On the basis of the smaller units that make up the crystals, cite three types of crystalline solids. For each type of crystalline solid, give an example of a substance that forms that type of solid.

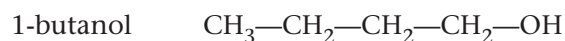
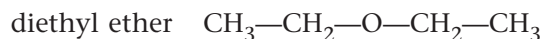
B. Bonding in Solids

35. How do *ionic* solids differ in structure from *molecular* solids? What are the fundamental particles in each? Give two examples of each type of solid and indicate the individual particles that make up the solids in each of your examples.
36. How do the physical properties of ionic solids, in general, differ from the properties of molecular solids? Give an example of each to illustrate your discussion.
37. Ionic solids are generally considerably harder than most molecular solids. Explain.
38. Ionic solids typically have melting points hundreds of degrees higher than the melting points of molecular solids. Explain.
39. What types of forces exist between the individual particles in an ionic solid? Are these forces relatively strong or relatively weak?
40. Compare the intermolecular forces that exist in a sample of solid krypton, Kr, with those that exist in diamond, C.
41. What is an *alloy*? Explain the differences in structure between substitutional and interstitial alloys. Give an example of each type.
42. Explain how the properties of a metal may be modified by alloying the metal with some other substance. Discuss, in particular, how the properties of iron are modified in producing the various types of steel.

Critical Thinking

For exercises 43–52 choose one of the following terms to match the definition or description given.

- a. alloy
 - b. specific heat
 - c. crystalline solid
 - d. dipole–dipole attraction
 - e. equilibrium vapor pressure
 - f. intermolecular
 - g. intramolecular
 - h. ionic solids
 - i. London dispersion forces
 - j. molar heat of fusion
 - k. molar heat of vaporization
 - l. molecular solids
 - m. normal boiling point
 - n. semiconductor
- 43.** boiling point at pressure of 1 atm
- 44.** energy required to melt 1 mol of a substance
- 45.** forces between atoms in a molecule
- 46.** forces between molecules in a solid
- 47.** instantaneous dipole forces for nonpolar molecules
- 48.** lining up of opposite charges on adjacent polar molecules
- 49.** maximum pressure of vapor that builds up in a closed container
- 50.** mixture of elements having metallic properties overall
- 51.** repeating arrangement of component species in a solid
- 52.** solids that melt at relatively low temperatures
- 53.** In carbon compounds a given group of atoms can often be arranged in more than one way. This means that more than one structure may be possible for the same atoms. For example, both the molecules diethyl ether and 1-butanol have the same number of each type of atom, but they have different structures and are said to be *isomers* of one another.



Which substance would you expect to have the larger vapor pressure? Why?

- 54.** Which of the substances in each of the following sets would be expected to have the highest boiling point? Explain why.
- a. Ga, KBr, O_2
 - b. Hg, NaCl, He
 - c. H_2 , O_2 , H_2O
- 55.** When a person has a severe fever, one therapy to reduce the fever is an “alcohol rub.” Explain how the evaporation of alcohol from the person’s skin removes heat energy from the body.
- 56.** What is steel? How do the properties of steel differ from the properties of its constituents?
- 57.** Which is stronger, a dipole–dipole attraction between two molecules or a covalent bond between two atoms within the same molecule? Explain.
- 58.** What are London dispersion forces and how do they arise in a nonpolar molecule? Are London forces typically stronger or weaker than dipole–dipole attractions between polar molecules? Are London forces stronger or weaker than covalent bonds? Explain.
- 59.** Discuss the types of intermolecular forces acting in the liquid state of each of the following substances.
- a. N_2
 - b. NH_3
 - c. He
 - d. CO_2 (linear, nonpolar)
- 60.** What do we mean when we say a liquid is *volatile*? Do volatile liquids have large or small vapor pressures? What types of intermolecular forces occur in highly volatile liquids?
- 61.** Although methane, CH_4 , and ammonia, NH_3 , differ in molar mass by only one unit, the boiling point of ammonia is over 100°C higher than that of methane (a nonpolar molecule). Explain.

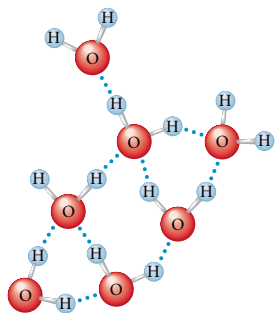
1 Which of the following correctly pairs the molecule with the strongest intermolecular attraction present for that case?

- A Li_2O : dipole–dipole; F_2 : London dispersion; SO_2 : dipole–dipole
- B Li_2O : dipole–dipole; F_2 : dipole–dipole; SO_2 : dipole–dipole
- C Li_2O : ion–ion; F_2 : London dispersion; SO_2 : dipole–dipole
- D Li_2O : ion–ion; F_2 : dipole–dipole; SO_2 : London dispersion

2 The difference in the strengths of the intermolecular forces among methane (CH_4) molecules and ammonia (NH_3) molecules is mainly due to

- A the fact that ammonia is polar and methane is nonpolar.
- B the fact that methane is polar and ammonia is nonpolar.
- C the fact that methane (with four hydrogens) has a greater chance to exhibit hydrogen bonding than ammonia (with three hydrogens).
- D the different molar masses.

3 In the diagram below, which lines represent the hydrogen bonds?



- A The dotted lines between the hydrogen atoms of one water molecule and the oxygen atoms of a different water molecule.
- B The solid lines between a hydrogen atom and oxygen atom in the same water molecule.
- C Both the solid lines and dotted lines represent the hydrogen bonds.
- D There are no hydrogen bonds represented in the diagram.

4 Which of the following decreases as the strength of intermolecular forces increases?

- A boiling point
- B melting point
- C vapor pressure
- D surface tension

5 The vapor pressure of water at $100.0\text{ }^\circ\text{C}$ is

- A 0.53 atm
- B 0.95 atm
- C 1.0 atm
- D 1.2 atm

6 Which of the following is not an example of an atomic solid?

- A iron
- B water
- C argon
- D graphite

7 Table sugar is an example of

- A an atomic solid.
- B an ionic solid.
- C a network solid.
- D a molecular solid.

For questions 8–11, identify the strongest intermolecular force exhibited by each of the following molecules.

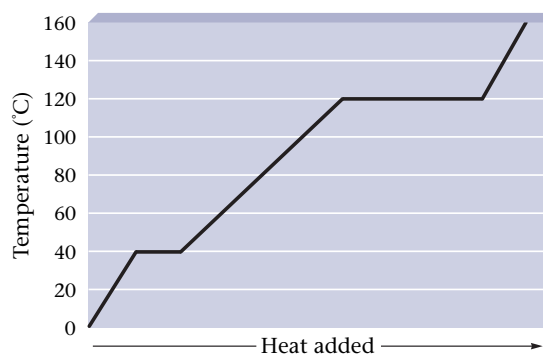
8 CH_4

9 O_2

10 CO

11 K_2O

12 Use the heating-cooling curve below to answer the following questions



- a. What is the freezing point of the liquid?
- b. What is the boiling point of the liquid?
- c. Which is greater, the heat of fusion or heat of vaporization? Explain.

LOOKING AHEAD

15.1 Forming Solutions

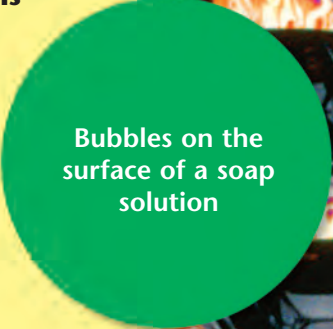
- A. Solubility
 - Solubility of Ionic Substances
 - Solubility of Polar Substances
 - Substances Insoluble in Water
 - How Substances Dissolve
- B. Solution Composition: An Introduction
- C. Factors Affecting the Rate of Dissolving

15.2 Describing Solution Composition

- A. Solution Composition: Mass Percent
- B. Solution Composition: Molarity
- C. Dilution

15.3 Properties of Solutions

- A. Stoichiometry of Solution Reactions
- B. Neutralization Reactions
- C. Normality
- D. Boiling Point and Freezing Point



Bubbles on the surface of a soap solution



IN YOUR LIFE

Solutions are very important to all of us. In fact we can walk, talk, laugh, and work only because millions of chemical reactions are occurring every second in the solutions in our bodies. We also encounter many solutions as we go through a typical day—the liquid soap we use in our morning shower, the cranberry juice we drink for breakfast, the sports drinks we have for lunch, the gasoline we put in our cars and the “energy drink” we may have in the middle of the afternoon. Even the “pure water” from the water fountain or a water bottle is a solution—it contains dissolved minerals from the earth, chlorine compounds to disinfect it, and traces of many other substances. A solution that we often take for granted is the one we live in—the air is a solution of nitrogen (N_2), oxygen (O_2), water vapor, carbon dioxide, and traces of many other gases. Solutions are clearly essential to our existence.



Brass, a solid solution of copper and zinc, is used to make musical instruments and many other objects.

WHAT DO YOU KNOW?

Prereading Questions

1. What is meant by the term *solution* (as in, a solution of salt and water)?
2. What happens when you mix water and oil?
3. What can you do to get sugar to dissolve more quickly in water?
4. A certain cookie is made by mixing 150.0 g of dough with 25.0 g of chocolate chips. What is the mass percent of chocolate chips in the cookie?
5. What is meant by a precipitation reaction?
6. What is the net ionic equation for all reactions of a strong acid and a strong base?
7. Why is salt scattered on roads during the winter in cold weather climates?

SECTION 15.1

Forming Solutions

Key Terms

- Solution
- Solvent
- Solute
- Aqueous solution
- Saturated
- Unsaturated
- Supersaturated
- Concentrated
- Dilute

Objectives

- To understand the process of dissolving
- To learn why certain substances dissolve in water
- To learn qualitative terms describing the concentration of a solution
- To understand the factors that affect the rate at which a solid dissolves

A **solution** is a homogeneous mixture, a mixture in which the components are uniformly intermingled. This means that a sample from one part is the same as a sample from any other part. For example, the first sip of coffee is the same as the last sip.

The atmosphere that surrounds us is a gaseous solution containing $O_2(g)$, $N_2(g)$, and other gases randomly dispersed. Solutions can also be solids. For example, brass is a homogeneous mixture—a solution—of copper and zinc.

These examples illustrate that a solution can be a gas, a liquid, or a solid (see **Table 15.1**). The substance present in the largest amount is called the **solvent**, and the other substance or substances are called **solutes**. For example, when we dissolve a teaspoon of sugar in a glass of water, the sugar is the solute and the water is the solvent.

Aqueous solutions are solutions with water as the solvent. Because they are so important, in this chapter we will concentrate on the properties of aqueous solutions.

Solution

A homogeneous mixture

Solvent

The dissolving medium in a solution

Solute

The substance dissolved in the solvent to make a solution

Aqueous solution

A solution with water as a solvent

Table 15.1

Various Types of Solutions

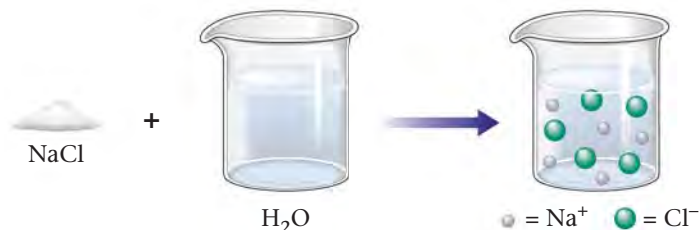
Example	State of Solution	Original State of Solute	State of Solvent
air, natural gas	gas	gas	gas
antifreeze in water	liquid	liquid	liquid
brass	solid	solid	solid
carbonated water (soda)	liquid	gas	liquid
seawater, sugar solution	liquid	solid	liquid

A. Solubility

What happens when you put a teaspoon of sugar in your iced tea and stir it, or when you add salt to water for cooking vegetables? Why do the sugar and salt “disappear” into the water? What does it mean when something dissolves—that is, when a solution forms?

Solubility of Ionic Substances

We saw in Chapter 8 that when sodium chloride dissolves in water, the resulting solution conducts an electric current. This convinces us that the solution contains *ions* that can move (this is how the electric current is conducted). The dissolving of solid sodium chloride in water is shown below.

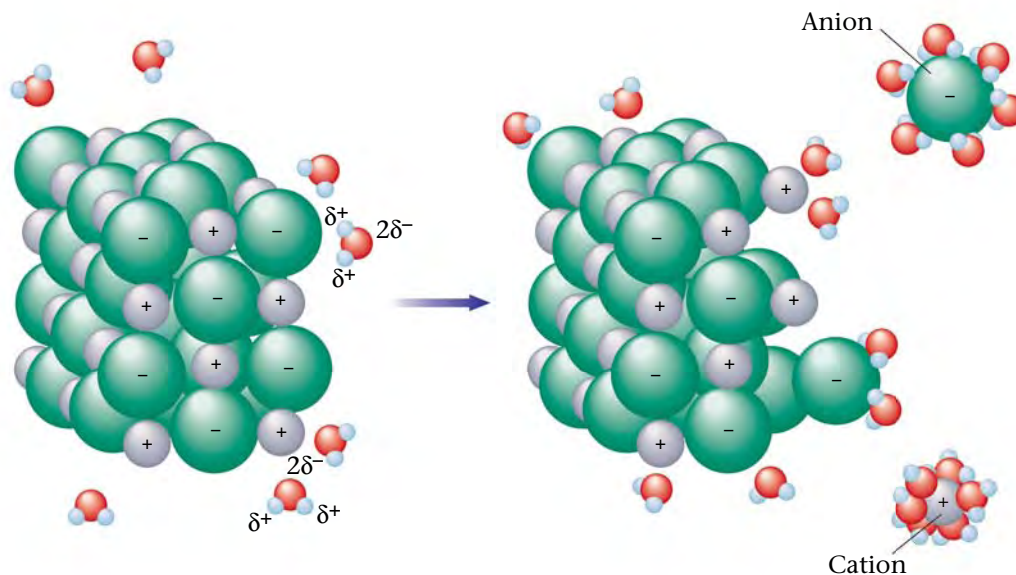


In the solid state the ions are packed closely together. However, when the solid dissolves, the ions are separated and dispersed throughout the solution. The strong ionic forces that hold the sodium chloride crystal together are overcome by the strong attractions between the ions and the polar water molecules. This process is represented in **Figure 15.1**. Notice that each polar water molecule orients itself in a way to maximize its attraction with a Cl^- or Na^+ ion. The negative end of a water molecule is attracted to a Na^+ ion, while the positive end is attracted to a Cl^- ion. The strong forces holding the positive and negative ions in the solid are replaced by strong water–ion interactions, and the solid dissolves (the ions disperse).

It is important to remember that when an ionic substance (such as a salt) dissolves in water, it breaks up into *individual* cations and anions, which are dispersed in the water. For instance, when ammonium nitrate, NH_4NO_3 , dissolves in water, the resulting solution contains NH_4^+ and NO_3^- ions, which move around independently. This process can be represented as



where *(aq)* indicates that the ions are surrounded by water molecules.



Information

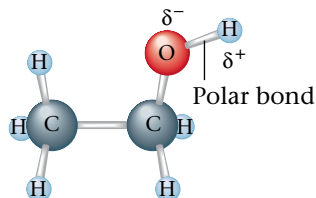
Cations are positive ions.
Anions are negative ions.

Figure 15.1

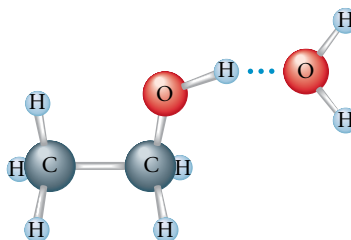
Polar water molecules interact with the positive and negative ions of a salt. These interactions replace the strong ionic forces holding the ions together in the undissolved solid, thus assisting in the dissolving process.

Solubility of Polar Substances

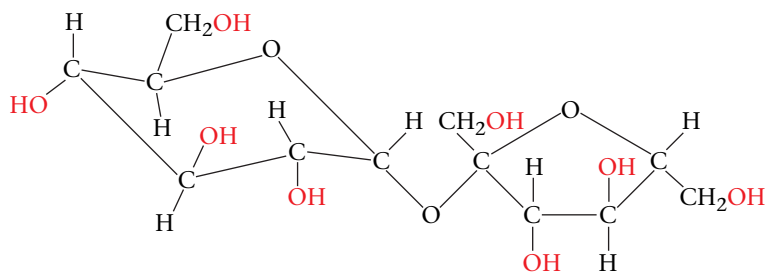
Water also dissolves many nonionic substances. Sugar is one example of a nonionic solute that is very soluble in water. Another example is ethanol, C_2H_5OH . Why is ethanol so soluble in water? The answer lies in the structure of the ethanol molecule.



The molecule contains a polar O—H bond like those in water, which makes it very compatible with water. Just as hydrogen bonds form among water molecules in pure water (see p. 490), ethanol molecules can form hydrogen bonds with water molecules in a solution of the two.



The sugar molecule is shown below. Common table sugar has the chemical name sucrose.



Notice that this molecule has many polar —OH groups, each of which can hydrogen-bond to a water molecule. Because of the attractions between sucrose and water molecules, solid sucrose is quite soluble in water.

Active Reading Questions

1. Does water generally dissolve polar or nonpolar substances? Explain.
2. What is similar about table salt and table sugar that allows them both to be soluble in water?

How Substances Dissolve

Notice in Figure 15.2 that the water molecules in liquid water are associated with each other by hydrogen-bonding interactions. For a solute to dissolve in water, a “hole” must be made in the water structure for each solute particle. This will occur only if the lost water–water interactions are replaced by similar water–solute interactions.

In the case of sodium chloride, strong interactions occur between the polar water molecules and the Na^+ and Cl^- ions. This allows the sodium chloride to dissolve.

In the case of ethanol or sucrose, hydrogen-bonding interactions can occur between the O—H groups on these molecules and water molecules, making these substances soluble as well.

But oil molecules are not soluble in water, because the many water–water interactions that would have to be broken to make “holes” for these large molecules are not replaced by favorable water–solute interactions.

Active Reading Question

Consider two one liter samples of solutions each containing 1.0 mol solute. One is a solution of table salt and one is a solution of table sugar. If you could see highly magnified views (molecular level), how would the solutions compare?

These considerations account for the observed behavior that “like dissolves like.” In other words, we observe that a given solvent usually dissolves solutes that have polarities similar to its own. For example, water dissolves most polar solutes, because the solute–solvent interactions formed in the solution are similar to the water–water interactions present in the pure solvent. Likewise, nonpolar solvents dissolve nonpolar solutes. For example, dry-cleaning solvents used for removing grease stains from clothes are nonpolar liquids. “Grease” is composed of nonpolar molecules, so a nonpolar solvent is needed to remove a grease stain.



HANDS-ON CHEMISTRY

• MINI-LAB •

Colors in Motion

Materials

- large clear container
- water
- water soluble markers
- filter paper

Procedure

1. Add water to a large clear container so that it covers the bottom of the container to a depth of about $\frac{1}{4}$ inch. Cover the container.
2. Obtain a piece of filter paper and use water soluble markers to place 4 dots at least $\frac{1}{2}$ inch apart and $\frac{1}{2}$ inch from the bottom of the paper.

3. Place the filter paper in the container so that the bottom of the paper is under water. The dots should not be submerged.
4. Make observations for the next several minutes.

Results/Analysis

1. Why do some colors separate while others do not?
2. Why do some colors move farther than others up the paper?
3. How do the separation of the colors and their rate of movement relate to solubility and intermolecular forces?

B. Solution Composition: An Introduction

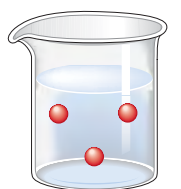
Even for very soluble substances, there is a limit to how much solute can be dissolved in a given amount of solvent. For example, when you add sugar to a glass of water, the sugar rapidly disappears at first. However, as you continue to add more sugar, at some point the solid no longer dissolves but collects at the bottom of the glass. When a solution contains as much solute as will dissolve at that temperature, we say it is **saturated**. If a solid solute is added to a solution already saturated with that solute, the added solid does not dissolve. A solution that has *not* reached the limit of solute that will dissolve in it is said to be **unsaturated**. When more solute is added to an unsaturated solution, it dissolves.

Sometimes when a solid is dissolved to the saturation limit at an elevated temperature and then allowed to cool, all of the solid may remain dissolved. This type of solution is called a **supersaturated** solution—it contains more dissolved solid than a saturated solution will hold at that temperature. A supersaturated solution is very unstable. Adding a crystal of the solid will cause immediate precipitation of solid until the solution reaches the saturation point.

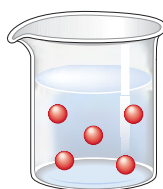
Although a chemical compound always has the same composition, a solution is a mixture and the amounts of the substances present can vary in different solutions. For example, coffee can be strong or weak. Strong coffee has more coffee dissolved in a given amount of water than weak coffee. To describe a solution completely, we must specify the amounts of solvent and solute. We sometimes use the qualitative terms *concentrated* and *dilute* to describe a solution. A relatively large amount of solute is dissolved in a **concentrated** solution (strong coffee is concentrated). A relatively small amount of solute is dissolved in a **dilute** solution (weak coffee is diluted).

As we will see in Section 15.2, the concentration of a solution is a measure of the amount of solute dissolved in a given volume of solution. We can use this idea to better understand the terms *concentrated* and *dilute*.

For example, compare the two solutions below in which each dot represents a solute molecule. Which solution is more concentrated?



Solution A
Volume = 1.0 L



Solution B
Volume = 1.0 L

The volume of each of the solutions is the same. However, the amount of solute in solution A is less than the amount of solute in solution B. Because of this, solution B is more concentrated than solution A.

When comparing the concentrations of two solutions you have to consider both the amount of solute and the volume of solution.

Saturated

Describes a solution that contains as much solute as will dissolve at that temperature

Unsaturated

Describes a solution in which more solute can dissolve than is dissolved already at that temperature

Supersaturated

Describes a solution that contains more solute than a saturated solution will hold at that temperature

Concentrated

Describes a solution in which a relatively large amount of solute is dissolved in a solution

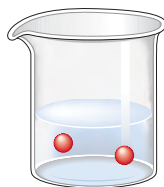
Dilute

Describes a solution in which a relatively small amount of solute is dissolved in a solution

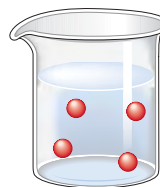


Sodium acetate crystallizing from a saturated solution

For example, the following two solutions have the same concentration.



Solution X
Volume = 1.0 L



Solution Y
Volume = 2.0 L

CRITICAL THINKING ?

One of the jobs performed by the U.S. Environmental Protection Agency (EPA) is to make sure our drinking water is safe.

What if the EPA insisted that all drinking water had to be 100% pure? How would it affect our drinking water supply?

Although the volume of solution Y is twice the volume of solution X the amount of solute in solution Y is also twice as much as that in Solution X.

Active Reading Question

Draw pictures of two solutions such that the solution with more solute is less concentrated than the solution with less solute.

Although these qualitative terms serve a useful purpose, we often need to know the exact amount of solute present in a given amount of solution. Now, we will consider various ways to describe the composition of a solution.

C. Factors Affecting the Rate of Dissolving

When a solid is being dissolved in a liquid to form a solution, the dissolving process may occur rapidly or slowly. Three factors affect the speed of the dissolving process: *surface area*, *stirring*, and *temperature*.

Because the dissolving process occurs at the surface of the solid being dissolved, the greater the amount of surface area exposed to the solvent, the faster the dissolving will occur. For example, if we want to dissolve a cube of sugar in water, how can we speed up the process? The answer is to grind up the cube into tiny crystals. Because the crystals from the ground-up cube expose much more surface area to the water than the original cube did, the sugar dissolves much more quickly.

The dissolving process is also increased by stirring the solution. Stirring removes newly dissolved particles from the solid surface and continuously exposes the surface to fresh solvent.

Finally, dissolving occurs more rapidly at higher temperatures. (Sugar dissolves much more rapidly in hot tea than iced tea.) Higher temperatures cause the solvent molecules to move more rapidly, thus increasing the rate of the dissolving process.

In addition to dissolving faster at higher temperatures, most solids are more soluble at higher temperatures. That is, in most cases more solid will dissolve in water at 90 °C than in water at 25 °C. The opposite is true for gases dissolved in water. The solubility of a gas in water typically decreases as the temperature increases.



Tea dissolves much faster in the hot water on the left than in the cold water on the right.

Let's Review

Factors Affecting Dissolving

- Surface area
- Stirring
- Temperature

CHEMISTRY EXPLORERS

Ellen Swallow Richards • 1842–1911

Ellen Swallow Richards was born in Massachusetts in 1842. She was the first woman in America to be accepted to a scientific school, the first woman to attend the Massachusetts Institute of Technology (MIT), and the first woman professional chemist in the nation. She spent her professional career in research and teaching at MIT, where she founded the field of home economics.

In the 1880s, Richards worked as water analyst for the Massachusetts State Board of Health while teaching at MIT. She conducted a two-year survey analyzing over 100,000 samples of drinking water in Massachusetts. From this study, she produced the first state water quality standards in the United States, which in turn led to the first modern sewage treatment plants. Her later studies in food quality led to state food and drug standards.

The next time you drink a glass of water, make a toast to Ellen Swallow Richards. It is her pioneering work that led to safe drinking water.



SECTION 15.1 REVIEW QUESTIONS

- 1 Draw a molecular (microscopic) level picture to show how salt and sugar look when dissolved in water.
- 2 Determine the number of moles of each type of ion in solution when 1 mole of each of the following dissolves in water:
 - a. sodium sulfate
 - b. ammonium acetate
 - c. calcium chloride
 - d. copper(II) sulfate
- 3 Chemists often say “like dissolves like.” What does this mean?
- 4 Why does water by itself not dissolve a grease stain?
- 5 Can a dilute solution also be saturated? Explain.
- 6 Increasing the number of collisions between solid and solvent particles increases the rate at which a solid dissolves. Explain how stirring, increasing surface area, and heating increase the number of collisions.



RESEARCH LINKS
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SECTION 15.2

Describing Solution Composition

Key Terms

- Mass percent
- Molarity (M)
- Standard solution
- Dilution

Objectives

- To understand mass percent and how to calculate it
- To understand and use molarity
- To learn to calculate the concentration of a solution made by diluting a stock solution

A. Solution Composition: Mass Percent

One way of describing a solution's composition is **mass percent**, which expresses the mass of solute present in a given mass of solution. The definition of mass percent follows:

$$\begin{aligned}\text{Mass percent} &= \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\% \\ &= \frac{\text{grams of solute}}{\text{grams of solute} + \text{grams of solvent}} \times 100\%\end{aligned}$$

For example, suppose a solution is prepared by dissolving 1.0 g of sodium chloride in 48 g of water. The solution has a mass of 49 g (48 g of H_2O plus 1.0 g of NaCl), and there is 1.0 g of solute (NaCl) present. The mass percent of solute, then, is

$$\frac{1.0 \text{ g solute}}{49 \text{ g solution}} \times 100\% = 0.020 \times 100\% = 2.0\% \text{ NaCl}$$

Information

The mass of the solution is the sum of the masses of the solute and the solvent.

HANDS-ON CHEMISTRY

• MINI-LAB •

Rainbow in a Straw

Materials

- five paper cups containing five different colored solutions of NaCl
- clear, colorless plastic straw

Procedure

1. Obtain five paper cups containing five different colored solutions. You will also need a clear, colorless plastic straw.
2. Each of your five solutions has a different concentration of NaCl. Your task is to order the solutions from least concentrated to most concentrated. You may only use the five solutions and the straw. You may need to experiment a little and you may make some mistakes along the way.

Results/Analysis

1. How does concentration of the solution relate to density of the solution?
2. List the order of solutions from least concentrated to most concentrated.

EXAMPLE 15.1**Solution Composition: Calculating Mass Percent**

A solution is prepared by mixing 1.00 g of ethanol, C_2H_5OH , with 100.0 g of water. Calculate the mass percent of ethanol in this solution.

Solution**Where do we want to go?**

Mass percent of ethanol = ? %

What do we know?

- 1.00 g ethanol (C_2H_5OH)
- 100.0 g H_2O
- mass percent = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$

How do we get there?

$$\begin{aligned} \text{Mass percent } C_2H_5OH &= \left(\frac{\text{grams of } C_2H_5OH}{\text{grams of solution}} \right) \times 100\% \\ &= \left(\frac{1.00 \text{ g } C_2H_5OH}{100.0 \text{ g } H_2O + 1.00 \text{ g } C_2H_5OH} \right) \times 100\% \\ &= \frac{1.00 \text{ g}}{101.0 \text{ g}} \times 100\% \\ &= 0.990\% C_2H_5OH \end{aligned}$$

Practice Problem • Exercise 15.1

A 135-g sample of seawater is evaporated to dryness, leaving 4.73 g of solid residue (the salts formerly dissolved in the seawater). Calculate the mass percent of solute present in the original seawater.

EXAMPLE 15.2**Solution Composition: Determining Mass of Solute**

Although milk is not a true solution (it is really a suspension of tiny globules of fat, protein, and other substrates in water), it does contain a dissolved sugar called lactose. Cow's milk typically contains 4.5% by mass of lactose, $C_{12}H_{22}O_{11}$. Calculate the mass of lactose present in 175 g of milk.

Solution**Where do we want to go?**

Mass lactose in 175 g milk = ? g

What do we know?

- 4.5% lactose ($C_{12}H_{22}O_{11}$) {solute}
- 175 g milk {solution}
- mass percent = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$

How do we get there?

$$\text{Mass percent} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 100\%$$

We now substitute the quantities we know:

$$\text{Mass percent} = \frac{\overset{\text{Mass of lactose}}{\text{grams of solute}}}{\underset{\text{Mass of milk}}{175 \text{ g}}} \times 100\% = \overset{\text{Mass percent}}{4.5\%}$$

We now solve for grams of solute by multiplying both sides by 175 g,

$$\cancel{175 \text{ g}} \times \frac{\text{grams of solute}}{\cancel{175 \text{ g}}} \times 100\% = 4.5\% \times 175 \text{ g}$$

and then dividing both sides by 100%,

$$\text{Grams of solute} \times \frac{100\%}{100\%} = \frac{4.5\%}{100\%} \times 175 \text{ g}$$

to give

$$\text{Grams of solute} = 0.045 \times 175 \text{ g} = 7.9 \text{ g lactose}$$

Practice Problem • Exercise 15.2

What mass of water must be added to 425 g of formaldehyde to prepare a 40.0% (by mass) solution of formaldehyde? This solution, called formalin, is used to preserve biological specimens.

Hint: Substitute the known quantities into the definition for mass percent, and then solve for the unknown quantity (mass of solvent).

B. Solution Composition: Molarity

When a solution is described in terms of mass percent, the amount of solution is given in terms of its mass. However, it is often more convenient to measure the volume of a solution than to measure its mass. Because of this, chemists often describe a solution in terms of concentration. We define the *concentration* of a solution as the amount of solute in a *given volume* of solution. The most commonly used expression of concentration is **molarity (M)**. Molarity describes the amount of solute in moles and the volume of the solution in liters.

Molarity is the number of moles of solute per volume of solution in liters.

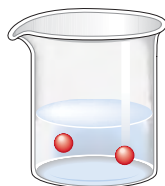
$$M = \text{molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{\text{mol}}{\text{L}}$$

A solution that is 1.0 molar (written as 1.0 M) contains 1.0 mol of solute per liter of solution.

Active Reading Question

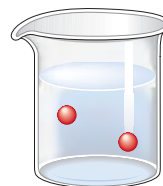
Adding 1.0 mole of solute to 1.0 L of water does not make a 1.0 M solution. Why not?

We need to consider both the amount of solute and the volume when determining the concentration of a solution. For example, consider the following two solutions where each dot represents 1.0 mol of solute.



Solution A
Volume = 1.0 L

$$\text{Concentration} = \frac{2.0 \text{ mol solute}}{1.0 \text{ L solution}} = 2.0 \text{ M}$$



Solution B
Volume = 2.0 L

$$\text{Concentration} = \frac{2.0 \text{ mol solute}}{2.0 \text{ L solution}} = 1.0 \text{ M}$$

EXAMPLE 15.3

Solution Composition: Calculating Molarity, I

Calculate the molarity of a solution prepared by dissolving 11.5 g of solid NaOH in enough water to make 1.50 L of solution.

Solution

Where do we want to go?

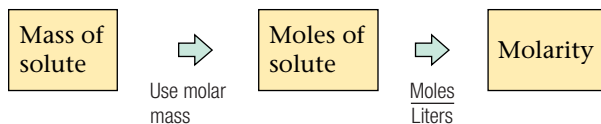
$$\text{molarity NaOH} = ? \text{ M}$$

What do we know?

- 11.5 g NaOH
- $V = 1.50 \text{ L}$
- $M = \frac{\text{moles of solute}}{\text{liters of solution}}$
- molar mass NaOH = 40.0 g/mol

How do we get there?

We have the mass (in grams) of solute, so we need to convert the mass of solute to moles (using the molar mass of NaOH). Then we can divide the number of moles by the volume in liters.



We compute the number of moles of solute, using the molar mass of NaOH (40.0 g).

$$11.5 \text{ g NaOH} \times \frac{1 \text{ mol NaOH}}{40.0 \text{ g NaOH}} = 0.288 \text{ mol NaOH}$$

Then we divide by the volume of the solution in liters.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} = \frac{0.288 \text{ mol NaOH}}{1.50 \text{ L solution}} = 0.192 \text{ M NaOH}$$

EXAMPLE 15.4

Solution Composition: Calculating Molarity, II

Calculate the molarity of a solution prepared by dissolving 1.56 g of gaseous HCl into enough water to make 26.8 mL of solution.

Solution

Where do we want to go?

$$\text{molarity HCl} = ? M$$

What do we know?

- 1.56 g HCl (g)
- $V = 26.8 \text{ mL}$
- $M = \frac{\text{moles of solute}}{\text{liters of solution}}$
- molar mass HCl = 36.5 g/mol

How do we get there?

First we calculate the number of moles of HCl (molar mass = 36.5 g).

$$\begin{aligned} 1.56 \text{ g HCl} \times \frac{1 \text{ mol HCl}}{36.5 \text{ g HCl}} &= 0.0427 \text{ mol HCl} \\ &= 4.27 \times 10^{-2} \text{ mol HCl} \end{aligned}$$

Next we change the volume of the solution from milliliters to liters, using the equivalence statement $1 \text{ L} = 1000 \text{ mL}$, which gives the appropriate conversion factor.

$$\begin{aligned} 26.8 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} &= 0.0268 \text{ L} \\ &= 2.68 \times 10^{-2} \text{ L} \end{aligned}$$

Finally, we divide the moles of solute by the liters of solution.

$$\text{Molarity} = \frac{4.27 \times 10^{-2} \text{ mol HCl}}{2.68 \times 10^{-2} \text{ L solution}} = 1.59 M \text{ HCl}$$

Practice Problem • Exercise 15.4

Calculate the molarity of a solution prepared by dissolving 1.00 g of ethanol, $\text{C}_2\text{H}_5\text{OH}$, in enough water to give a final volume of 101 mL.

It is important to realize that the description of a solution's composition may not accurately reflect the true chemical nature of the solute as it is present in the dissolved state. Solute concentration is always written in terms of the form of the solute *before* it dissolves. For example, describing a solution as 1.0 M NaCl means that the solution was prepared by dissolving 1.0 mol of solid NaCl in enough water to make 1.0 L of solution; it does not mean that the solution contains 1.0 mol of NaCl units. Actually the solution contains 1.0 mole of Na^+ ions and 1.0 mole of Cl^- ions. That is, it contains 1.0 M Na^+ and 1.0 M Cl^- .

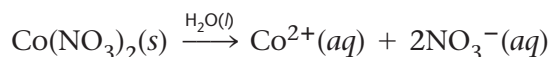
EXAMPLE 15.5**Solution Composition: Calculating Ion Concentration from Molarity**

Give the concentrations of all the ions in each of the following solutions:

- 0.50 M $\text{Co}(\text{NO}_3)_2$
- 1 M FeCl_3

Solution

- When solid $\text{Co}(\text{NO}_3)_2$ dissolves, it produces ions as follows:

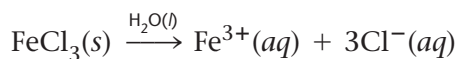


which we can represent as



Therefore, a solution that is 0.50 M $\text{Co}(\text{NO}_3)_2$ contains 0.50 M Co^{2+} and (2×0.50) M NO_3^{-} , or 1.0 M NO_3^{-} .

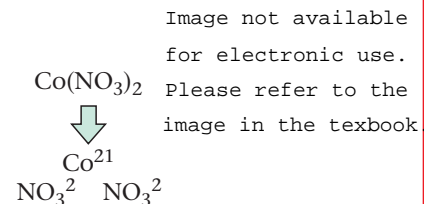
- When solid FeCl_3 dissolves, it produces ions as follows:



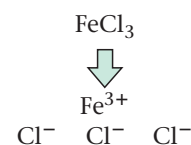
or



A solution that is 1 M FeCl_3 contains 1 M Fe^{3+} ions and 3 M Cl^{-} ions.



A solution of cobalt(II) nitrate

**Practice Problem • Exercise 15.5**

Give the concentrations of the ions in each of the following solutions:

- 0.10 M Na_2CO_3
- 0.010 M $\text{Al}_2(\text{SO}_4)_3$

Often we need to determine the number of moles of solute present in a given volume of a solution of known molarity. To do this, we use the definition of molarity. When we multiply the molarity of a solution by the volume (in liters), we get the moles of solute present in that sample:

$$\begin{aligned} \text{Liters of solution} \times \text{molarity} &= \text{liters of solution} \times \frac{\text{moles of solute}}{\text{liters of solution}} \\ &= \text{moles of solute} \end{aligned}$$

EXAMPLE 15.6

Solution Composition: Calculating Number of Moles from Molarity

How many moles of Ag^+ ions are present in 25 mL of a 0.75 M AgNO_3 solution?

Solution

Where do we want to go?

Moles Ag^+ in sample solution = ? mol

What do we know?

- Molarity of solution = 0.75 M
- $V = 25 \text{ mL}$
- Moles = $M \times V$

MATH



$$M = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$\text{Liters} \times M \rightarrow \text{Moles of solute}$$

How do we get there?

A 0.75 M AgNO_3 solution contains 0.75 M Ag^+ ions and 0.75 M NO_3^- ions. Next we must convert the volume given from mL to L.

$$25 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.025 \text{ L} = 2.5 \times 10^{-2} \text{ L}$$

Now we multiply the volume times the molarity.

$$2.5 \times 10^{-2} \text{ L}_{\text{solution}} \times \frac{0.75 \text{ mol Ag}^+}{\text{L}_{\text{solution}}} = 1.9 \times 10^{-2} \text{ mol Ag}^+$$

Practice Problem • Exercise 15.6

Calculate the number of moles of Cl^- ions in 1.75 L of $1.0 \times 10^{-3} \text{ M AlCl}_3$.

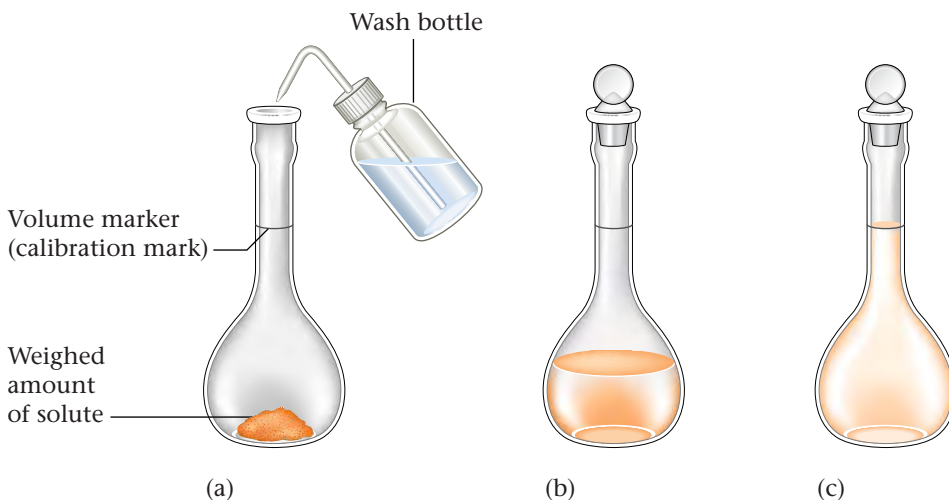
Standard solution

A solution in which the concentration is accurately known

A **standard solution** is a solution whose *concentration is accurately known*. When the appropriate solute is available in pure form, a standard solution can be prepared by weighing out a sample of solute, transferring it completely to a *volumetric flask* (a flask of accurately known volume), and adding enough solvent to bring the volume up to the mark on the neck of the flask. This procedure is illustrated in **Figure 15.3**.

Figure 15.3

Steps involved in the preparation of a standard aqueous solution. (a) Put a weighed amount of a substance (the solute) into the volumetric flask, and add a small quantity of water. (b) Dissolve the solid in the water by gently swirling the flask (*with the stopper in place*). (c) Add more water (with gentle swirling) until the level of the solution just reaches the mark etched on the neck of the flask. Then mix the solution thoroughly by inverting the flask several times.



EXAMPLE 15.7

To analyze the alcohol content of a certain substance, a chemist needs 1.00 L of an aqueous 0.200 M $\text{K}_2\text{Cr}_2\text{O}_7$ (potassium dichromate) solution. How much solid $\text{K}_2\text{Cr}_2\text{O}_7$ (molar mass = 294.2 g/mol) must be weighed out to make this solution?

Solution**Where do we want to go?**

Mass of $\text{K}_2\text{Cr}_2\text{O}_7$ needed to make solution = ? g

What do we know?

- Molarity of solution = 0.200 M
- Molar mass $\text{K}_2\text{Cr}_2\text{O}_7$ = 294.2 g/mol
- $V = 1.00$ L
- Moles = $M \times V$

How do we get there?

We need to calculate the number of grams of solute ($\text{K}_2\text{Cr}_2\text{O}_7$) present (and thus the mass needed to make the solution). First we determine the number of moles of $\text{K}_2\text{Cr}_2\text{O}_7$ present by multiplying the volume (in liters) by the molarity.

$$1.00 \text{ L solution} \times \frac{0.200 \text{ mol } \text{K}_2\text{Cr}_2\text{O}_7}{\text{L solution}} = 0.200 \text{ mol } \text{K}_2\text{Cr}_2\text{O}_7$$

Then we convert the moles of $\text{K}_2\text{Cr}_2\text{O}_7$ to grams, using the molar mass of $\text{K}_2\text{Cr}_2\text{O}_7$ (294.2 g).

$$0.200 \text{ mol } \text{K}_2\text{Cr}_2\text{O}_7 \times \frac{294.2 \text{ g } \text{K}_2\text{Cr}_2\text{O}_7}{\text{mol } \text{K}_2\text{Cr}_2\text{O}_7} = 58.8 \text{ g } \text{K}_2\text{Cr}_2\text{O}_7$$

Therefore, to make 1.00 L of 0.200 M $\text{K}_2\text{Cr}_2\text{O}_7$, the chemist must weigh out 58.8 g of $\text{K}_2\text{Cr}_2\text{O}_7$ and dissolve it in enough water to make 1.00 L of solution. This is most easily done by using a 1.00-L volumetric flask.

Practice Problem • Exercise 15.7

Formalin is an aqueous solution of formaldehyde, HCHO, used as a preservative for biological specimens. How many grams of formaldehyde must be used to prepare 2.5 L of 12.3 M formalin?

MATH

Liters \times M \Rightarrow Moles of solute

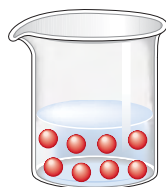
C. Dilution

To save time and space in the laboratory, solutions that are routinely used are often purchased or prepared in concentrated form (called *stock solutions*). Water (or another solvent) is then added to achieve the molarity desired for a particular solution. The process of adding more solvent to a solution is called **dilution**. For example, the common laboratory acids are purchased as concentrated solutions and diluted with water as they are needed. A typical dilution calculation involves determining how much water must be added to an amount of stock solution to achieve a solution of the desired concentration. The key to doing these calculations is to remember that *only water is added in the dilution*. The amount of solute in the final, more dilute, solution is the *same* as the amount of solute in the original concentrated stock solution.

Dilution

The process of adding solvent to a solution to lower the concentration of solute

For example, consider the following solutions in which each dot represents 1.0 mol of solute particles.



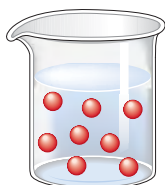
$$\text{Concentration} = \frac{8.0 \text{ mol solute}}{1.0 \text{ L solution}} = 8.0 \text{ M}$$

Solution 1
Volume = 1.0 L

information

Dilution with water doesn't alter the number of moles of solute present.

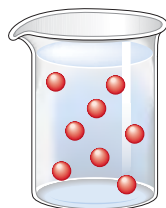
If we double the volume of the solution by adding water, we do not change the moles of solute, but the concentration becomes half of its original value.



$$\text{Concentration} = \frac{8.0 \text{ mol solute}}{2.0 \text{ L solution}} = 4.0 \text{ M}$$

Solution 2
Volume = 2.0 L

If the volume is doubled again, the concentration is again cut by 2, becoming one-fourth of the original concentration.



$$\text{Concentration} = \frac{8.0 \text{ mol solute}}{4.0 \text{ L solution}} = 2.0 \text{ M}$$

Solution 3
Volume = 4.0 L

information

The molarities of stock solutions of the common concentrated acids are:

Sulfuric (H_2SO_4)	18 M
Nitric (HNO_3)	16 M
Hydrochloric (HCl)	12 M

Notice in all three solutions,

Moles of solute after dilution = moles of solute before dilution

The number of moles of solute stays the same but more water is added, increasing the volume, so the molarity decreases.

$$M = \frac{\text{moles of solute}}{\text{volume (L)}}$$

↑ Decreases (under moles of solute)
↑ Increases (water added) (under volume (L))
Remains constant (under the fraction)

For example, suppose we want to prepare 500. mL of 1.00 M acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, from a 17.5 M stock solution of acetic acid. What volume of the stock solution is required?

Where do we want to go?

Volume of stock solution needed = ? L

What do we know?

- Molarity of stock solution = 17.5 M
- Solution required
 $V_{\text{dilute}} = 500.0 \text{ mL}$
 $M_{\text{dilute}} = 1.00 \text{ M}$
- Moles = $M \times V$

How do we get there?

Find the moles of acetic acid needed in the dilute solution.

- Convert the volume of the dilute solution to L.

$$500. \text{ mL solution} \times \frac{1 \text{ L solution}}{1000 \text{ mL solution}} = 0.500 \text{ L solution}$$

$V_{\text{dilute solution}}$ (in mL) Convert mL to L

- Moles acetic acid needed = $M \times V$

$$0.500 \text{ L solution} \times \frac{1.00 \text{ mol HC}_2\text{H}_3\text{O}_2}{1 \text{ L solution}} = 0.500 \text{ mol HC}_2\text{H}_3\text{O}_2$$

$M_{\text{dilute solution}}$

Find the volume of the stock (concentrated) solution that contains 0.500 mol acetic acid.

Because volume \times molarity = moles, we have

$$V \text{ (in liters)} \times \frac{17.5 \text{ mol HC}_2\text{H}_3\text{O}_2}{\text{L solution}} = 0.500 \text{ mol HC}_2\text{H}_3\text{O}_2$$

Solving for V (by dividing both sides by $\frac{17.5 \text{ mol}}{\text{L solution}}$) gives

$$V = \frac{0.500 \text{ mol HC}_2\text{H}_3\text{O}_2}{\frac{17.5 \text{ mol HC}_2\text{H}_3\text{O}_2}{\text{L solution}}} = 0.0286 \text{ L, or } 28.6 \text{ mL, of solution}$$

Top Ten Elements in Seawater

Element	Amount (kg/m ³)
Oxygen	3.9×10^{12}
Hydrogen	5.0×10^{11}
Chlorine	9.2×10^{10}
Sodium	5.3×10^{10}
Magnesium	6.1×10^9
Sulfur	4.3×10^9
Calcium	1.9×10^9
Potassium	1.9×10^9
Bromine	3.1×10^8
Carbon	1.3×10^8

MATH

Liters \times M \Rightarrow Moles of solute

HANDS-ON CHEMISTRY**• MINI-LAB •****Can We Add Concentrations?****Materials**

- water
- food coloring
- graduated cylinder

Procedure

- Place 10.0 mL of water and add 10 drops of food coloring. Label the concentration of this solution 1.0 D (where D has units of drops/mL). This is solution A.
- Add 10.0 mL of water to another cup and add 5 drops of food coloring to the water. This is solution B.
- Add half of solution A and half of solution B to a third cup to make solution C.

Results/Analysis

- What is the concentration of solution B in units of D ?
- How does this solution C look compared to solution A and solution B?
- What is the concentration of solution C in units of D ?
- Can you add the concentrations of solution A and solution B to get the concentration of solution C? Why or why not? How does your answer to question 2 help you answer this question?

Therefore, to make 500. mL of a 1.00 M acetic acid solution, we take 28.6 mL of 17.5 M acetic acid and dilute it to a total volume of 500. mL. This process is illustrated in **Figure 15.4**. Because *the moles of solute remain the same* before and after dilution, we can write

Initial conditions		= moles of solute =		Final conditions
M_1	×		M_2	×
Molarity before dilution			Molarity after dilution	
V_1			V_2	
Volume before dilution			Volume after dilution	

Does it make sense?

We can check our calculations on acetic acid by showing that $M_1 \times V_1 = M_2 \times V_2$. In the above example, $M_1 = 17.5$, $V_1 = 0.0286$ L, $V_2 = 0.500$ L, and $M_2 = 1.00$ M, so

$$M_1 \times V_1 = 17.5 \frac{\text{mol}}{\text{L}} \times 0.0286 \text{ L} = 0.500 \text{ mol}$$

$$M_2 \times V_2 = 1.00 \frac{\text{mol}}{\text{L}} \times 0.500 \text{ L} = 0.500 \text{ mol}$$

and therefore

$$M_1 \times V_1 = M_2 \times V_2$$

This shows that the volume (V_2) we calculated is correct.

Active Reading Question

You add 95.0 mL of water to 5.0 mL of 5.0 M acetic acid. Compare the concentration and number of moles of acetic acid in the concentrated acid to the concentration and number of moles of acetic acid in the dilute solution.

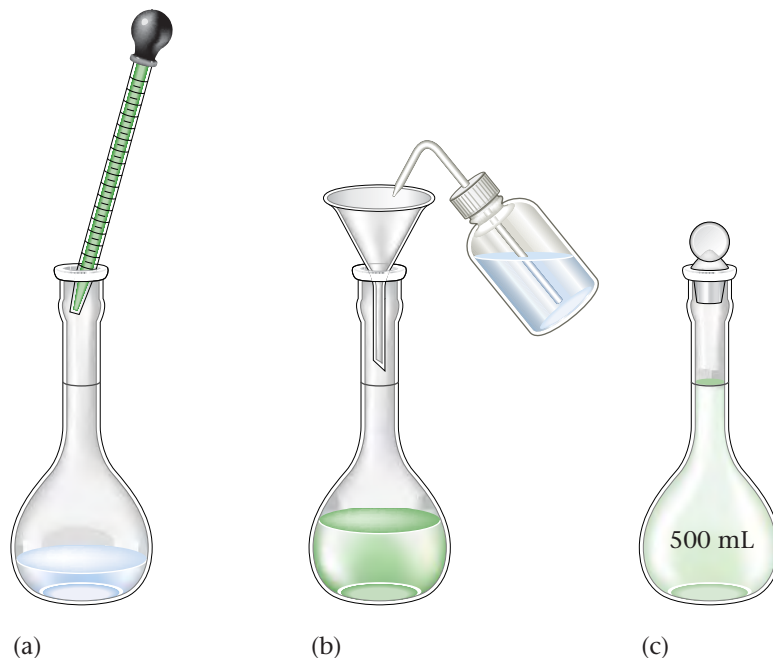


Figure 15.4

- (a) 28.6 mL of 17.5 M acetic acid solution is transferred to a volumetric flask that already contains some water.
 (b) Water is added to the flask (with swirling) to bring the volume to the calibration mark, and the solution is mixed by inverting the flask several times.
 (c) The resulting solution is 1.00 M acetic acid.

EXAMPLE 15.8**Calculating Concentrations of Diluted Solutions**

What volume of 16 M sulfuric acid must be used to prepare 1.5 L of a 0.10 M H_2SO_4 solution?

Solution**Where do we want to go?**

Volume of H_2SO_4 needed to make solution = ? mL

What do we know?

- Moles solute = $M_1 \times V_1 = M_2 \times V_2$

Initial Conditions (concentrated)	Final Conditions (dilute)
$M_1 = 16 \text{ M}$	$M_2 = 0.10 \text{ M}$
$V_1 = ?$	$V_2 = 1.5 \text{ L}$

How do we get there?

We can solve the equation

$$M_1 \times V_1 = M_2 \times V_2$$

for V_1 by dividing both sides by M_1

$$\frac{M_1 \times V_1}{M_1} = \frac{M_2 \times V_2}{M_1}$$

to give

$$V_1 = \frac{M_2 \times V_2}{M_1}$$

Now we substitute the known values of M_2 , V_2 , and M_1 .

$$V_1 = \frac{(0.10 \frac{\text{mol}}{\text{L}}) (1.5 \text{ L})}{16 \frac{\text{mol}}{\text{L}}} = 9.4 \times 10^{-3} \text{ L}$$

$$9.4 \times 10^{-3} \text{ L} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 9.4 \text{ mL}$$

Therefore, $V_1 = 9.4 \times 10^{-3} \text{ L}$, or 9.4 mL. To make 1.5 L of 0.10 M H_2SO_4 using 16 M H_2SO_4 , we must take 9.4 mL of the concentrated acid and dilute it with water to a final volume of 1.5 L. The correct way to do this is to add the 9.4 mL of acid to about 1 L of water and then dilute to 1.5 L by adding more water.

Image not available for electronic use. Please refer to the image in the textbook.

Approximate dilutions can be carried out using a calibrated beaker. Here concentrated sulfuric acid is being added to water to make a dilute solution.

DID YOU KNOW

It is always best to add concentrated acid to water, not water to the acid. That way, if any splashing occurs accidentally, it is dilute acid that splashes.

Practice Problem • Exercise 15.8

What volume of 12 M HCl must be taken to prepare 0.75 L of 0.25 M HCl?

**Good to the Last Drop!****Materials**

- water
- food coloring
- graduated cylinder

Procedure

1. Place 10.0 mL of water and add 20 drops of food coloring. Label the concentration of this solution $2.0 D$ (where D has units of drops/mL). This is solution A.
2. Take 1.0 mL of solution A and place it in an empty cup. Add 9.0 mL of water. This is solution B.
3. Take 1.0 mL of solution B and place it in an empty cup. Add 9.0 mL of water. This is solution C.
4. Continue this successive dilution until you can no longer see any color.

Results/Analysis

1. What is the concentration of solution B in units of D ?
2. What is the concentration of solution C in units of D ?
3. What is the concentration of the solution in step 4 in units of D ?

**SECTION 15.2****REVIEW QUESTIONS**

1. Vinegar is made by adding 33 g of acetic acid to 625 g of water. What is the percent by mass of acetic acid in this solution of vinegar?
2. Assuming the density of vinegar is 1.0 g/mL, what is the molarity of vinegar? (Use the percent by mass of acetic acid in vinegar from question 1.) The molar mass of acetic acid is 60.05 g/mol.
3. A solution is labeled “0.450 M magnesium nitrate.” Calculate the concentration of each ion present in solution.
4. Calculate the number of moles of KOH in 150.0 mL of a 0.500 M KOH solution.
5. If 4.25 g of CaBr_2 is dissolved in enough water to make 125 mL of solution, what is the molarity of the solution?
6. What is the key idea to remember when working with the dilution of a solution?
7. Calculate the new concentration when 50.0 mL of water is added to 725.0 mL of 1.25 M NaCl.

**RESEARCH LINKS**
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Properties of Solutions

Objectives

- To learn to solve stoichiometric problems involving solution reactions
- To do calculations involving acid–base reactions
- To learn about normality and equivalent weight
- To use normality in stoichiometric calculations
- To understand the effect of a solute on solution properties

A. Stoichiometry of Solution Reactions

Because so many important reactions occur in solution, it is important to be able to do stoichiometric calculations for solution reactions. The principles needed to perform these calculations are similar to those developed in Chapter 9. It is helpful to think in terms of the following steps:

Steps for Solving Stoichiometric Problems Involving Solutions

- Step 1** Write the balanced equation for the reaction. For reactions involving ions, it is best to write the net ionic equation.
- Step 2** Calculate the moles of reactants.
- Step 3** Determine which reactant is limiting.
- Step 4** Calculate the moles of other reactants or products, as required.
- Step 5** Convert to grams or other units, if required.

Key Terms

- Neutralization reaction
- Equivalent of an acid
- Equivalent of a base
- Equivalent weight
- Colligative property

CRITICAL THINKING ?

What if all ionic solids were soluble in water? How would it affect reactions in aqueous solution?

Active Reading Question

When solving stoichiometry problems for solution reactions, what type of chemical equation is most convenient to use?

EXAMPLE 15.9

Solution Stoichiometry: Calculating Mass of Reactants and Products

Calculate the mass of solid NaCl that must be added to 1.50 L of a 0.100 M AgNO₃ solution to precipitate all of the Ag⁺ ions in the form of AgCl. Calculate the mass of AgCl formed.

Solution

Where do we want to go?

Mass of AgCl formed = ? g

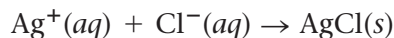
What do we know?

- $V = 1.50 \text{ L}$
- $M = 0.100 \text{ M AgNO}_3$
- $\text{NaCl}(s) + \text{AgNO}_3(aq) \rightarrow \text{AgCl}(s) + \text{NaNO}_3(aq)$

How do we get there?

Step 1 Write the balanced equation for the reaction.

When added to the AgNO_3 solution (which contains Ag^+ and NO_3^- ions), the solid NaCl dissolves to yield Na^+ and Cl^- ions. Solid AgCl forms according to the following balanced net ionic reaction:



Step 2 Calculate the moles of reactants.

In this case we must add just enough Cl^- ions to react with all the Ag^+ ions present, so we must calculate the moles of Ag^+ ions present in 1.50 L of a 0.100 M AgNO_3 solution. (Remember that a 0.100 M AgNO_3 solution contains 0.100 M Ag^+ ions and 0.100 M NO_3^- ions.)

$$1.50 \cancel{\text{L}} \times \frac{0.100 \text{ mol Ag}^+}{\cancel{\text{L}}} = 0.150 \text{ mol Ag}^+$$

Moles of Ag^+ present
in 1.5 L of 0.100 M AgNO_3

MATH



Liters \times M \Rightarrow Moles of solute

Step 3 Determine which reactant is limiting.

In this situation we want to add just enough Cl^- to react with the Ag^+ present. That is, we want to precipitate *all* the Ag^+ in the solution. Thus the Ag^+ present determines the amount of Cl^- needed.

Step 4 Calculate the moles of Cl^- required.

We have 0.150 mol of Ag^+ ions and, because one Ag^+ ion reacts with one Cl^- ion, we need 0.150 mol of Cl^- ,

$$0.150 \text{ mol Ag}^+ \times \frac{1 \text{ mol Cl}^-}{1 \text{ mol Ag}^+} = 0.150 \text{ mol Cl}^-$$

so 0.150 mol of AgCl will be formed.



Step 5 Convert to grams of NaCl required.

To produce 0.150 mol Cl^- , we need 0.150 mol NaCl . We calculate the mass of NaCl required as follows:

$$0.150 \text{ mol NaCl} \times \frac{58.4 \text{ g NaCl}}{1 \text{ mol NaCl}} = 8.76 \text{ g NaCl}$$

Moles



Times molar mass

Mass

The mass of AgCl formed is

$$0.150 \text{ mol AgCl} \times \frac{143.3 \text{ g AgCl}}{1 \text{ mol AgCl}} = 21.5 \text{ g AgCl}$$

When aqueous sodium chloride is added to a solution of silver nitrate, a white silver chloride precipitate forms.



EXAMPLE 15.10**Solution Stoichiometry: Determining Limiting Reactants and Calculating Mass of Products**

When $\text{Ba}(\text{NO}_3)_2$ and K_2CrO_4 react in aqueous solution, the yellow solid BaCrO_4 is formed. Calculate the mass of BaCrO_4 that forms when 3.50×10^{-3} mol of solid $\text{Ba}(\text{NO}_3)_2$ is dissolved in 265 mL of 0.0100 M K_2CrO_4 solution.

Solution**Where do we want to go?**

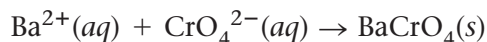
Mass of BaCrO_4 formed in the reaction = ? g

What do we know?

- Amount $\text{Ba}(\text{NO}_3)_2(\text{s}) = 3.50 \times 10^{-3}$ mol
- K_2CrO_4
 - 0.100 M
 - $V = 265$ mL
- $\text{Ba}(\text{NO}_3)_2(\text{s}) + \text{K}_2\text{CrO}_4(\text{aq}) \rightarrow \text{BaCrO}_4(\text{s}) + 2\text{KNO}_3(\text{aq})$

How do we get there?

Step 1 The original K_2CrO_4 solution contains the ions K^+ and CrO_4^{2-} . When the $\text{Ba}(\text{NO}_3)_2$ is dissolved in this solution, Ba^{2+} and NO_3^- ions are added. The Ba^{2+} and CrO_4^{2-} ions react to form solid BaCrO_4 . The balanced net ionic equation is



Step 2 Next we determine the moles of reactants. We are told that 3.50×10^{-3} mol of $\text{Ba}(\text{NO}_3)_2$ is added to the K_2CrO_4 solution. Each formula unit of $\text{Ba}(\text{NO}_3)_2$ contains one Ba^{2+} ion, so 3.50×10^{-3} mol of $\text{Ba}(\text{NO}_3)_2$ gives 3.50×10^{-3} mol of Ba^{2+} ions in solution.



Because $V \times M = \text{moles of solute}$, we can compute the moles of K_2CrO_4 in the solution from the volume and molarity of the original solution. First we must convert the volume of the solution (265 mL) to liters.

$$265 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.265 \text{ L}$$

Next we determine the number of moles of K_2CrO_4 using the molarity of the K_2CrO_4 solution (0.0100 M).

$$0.265 \text{ L} \times \frac{0.0100 \text{ mol } \text{K}_2\text{CrO}_4}{\text{L}} = 2.65 \times 10^{-3} \text{ mol } \text{K}_2\text{CrO}_4$$

We know that

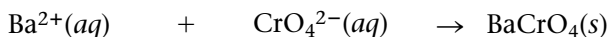


so the solution contains 2.65×10^{-3} mol of CrO_4^{2-} ions.



Barium chromate precipitating

Step 3 The balanced equation tells us that one Ba^{2+} ion reacts with one CrO_4^{2-} . Because the number of moles of CrO_4^{2-} ions (2.65×10^{-3}) is smaller than the number of moles of Ba^{2+} ions (3.50×10^{-3}), the CrO_4^{2-} will run out first.



$3.50 \times 10^{-3} \text{ mol}$

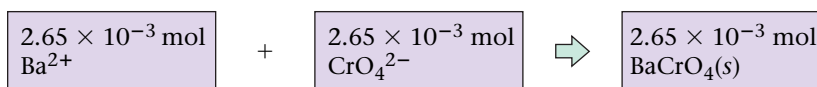
$2.65 \times 10^{-3} \text{ mol}$

Smaller
(runs out first)

Therefore, the CrO_4^{2-} is limiting.



Step 4 The $2.65 \times 10^{-3} \text{ mol}$ of CrO_4^{2-} ions will react with $2.65 \times 10^{-3} \text{ mol}$ of Ba^{2+} ions to form $2.65 \times 10^{-3} \text{ mol}$ of BaCrO_4 .



Step 5 The mass of BaCrO_4 formed is obtained from its molar mass (253.3 g) as follows:

$$2.65 \times 10^{-3} \text{ mol BaCrO}_4 \times \frac{253.3 \text{ g BaCrO}_4}{\text{mol BaCrO}_4} = 0.671 \text{ g BaCrO}_4$$

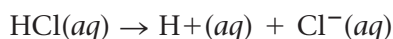
Practice Problem • Exercise 15.10

When aqueous solutions of Na_2SO_4 and $\text{Pb}(\text{NO}_3)_2$ are mixed, PbSO_4 precipitates. Calculate the mass of PbSO_4 formed when 1.25 L of 0.0500 M $\text{Pb}(\text{NO}_3)_2$ and 2.00 L of 0.0250 M Na_2SO_4 are mixed.

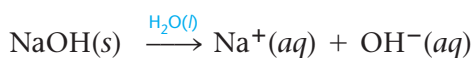
Hint: Calculate the moles of Pb^{2+} and SO_4^{2-} in the mixed solution, decide which ion is limiting, and calculate the moles of PbSO_4 formed.

B. Neutralization Reactions

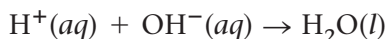
So far we have considered the stoichiometry of reactions in solution that result in the formation of a precipitate. Another common type of solution reaction occurs between an acid and a base. We introduced these reactions in Chapter 8. Recall from that discussion that an acid is a substance that furnishes H^+ ions. A strong acid, such as hydrochloric acid, HCl , dissociates (ionizes) completely in water.



Strong bases are water-soluble metal hydroxides, which are completely dissociated in water. An example is NaOH , which dissolves in water to give Na^+ and OH^- ions.



When a strong acid and strong base react, the net ionic reaction is



An acid–base reaction is often called a **neutralization reaction**. When just enough strong base is added to react exactly with the strong acid in a solution, we say the acid has been *neutralized*. One product of this reaction is always water. The steps in dealing with the stoichiometry of any neutralization reaction are the same as those we followed previously.

Neutralization reaction

An acid–base reaction

Active Reading Question

What does it mean to say an acid has been neutralized?

EXAMPLE 15.11

Solution Stoichiometry: Calculating Volume in Neutralization Reactions

What volume of a 0.100 M HCl solution is needed to neutralize 25.0 mL of a 0.350 M NaOH solution?

Solution

Where do we want to go?

Volume of HCl required for neutralization = ? L

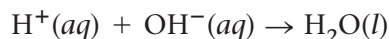
What do we know?

- HCl molarity = 0.100 M
- NaOH sample
 - 0.350 M
 - $V = 25.0$ mL
- $\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq)$

How do we get there?

Step 1 Write the balanced equation for the reaction.

Hydrochloric acid is a strong acid, so all the HCl molecules dissociate to produce H^+ and Cl^- ions. Also, when the strong base NaOH dissolves, the solution contains Na^+ and OH^- ions. When these two solutions are mixed, the H^+ ions from the hydrochloric acid react with the OH^- ions from the sodium hydroxide solution to form water. The balanced net ionic equation for the reaction is



Step 2 Calculate the moles of reactants.

In this problem we are given a volume (25.0 mL) of 0.350 M NaOH, and we want to add just enough 0.100 M HCl to provide just enough H^+ ions to react with all the OH^- . Therefore, we must calculate the number of moles of OH^- ions in the 25.0-mL sample of 0.350 M NaOH. To do this, we first change the volume to liters and multiply by the molarity.

$$25.0 \text{ mL NaOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.350 \text{ mol OH}^-}{\text{L NaOH}} = 8.75 \times 10^{-3} \text{ mol OH}^-$$

Moles of OH^- present
in 25.0 mL of
0.350 M NaOH

Step 3 Determine which reactant is limiting.

This problem requires the addition of just enough H^+ ions to react exactly with the OH^- ions present, so the number of moles of OH^- ions present determines the number of moles of H^+ that must be added. The OH^- ions are limiting.

Step 4 Calculate the moles of H^+ required.

The balanced equation tells us that the H^+ and OH^- ions react in a 1:1 ratio, so 8.75×10^{-3} mol of H^+ ions is required to neutralize (exactly react with) the 8.75×10^{-3} mol of OH^- ions present.

Step 5 Calculate the volume of 0.100 M HCl required.

Next we must find the volume (V) of 0.100 M HCl required to furnish this amount of H^+ ions. Because the volume (in liters) times the molarity gives the number of moles, we have

$$V \times \frac{0.100 \text{ mol H}^+}{\text{L}} = 8.75 \times 10^{-3} \text{ mol H}^+$$

Unknown volume (in liters) Moles of H^+ needed

Now we must solve for V by dividing both sides of the equation by $\frac{0.100 \text{ mole H}^+}{\text{L}}$.

$$V \times \frac{\cancel{0.100 \text{ mol H}^+}}{\cancel{\text{L}}} = \frac{8.75 \times 10^{-3} \cancel{\text{ mol H}^+}}{\cancel{0.100 \text{ mol H}^+} \text{ L}}$$
$$V = 8.75 \times 10^{-2} \text{ L}$$

Changing liters to milliliters, we have

$$V = 8.75 \times 10^{-2} \cancel{\text{ L}} \times \frac{1000 \text{ mL}}{\cancel{\text{ L}}} = 87.5 \text{ mL}$$

Therefore, 87.5 mL of 0.100 M HCl is required to neutralize 25.0 mL of 0.350 M NaOH.

 **Practice Problem • Exercise 15.11**

Calculate the volume of 0.10 M HNO_3 needed to neutralize 125 mL of 0.050 M KOH.

Equivalent of an acid

The amount of acid that can furnish one mole of hydrogen ions (H^+)

Equivalent of a base

The amount of base that can furnish one mole of hydroxide ions (OH^-)

Equivalent weight

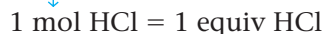
The mass (in grams) of one equivalent of an acid or a base

C. Normality

Normality is another unit of concentration that is sometimes used, especially when dealing with acids and bases. The use of normality focuses mainly on the H^+ and OH^- available in an acid–base reaction. Before we discuss normality, however, we need to define some terms. One **equivalent of an acid** is the *amount of that acid that can furnish 1 mol of H^+ ions*. Similarly, one **equivalent of a base** is defined as the *amount of that base that can furnish 1 mol of OH^- ions*. The **equivalent weight** of an acid or a base is the mass in grams of 1 equivalent (equiv) of that acid or base.

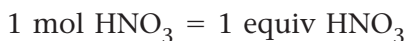
The common strong acids are HCl, HNO₃, and H₂SO₄. For HCl and HNO₃ each molecule of acid furnishes one H⁺ ion, so 1 mol of HCl can furnish 1 mol of H⁺ ions. This means that

Furnishes 1 mol of H⁺



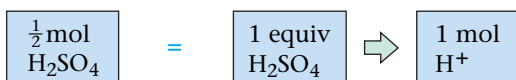
Molar mass (HCl) = equivalent weight (HCl)

Likewise, for HNO₃,

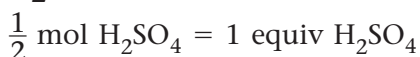


Molar mass (HNO₃) = equivalent weight (HNO₃)

However, H₂SO₄ can furnish *two* H⁺ ions per molecule, so 1 mol of H₂SO₄ can furnish *two* mol of H⁺. This means that



Because each mole of H₂SO₄ can furnish 2 mol of H⁺, we need to take only $\frac{1}{2}$ mol of H₂SO₄ to get 1 equiv of H₂SO₄. Therefore,

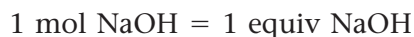


and

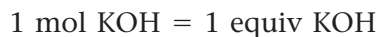
$$\begin{aligned} \text{Equivalent weight (H}_2\text{SO}_4) &= \frac{1}{2} \text{ molar mass (H}_2\text{SO}_4) \\ &= \frac{1}{2} (98 \text{ g}) = 49 \text{ g} \end{aligned}$$

The equivalent weight of H₂SO₄ is 49 g.

The common strong bases are NaOH and KOH. For NaOH and KOH, each formula unit furnishes one OH⁻ ion, so we can say



Molar mass (NaOH) = equivalent weight (NaOH)



Molar mass (KOH) = equivalent weight (KOH)

These ideas are summarized in **Table 15.2**.

Information



Table 15.2

The Molar Masses and Equivalent Weights of the Common Strong Acids and Bases

	Molar Mass (g)	Equivalent Weight (g)
<i>Acid</i>		
HCl	36.5	36.5
HNO ₃	63.0	63.0
H ₂ SO ₄	98.0	49.0 = $\frac{98.0}{2}$
<i>Base</i>		
NaOH	40.0	40.0
KOH	56.1	56.1

EXAMPLE 15.12

Solution Stoichiometry: Calculating Equivalent Weight

Phosphoric acid, H_3PO_4 , can furnish three H^+ ions per molecule. Calculate the equivalent weight of H_3PO_4 .

Solution

Where do we want to go?

Equivalent weight for $\text{H}_3\text{PO}_4 = ? \text{ g/equivalent}$

What do we know?

- Molar mass = 98.0 g/mol

How do we get there?

The key point here involves how many protons (H^+ ions) each molecule of H_3PO_4 can furnish.



Because each H_3PO_4 can furnish three H^+ ions, 1 mol of H_3PO_4 can furnish 3 mol of H^+ ions:



So 1 equiv of H_3PO_4 (the amount that can furnish 1 mol of H^+) is one-third of a mole.



This means the equivalent weight of H_3PO_4 is one-third its molar mass.

Equivalent weight	=	$\frac{\text{Molar mass}}{3}$
----------------------	---	-------------------------------

$$\text{Equivalent weight } (\text{H}_3\text{PO}_4) = \frac{\text{molar mass } (\text{H}_3\text{PO}_4)}{3} = \frac{98.0 \text{ g}}{3} = 32.7 \text{ g}$$

Normality (N) is defined as the number of equivalents of solute per liter of solution.

$$\text{Normality} = N = \frac{\text{number of equivalents}}{\text{liter of solution}} = \frac{\text{equivalents}}{\text{liter}} = \frac{\text{equiv}}{\text{L}}$$

This means that a 1 N solution contains 1 equivalent of solute per liter of solution. Notice that when we multiply the volume of a solution in liters by the normality, we get the number of equivalents.

$$N \times V = \frac{\text{equiv}}{\text{L}} \times \mathcal{L} = \text{equiv}$$

 **Active Reading Question**

What is the normality of a 4.0 M H₂SO₄ solution?

EXAMPLE 15.13

Solution Stoichiometry: Calculating Normality

A solution of sulfuric acid contains 86 g of H₂SO₄ per liter of solution. Calculate the normality of this solution.

Solution

Where do we want to go?

Normality = ? equivalent/L

What do we know?

- Molar mass = 98.0 g/mol
- H₂SO₄
 - 1.00 L
 - 86 g
- $N = \frac{\text{equivalents}}{L}$

How do we get there?

To find the number of equivalents present, we must calculate the number of equivalents represented by 86 g of H₂SO₄. To do this calculation, we focus on the definition of the equivalent: it is the amount of acid that furnishes 1 mol of H⁺. Because H₂SO₄ can furnish two H⁺ ions per molecule, 1 equiv of H₂SO₄ is $\frac{1}{2}$ mol of H₂SO₄, so

$$\text{Equivalent weight (H}_2\text{SO}_4) = \frac{\text{molar mass (H}_2\text{SO}_4)}{2} = \frac{98.0 \text{ g}}{2} = 49.0 \text{ g}$$

We have 86 g of H₂SO₄

$$86 \text{ g H}_2\text{SO}_4 \times \frac{1 \text{ equiv H}_2\text{SO}_4}{49.0 \text{ g H}_2\text{SO}_4} = 1.8 \text{ equiv H}_2\text{SO}_4$$

$$N = \frac{\text{equiv}}{L} = \frac{1.8 \text{ equiv H}_2\text{SO}_4}{1.0 \text{ L}} = 1.8 \text{ N H}_2\text{SO}_4$$

Does it make sense?

We know that 86 g is more than 1 equivalent of H₂SO₄ (49 g) so this answer makes sense.

Practice Problem • Exercise 15.13

Calculate the normality of a solution containing 23.6 g of KOH in 755 mL of solution.

information

To calculate the concentration of a solution, first write the appropriate definition. Then decide how to calculate the quantities shown in the definition.

The main advantage of using equivalents is that 1 equiv of acid contains the same number of available H^+ ions as the number of OH^- ions present in 1 equiv of base. That is,

0.75 equiv (base) will react exactly with 0.75 equiv (acid).

0.23 equiv (base) will react exactly with 0.23 equiv (acid).

And so on.

In each of these cases, the *number of* H^+ ions furnished by the sample of acid is the same as the *number of* OH^- ions furnished by the sample of base.

n equivalents of any acid will exactly neutralize n equivalents of any base.



Because we know that equal equivalents of acid and base are required for neutralization, we can say that

$$\text{equiv (acid)} = \text{equiv (base)}$$

That is,

$$N_{\text{acid}} \times V_{\text{acid}} = \text{equiv (acid)} = \text{equiv (base)} = N_{\text{base}} \times V_{\text{base}}$$

For any neutralization reaction,

$$N_{\text{acid}} \times V_{\text{acid}} = N_{\text{base}} \times V_{\text{base}}$$

CELEBRITY CHEMICAL Mercury (Hg)

Mercury was isolated from its principal ore, cinnabar (HgS), as early as 500 B.C. Cinnabar was widely used in the ancient world as a red pigment (vermilion). Mercury, whose mobility inspired its naming after the messenger of the gods in Roman mythology, was of special interest to alchemists. Alchemists hoped to find a way to change cheap metals into gold, and it was generally thought that mercury, which easily forms alloys with many metals, was the key to this process. Mercury has a common name of quicksilver.

Mercury is a very dense (13.6 times denser than water), shiny liquid that has a surprisingly high vapor pressure for a heavy metal. Because mercury vapor is quite toxic, it must be stored in stoppered containers and handled in well-ventilated areas. Many instances of suspected mercury poisoning occurred in the days before its hazards were understood.

For example, it is very possible that Sir Isaac Newton, the renowned physicist, suffered from mercury poisoning at one time in his career and

had to "retire" to the country away from his laboratory to regain his strength. Some have speculated that Mozart, the famous composer, may have died of mercury poisoning due to the mercury salts that were used to treat an illness. Also, the "mad hatters" of *Alice in Wonderland* fame have a historical connection to mercury. Apparently, the mercury salts used by workers in London to treat felt used in hats caused the hatters to become "mad" over the years.

It turns out that a change in personality is a characteristic of heavy metal poisoning.

The Mad Hatter's Tea Party, an illustration by Sir John Tenniel, from Alice's Adventures in Wonderland



EXAMPLE 15.14**Solution Stoichiometry: Using Normality in Calculations**

What volume of a 0.075 N KOH solution is required to react exactly with 0.135 L of 0.45 N H_3PO_4 ?

Solution**Where do we want to go?**

Volume KOH required to react = ? L

What do we know?

- 0.075 N KOH
- H_3PO_4
 - 0.135 L
 - 0.45 N
- In neutralization \rightarrow equivalents_{acid} = equivalents_{base}
- $N_{\text{acid}} \times V_{\text{acid}} = N_{\text{base}} \times V_{\text{base}}$

How do we get there?

We know that for neutralization, equiv (acid) = equiv (base), or

$$N_{\text{acid}} \times V_{\text{acid}} = N_{\text{base}} \times V_{\text{base}}$$

We want to calculate for the volume of base, V_{base} , so we solve for V_{base} by dividing both sides by N_{base} .

$$\frac{N_{\text{acid}} \times V_{\text{acid}}}{N_{\text{base}}} = \frac{\cancel{N_{\text{base}}} \times V_{\text{base}}}{\cancel{N_{\text{base}}}} = V_{\text{base}}$$

Now we can substitute the given values $N_{\text{acid}} = 0.45 \text{ N}$, $V_{\text{acid}} = 0.135 \text{ L}$, and $N_{\text{base}} = 0.075 \text{ N}$ into the equation.

$$V_{\text{base}} = \frac{N_{\text{acid}} \times V_{\text{acid}}}{N_{\text{base}}} = \frac{(0.45 \frac{\text{equiv}}{\text{L}}) (0.135 \text{ L})}{0.075 \frac{\text{equiv}}{\text{L}}} = 0.81 \text{ L}$$

This gives $V_{\text{base}} = 0.81 \text{ L}$, so 0.81 L of 0.075 N KOH is required to react exactly with 0.135 L of 0.45 N H_3PO_4 .

Practice Problem • Exercise 15.14

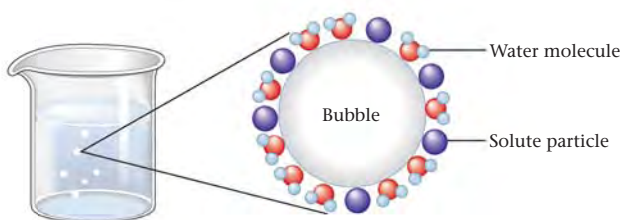
What volume of 0.50 N H_2SO_4 is required to react exactly with 0.250 L of 0.80 N KOH?

D. Boiling Point and Freezing Point

We saw in Chapter 14 that at an atmospheric pressure of 1 atm water freezes at 0 °C and boils at 100 °C. If we dissolve a solute such as NaCl in water, does the presence of the solute affect the freezing point of the water? Yes. A 1.0 M NaCl solution freezes at about -1 °C and boils at about 104 °C. The presence of solute “particles” in the water extends the liquid range of water—that is, water containing a solute exists as a liquid over a wider temperature range than does pure water.

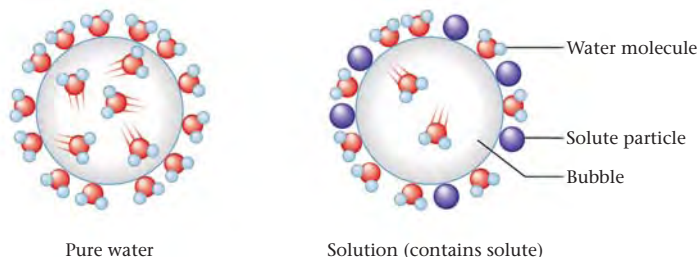
Why does this happen? For example, what effect do the particles have that causes the boiling point of water to increase? To answer this question we need to reconsider the boiling process. Recall that for water to boil, bubbles must be able to form in the interior of the liquid. That is, energetic water molecules in the bubble must be able to exert an internal pressure large enough to push back the atmospheric pressure.

Now consider the situation in a solution that contains a solute.



The forming bubble is now surrounded by solute particles as well as water molecules. As a result, solute particles may block the pathway of some of the water molecules trying to enter the bubble. Because fewer H₂O molecules can enter the bubble, the molecules that do get inside it must have relatively high energies to produce the internal pressure necessary to maintain the bubble.

Compare the situation with and without a solute shown below.



To produce the same pressure, the smaller number of molecules found in the bubble in the solution must have higher energies (speeds) than those in the bubble in pure water.

Thus, to cause the water in the solution to boil, we must heat the water to a temperature higher than 100 °C to produce the higher-energy water molecules needed. In other words, the boiling point of the solution occurs at a higher temperature than does the boiling point of pure water. The presence of the solute increases the boiling point of water. The more solute present, the higher the boiling point, because the smaller number of H₂O molecules able to enter the bubble must have increasing energies.

Note that the raising of the boiling point depends on the number and not the specific identity of the solute particles. It's the number that matters—the more particles present, the more those particles block water molecules from entering the bubble. A solution property that depends on the *number* of solute particles present is called a **colligative property**. Raising of the boiling point by a solute is one of the colligative properties of solutions.

Another colligative property of solutions is the lowering of the freezing point. (As mentioned before, a 1 M NaCl solution freezes at approximately $-1\text{ }^{\circ}\text{C}$.) This property is particularly important in cold areas of the world where salt is applied to icy roads. When the applied salt dissolves in the thin layer of water on the surface of ice, forming a very concentrated solution, it causes the ice to melt (it lowers the freezing point of water). The same property is also used to protect automobile engines. The coolant used

in engines is a solution containing ethylene glycol $\left(\begin{array}{c} \text{H} \quad \text{H} \\ | \quad | \\ \text{H}-\text{C}-\text{C}-\text{H} \\ | \quad | \\ \text{OH} \quad \text{OH} \end{array} \right)$ dissolved

in water. The solute (ethylene glycol) raises the boiling point and lowers the freezing point of the water, thereby protecting the engine from overheating and freezing.

Colligative property

A property that is dependent only on the number of solute particles present in solution

Active Reading Question

When salt is added to water, what is true about the freezing point of the solution compared to the freezing point of pure water? What is true about the boiling point of the solution compared to the boiling point of water?

SECTION 15.3 REVIEW QUESTIONS

- The key to solving stoichiometry problems is the mole. How do we find the number of moles when solutions are mixed to produce a reaction?
- How many milliliters of 0.10 M $\text{Pb}(\text{NO}_3)_2$ solution are required to precipitate all the lead, as PbI_2 , from 125.0 mL of 0.10 M NaI solution?
- Determine the mass of Ag_2CrO_4 produced when 100.0 mL of 0.500 M silver nitrate is added to 100.0 mL of 0.500 M potassium chromate.
- Once you have determined moles of H^+ or OH^- in a neutralization reaction how can you find the volume of the substance? What is the critical information you need to determine volume?
- What volume of 0.25 M NaOH is required to neutralize 125.0 mL of 0.15 M HCl?
- How is normality different from molarity?
- Is a 1 M H_2SO_4 solution the same as a 1 N H_2SO_4 solution? Explain.
- Other than for taste, why is salt added to water when cooking pasta?



Key Terms

15.1 Solution
Solvent
Solute
Aqueous solution
Saturated
Unsaturated
Supersaturated
Concentrated
Dilute

15.2 Mass percent
Molarity (M)
Standard solution
Dilution

15.3 Neutralization reaction
Equivalent of an acid
Equivalent of a base
Equivalent weight
Colligative property

Key Ideas

15.1 Forming Solutions

- A solution is a homogeneous mixture of a solute dissolved in a solvent.
- Substances with similar polarities tend to dissolve with each other to form a solution.
- Water is a very polar substance and tends to dissolve ionic solids or other polar substances.
- Various terms are used to describe solutions:
 - Saturated—contains the maximum possible dissolved solid
 - Unsaturated—not saturated
 - Supersaturated—contains more dissolved solid than should dissolve at a given temperature
 - Concentrated—contains a relatively large amount of solute
 - Dilute—contains a relatively small amount of solute
- The rate of dissolution is affected by
 - Surface area of solute
 - Stirring
 - Temperature

15.2 Describing Solution Composition

- Descriptions of solution composition:
 - Mass percent of solute = $\frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$
 - Molarity = $\frac{\text{moles of solute}}{\text{liters of solution}}$
 - Normality = $\frac{\text{equivalents of solute}}{\text{liters of solution}}$
- Dilution of a solution occurs when additional solvent is added to lower the concentration of a solution.
 - No solute is added so
 $\text{mol solute (before dilution)} = \text{mol solute (after dilution)}$
- A standard solution can be diluted to produce solutions with appropriate concentrations for various laboratory procedures.

15.3 Properties of Solutions

- Moles of a dissolved reactant or product can be calculated from the known concentration and volume of the substance
$$\text{Mol} = (\text{concentration})(\text{volume})$$
- The properties of a solvent are affected by dissolving a solute.
- The boiling point of a solvent increases as the amount of dissolved solute increases.
- The melting point of a solvent decreases as the amount of dissolved solute increases.

All exercises with blue numbers have answers in the back of this book.

15.1 Forming Solutions

A. Solubility

1. Explain why a solution is a *homogeneous* mixture. Give two examples of homogeneous mixtures.
2. In a solution, the substance present in the largest amount is called the _____, whereas the other substances present are called the _____.
3. Discuss how an *ionic* solute dissolves in water. How are the strong interionic forces in the solid overcome to permit the solid to dissolve? How are the dissolved positive and negative ions shielded from one another, preventing them from recombining to form the solid?
4. Why are some molecular solids (such as sugar or ethyl alcohol) soluble in water, while other molecular solids (such as petroleum) are insoluble in water? What structural features(s) of *some* molecular solids may tend to make them soluble in water?

B. Solution Composition: An Introduction

5. A solution that contains as much solute as will dissolve at a given temperature is said to be _____.
6. A solution that has not reached its limit of dissolved solute is said to be _____.
7. A solution is a homogeneous mixture and, unlike a compound, has _____ composition.
8. The label “concentrated H_2SO_4 ” on a bottle means that there is a relatively _____ amount of H_2SO_4 present in the solution.

C. Factors Affecting the Rate of Dissolving

9. What does it mean to increase the surface area of a solid? Explain why this change causes an increase in the rate of dissolving.
10. Use a molecular explanation to explain why increasing the temperature speeds up the rate of dissolving a solid in a liquid.
11. Explain why the solubility of a gas generally decreases with an increase in temperature.

15.2 Describing Solution Composition

A. Solution Composition: Mass Percent

12. Calculate the mass percent of calcium chloride in each of the following solutions.
 - a. 5.00 g of calcium chloride in 95.0 g of water
 - b. 1.00 g of calcium chloride in 19.0 g of water
 - c. 15.0 g of calcium chloride in 285 g of water
 - d. 2.00 mg of calcium chloride in 0.0380 g of water
13. Calculate the mass, in grams, of NaCl present in each of the following solutions.
 - a. 11.5 g of 6.25% NaCl solution
 - b. 6.25 g of 11.5% NaCl solution
 - c. 54.3 g of 0.91% NaCl solution
 - d. 452 g of 12.3% NaCl solution
14. A laboratory assistant prepared a potassium chloride solution for her class by dissolving 5.34 g of KCl in 152 g of water. What is the mass percent of the solution she prepared?
15. If 67.1 g of CaCl_2 is added to 275 g of water, calculate the mass percent of CaCl_2 in the solution.
16. What mass of each solute is present in 285 g of a solution that contains 5.00% by mass NaCl and 7.50% by mass Na_2CO_3 ?
17. A hexane solution contains as impurities 5.2% (by mass) heptane and 2.9% (by mass) pentane. Calculate the mass of each component present in 93 g of the solution.

B. Solution Composition: Molarity

18. A solution that is labeled “0.105 M NaOH” would contain _____ mol of NaOH per liter of solution.
19. How many moles of each ion are present, per liter, in a solution that is labeled “0.221 M CaCl_2 ”?
20. If you were to prepare exactly 1.00 L of a 5 M NaCl solution, you would *not* need exactly 1.00 L of water. Explain.

21. For each of the following solutions, the number of moles of solute is given, followed by the total volume of solution prepared. Calculate the molarity.

- 0.50 mol KBr; 250 mL
- 0.50 mol KBr; 500. mL
- 0.50 mol KBr; 750 mL
- 0.50 mol KBr; 1.0 L

22. For each of the following solutions, the mass of the solute is given, followed by the total volume of the solution prepared. Calculate the molarity.

- 4.25 g CuCl₂; 125 mL
- 0.101 g NaHCO₃; 11.3 mL
- 52.9 g Na₂CO₃; 1.15 L
- 0.14 mg KOH; 1.5 mL

23. If a 45.3-g sample of potassium nitrate is dissolved in enough water to make 225 mL of solution, what will be the molarity?

24. An alcoholic iodine solution (“tincture” of iodine) is prepared by dissolving 5.15 g of iodine crystals in enough alcohol to make a volume of 225 mL. Calculate the molarity of iodine in the solution.

25. Suppose 1.01 g of FeCl₃ is placed in a 10.0-mL volumetric flask, water is added, the mixture is shaken to dissolve the solid, and then water is added to the calibration mark of the flask. Calculate the molarity of each ion present in the solution.

26. If 495 g of NaOH is dissolved to a final total volume of 20.0 L, what is the molarity of the solution?

27. Calculate the number of *moles* and the number of *grams* of the indicated solutes present in each of the following solution samples.

	mol	g
a. 127 mL of 0.105 M HNO ₃		
b. 155 mL of 15.1 M NH ₃		
c. 2.51 L of 2.01 × 10 ⁻³ M KSCN		
d. 12.2 mL of 2.45 M HCl		

28. What mass (in grams) of NH₄Cl is needed to prepare 450. mL of 0.251 M NH₄Cl solution?

29. Calculate the number of moles of *each* ion present in each of the following solutions.

- 10.2 mL of 0.451 M AlCl₃ solution
- 5.51 L of 0.103 M Na₃PO₄ solution
- 1.75 mL of 1.25 M CuCl₂ solution
- 25.2 mL of 0.00157 M Ca(OH)₂ solution

30. Standard silver nitrate solutions are used in the analysis of samples containing chloride ion. How many grams of silver nitrate are needed to prepare 250. mL of a 0.100 M AgNO₃ solution?

C. Dilution

31. Calculate the new molarity that results when 250. mL of water is added to each of the following solutions.

- 125 mL of 0.251 M HCl
- 445 mL of 0.499 M H₂SO₄
- 5.25 L of 0.101 M HNO₃
- 11.2 mL of 14.5 M HC₂H₃O₂

32. Many laboratories keep bottles of 3.0 M solutions of the common acids on hand. Given the following molarities of the concentrated acids, determine how many milliliters of each concentrated acid would be required to prepare 225 mL of a 3.0 M solution of the acid.

Acid	Molarity of Concentrated Reagent
HCl	12.1 M
HNO ₃	15.9 M
H ₂ SO ₄	18.0 M
HC ₂ H ₃ O ₂	17.5 M
H ₃ PO ₄	14.9 M

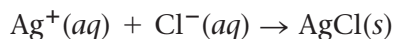
33. A chemistry student needs 125 mL of 0.150 M NaOH solution for an experiment, but the only solution available in the laboratory is 3.02 M. Describe how the student could prepare the solution he needs.

34. How much *water* must be added to 500. mL of 0.200 M HCl to produce a 0.150 M solution? (Assume that the volumes are additive.)

15.3 Properties of Solutions

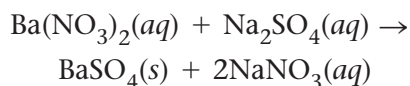
A. Stoichiometry of Solution Reactions

35. One way to determine the amount of chloride ion in a water sample is to react the sample with standard AgNO_3 solution to produce solid AgCl .

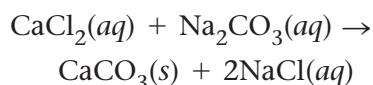


If a 25.0-mL water sample requires 27.2 mL of 0.104 M AgNO_3 in such a reaction, what is the concentration of Cl^- in the sample?

36. What volume (in mL) of 0.25 M Na_2SO_4 solution is needed to precipitate all the barium, as $\text{BaSO}_4(s)$, from 12.5 mL of 0.15 M $\text{Ba}(\text{NO}_3)_2$ solution?



37. If 36.2 mL of 0.158 M CaCl_2 solution is added to 37.5 mL of 0.149 M Na_2CO_3 , what mass of calcium carbonate, CaCO_3 , will be precipitated? The reaction is



38. When aqueous solutions of lead(II) ion are treated with potassium chromate solution, a bright yellow precipitate of lead(II) chromate, PbCrO_4 , forms. How many grams of lead chromate form when a 1.00-g sample of $\text{Pb}(\text{NO}_3)_2$ is added to 25.0 mL of 1.00 M K_2CrO_4 solution?

B. Neutralization Reactions

39. What volume of 0.200 M HCl solution is needed to neutralize 25.0 mL of 0.150 M NaOH solution?
40. The concentration of a sodium hydroxide solution is to be determined. A 50.0-mL sample of 0.104 M HCl solution requires 48.7 mL of the sodium hydroxide solution to reach the point of neutralization. Calculate the molarity of the NaOH solution.
41. What volume of 1.00 M NaOH is required to neutralize each of the following solutions?
- 25.0 mL of 0.154 M acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$
 - 35.0 mL of 0.102 M hydrofluoric acid, HF
 - 10.0 mL of 0.143 M phosphoric acid, H_3PO_4
 - 35.0 mL of 0.220 M sulfuric acid, H_2SO_4

C. Normality

42. Explain why the equivalent weight of H_2SO_4 is half the molar mass of this substance. How many hydrogen ions does each H_2SO_4 molecule produce when reacting with an excess of OH^- ions?
43. How many equivalents of hydroxide ion are needed to react with 1.53 equivalents of hydrogen ion? How can you tell when no balanced chemical equation is given for the reaction?
44. For each of the following solutions, the mass of solute taken is indicated, along with the total volume of solution prepared. Calculate the normality of each solution.
- 0.113 g NaOH ; 10.2 mL
 - 12.5 mg $\text{Ca}(\text{OH})_2$; 100. mL
 - 12.4 g H_2SO_4 ; 155 mL
45. Calculate the normality of each of the following solutions.
- 0.134 M NaOH
 - 0.00521 M $\text{Ca}(\text{OH})_2$
 - 4.42 M H_3PO_4
46. A solution of phosphoric acid, H_3PO_4 , is found to contain 35.2 g of H_3PO_4 per liter of solution. Calculate the molarity and normality of the solution.
47. What volume of 0.172 N H_2SO_4 is required to neutralize 56.2 mL of 0.145 M NaOH ?
48. What volume of 0.151 N NaOH is required to neutralize 24.2 mL of 0.125 N H_2SO_4 ? What volume of 0.151 N NaOH is required to neutralize 24.1 mL of 0.125 M H_2SO_4 ?

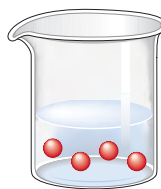
D. Boiling Point and Freezing Point

49. What is meant by the term "colligative property"?
50. Explain on a molecular level why the increase in boiling point is a colligative property.
51. Antifreeze that you use in your car could also be called "antiboil." Explain why.

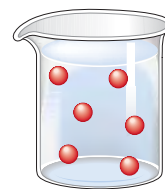
Critical Thinking

- 52.** Suppose 50.0 mL of 0.250 M CoCl_2 solution is added to 25.0 mL of 0.350 M NiCl_2 solution. Calculate the concentration, in moles per liter, of each of the ions present after mixing. Assume that the volumes are additive.
- 53.** Calculate the mass of AgCl formed, and the concentration of silver ion remaining in solution, when 10.0 g of solid AgNO_3 is added to 50. mL of 1.0×10^{-2} M NaCl solution. Assume there is no volume change upon addition of the solid.
- 54.** What mass of BaSO_4 will be precipitated from a large container of concentrated $\text{Ba}(\text{NO}_3)_2$ solution if 37.5 mL of 0.221 M H_2SO_4 is added?
- 55.** Strictly speaking, the solvent is the component of a solution that is present in the largest amount on a *mole* basis. For solutions involving water, water is almost always the solvent because there tend to be many more water molecules present than molecules of any conceivable solute. To see why this is so, calculate the number of moles of water present in 1.0 L of water. Recall that the density of water is very nearly 1.0 g/mL under most conditions.
- 56.** A 14.2-g sample of CaCl_2 is added to a 50.0-mL volumetric flask. After dissolving the salt, water is added to the calibration mark of the flask. Calculate the molarity of the solution.
- 57.** Calculate the new molarity when 150. mL of water is added to each of the following solutions.
- 125 mL of 0.200 M HBr
 - 155 mL of 0.250 M $\text{Ca}(\text{C}_2\text{H}_3\text{O}_2)_2$
 - 0.500 L of 0.250 M H_3PO_4
 - 15 mL of 18.0 M H_2SO_4

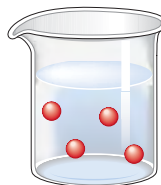
- 58.** Calculate the normality of each of the following solutions.
- 0.50 M acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$
 - 0.00250 M sulfuric acid, H_2SO_4
 - 0.10 M potassium hydroxide, KOH
- 59.** If 27.5 mL of 3.5×10^{-2} N $\text{Ca}(\text{OH})_2$ solution is needed to neutralize 10.0 mL of nitric acid solution of unknown concentration, what is the normality of the nitric acid?
- 60.** The figures below are molecular-level representations of four aqueous solutions of the same solute. Arrange the solutions from most to least concentrated.



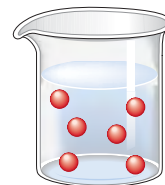
Solution A
Volume = 1.0 L



Solution B
Volume = 4.0 L



Solution C
Volume = 2.0 L



Solution D
Volume = 2.0 L

- 1 Which of the following aqueous solutions has the greatest number of ions in solution?
- A 2.0-L of 1.50 M sodium phosphate
 - B 3.0-L of 1.50 M sodium chloride
 - C 2.0-L of 2.00 M potassium fluoride
 - D 1.0-L of 3.00 M sodium sulfate
- 2 When a solvent contains as much of the solute as it can hold, the solution is said to be
- A unsaturated
 - B supersaturated
 - C saturated
 - D dilute
- 3 Which of the following actions does not generally increase the rate of a solute dissolving in a solvent?
- A Increasing the surface area of the solute
 - B Stirring the solution
 - C Raising the temperature of the solution
 - D Adding more solute
- 4 Determine the concentration of a solution made by dissolving 10.0 g of sodium chloride in 750.0 mL of solution.
- A 0.133 M
 - B 0.171 M
 - C 0.228 M
 - D 0.476 M
- 5 You have two solutions of sodium chloride. One is a 2.00 M solution and the other is a 4.00 M solution. You have much more of the 4.00 M solution and you add the solutions together. Which of the following could be the concentration of the final solution?
- A 3.00 M
 - B 3.80 M
 - C 6.00 M
 - D 8.00 M
- 6 What is the minimum volume of a 2.00 M NaOH (aq) solution needed to make 150.0 mL of a 0.800 M NaOH (aq) solution?
- A 60.0 mL
 - B 90.0 mL
 - C 120. mL
 - D 150. mL
- 7 You have equal masses of different solutes dissolved in equal volumes of solution. Which of the solutes would make the solution with the highest concentration measured in molarity?
- A NaOH
 - B KCl
 - C KOH
 - D LiOH
- 8 You react 250.0 mL of 0.10 M barium nitrate with 200.0 mL of 0.10 M potassium phosphate. What ions are left in solution after the reaction is complete?
- A potassium ion, nitrate ion, barium ion
 - B potassium ion, nitrate ion, phosphate ion
 - C potassium ion, nitrate ion
 - D barium ion, nitrate ion, phosphate ion
- 9 What volume of 0.500 M NaOH is needed to neutralize 45.0 mL of 0.400 M HCl?
- A 36.0 mL
 - B 45.0 mL
 - C 56.3 mL
 - D 71.3 mL
- 10 Which of the following has the largest equivalent weight?
- A HCl
 - B H₂SO₄
 - C H₃PO₄
 - D HNO₃
- 11 When a solute is dissolved in a solvent, the boiling point of the solution is higher than the boiling point of the pure solvent. Answer the following questions about this phenomenon.
- a. Explain why adding a solute to a solvent raises the boiling point.
 - b. Explain why adding 1 mole of table salt to water raises the boiling point more than adding 1 mole of table sugar to the same amount of water.

LOOKING AHEAD

16.1 Properties of Acids and Bases

- A. Acids and Bases
 - The Arrhenius Model
 - The Brønsted–Lowry Model
- B. Acid Strength
- C. Water as an Acid and a Base

16.2 Determining the Acidity of a Solution

- A. The pH Scale
- B. Measuring pH
- C. Calculating the pH of Strong Acid Solutions

16.3 Titrations and Buffers

- A. Acid–Base Titrations
- B. Buffered Solutions

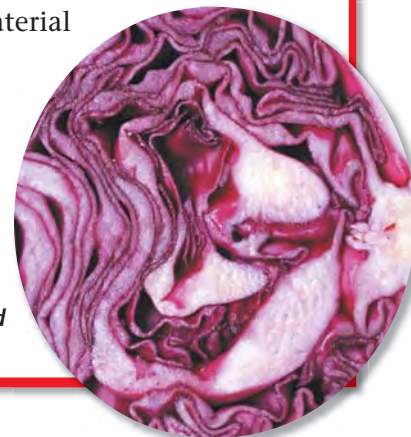


Fruits contain various amounts of citric acid.

IN YOUR LIFE

Acids are very important substances. They cause lemons to be sour, digest food in the stomach (and sometimes cause heartburn), dissolve rock to make fertilizer, dissolve your tooth enamel to form cavities. Acids are essential industrial chemicals. In fact, the chemical in first place in terms of the amount manufactured in the United States is sulfuric acid, H_2SO_4 . Eighty *billion* pounds of this material are used every year in the manufacture of fertilizers, detergents, plastics, pharmaceuticals, storage batteries, and metals. The acid-base properties of substances also can be used to make interesting novelties such as the foaming chewing gum described in Chemistry in Your World.

Red cabbage juice can be used as an indicator of acidity.

**WHAT DO YOU KNOW?****Prereading Questions**

1. What is meant by the term *acid*? Name two products you think are acidic.
2. What is meant by the term *base*?
3. What do you know about the pH scale?
4. What is the net ionic equation for the reaction between a strong acid and a strong base?

SECTION 16.1

Properties of Acids and Bases

Key Terms

- Acids
- Bases
- Arrhenius concept of acids and bases
- Brønsted–Lowry model
- Conjugate acid
- Conjugate base
- Conjugate acid–base pair
- Hydronium ion
- Completely ionized (completely dissociated)
- Strong acid
- Weak acid
- Diprotic acid
- Oxyacids
- Organic acids
- Carboxyl group
- Amphoteric substance
- Ionization of water
- Ion-product constant

Objectives

- To learn about two models of acids and bases
- To understand the relationship of conjugate acid–base pairs
- To understand the concept of acid strength
- To understand the relationship between acid strength and the strength of the conjugate base
- To learn about the ionization of water

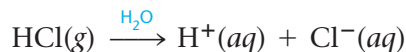
A. Acids and Bases

Acids were first recognized as substances that taste sour. Vinegar tastes sour because it is a dilute solution of acetic acid; citric acid is responsible for the sour taste of a lemon. Bases, sometimes called *alkalis*, are characterized by their bitter taste and slippery feel. Most hand soaps and commercial preparations for unclogging drains are highly basic.

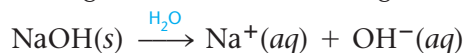
The Arrhenius Model

The first person to recognize the essential nature of acids and bases was Svante Arrhenius. On the basis of his experiments with electrolytes, Arrhenius postulated that **acids** produce hydrogen ions in aqueous solution, whereas **bases** produce hydroxide ions.

For example, when hydrogen chloride gas is dissolved in water each molecule produces ions as follows:



This solution is the strong acid known as hydrochloric acid. On the other hand, when solid sodium hydroxide is dissolved in water, its ions separate producing a solution containing Na^+ and OH^- ions.



This solution is called a strong base.

The Brønsted–Lowry Model

Although the **Arrhenius concept of acids and bases** was a major step forward in understanding acid–base chemistry, this concept is limited because it allows for only one kind of base—the hydroxide ion. A more general definition of acids and bases was suggested by the Danish chemist Johannes Brønsted and the English chemist Thomas Lowry.

In the **Brønsted–Lowry model**, an acid is a proton (H^+) donor, and a base is a proton acceptor.

Acid

A substance that produces hydrogen ions (H^+) in solution; a proton donor

Base

A substance that produces hydroxide ions (OH^-) in solution; a substance that accepts a proton

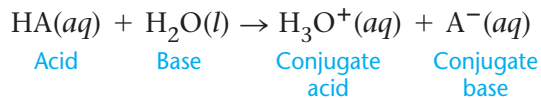
Arrhenius concept of acids and bases

Acids produce hydrogen ions (H^+) in solution; bases produce hydroxide ions (OH^-) in solution.

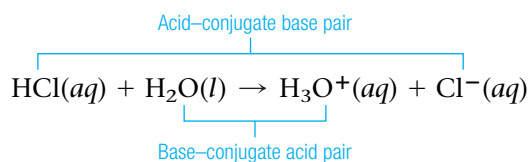


The label on a bottle of concentrated hydrochloric acid

According to the Brønsted–Lowry model, the general reaction that occurs when an acid is dissolved in water can best be represented as an acid (HA) donating a proton to a water molecule to form a new acid (the **conjugate acid**) and a new base (the **conjugate base**).

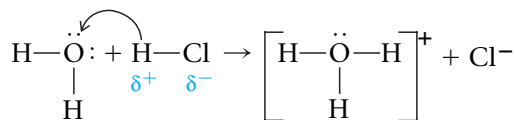


This model emphasizes the significant role of the polar water molecule in pulling the proton from the acid. Note that the conjugate base is everything that remains of the acid molecule after a proton is lost. The conjugate acid is formed when the proton is transferred to the base. A **conjugate acid–base pair** consists of two substances related to each other by the donating and accepting of a *single proton*. In the above equation there are two conjugate acid–base pairs: HA (acid) and A[−] (base), and H₂O (base) and H₃O⁺ (acid). For example, when hydrogen chloride is dissolved in water it behaves as an acid.



In this case HCl is the acid that loses an H⁺ ion to form Cl[−], its conjugate base. On the other hand, H₂O (behaving as a base) gains an H⁺ ion to form H₃O⁺ (the conjugate acid).

How can water act as a base? Remember that the oxygen of the water molecule has two unshared electron pairs, either of which can form a covalent bond with an H⁺ ion. When gaseous HCl dissolves in water, the following reaction occurs.



Note that an H⁺ ion is transferred from the HCl molecule to the water molecule to form H₃O⁺, which is called the **hydronium ion**.



Active Reading Question

How are the substances in a conjugate acid–base pair related to each other?

Information

Recall that (aq) means the substance is hydrated—it has water molecules clustered around it.

Conjugate acid

The substance formed when a proton is added to a base

Conjugate base

The remaining substance when a proton is lost from an acid

Conjugate acid–base pair

Two substances related to each other by the donating and accepting of a single proton

Hydronium ion

H₃O⁺

EXAMPLE 16.1

Identifying Conjugate Acid–Base Pairs

Which of the following represent conjugate acid–base pairs?

- HF, F⁻
- NH₄⁺, NH₃
- HCl, H₂O

Solution

- a. and b. HF, F⁻ and NH₄⁺, NH₃ are conjugate acid–base pairs because the two species differ by one H⁺.



- c. HCl and H₂O are not a conjugate acid–base pair because they are not related by the removal or addition of one H⁺. The conjugate base of HCl is Cl⁻. The conjugate acid of H₂O is H₃O⁺.

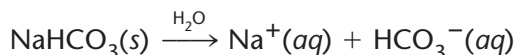
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Science, Technology, and Society

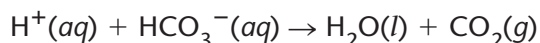
Gum That Foams

Mad Dawg chewing gum is a practical joker's dream come true. It is noticeably sour when someone first starts to chew it, but the big surprise comes about ten chews later when brightly colored foam oozes from the person's mouth. Although the effect is dramatic, the cause is simple acid–base chemistry.

The foam consists of sugar and saliva churned into a bubbling mess by carbon dioxide released from the gum. The carbon dioxide is formed when sodium bicarbonate (NaHCO₃) present in the gum is mixed with citric acid and malic acid (also present in the gum) in the moist environment of the mouth. As NaHCO₃ dissolves in the water of the saliva, it separates into its ions:



The bicarbonate ion, when exposed to H⁺ ions from acids, decomposes to carbon dioxide and water.



The acids present in the gum also cause it to be sour, stimulating extra salivation and thus extra foam.

Although the chemistry behind Mad Dawg is well understood, the development of the gum into a safe, but fun, product was not so easy. In fact, early versions of the gum exploded because the acids and the sodium bicarbonate mixed prematurely. As solids, citric and malic acids and sodium bicarbonate do not react with each other. However, the presence of water frees the ions to move and react. In the manufacture of the gum, colorings and flavorings are applied as aqueous solutions. The water caused the gum to explode in early attempts to manufacture it. The makers of Mad Dawg obviously solved the problem.



Chewing Mad Dawg gum

EXAMPLE 16.2**Writing Conjugate Bases**

Write the conjugate base for each of the following:

- HClO_4
- H_3PO_4
- CH_3NH_3^+

Solution

To get the conjugate base for an acid, we must remove an H^+ ion.

- $$\text{HClO}_4 \rightarrow \text{H}^+ + \text{ClO}_4^-$$

Acid Conjugate base
- $$\text{H}_3\text{PO}_4 \rightarrow \text{H}^+ + \text{H}_2\text{PO}_4^-$$

Acid Conjugate base
- $$\text{CH}_3\text{NH}_3^+ \rightarrow \text{H}^+ + \text{CH}_3\text{NH}_2$$

Acid Conjugate base

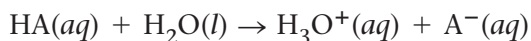
Practice Problem • Exercise 16.2

Which of the following represent conjugate acid–base pairs?

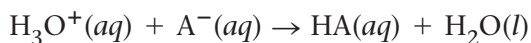
- $\text{H}_2\text{O}, \text{H}_3\text{O}^+$
- $\text{OH}^-, \text{HNO}_3$
- $\text{H}_2\text{SO}_4, \text{SO}_4^{2-}$
- $\text{HC}_2\text{H}_3\text{O}_2, \text{C}_2\text{H}_3\text{O}_2^-$

B. Acid Strength

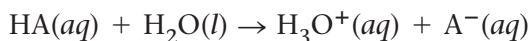
We have seen that when an acid dissolves in water, a proton is transferred from the acid to water:



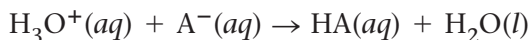
In this reaction a new acid, H_3O^+ (called the conjugate acid), and a new base, A^- (the conjugate base), are formed. The conjugate acid and base can react with one another,



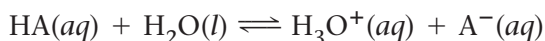
to re-form the parent acid and a water molecule. Therefore, this reaction can occur “in both directions.” The forward reaction is



and the reverse reaction is



Note that the products in the forward reaction are the reactants in the reverse reaction. We usually represent the situation in which the reaction can occur in both directions by double arrows:



Strong acid

An acid that completely dissociates to produce H^+ ions in solution

Weak acid

An acid that dissociates to a slight extent in aqueous solution

CRITICAL THINKING ?

Vinegar contains acetic acid and is used in salad dressings.

What if acetic acid was a strong acid instead of a weak acid? Would it be safe to use vinegar as a salad dressing?

Information

A strong acid is completely dissociated in water. No HA molecules remain. Only H_3O^+ , H_2O , and A^- are present.



A hydrochloric acid solution readily conducts electric current, as shown by the brightness of the bulb.

This situation represents a competition for the H^+ ion between H_2O (in the forward reaction) and A^- (in the reverse reaction). If H_2O “wins” this competition—that is, if H_2O has a very high attraction for H^+ compared to A^- , then the solution will contain mostly H_3O^+ and A^- . We describe this situation by saying that the H_2O molecule is a much stronger base (more attraction for H^+) than A^- . In this case the forward reaction predominates:



We say that the acid HA is **completely ionized** or **completely dissociated**. This situation represents a **strong acid**.

The opposite situation can also occur. Sometimes A^- “wins” the competition for the H^+ ion. In this case A^- is a much stronger base than H_2O and the reverse reaction predominates:



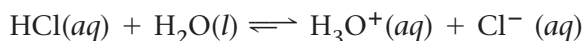
Here A^- has a much larger attraction for H^+ than does H_2O , and most of the HA molecules remain intact. This situation represents a **weak acid**.

Active Reading Question

Why is it not correct to say that a weak acid is a diluted strong acid?

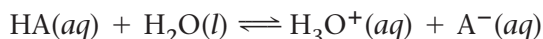
Strong Acids We can determine what is actually going on in a solution by measuring its ability to conduct an electric current. Recall from Chapter 8 that a solution can conduct a current in proportion to the number of ions that are present. When 1 mole of solid sodium chloride is dissolved in 1 L of water, the resulting solution is an excellent conductor of an electric current because the Na^+ and Cl^- ions separate completely. We call NaCl a strong electrolyte.

Similarly, when 1 mole of hydrogen chloride is dissolved in 1 L of water, the resulting solution is an excellent conductor. Therefore, hydrogen chloride is also a strong electrolyte, which means that each HCl molecule must produce H^+ and Cl^- ions. This tells us that the forward reaction predominates:



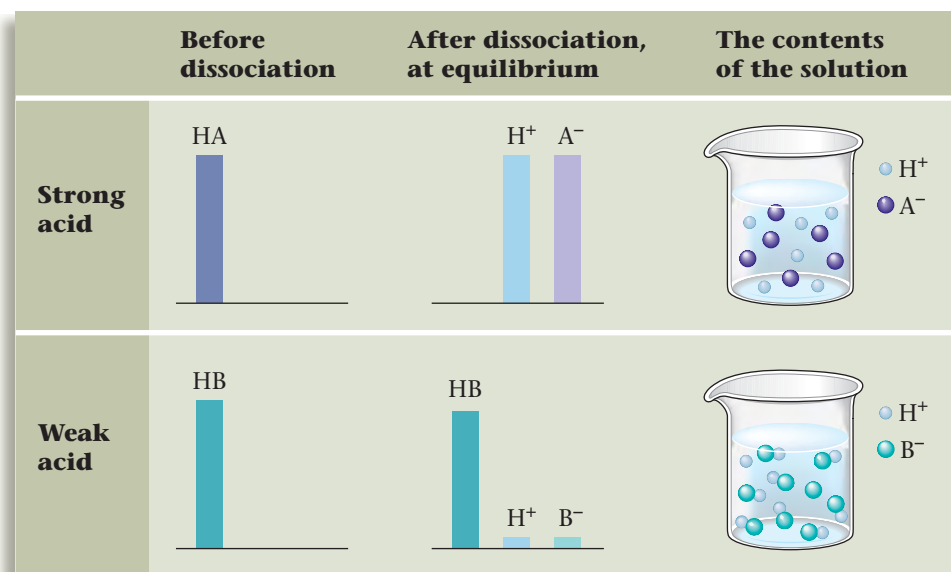
(Accordingly, the arrow pointing right is longer than the arrow pointing left.) In solution there are virtually no HCl molecules, only H^+ and Cl^- ions. This shows that Cl^- is a very poor base compared to the H_2O molecule; it has virtually no ability to attract H^+ ions in water. This aqueous solution of hydrogen chloride (called *hydrochloric acid*) is a strong acid.

In general, the strength of an acid is defined by the position of its ionization (dissociation) reaction:



A strong acid is one for which *the forward reaction predominates*. This means that almost all the original HA is dissociated (ionized) (see **Figure 16.1**). There is an important connection between the strength of an acid and that of its conjugate base.

A strong acid contains a relatively weak conjugate base—one that has a low attraction for protons.

**Figure 16.1**

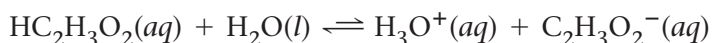
Graphical representation of the behavior of acids of different strengths in aqueous solution. A strong acid is completely dissociated. In contrast, only a small fraction of the molecules of a weak acid are dissociated.

A strong acid can be described as an acid whose conjugate base is a much weaker base than water (see **Figure 16.2**). In this case the water molecules win the competition for the H⁺ ions.

Active Reading Question

Which is a stronger base, a conjugate base from a strong acid or a conjugate base from a weak acid?

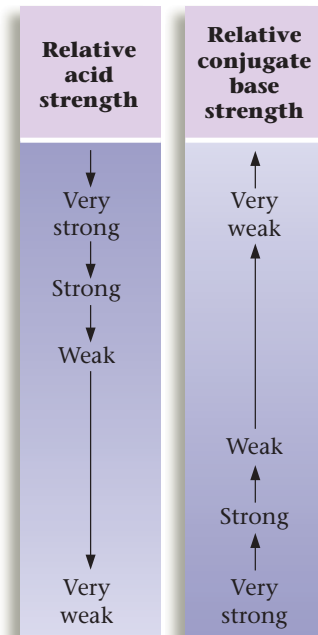
Weak Acids In contrast to hydrochloric acid, when acetic acid, HC₂H₃O₂, is dissolved in water, the resulting solution conducts an electric current only weakly. That is, acetic acid is a weak electrolyte, which means that only a few ions are present. In other words, for the following reaction, the reverse reaction predominates (thus the arrow pointing left is longer).



In fact, measurements show that only about one in one hundred (1%) of the HC₂H₃O₂ molecules is dissociated (ionized) in a 0.1 M solution of acetic acid. Thus acetic acid is a weak acid. When acetic acid molecules are placed in water, almost all of the molecules remain undissociated. This tells us that the acetate ion, C₂H₃O₂⁻, is an effective base—it very successfully attracts H⁺ ions in water. This means that acetic acid remains largely in the form of HC₂H₃O₂ molecules in solution. A weak acid is one for which the *reverse reaction predominates*.



Most of the acid originally placed in the solution is still present as HA at equilibrium. That is, a weak acid dissociates (ionizes) only to a very small extent in aqueous solution (Figure 16.1). In contrast to a strong acid, a weak acid has a conjugate base that is a much stronger base than water. In this case a water molecule is not very successful in pulling an H⁺ ion away from the conjugate base. A *weak acid contains a relatively strong conjugate base* (Figure 16.2).

**Figure 16.2**

The relationship of acid strength and conjugate base strength for the dissociation reaction

information

A weak acid is mostly undissociated in water.



An acetic acid solution conducts only a small amount of current as shown by the dimly lit bulb.

Diprotic acid

An acid that can furnish two protons

Oxyacid

An acid in which the acidic proton is attached to an oxygen atom

Organic acid

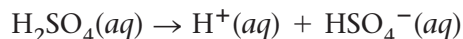
An acid with a carbon-atom backbone and a carboxyl group

The various ways of describing the strength of an acid are summarized in **Table 16.1**.

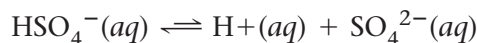
Strong Acids The common strong acids are:

- sulfuric acid, $\text{H}_2\text{SO}_4(aq)$
- hydrochloric acid, $\text{HCl}(aq)$
- nitric acid, $\text{HNO}_3(aq)$
- perchloric acid, $\text{HClO}_4(aq)$

Sulfuric acid is actually a **diprotic acid**, an acid that can furnish two protons. The acid H_2SO_4 is a strong acid that is virtually 100% dissociated in water:

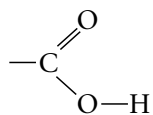


The HSO_4^- ion is also an acid but it is a weak acid:

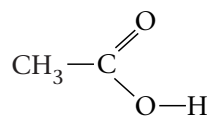


Most of the HSO_4^- ions remain undissociated.

Oxyacids Most acids are **oxyacids**, in which the acidic hydrogen is attached to an oxygen atom. The strong acids we have mentioned, except hydrochloric acid, are typical examples. **Organic acids**, those with a carbon-atom backbone, commonly contain the **carboxyl group**:



Acids of this type are usually weak. An example is acetic acid, CH_3COOH , which is often written as $\text{HC}_2\text{H}_3\text{O}_2$.



Other Acids There are some important acids in which the acidic proton is attached to an atom other than oxygen. The most significant of these are the hydrohalic acids HX , where X represents a halogen atom. Examples are $\text{HCl}(aq)$, a strong acid, and $\text{HF}(aq)$, a weak acid.

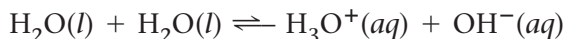
Table 16.1

Ways to Describe Acid Strength

Property	Strong Acid	Weak Acid
the acid ionization (dissociation) reaction	forward reaction predominates	reverse reaction predominates
strength of the conjugate base compared with that of water	A^- is a much weaker base than H_2O	A^- is a much stronger base than H_2O

C. Water as an Acid and a Base

A substance is said to be *amphoteric* if it can behave either as an acid or as a base. Water is the most common **amphoteric substance**. We can see this clearly in the **ionization of water**, which involves the transfer of a proton from one water molecule to another to produce a hydroxide ion and a hydronium ion.



In this reaction one water molecule acts as an acid by furnishing a proton, and the other acts as a base by accepting the proton. The forward reaction for this process does not occur to a very great extent. That is, in pure water only a tiny amount of H_3O^+ and OH^- exist. At 25 °C the actual concentrations are

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$$

Notice that in pure water the concentrations of $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ are equal because they are produced in equal numbers in the ionization reaction.

One of the most interesting and important things about water is that the mathematical *product* of the H_3O^+ and OH^- concentrations is always constant. We can find this constant by multiplying the concentrations of H_3O^+ and OH^- at 25 °C:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

We call this constant K_w . Thus at 25 °C:

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} = K_w$$

To simplify the notation we often write H_3O^+ as just H^+ . Thus we would write the K_w expression as follows:

$$[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} = K_w \text{ at } 25 \text{ }^\circ\text{C}$$

K_w is called the **ion-product constant** for water. The units are customarily omitted when the value of the constant is given and used.

It is important to recognize the meaning of K_w . In any aqueous solution at 25 °C, *no matter what it contains*, the product of $[\text{H}^+]$ and $[\text{OH}^-]$ must always equal 1.0×10^{-14} . This means that if the $[\text{H}^+]$ goes up, the $[\text{OH}^-]$ must go down so that the product of the two is still 1.0×10^{-14} . For example, if HCl gas is dissolved in water, increasing the $[\text{H}^+]$, the $[\text{OH}^-]$ must decrease.

There are three possible situations we might encounter in an aqueous solution.

- If we add an acid to water (an H^+ donor), we get an *acidic solution*. In this case, because we have added a source of H^+ , the $[\text{H}^+]$ will be greater than the $[\text{OH}^-]$.
- On the other hand, if we add a base (a source of OH^-) to water, the $[\text{OH}^-]$ will be greater than the $[\text{H}^+]$. This is a *basic solution*.
- Finally, we might have a situation in which $[\text{H}^+] = [\text{OH}^-]$. This is called a *neutral solution*. Pure water is automatically neutral but we can also obtain a neutral solution by adding equal amounts of H^+ and OH^- .

Amphoteric substance

A substance that can behave either as an acid or as a base

Information

Square brackets ($[]$) are used to indicate the concentration of a substance in moles per liter (M).

MATH

$$\begin{aligned} K_w &= [\text{H}^+][\text{OH}^-] \\ &= 1.0 \times 10^{-14} \end{aligned}$$

CRITICAL THINKING

What if the K for the auto-ionization of water were 1×10^{14} instead of 1×10^{-14} ?

Information

Remember that H^+ represents H_3O^+ .

Let's Review

1. An *acidic solution*, where $[H^+] > [OH^-]$
 2. A *basic solution*, where $[OH^-] > [H^+]$
 3. A *neutral solution*, where $[H^+] = [OH^-]$
- In each case, however, $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$.

EXAMPLE 16.3

Calculating Ion Concentrations in Water

Calculate $[H^+]$ or $[OH^-]$ as required for each of the following solutions at $25^\circ C$, and state whether the solution is neutral, acidic, or basic.

- a. $1.0 \times 10^{-5} M OH^-$
- b. $1.0 \times 10^{-7} M OH^-$
- c. $10.0 M H^+$

Solution

- a. $1.0 \times 10^{-5} M OH^-$

Where do we want to go?

$$[H^+] = ? M$$

What do we know?

- $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$
- $[OH^-] = 1.0 \times 10^{-5} M$

How do we get there?

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$

MATH



The more negative the exponent, the smaller the number.

Dividing both sides of the equation by $[OH^-]$ we get

$$[H^+] = \frac{1.0 \times 10^{-14}}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-5}} = 1.0 \times 10^{-9} M$$

Since $[OH^-] = 1.0 \times 10^{-5} M$ is greater than $[H^+] = 1.0 \times 10^{-9} M$ the solution is basic.

- b. $1.0 \times 10^{-7} M OH^-$

Where do we want to go?

$$[H^+] = ? M$$

What do we know?

- $K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$
- $[OH^-] = 1.0 \times 10^{-7} M$

How do we get there?

$$K_w = [\text{H}^+] [\text{OH}^-] = 1.0 \times 10^{-14}$$

Dividing both sides of the equation by $[\text{OH}^-]$ we get

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-7}} = 1.0 \times 10^{-7} \text{ M}$$

Since $[\text{OH}^-] = [\text{H}^+] = 1.0 \times 10^{-7} \text{ M}$ the solution is neutral

c. 10.0 M H^+

Where do we want to go?

$$[\text{OH}^-] = ? \text{ M}$$

What do we know?

- $K_w = [\text{H}^+] [\text{OH}^-] = 1.0 \times 10^{-14}$
- $[\text{H}^+] = 10.0 \text{ M}$

How do we get there?

$$K_w = [\text{H}^+] [\text{OH}^-] = 1.0 \times 10^{-14}$$

Dividing both sides of the equation by $[\text{H}^+]$ we get

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}^+]} = \frac{1.0 \times 10^{-14}}{10.0} = 1.0 \times 10^{-15} \text{ M}$$

Since $[\text{H}^+] = 10.0 \text{ M}$ is greater than $[\text{OH}^-] = 1.0 \times 10^{-15} \text{ M}$ the solution is acidic.

Practice Problem • Exercise 16.3

Calculate $[\text{H}^+]$ in a solution in which $[\text{OH}^-] = 2.0 \times 10^{-2} \text{ M}$.
Is this solution acidic, neutral, or basic?

EXAMPLE 16.4

Using the Ion-Product Constant in Calculations

Is it possible for an aqueous solution at 25 °C to have $[H^+] = 0.010 M$ and $[OH^-] = 0.010 M$?

Solution

The concentration 0.010 M can also be expressed as $1.0 \times 10^{-2} M$. Thus, if $[H^+] = [OH^-] = 1.0 \times 10^{-2} M$, the product $[H^+][OH^-] = (1.0 \times 10^{-2})(1.0 \times 10^{-2}) = 1.0 \times 10^{-4}$.

This is not possible. The product of $[H^+]$ and $[OH^-]$ must always be 1.0×10^{-14} in water at 25 °C, so a solution could not have $[H^+] = [OH^-] = 0.010 M$.

If H^+ and OH^- are added to water in these amounts, they will react with each other to form H_2O ,



until the product $[H^+][OH^-] = 1.0 \times 10^{-14}$.

This is a general result. When H^+ and OH^- are added to water in amounts such that the product of their concentrations is greater than 1.0×10^{-14} , they will react to form water until enough H^+ and OH^- are consumed (until $[H^+][OH^-] = 1.0 \times 10^{-14}$).



SECTION 16.1

REVIEW QUESTIONS

- How is the Arrhenius concept of an acid different from the Brønsted–Lowry model of an acid?
- Which of the following represent conjugate acid–base pairs? For those pairs that are not conjugates, write the correct conjugate acid or base for each species in the pair.
 - HSO_4^- , SO_4^{2-}
 - HBr, BrO^-
 - $H_2PO_4^-$, PO_4^{3-}
 - HNO_3 , NO_2^-
- Draw a molecular-level view of an aqueous solution of the strong acid HCl.
- Draw a molecular-level view of an aqueous solution of the weak acid HF.
- Use the Brønsted–Lowry model to label the acid–base pairs in the following equation for the ionization of water:
$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$
- What does it mean to say a solution is acidic? Basic? Neutral?



Determining the Acidity of a Solution

Objectives

- To understand pH and pOH
- To learn to find pH and pOH for various solutions
- To use a calculator to find pH
- To learn methods for measuring pH of a solution
- To learn to calculate the pH of strong acids

A. The pH Scale

To express small numbers conveniently, chemists often use the “p scale,” which is based on common logarithms (base 10 logs). In this system, if N represents some number, then

$$pN = -\log N = (-1) \times \log N$$

That is, the p means to take the log of the number that follows and multiply the result by -1 . For example, to express the number 1.0×10^{-7} on the p scale.

1. Enter (-).
2. Press the \log key.
3. Enter 1.0×10^{-7} .

$$p(1.0 \times 10^{-7}) = -\log(1.0 \times 10^{-7}) = 7.00$$

Because the $[H^+]$ in an aqueous solution is typically quite small, using the p scale in the form of the **pH scale** provides a convenient way to represent solution acidity. The pH is defined as

$$pH = -\log[H^+]$$

To obtain the pH value of a solution, we must compute the negative log of the $[H^+]$. On a typical calculator, this involves the following steps:

Steps for Calculating pH on a Calculator

Step 1 Enter (-).

Step 2 Press the \log key.

Step 3 Enter the $[H^+]$.

In the case where $[H^+] = 1.0 \times 10^{-5} M$, following the above steps gives a pH value of 5.00.

To represent pH to the appropriate number of significant figures, you need to know the following rule for logarithms:

The number of decimal places for a log must be equal to the number of significant figures in the original number.

$$[H^+] = \overset{\substack{2 \text{ significant} \\ \text{figures}}}{\boxed{1.0}} \times 10^{-5} M \quad \text{and} \quad pH = \underset{\substack{2 \text{ decimal} \\ \text{places}}}{\boxed{5.00}}$$

Key Terms

- pH scale
- Indicators
- Indicator paper
- pH meter

Information

The pH scale provides a compact way to represent solution acidity.

CRITICAL THINKING ?

What if an elected official decided to ban all products with a pH outside of the 6–8 range? Would the ban affect products you buy? Give some examples of products that would no longer be available.

EXAMPLE 16.5

Calculating pH

Calculate the pH value for each of the following solutions at 25 °C.

- A solution in which $[\text{H}^+] = 1.0 \times 10^{-9} \text{ M}$
- A solution in which $[\text{OH}^-] = 1.0 \times 10^{-6} \text{ M}$

Solution (a)

For this solution $[\text{H}^+] = 1.0 \times 10^{-9}$.

$$-\log 1.0 \times 10^{-9} = 9.00$$

$$\text{pH} = 9.00$$

Solution (b)

In this case we are given the $[\text{OH}^-]$. Thus we must first calculate $[\text{H}^+]$ from the K_w expression. We solve

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

for $[\text{H}^+]$ by dividing both sides by $[\text{OH}^-]$.

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-6}} = 1.0 \times 10^{-8}$$

Now that we know the $[\text{H}^+]$, we can calculate the pH by following the example above. Doing so yields $\text{pH} = 8.00$.

Practice Problem • Exercise 16.5

Calculate the pH value for each of the following solutions at 25 °C.

- A solution in which $[\text{H}^+] = 1.0 \times 10^{-3} \text{ M}$
- A solution in which $[\text{OH}^-] = 5.0 \times 10^{-5} \text{ M}$

Active Reading Question

Why do we use the “p scale” in acid–base chemistry?

Because the pH scale is a log scale based on 10, the pH changes by 1 for every power-of-10 change in the $[\text{H}^+]$.

information

The pH decreases as $[\text{H}^+]$ increases, and vice versa.

For example, a solution of pH 3 has an H^+ concentration of 10^{-3} M , which is 10 times that of a solution of pH 4 ($[\text{H}^+] = 10^{-4} \text{ M}$) and 100 times that of a solution of pH 5. This is illustrated in **Table 16.2**. Also note from Table 16.2 that *the pH decreases as the $[\text{H}^+]$ increases*. That is, a lower pH means a more acidic solution. The pH scale and the pH values for several common substances are shown in **Figure 16.3**.

Active Reading Question

Solution A has a pH of 4.00. Solution B has a pH of 5.00. Which solution, A or B, is more acidic? How many times more acidic is it?

Table 16.2

The Relationship of the H^+ Concentration of a Solution to Its pH

$[H^+]$	pH
1.0×10^{-1}	1.00
1.0×10^{-2}	2.00
1.0×10^{-3}	3.00
1.0×10^{-4}	4.00
1.0×10^{-5}	5.00
1.0×10^{-6}	6.00
1.0×10^{-7}	7.00

Log scales similar to the pH scale are used for representing other quantities. For example,

$$pOH = -\log[OH^-]$$

Therefore, in a solution in which

$$[OH^-] = 1.0 \times 10^{-12} M$$

the pOH is

$$-\log[OH^-] = -\log(1.0 \times 10^{-12}) = 12.00$$

EXAMPLE 16.6

Calculating pH and pOH

Calculate the pH and pOH for each of the following solutions at 25 °C.

- a. $1.0 \times 10^{-3} M OH^-$ b. $1.0 M H^+$

Solution

- a. We are given the $[OH^-]$, so we can calculate the pOH value by taking $-\log[OH^-]$.

$$pOH = -\log[OH^-] = -\log(1.0 \times 10^{-3}) = 3.00$$

To calculate the pH, we must first solve the K_w expression for $[H^+]$.

$$[H^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-3}} = 1.0 \times 10^{-11} M$$

Now we compute the pH.

$$pH = -\log[H^+] = -\log(1.0 \times 10^{-11}) = 11.00$$

- b. In this case we are given the $[H^+]$ and we can compute the pH.

$$pH = -\log[H^+] = -\log(1.0) = 0$$

We next solve the K_w expression for $[OH^-]$

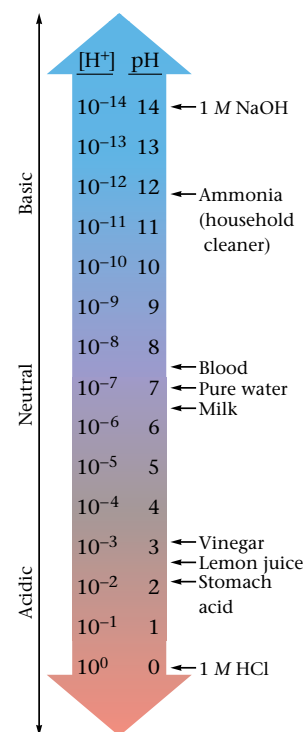
$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{1.0} = 1.0 \times 10^{-14} M$$

Now we compute the pOH.

$$pOH = -\log[OH^-] = -\log(1.0 \times 10^{-14}) = 14.00$$

Figure 16.3

The pH scale and pH values of some common substances



We can obtain a convenient relationship between pH and pOH by starting with the K_w expression

$$[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$$

and taking the negative log of both sides.

$$-\log([\text{H}^+][\text{OH}^-]) = -\log(1.0 \times 10^{-14})$$

Because the log of a product equals the sum of the logs of the terms, that is, $\log(A \times B) = \log A + \log B$, we have

$$\underbrace{-\log[\text{H}^+]}_{\text{pH}} + \underbrace{(-\log[\text{OH}^-])}_{\text{pOH}} = -\log(1.0 \times 10^{-14}) = 14.00$$

i nformation

The symbol p means $-\log$.

which gives the equation

$$\text{pH} + \text{pOH} = 14.00$$

This means that once we know either the pH or the pOH for a solution, we can calculate the other. For example, if a solution has a pH of 6.00, the pOH is calculated as follows:

$$\begin{aligned}\text{pH} + \text{pOH} &= 14.00 \\ \text{pOH} &= 14.00 - \text{pH} \\ \text{pOH} &= 14.00 - 6.00 = 8.00\end{aligned}$$

EXAMPLE 16.7

Calculating pOH from pH



Red blood cells can exist only over a narrow range of pH.

The pH of blood is about 7.4. What is the pOH of blood?

Solution

$$\begin{aligned}\text{pH} + \text{pOH} &= 14.00 \\ \text{pOH} &= 14.00 - \text{pH} \\ &= 14.00 - 7.4 \\ &= 6.6\end{aligned}$$

The pOH of blood is 6.6.

Practice Problem • Exercise 16.7

A sample of rain in an area with severe air pollution has a pH of 3.5. What is the pOH of this rainwater?

It is also possible to find the $[\text{H}^+]$ or $[\text{OH}^-]$ from the pH or pOH. To find the $[\text{H}^+]$ from the pH, we must go back to the definition of pH:

$$\text{pH} = -\log[\text{H}^+]$$

or

$$-\text{pH} = \log[\text{H}^+]$$

To arrive at $[\text{H}^+]$ on the right-hand side of this equation we must “undo” the log operation. This is called taking the *antilog* or the *inverse log*.

$$\text{Inverse log } (-\text{pH}) = [\text{H}^+]$$

There are different methods for carrying out the inverse log operation on various calculators. One common method is the two-key 2nd log sequence. (Consult your calculator's manual to find out how to do the antilog or inverse log operation.) The steps in going from pH to $[\text{H}^+]$ are as follows:

Steps for Calculating $[\text{H}^+]$ from pH

Step 1 Take the inverse log (antilog) by using the 2nd log keys in that order. (Your calculator may require different keys for this operation.)

Step 2 Enter (-).

Step 3 Enter the pH.

For practice, we will convert $\text{pH} = 7.0$ to $[\text{H}^+]$.

$$\text{pH} = 7.0$$

$$-\text{pH} = -7.0$$

The inverse log of -7.0 gives 1×10^{-7}

$$[\text{H}^+] = 1 \times 10^{-7} \text{ M}$$

This process is illustrated further in Example 16.8.

EXAMPLE 16.8

Calculating $[\text{H}^+]$ from pH

The pH of a human blood sample was measured to be 7.41. What is the $[\text{H}^+]$ in this blood?

Solution

$$\text{pH} = 7.41$$

$$-\text{pH} = -7.41$$

$$[\text{H}^+] = \text{inverse log } (-7.41) = 3.9 \times 10^{-8}$$

so,

$$[\text{H}^+] = 3.9 \times 10^{-8} \text{ M.}$$

Notice that because the pH has two decimal places, we need two significant figures for $[\text{H}^+]$.

Practice Problem • Exercise 16.8

The pH of rainwater in a polluted area was found to be 3.50. What is the $[\text{H}^+]$ for this rainwater?

MATH



This operation may involve a 10^x key on some calculators.



Measuring the pH of the water in a river

A similar procedure is used to change from pOH to $[\text{OH}^-]$, as shown in Example 16.9.

EXAMPLE 16.9

Calculating $[\text{OH}^-]$ from pOH

The pOH of the water in a fish tank is found to be 6.59. What is the $[\text{OH}^-]$ for this water?

Solution

We use the same steps as for converting pH to $[\text{H}^+]$, except that we use the pOH to find $[\text{OH}^-]$.

$$\text{pOH} = 6.59$$

$$-\text{pOH} = -6.59$$

$$\text{The inverse log } (-6.59) = 2.6 \times 10^{-7}$$

$$[\text{OH}^-] = 2.6 \times 10^{-7} \text{ M}$$

Note that two significant figures are required.

Practice Problem • Exercise 16.9

The pOH of a liquid drain cleaner was found to be 10.50. What is the $[\text{OH}^-]$ for this cleaner?

HANDS-ON CHEMISTRY

• MINI-LAB •

Cabbage Juice Indicator

Materials

- samples to test
- cabbage juice indicator
- water
- reference solutions
- droppers

Procedure

1. Bring some samples from home to test for pH levels. If they are liquids, the samples should be clear, white, or light-colored. If they are solids, they should be at least a little soluble in water.
2. Obtain cabbage juice indicator from your teacher. The pH indicator is made by boiling red cabbage leaves in water.

3. Take a small amount of your sample (about 5 mL of the liquid or a small amount of solid dissolved in 5 mL of water) and add a few of drops of cabbage juice indicator to the sample.
4. Use the reference solutions provided by your teacher to determine the approximate pH of each of your samples.
5. Compare your results with those of your classmates.

Results/Analysis

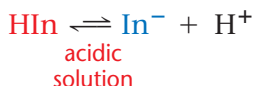
1. Do you see any patterns in pH values? For example, are most cleaning supplies acidic, basic, or neutral? What about most foods?
2. Do any of the pH values surprise you? Why?

B. Measuring pH

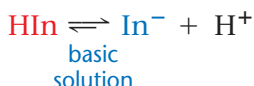
The traditional way of determining the pH of a solution is by using **indicators**—substances that exhibit different colors in acidic and basic solutions. We can represent a generic indicator as a weak acid HIn. Let's assume that for a particular indicator HIn is red and the anion In^- is blue.



Thus, in a solution containing lots of H^+ , the indicator will be red (the HIn form).



In a basic solution, the H^+ will be removed from HIn to form the blue In^- ion and the solution will be blue.



In intermediate pH regions, significant amounts of both HIn and In^- will be present and the solution will be some shade of purple (red plus blue).

Many substances exist that turn different colors when H^+ is present (HIn) than when H^+ is absent (In^-). Some of these indicators are shown in **Figure 16.4** along with the colors of their HIn and In^- forms. Note that different indicators change at different pH values depending on the acid strength of HIn for a particular indicator.

A convenient way to measure the approximate pH of a solution is by using **indicator paper**—a strip of paper coated with a combination of indicators. Indicator paper turns a specific color for each pH value (see **Figure 16.5**).

The pH value of a solution can be measured electronically using a **pH meter** (see **Figure 16.6**). A pH meter contains a probe that is very sensitive to the $[\text{H}^+]$ in a solution. When the probe is inserted into a solution, the $[\text{H}^+]$ in the solution produces a voltage that appears as a pH reading on the meter. Because a pH meter can determine the pH value of a solution so quickly and accurately, these devices are now used almost universally.

C. Calculating the pH of Strong Acid Solutions

We will now learn to calculate the pH for a solution containing a strong acid of known concentration. For example, if we know a solution contains 1.0 M HCl, how can we find the pH of the solution? To answer this question we must know that when HCl dissolves in water, each molecule dissociates (ionizes) into H^+ and Cl^- ions. That is, we must know that HCl is a strong acid. Thus, although the label on the bottle says 1.0 M HCl, the solution contains virtually no HCl molecules. A 1.0 M HCl solution contains H^+ and Cl^- ions rather than HCl molecules. Typically, container labels indicate the substance(s) used to make up the solution but do not necessarily describe the solution components after dissolution. In this case,



Therefore, the $[\text{H}^+]$ in the solution is 1.0 M. The pH is then

$$\text{pH} = -\log[\text{H}^+] = -\log(1.0) = 0$$

Indicator (acid-base)

A chemical that changes color depending on the pH of a solution

Indicator paper

A strip of paper coated with a combination of acid-base indicators

pH meter

A device used to measure the pH of a solution



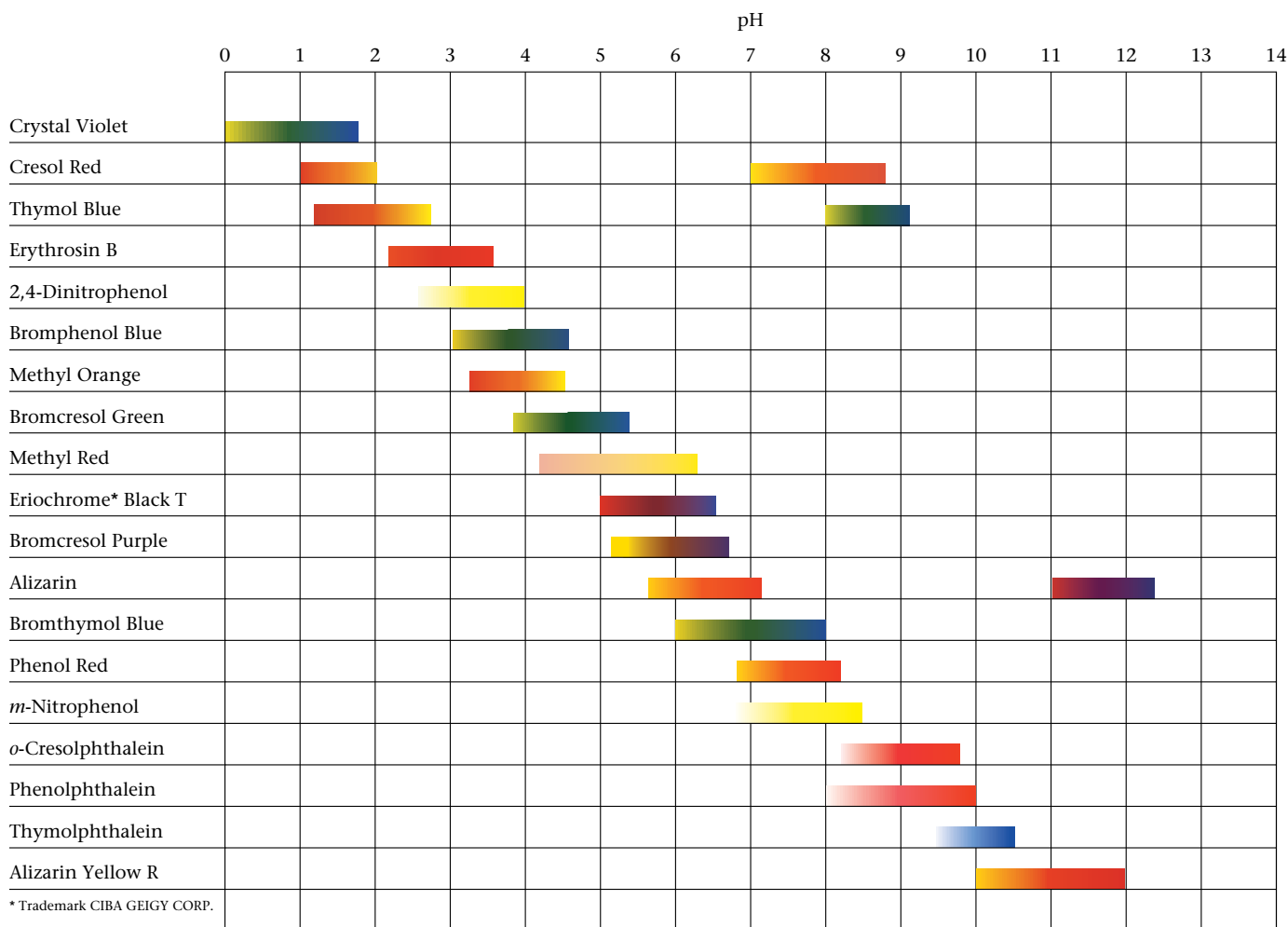
Figure 16.5

Indicator paper being used to measure the pH of a solution. The pH is determined by comparing the color that the solution turns the paper to the color chart.



Figure 16.6

A pH meter



The pH ranges shown are approximate. Specific transition ranges depend on the indicator solvent chosen.

Figure 16.4

The useful pH ranges for several common indicators. Note that most indicators have a useful range of about two pH units.

EXAMPLE 16.10

Calculating the pH of Strong Acid Solutions

Calculate the pH of 0.10 M HNO_3 .

Solution

HNO_3 is a strong acid, so the ions in solution are H^+ and NO_3^- . In this case,



Thus

$$[\text{H}^+] = 0.10 \text{ M} \text{ and } \text{pH} = -\log(0.10) = 1.00$$

Practice Problem • Exercise 16.10

Calculate the pH of a solution of 5.0×10^{-3} M HCl.

CHEMISTRY EXPLORERS

Arnold Beckman • 1900–2004

Arnold Beckman was born in 1900 in Cullom, Illinois, a town of 500 people that had no electricity or telephones. Beckman says, “In Cullom we were forced to improvise. I think it was a good thing.”

The son of a blacksmith, Beckman became so fascinated with chemistry that his father built him a small “chemistry shed” in the backyard for his tenth birthday.

Beckman’s interest in chemistry was fostered by his high school teachers, and he eventually attended the University of Illinois, Urbana-Champaign where he earned a bachelor’s degree in chemical engineering and a master’s degree in physical chemistry. He then went to the California Institute of Technology where he earned a Ph.D. and became a faculty member.

In 1935 Beckman invented something that would cause a revolution in scientific instrumentation. A college friend who worked in a laboratory in the California citrus industry needed an accurate, convenient way to measure the acidity of lemon juice. In response, Beckman invented the pH meter. This compact, sturdy device was an immediate hit. In fact, business

was so good that Beckman left Caltech to head his own company.

Over the years Beckman invented many other devices, including an instrument for measuring the light absorbed by molecules. At age 65, after he retired as president of Beckman Instruments, Beckman began a new career—heading a foundation which shared his wealth for the support of scientific research and education. In 1984 he and Mabel, his wife of 58 years, donated \$40 million to his alma mater—the University of Illinois—to fund the Beckman Institute. The Beckmans’ foundation has also funded many other research institutes, including one at Caltech, and currently gives \$20 million each year to various scientific endeavors.

Arnold Beckman was a man known for his incredible creativity but even more he was recognized as a man of absolute integrity. Dr. Beckman has important words for us: “Whatever you do, be enthusiastic about it.” Dr. Beckman passed away on May 18, 2004, at the age of 104.

Image not available for electronic use. Please refer to the image in the textbook.



SECTION 16.2

REVIEW QUESTIONS

- What is the importance of pH?
- Calculate the pH of an aqueous solution that has a concentration of H^+ equal to
 - $2.30 \times 10^{-4} M$
 - $4.37 \times 10^{-11} M$
- Calculate the concentration of H^+ in an aqueous solution with a pH of
 - 3.40
 - 9.20
- Determine the pOH of a solution that has a concentration of H^+ equal to
 - $1.85 \times 10^{-10} M$
 - $9.23 \times 10^{-5} M$
- Determine the pH of a solution that has a concentration of OH^- equal to
 - $2.15 \times 10^{-3} M$
 - $7.24 \times 10^{-9} M$
- If you placed phenol red in an acid solution what color would the solution turn?
- Calculate the pH of 100.0 mL of a 0.150 M aqueous solution of HNO_3 .

Titrations and Buffers

Key Terms

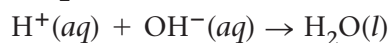
- Neutralization reaction
- Titration
- Standard solution
- Buret
- Stoichiometric point (equivalence point)
- Titration curve (pH curve)
- Buffered solution

Objectives

- To learn about acid–base titrations
- To understand the general characteristics of buffered solutions

A. Acid–Base Titrations

By this time you know that a strong acid solution contains H^+ ions and a solution of strong base contains OH^- ions. As we saw earlier, when a strong acid and a strong base are mixed, the H^+ and OH^- react to form H_2O :



This reaction is called a **neutralization reaction** because if equal amounts of H^+ and OH^- are available for reaction, a neutral solution ($\text{pH} = 7$) will result.

To analyze the acid or base content of a solution, chemists often perform a titration. A **titration** involves the delivery of a measured volume of a solution of known concentration (the *titrant*) into the solution being analyzed (the *analyte*). The titrant contains a substance that reacts in a known way with the analyte. For example, if the analyte contains a base, the titrant would be a **standard solution** (a solution of known concentration) of a strong acid. To run the titration the standard solution of titrant is loaded into a buret. A **buret** is a cylindrical device with a stopcock at the bottom that allows accurate measurement of the delivery of a given volume of liquid (see **Figure 16.7**). The titrant is added slowly to the analyte until exactly enough has been added to just react with all of the analyte. This point is called the **stoichiometric point** or **equivalence point** for the titration. For an acid–base titration the equivalence point can be determined by using a pH meter or indicator. In the titration of a strong acid and a strong base, the equivalence point occurs when an equal amount of H^+ and OH^- have reacted so that the solution is neutral ($\text{pH} = 7$). Thus the titration would stop when the pH meter shows $\text{pH} = 7$. Alternatively, an indicator could be used that changes color near $\text{pH} = 7$.

Neutralization reaction

An acid–base reaction

Titration

A technique in which a solution of known concentration is used to determine the concentration of another solution

Standard solution

A solution in which the concentration is accurately known

Buret

A device used for the accurate measurement of the delivery of a given volume of a liquid or solution

Equivalence point (stoichiometric point)

The point in a titration when enough titrant has been added to react exactly with the substance in solution that is being titrated

Active Reading Question

Why can a strong acid–strong base titration be called a neutralization reaction? What is meant by a neutral solution?

Figure 16.7 shows three steps in the titration of a solution of strong acid (the analyte) with a solution of sodium hydroxide (the titrant). The titration is stopped (equivalence point reached) after the drop of $\text{NaOH}(aq)$ that first changes the indicator permanently from colorless to red.

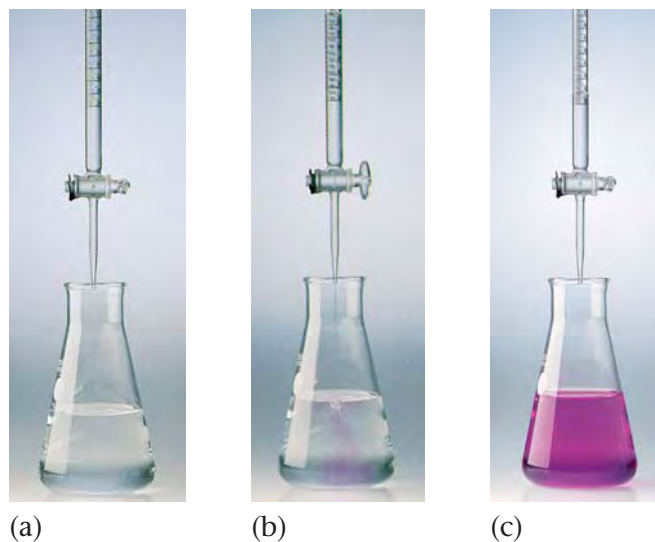
Figure 16.7

The titration of an acid with a base.

(a) The titrant (the base) is in the buret, and the flask contains the acid solution along with a small amount of indicator.

(b) As base is added drop by drop to the acid solution in the flask during the titration, the indicator changes color, but the color disappears on mixing.

(c) The stoichiometric (equivalence) point is marked by a permanent indicator color change. The volume of base added is the difference between the final and initial buret readings.

**EXAMPLE 16.11****The Titration of a Strong Acid with a Strong Base**

Determine the volume of 0.100 M NaOH needed to titrate 50.0 mL of 0.200 M HNO₃.

Solution**Where do we want to go?**

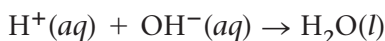
Volume of 0.100 M NaOH = ? mL

What do we know?

- 50.0 mL 0.200 M HNO₃
- 0.100 M NaOH
- $\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l)$
- HNO₃ is a strong acid containing H⁺ and NO₃⁻ ions.
- NaOH is a strong base containing Na⁺ and OH⁻ ions.

How do we get there?

The titration is shown in **Figure 16.8**. To reach the equivalence point we must add enough OH⁻ to just react with the H⁺ ions originally present in the HNO₃(aq) solution. The neutralization reaction is:

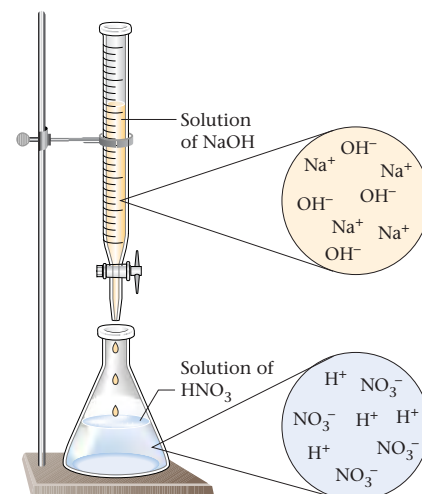


That is, at the equivalence point:

$$\begin{array}{l} \text{moles OH}^- \text{ added} = \text{moles H}^+ \text{ originally present} \\ \text{From } 0.100 \text{ M NaOH} \quad \text{From } 50.0 \text{ mL } 0.200 \text{ M HNO}_3 \end{array}$$

We can determine the moles of H⁺ originally present in the 50.0-mL sample of 0.200 M HNO₃ by multiplying the volume (in L) times the molarity:

$$\begin{aligned} 50.0 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{200 \text{ mol H}^+}{\text{L}} \\ = 0.0500 \cancel{\text{L}} \times 0.200 \frac{\text{mol H}^+}{\cancel{\text{L}}} = 1.00 \times 10^{-2} \text{ mol} \end{aligned}$$

**Figure 16.8**

A molecular-level representation of the solutions in the titration of 0.200 M HNO₃ with 0.100 M NaOH. Note that the solutions contain ions because HNO₃ is a strong acid and NaOH is a strong base.

Thus, to complete the titration (give a neutral solution), we must add 1.00×10^{-2} mol OH^- . What volume of 0.100 M NaOH contains 1.00×10^{-2} mol OH^- ?

$$? \text{ vol} \times 0.100 \frac{\text{mol OH}^-}{\text{L}} = 1.00 \times 10^{-2} \text{ mol OH}^-$$

Rearranging the preceding equation gives

$$? \text{ vol} = \frac{1.00 \times 10^{-2}}{0.100} \text{ L} = 1.00 \times 10^{-1} \text{ L}$$

Thus

$$? \text{ vol} = 0.100 \text{ L} = 100. \text{ mL}$$

Thus it takes 100. mL of 0.100 M NaOH to neutralize 50.0 mL of 0.200 M HNO_3 .

Practice Problem • Exercise 16.11

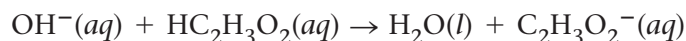
Calculate the volume of 0.3000 M HCl needed to titrate 75.00 mL of 0.1500 M $\text{KOH}(aq)$.

Titration curve (pH curve)

A plot of pH of solution versus volume of titrant added to a given solution

In a modern acid–base titration, typically the probe of a pH meter is inserted in the solution being analyzed and the pH is monitored during the entire titration. A plot of the resulting data (pH versus volume of titrant added) is called the **titration curve** (or **pH curve**). The pH curve for the titration of 0.200 M HNO_3 with 0.100 M NaOH (see Example 16.11) is shown in **Figure 16.9**. Notice from this pH curve that the pH changes very rapidly when the titration nears the equivalence point.

So far we have discussed the titration of a strong acid with a strong base such as $\text{NaOH}(aq)$. Weak acids can also be titrated with $\text{NaOH}(aq)$. For example, acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$), the weak acid found in vinegar, can be titrated according to the following reaction:



Here the formula for the entire acid is written in the neutralization reaction instead of just H^+ , as is the case for strong acids. This is done because acetic acid is a weak acid and retains its H^+ in water. In contrast, OH^- is such a strong base that it can easily extract the H^+ from $\text{HC}_2\text{H}_3\text{O}_2$.

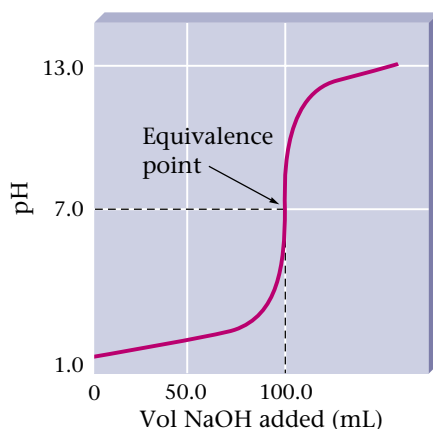


Figure 16.9

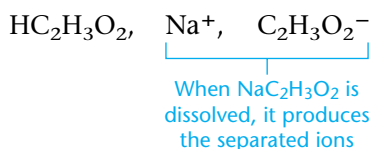
The pH curve for the titration of 50.0 mL of 0.200 M HNO_3 with 0.100 M NaOH. Note that the equivalence point occurs at 100.0 mL of NaOH added, the point where exactly enough OH^- has been added to react with all the H^+ originally present. The pH of 7 at the equivalence point is characteristic of a strong acid–strong base titration.

B. Buffered Solutions

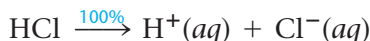
A **buffered solution** is one that resists a change in its pH even when a strong acid or base is added to it. For example, when 0.01 mol of HCl is added to 1 L of pure water, the pH changes from its initial value of 7 to a final value of 2, a change of 5 pH units. However, when 0.010 mol of HCl is added to 1.0 L of a solution containing both 0.10 M acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) and 0.10 M sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$), the pH changes from an initial value of 4.74 to a final value of 4.66, a change of only 0.08 pH units. The latter solution is buffered—it undergoes only a very slight change in pH when a strong acid or base is added to it.

Buffered solutions are vitally important to living organisms whose cells can survive only in a very narrow pH range. Many goldfish have died because their owners did not realize the importance of buffering the aquarium water at an appropriate pH. For humans to survive, the pH of the blood must be maintained between 7.35 and 7.45. This narrow range is maintained by several different buffering systems.

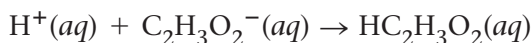
A solution is **buffered** by the *presence of a weak acid and its conjugate base*. An example of a buffered solution is an aqueous solution that contains acetic acid and sodium acetate. The sodium acetate is a salt that furnishes acetate ions (the conjugate base of acetic acid) when it dissolves. To see how this system acts as a buffer, we must recognize that the species present in this solution are



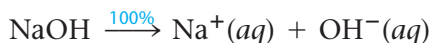
What happens in this solution when a strong acid such as HCl is added? In pure water, the H^+ ions from the HCl accumulate, thus lowering the pH.



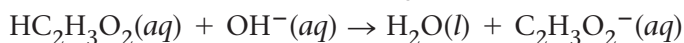
However, this buffered solution contains $\text{C}_2\text{H}_3\text{O}_2^-$ ions, which are basic. That is, $\text{C}_2\text{H}_3\text{O}_2^-$ has a strong affinity for H^+ , as evidenced by the fact that $\text{HC}_2\text{H}_3\text{O}_2$ is a weak acid. This means that the $\text{C}_2\text{H}_3\text{O}_2^-$ and H^+ ions do not exist together in large numbers. Because the $\text{C}_2\text{H}_3\text{O}_2^-$ ion has a high affinity for H^+ , these two combine to form $\text{HC}_2\text{H}_3\text{O}_2$ molecules. Thus the H^+ from the added HCl does not accumulate in solution but reacts with the $\text{C}_2\text{H}_3\text{O}_2^-$ as follows:



Next consider what happens when a strong base such as sodium hydroxide is added to the buffered solution. If this base were added to pure water, the OH^- ions from the solid would accumulate and greatly change (raise) the pH.



However, in the buffered solution the OH^- ion, which has a *very strong* affinity for H^+ , reacts with $\text{HC}_2\text{H}_3\text{O}_2$ molecules as follows:



Information

Water: pH = 7
0.01 M HCl: pH = 2

Buffered solution

A solution that resists a change in pH when either an acid or a base are added



For goldfish to survive, the pH of the water must be carefully controlled.

Active Reading Question

What species must be present in a buffered solution?

This happens because, although $C_2H_3O_2^-$ has a strong affinity for H^+ , OH^- has a much stronger affinity for H^+ and thus can remove H^+ ions from acetic acid molecules.

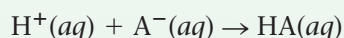
Note that the buffering materials dissolved in the solution prevent added H^+ or OH^- from building up in the solution. Any added H^+ is trapped by $C_2H_3O_2^-$ to form $HC_2H_3O_2$. Any added OH^- reacts with $HC_2H_3O_2$ to form H_2O and $C_2H_3O_2^-$.

The general properties of a buffered solution are summarized in **Table 16.3**.

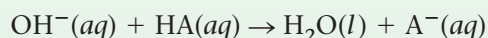
Table 16.3

The Characteristics of a Buffer

1. The solution contains a weak acid HA and its conjugate base A^- .
2. The buffer resists changes in pH by reacting with any added H^+ or OH^- so that these ions do not accumulate.
3. Any added H^+ reacts with the base A^- .



4. Any added OH^- reacts with the weak acid HA.



SECTION 16.3 REVIEW QUESTIONS

1. What is true about the amounts of H^+ and OH^- at the stoichiometric point in a titration of a strong acid with a strong base?
2. At the equivalence point in a titration you know the volume (x mL) of a standard solution of acid (2.5 M) which has been added to a known volume of base (y mL). Show how you can determine the concentration for the base.
3. List two ways to identify the equivalence point in a titration.
4. Determine the volume of 0.250 M NaOH needed to titrate 125.0 mL of 0.150 M HCl.
5. What is the importance of a buffered solution?
6. How does a buffer keep the pH from changing greatly when an acid is added to the solution?



Key Terms

16.1 Acids

Bases

Arrhenius concept of acids and bases

Brønsted–Lowry model

Conjugate acid

Conjugate base

Conjugate acid–conjugate base pair

Hydronium ion

Completely ionized (completely dissociated)

Strong acid

Weak acid

Diprotic acid

Oxyacids

Organic acids

Carboxyl group

Amphoteric substance

Ionization of water

Ion-product constant

Key Ideas

16.1 Properties of Acids and Bases

■ The Arrhenius model

- Acids produce H^+ in aqueous solution.
- Bases produce OH^- in aqueous solution.

■ Brønsted–Lowry model

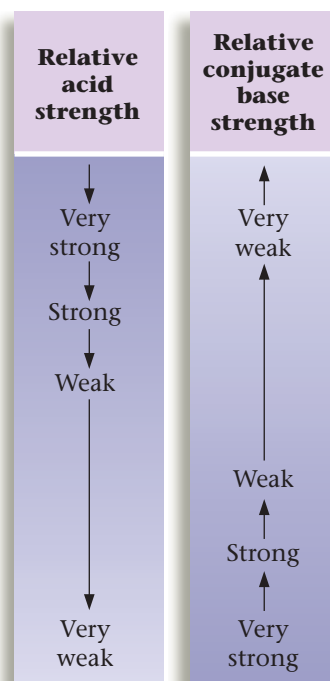
- Acid \rightarrow proton donor
- Base \rightarrow proton acceptor
- $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$

■ Acid strength

- A strong acid is completely dissociated (ionized).
 - ◆ $HA(aq) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq)$
 - ◆ Has a weak conjugate base
- A weak acid is dissociated (ionized only to a very slight extent).
 - ◆ Has a strong conjugate base

■ Water is an acid and a base (amphoteric).

- $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$
- $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$ at 25 °C



16.2 pH scale

Indicators

Indicator paper

pH meter

16.3 Neutralization reaction

Titration

Standard solution

Buret

Stoichiometric point
(equivalence point)

Titration curve (pH curve)

Buffered solution

16.2 Determining the Acidity of a Solution

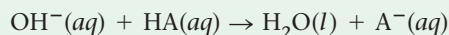
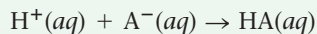
- The acidity of a solution is expressed in terms of $[H^+]$.
- $pH = -\log [H^+]$
 - Lower pH means greater acidity ($[H^+]$).
- $pOH = -\log [OH^-]$
- For a neutral solution $[H^+] = [OH^-]$.
- For an acidic solution $[H^+] > [OH^-]$.
- For a basic solution $[H^+] < [OH^-]$.
- The pH of a solution can be measured by
 - Indicators
 - pH meter
- For a strong acid the concentration of H^+ in the solution equals the initial acid concentration.

16.3 Titrations and Buffers

- An acid–base titration is called a neutralization reaction.
- In a titration a measured volume of a standard solution (titrant) is added to the solution being analyzed (analyte).
- The titration is stopped as close to the stoichiometric point (equivalence point) as possible as marked by an indicator or pH meter.
- A buffered solution is a solution that resists a change in its pH when an acid or base is added.
- Buffered solutions contain a weak acid and its conjugate base.

The Characteristics of a Buffer

1. The solution contains a weak acid HA and its conjugate base A^- .
2. The buffer resists changes in pH by reacting with any added H^+ or OH^- so that these ions do not accumulate.
3. Any added H^+ reacts with the base A^- .
4. Any added OH^- reacts with the weak acid HA.



All exercises with **blue numbers** have answers in the back of this book.

16.1 Properties of Acids and Bases

A. Acids and Bases

- What are some physical properties that historically led chemists to classify various substances as acids and bases?
- In the Arrhenius definition, what characterizes an acid? What characterizes a base? Why are the Arrhenius definitions too restrictive?
- What is an acid in the Brønsted–Lowry model? What is a base?
- How do the components of a conjugate acid–base pair differ from one another? Give an example of a conjugate acid–base pair to illustrate your answer.
- How does a Brønsted–Lowry acid form its conjugate base when dissolved in water? How is the water involved in this process?
- When an acid is dissolved in water, what ion does the water form? What is the relationship of this ion to water itself?
- Which of the following represent conjugate acid–base pairs? For those pairs that are not conjugates, write the correct conjugate acid or base for each species in the pair.
 - H_2SO_4 , SO_4^{2-}
 - H_2PO_4^- , HPO_4^{2-}
 - HClO_4 , Cl^-
 - NH_4^+ , NH_2^-
- In each of the following chemical equations, identify the conjugate acid–base pairs.
 - $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$
 - $\text{PO}_4^{3-} + \text{H}_2\text{O} \rightleftharpoons \text{HPO}_4^{2-} + \text{OH}^-$
 - $\text{C}_2\text{H}_3\text{O}_2^- + \text{H}_2\text{O} \rightleftharpoons \text{HC}_2\text{H}_3\text{O}_2 + \text{OH}^-$
- Write the conjugate *acid* for each of the following:
 - HSO_4^-
 - SO_3^{2-}
 - ClO_4^-
 - H_2PO_4^-
- Write the conjugate *base* for each of the following:
 - HCO_3^-
 - H_2PO_4^-
 - HCl
 - HSO_4^-
- Write a chemical equation showing how each of the following species behaves as a *base* when dissolved in water.
 - NH_3
 - NH_2^-
 - O_2^-
 - F^-

- Write a chemical equation showing how each of the following species behaves as an *acid* when dissolved in water.

- HClO_4
- $\text{HC}_2\text{H}_3\text{O}_2$
- HSO_3^-
- HBr

B. Acid Strength

- What does it mean to say that an acid is strong in aqueous solution? What does this reveal about the ability of the acid's anion to attract protons?
- What does it mean to say that an acid is weak in aqueous solution? What does this reveal about the ability of the acid's anion to attract protons?
- A strong acid has a weak conjugate base, whereas a weak acid has a relatively strong conjugate base. Explain.
- Name four strong acids. For each acid, write the equation showing the acid dissociating in water.
- Which of the following acids have relatively strong conjugate bases?
 - CH_3COOH ($\text{HC}_2\text{H}_3\text{O}_2$)
 - HF
 - H_2S
 - HCl
- Which of the following bases have relatively strong conjugate acids?
 - SO_4^{2-}
 - Br^-
 - CN^-
 - CH_3COO^- ($\text{C}_2\text{H}_3\text{O}_2^-$)
- Water as an Acid and a Base**
 - Anions containing hydrogen (for example, HCO_3^- and $\text{H}_2\text{PO}_4^{2-}$) show amphoteric behaviour when reacting with other acids or bases. Write equations illustrating the amphotericism of these anions.
 - Write an equation showing the auto-ionization of water. Write the expression for the ion-product constant, K_w , for water.
 - What happens to the hydroxide ion concentration in aqueous solutions when we increase the hydrogen ion concentration by adding an acid? What happens to the hydrogen ion concentration in aqueous solutions when we increase the hydroxide ion concentration by adding a base? Explain.

- 22.** Calculate the $[H^+]$ in each of the following solutions, and indicate whether the solution is acidic, basic, or neutral.
- $[OH^-] = 3.99 \times 10^{-5} M$
 - $[OH^-] = 2.91 \times 10^{-9} M$
 - $[OH^-] = 7.23 \times 10^{-2} M$
 - $[OH^-] = 9.11 \times 10^{-7} M$
- 23.** Calculate the $[OH^-]$ in each of the following solutions, and indicate whether the solution is acidic, basic, or neutral.
- $[H^+] = 1.00 \times 10^{-7} M$
 - $[H^+] = 7.00 \times 10^{-7} M$
 - $[H^+] = 7.00 \times 10^{-1} M$
 - $[H^+] = 5.99 \times 10^{-6} M$
- 24.** For each pair of concentrations, tell which represents the more acidic solution.
- $[H^+] = 1.2 \times 10^{-3} M$ or $[H^+] = 4.5 \times 10^{-4} M$
 - $[H^+] = 2.6 \times 10^{-6} M$ or $[H^+] = 4.3 \times 10^{-8} M$
 - $[H^+] = 0.000010 M$ or $[H^+] = 0.0000010 M$
- 25.** For each pair of concentrations, tell which represents the more basic solution.
- $[H^+] = 1.59 \times 10^{-7} M$ or $[H^+] = 1.04 \times 10^{-8} M$
 - $[H^+] = 5.69 \times 10^{-8} M$ or $[OH^-] = 4.49 \times 10^{-6} M$
 - $[H^+] = 5.99 \times 10^{-8} M$ or $[OH^-] = 6.01 \times 10^{-7} M$

16.2 Determining the Acidity of a Solution

A. The pH Scale

- 26.** Why do scientists tend to express the acidity of a solution in terms of its pH, rather than in terms of the molarity of hydrogen ion present? How is pH defined mathematically?
- 27.** Using Figure 16.3, list the approximate pH of five “everyday” solutions. How do the familiar properties (such as sour taste for acids) of these solutions correspond to their pH?
- 28.** For a hydrogen ion concentration of $2.33 \times 10^{-6} M$, how many *decimal places* should we give when expressing the pH of the solution?

- 29.** As the hydrogen ion concentration of a solution *increases*, does the pH of the solution increase or decrease? Explain.
- 30.** Calculate the pH corresponding to each of the hydrogen ion concentrations given below. Tell whether each solution is acidic, basic, or neutral.
- $[H^+] = 0.00100 M$
 - $[H^+] = 2.19 \times 10^{-4} M$
 - $[H^+] = 9.18 \times 10^{-11} M$
 - $[H^+] = 4.71 \times 10^{-7} M$
- 31.** Calculate the pH corresponding to each of the hydroxide ion concentrations given below. Tell whether each solution is acidic, basic, or neutral.
- $[OH^-] = 1.00 \times 10^{-7} M$
 - $[OH^-] = 4.59 \times 10^{-13} M$
 - $[OH^-] = 1.04 \times 10^{-4} M$
 - $[OH^-] = 7.00 \times 10^{-1} M$
- 32.** Calculate the pH corresponding to each of the pOH values listed, and indicate whether each solution is acidic, basic, or neutral.
- pOH = 4.32
 - pOH = 8.90
 - pOH = 1.81
 - pOH = 13.1
- 33.** For each hydrogen or hydroxide ion concentration listed, calculate the concentration of the complementary ion and the pH and pOH of the solution.
- $[H^+] = 1.00 \times 10^{-7} M$
 - $[OH^-] = 4.39 \times 10^{-5} M$
 - $[H^+] = 4.29 \times 10^{-11} M$
 - $[OH^-] = 7.36 \times 10^{-2} M$
- 34.** Calculate the hydrogen ion concentration, in moles per liter, for solutions with each of the following pH values.
- pH = 1.04
 - pH = 13.1
 - pH = 5.99
 - pH = 8.62

35. Calculate the hydrogen ion concentration, in moles per liter, for solutions with each of the following pOH values.

- a. $\text{pOH} = 3.91$
- b. $\text{pOH} = 12.56$
- c. $\text{pOH} = 1.15$
- d. $\text{pOH} = 8.77$

36. Calculate the hydrogen ion and hydroxide ion concentrations, in moles per liter, for solutions with each of the following pH or pOH values.

- a. $\text{pH} = 5.12$
- b. $\text{pOH} = 5.12$
- c. $\text{pH} = 7.00$
- d. $\text{pOH} = 13.00$

B. Measuring pH

37. Name three ways we can measure the pH of a solution. Which is the most accurate? Explain.

38. Describe how an indicator works. Use a chemical equation in your explanation.

C. Calculating the pH of Strong Acid Solutions

39. When 1 mol of gaseous hydrogen chloride is dissolved in enough water to make 1 L of solution, approximately how many HCl molecules remain in the solution? Explain.

40. A bottle of acid solution is labeled “3 M HNO_3 .” What are the substances that are actually present in the solution? Are any HNO_3 molecules present? Why or why not?

41. Calculate the hydrogen ion concentration and the pH of each of the following solutions of strong acids.

- a. $1.04 \times 10^{-4} \text{ M HCl}$
- b. 0.00301 M HNO_3
- c. $5.41 \times 10^{-4} \text{ M HClO}_4$
- d. $6.42 \times 10^{-2} \text{ M HNO}_3$

42. Calculate the hydrogen ion concentration and the pH of each of the following solutions of strong acids.

- a. 0.00010 M HCl
- b. 0.0050 M HNO_3
- c. $4.21 \times 10^{-5} \text{ M HClO}_4$
- d. $6.33 \times 10^{-3} \text{ M HNO}_3$

16.3 Titrations and Buffers

A. Acid-Base Titrations

43. A 100.0-mL sample of 0.50 M $\text{HCl}(aq)$ is titrated with 0.10 M $\text{NaOH}(aq)$. What volume of the NaOH solution is required to reach the endpoint of the titration?

44. If 26.5 mL of a 0.20 M aqueous solution of NaOH is required to titrate 50.0 mL of an aqueous solution of HNO_3 , what is the concentration of the HNO_3 solution?

B. Buffered Solutions

45. What characteristic properties do buffered solutions possess?

46. What two components make up a buffered solution? Give an example of a combination that would serve as a buffered solution.

47. Which component of a buffered solution is capable of combining with an added strong acid? Using your example from exercise 46, show how this component would react with added HCl.

48. Which component of a buffered solution consumes added strong base? Using your example from exercise 46, show how this component would react with added NaOH.

49. Which of the following combinations would act as buffered solutions?

- a. HCl and NaCl
- b. CH_3COOH and CH_3COOK
- c. H_2S and NaHS
- d. H_2S and Na_2S

50. For those combinations in exercise 49 that behave as buffered solutions, write equations showing how the components of the buffer consume added strong acid (HCl) or strong base (NaOH).

Critical Thinking

- 51.** The concepts of acid–base equilibria were developed in this chapter for aqueous solutions (in aqueous solutions, water is the solvent and is intimately involved in the equilibria). However, the Brønsted–Lowry acid–base theory can be extended easily to other solvents. One such solvent that has been investigated in depth is liquid ammonia, NH_3 .
- Write a chemical equation indicating how HCl behaves as an acid in liquid ammonia.
 - Write a chemical equation indicating how OH^- behaves as a base in liquid ammonia.
- 52.** Which of the following conditions indicate an acidic solution?
- $\text{pH} < 7.0$
 - $\text{pOH} < 7.0$
 - $[\text{H}^+] > [\text{OH}^-]$
 - $[\text{H}^+] > 1.0 \times 10^{-7} \text{ M}$
- 53.** Which of the following conditions indicate a basic solution?
- $\text{pOH} < 7.0$
 - $\text{pH} > 7.0$
 - $[\text{OH}^-] < [\text{H}^+]$
 - $[\text{H}^+] < 1.0 \times 10^{-7} \text{ M}$
- 54.** Which of the following acids are classified as strong acids?
- HNO_3
 - CH_3COOH ($\text{HC}_2\text{H}_3\text{O}_2$)
 - HCl
 - HF
 - HClO_4
- 55.** Which of the following represent conjugate acid–base pairs? For those pairs that are not conjugates, write the correct conjugate acid or base for each species in the pair.
- H_2O , OH^-
 - H_2SO_4 , SO_4^{2-}
 - H_3PO_4 , H_2PO_4^-
 - $\text{HC}_2\text{H}_3\text{O}_2$, $\text{C}_2\text{H}_3\text{O}_2^-$
- 56.** In each of the following chemical equations, identify the conjugate acid–base pairs.
- $\text{CH}_3\text{NH}_2 + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{NH}_3^+ + \text{OH}^-$
 - $\text{CH}_3\text{COOH} + \text{NH}_3 \rightleftharpoons \text{CH}_3\text{COO}^- + \text{NH}_4^+$
 - $\text{HF} + \text{NH}_3 \rightleftharpoons \text{F}^- + \text{NH}_4^+$
- 57.** Which of the following bases have relatively strong conjugate acids?
- F^-
 - Cl^-
 - HSO_4^-
 - NO_3^-
- 58.** For each pair of concentrations, tell which represents the more basic solution.
- $[\text{H}^+] = 0.000013 \text{ M}$ or $[\text{OH}^-] = 0.0000032 \text{ M}$
 - $[\text{H}^+] = 1.03 \times 10^{-6} \text{ M}$ or $[\text{OH}^-] = 1.54 \times 10^{-8} \text{ M}$
 - $[\text{OH}^-] = 4.02 \times 10^{-7} \text{ M}$ or $[\text{OH}^-] = 0.0000001 \text{ M}$
- 59.** A poor, misguided chemistry student is working on a laboratory experiment. He has 90.0 mL of a strong base, 0.400 M NaOH . To this he adds 30.0 mL of a 1.75 M solution of HNO_3 . He then adds 240.0 mL of water. Will the resulting solution be neutral? What will the $[\text{H}^+]$ and pH be? What is the pOH? Would the solution turn litmus pink?
- 60.** An eyedropper is calibrated by counting the number of drops required to deliver exactly 1.00 mL. Twenty drops are required. What is the volume of one drop? Suppose one such drop of 0.200 M HCl is added to 1.00 mL of water. What is the $[\text{H}^+]$? By what factor did the $[\text{H}^+]$ change when the one drop was added?
- 61.** Determine the pH of 100.0 mL of a 1.00 M HCl solution. Suppose you double the volume of the solution by adding water. What is the new pH? You double the volume once more with water. What is the new pH? If you continue to add more water, what is the maximum pH that can be reached? Explain your answer.
- 62.** You dissolve 0.500 g of a solid acid (HA) in 100. mL of water. You titrate this solution with 0.150 M NaOH . You find that 29.7 mL of the NaOH solution is required to reach the endpoint of the titration. What is the molar mass of the acid? The equation for the reaction between the acid and base is
- $$\text{HA}(aq) + \text{NaOH}(aq) \rightarrow \text{NaA}(aq) + \text{H}_2\text{O}(l)$$

1 Which of the following is not a conjugate acid-base pair?

- A H_2O , H_3O^+
- B H_2CO_3 , CO_3^{2-}
- C HCl , Cl^-
- D $\text{HC}_2\text{H}_3\text{O}_2$, $\text{C}_2\text{H}_3\text{O}_2^-$

2 Which of the following statements about a strong acid is incorrect?

- A A strong acid is considered to be completely ionized in water.
- B In a reaction between a strong acid and water, the forward reaction predominates.
- C A strong acid has a relatively strong conjugate base.
- D A strong acid is also a strong electrolyte.

3 Which of the following is true about a solution in which $[\text{OH}^-] = 4.8 \times 10^{-4} \text{ M}$?

- A $[\text{H}^+] = 2.1 \times 10^{-11} \text{ M}$ and the solution is basic
- B $[\text{H}^+] = 2.1 \times 10^{-11} \text{ M}$ and the solution is acidic
- C $[\text{H}^+] = 4.8 \times 10^{-4} \text{ M}$ and the solution is acidic
- D $[\text{H}^+] = 4.8 \times 10^{-4} \text{ M}$ and the solution is basic

4 Calculate the pH of a solution in which $[\text{OH}^-] = 2.3 \times 10^{-3} \text{ M}$.

- A 2.64
- B 3.58
- C 10.42
- D 11.36

5 Solution A has a pH of 2.00. Solution B has a pH of 4.00. Which of the following statements is true?

- A Solution B is two times as acidic as solution A.
- B Solution A is one hundred times as acidic as solution B.
- C Solution A is two times as acidic as solution B.
- D Solution B is one hundred times as acidic as solution A.

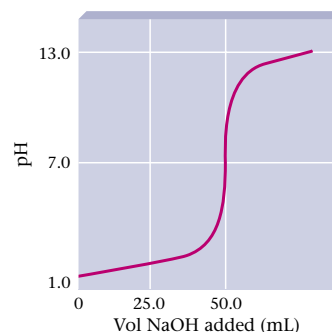
6 Calculate the pH of a 0.045 M solution of nitric acid.

- A 0.045
- B 1.00
- C 1.35
- D 1.65

7 Which of the following is another term for a solution of known concentration?

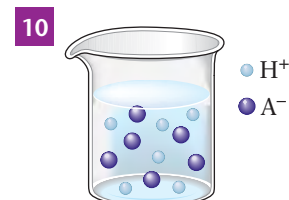
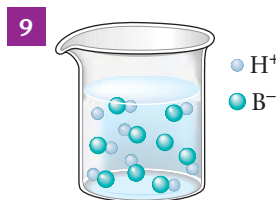
- A A standard solution
- B A buffered solution
- C An analyte
- D An indicator solution

8 According to the pH curve, what volume of base is required to reach the endpoint in the titration?



- A 0 mL
- B 25.0 mL
- C 50.0 mL
- D 100.0 mL

The following figures are molecular-level representations of acid solutions. Label each as a strong acid or a weak acid.



11 Answer the following questions concerning buffered solutions.

- a. Explain what a buffered solution does.
- b. Describe the substances that make up a buffered solution.
- b. Explain how a buffered solution works.

LOOKING AHEAD

17.1 Reaction Rates and Equilibrium

- A. How Chemical Reactions Occur
- B. Conditions That Affect Reaction Rates
- C. Heterogeneous Reactions
- D. The Equilibrium Condition
- E. Chemical Equilibrium: A Dynamic Condition

17.2 Characteristics of Equilibrium

- A. The Equilibrium Constant: An Introduction
- B. Heterogeneous Equilibria

17.3 Application of Equilibria

- A. Le Châtelier's Principle
 - The Effect of a Change in Concentration
 - The Effect of a Change in Volume
 - The Effect of a Change in Temperature
- B. Applications Involving the Equilibrium Constant
- C. Solubility Equilibria

Equilibrium can be analogous to traffic flowing both ways on the George Washington bridge between New Jersey and New York.



IN YOUR LIFE

Chemistry is mostly about reactions—processes in which groups of atoms are reorganized. So far we have learned to describe chemical reactions by using balanced equations and to calculate amounts of reactants and products. However, there are many important characteristics of reactions that we have not yet considered.

For example, why does refrigeration prevent food from spoiling? That is, why do the chemical reactions that cause food to decompose occur more slowly at lower temperatures? On the other hand, how can a chemical manufacturer speed up a chemical reaction that runs too slowly to be economical?

It turns out that chemical reactions are also reversible. That is, the reaction can run from “right to left” as written. This has real-life consequences. For example, if we exert ourselves (by running or bike riding, for example) while we are at high elevations, we feel lightheaded. It is also true that athletes who train at high elevations are better able to compete at low elevations, but this effect does not last for long. Why is this?



Refrigeration prevents food spoilage.

WHAT DO YOU KNOW?

Prereading Questions

1. What could you do to speed up a chemical reaction?
2. What do you think of when you hear the word *equilibrium*?
3. When an ionic solid dissolves, what is present in the solution?
4. What does it mean for a solution to be saturated?

SECTION 17.1

Reaction Rates and Equilibrium

Key Terms

- Collision model
- Activation energy (E_a)
- Catalyst
- Enzymes
- Homogeneous reactions
- Heterogeneous reactions
- Equilibrium
- Chemical equilibrium

Objectives

- To understand the collision model of chemical reactions
- To understand activation energy
- To understand how a catalyst speeds up a chemical reaction
- To explore reactions with reactants or products in different phases
- To learn how equilibrium is established
- To learn about the characteristics of chemical equilibrium

A. How Chemical Reactions Occur

In writing the equation for a chemical reaction, we put the reactants on the left and the products on the right with an arrow between them. But how do the atoms in the reactants reorganize to form the products?

Chemists believe that molecules react by colliding with each other. Some collisions are violent enough to break bonds, allowing the reactants to rearrange to form the products. For example, consider the reaction



which we think occurs as shown in **Figure 17.1**. Notice that the Br—N bonds in the two BrNO molecules must be broken and a new Br—Br bond must be formed during a collision for the reactants to become products.

The idea that reactions occur during molecular collisions, which is called the **collision model**, explains many characteristics of chemical reactions. For example, it explains why a reaction proceeds faster if the concentrations of the reacting molecules are increased (higher concentrations lead to more collisions and therefore to more reaction events). The collision model also explains why reactions go faster at higher temperatures.

Collision model

Molecules must collide in order to react; used to account for the fact that reaction rate depends on concentrations of reactants and temperature

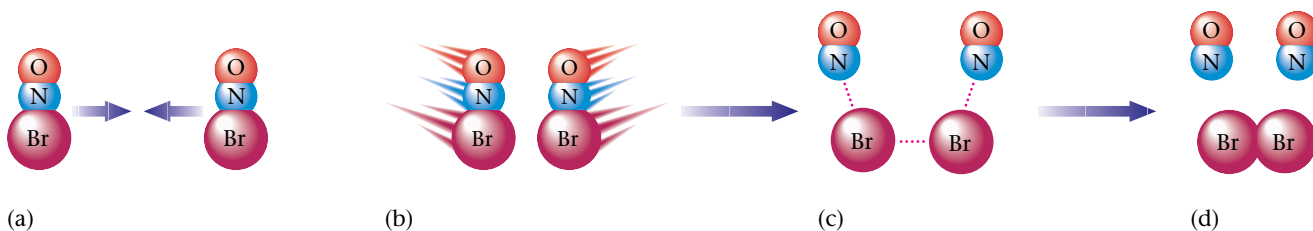


Figure 17.1

Visualizing the reaction $2\text{BrNO}(g) \rightarrow 2\text{NO}(g) + \text{Br}_2(g)$. (a) Two BrNO molecules approach each other at high speeds. (b) The collision occurs. (c) The energy of the collision causes the Br—N bonds to break and allows the Br—Br bond to form. (d) The products: one Br₂ and two NO molecules.

B. Conditions That Affect Reaction Rates

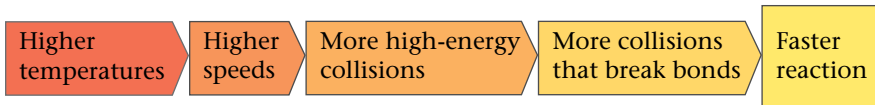
It is easy to see why reactions speed up when the *concentrations* of reacting molecules are increased: higher concentrations (more molecules per unit volume) lead to more collisions and so to more reaction events. But reactions also speed up when the *temperature* is increased. Why? The answer lies in the fact that not all collisions possess enough energy to break bonds. A minimum energy called the **activation energy (E_a)** is needed for a

Activation energy (E_a)

The minimum energy required in order to cause a chemical reaction

reaction to occur (see **Figure 17.2**). If a given collision possesses an energy greater than E_a , that collision can result in a reaction. If a collision has an energy less than E_a , the molecules will bounce apart unchanged.

The reason that a reaction occurs faster as the temperature is increased is that the speeds of the molecules increase with temperature. So at higher temperatures, the average collision is more energetic. This makes it more likely that a given collision will possess enough energy to break bonds and to produce the molecular rearrangements needed for a reaction to occur.



Active Reading Question

Why does increasing the temperature speed up a chemical reaction?

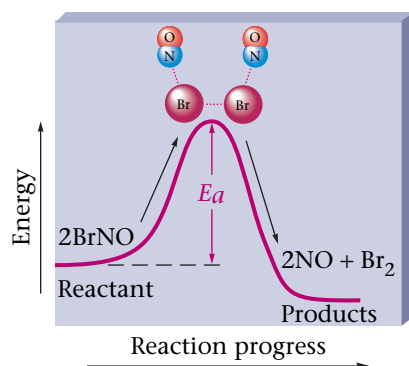
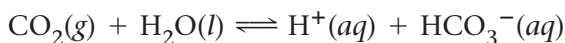


Figure 17.2

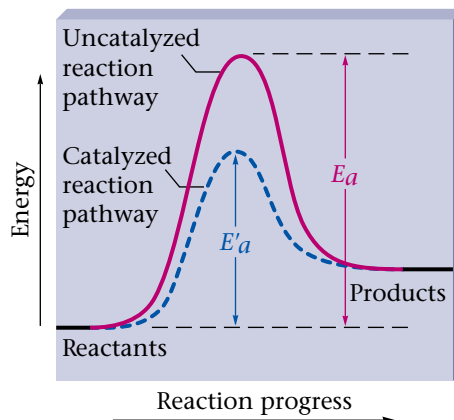
When molecules collide, a certain minimum energy called the activation energy (E_a) is needed for a reaction to occur. If the energy contained in a collision of two BrNO molecules is greater than E_a , the reaction can go “over the hump” to form products. If the collision energy is less than E_a , the colliding molecules bounce apart unchanged.

Is it possible to speed up a reaction without changing the temperature or the reactant concentrations? Yes, by using something called a **catalyst**, a substance that speeds up a reaction without being consumed. This may sound too good to be true, but it is a very common occurrence. In fact, you would not be alive now if your body did not contain thousands of catalysts called **enzymes**. Enzymes allow our bodies to speed up complicated reactions that would be too slow to sustain life at normal body temperatures. For example, the enzyme carbonic anhydrase speeds up the reaction between carbon dioxide and water



to help prevent an excess accumulation of carbon dioxide in our blood.

Although we cannot consider the details here, a catalyst works because it provides a new pathway for the reaction—a pathway that has a lower activation energy than the original pathway, as illustrated in **Figure 17.3**. Because of the lower activation energy, more collisions will have enough energy to allow a reaction. This in turn leads to a faster reaction.



CRITICAL THINKING ?

Many conditions need to be met to produce a chemical reaction between molecules.

What if all collisions between molecules resulted in a chemical reaction? How would life be different?

CRITICAL THINKING ?

Most modern refrigerators have an internal temperature of 45 °F. What if refrigerators were set at 55 °F in the factory? How would this change affect our lives?

Catalyst

A substance that speeds up a reaction without being consumed

Enzyme

A large molecule, usually a protein, which catalyzes biological reactions

Figure 17.3

Comparison of the activation energies for an uncatalyzed reaction (E_a) and for the same reaction with a catalyst present (E'_a). Note that a catalyst works by lowering the activation energy for a reaction.

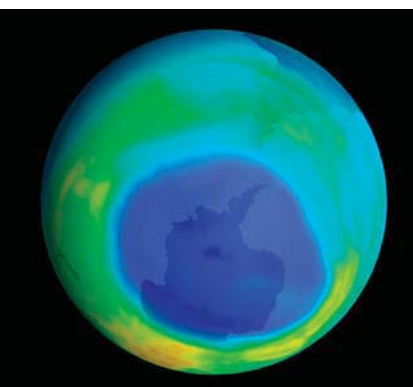
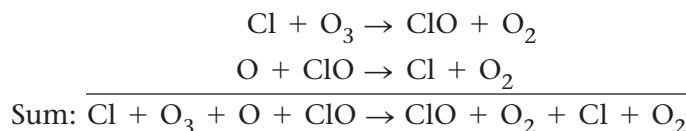
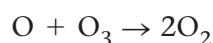


Figure 17.4
Data from the Total Ozone Mapping Spectrometer (TOMS) Earth Probe for the month of September, 2003. Areas of depleted ozone over the Antarctic are shown in dark blue.

Ozone Destruction A very important example of a reaction involving a catalyst occurs in our atmosphere; it is the breakdown of ozone, O_3 , which is catalyzed by chlorine atoms. Ozone is one constituent of the earth's upper atmosphere that is especially crucial, because it absorbs harmful high-energy radiation from the sun. There are natural processes that result in both the formation and the destruction of ozone in the upper atmosphere. The natural balance of all these opposing processes has resulted in an amount of ozone that has been relatively constant over the years. However, the ozone level now seems to be decreasing, especially over Antarctica (see **Figure 17.4**), apparently because chlorine atoms act as catalysts for the decomposition of ozone to oxygen by the following pair of reactions:



When species that appear on both sides of the equation are canceled, the end result is the reaction



CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Protecting the Ozone

Chlorofluorocarbons (CFCs) are ideal compounds for refrigerators and air conditioners because they are nontoxic and noncorrosive. However, the chemical inertness of these substances, once thought to be their major virtue, turns out to be their fatal flaw. When these compounds leak into the atmosphere, as they inevitably do, they are so unreactive they persist there for decades. Eventually these CFCs reach altitudes where ultraviolet light causes them to decompose, producing chlorine atoms that promote the destruction of the ozone in the stratosphere (see discussion above). Because of this problem, the world's industrialized nations have signed an agreement (called the Montreal Protocol) that banned CFCs in 1996 (with a 10-year grace period for developing nations). So we must find substitutes for the CFCs—and fast.

In fact, the search for substitutes is now well under way. Worldwide production of CFCs has already decreased to half of the 1986 level of 1.13 million metric tons. One strategy for replacing the CFCs has been to switch to similar compounds that contain carbon and hydrogen atoms substituted for chlorine

atoms. For example, the United States appliance industry has switched from Freon-12 (CF_2Cl_2) to the compound CH_2FCH_3 (called HFC-134a) for home refrigerators, and most of the new cars and trucks sold in the United States have air conditioners that employ HFC-134a. Converting the 140 million autos currently on the road in the United States that use CF_2Cl_2 will pose a major headache, but experience suggests that replacement of Freon-12 with HFC-134a is less expensive than was originally feared.

The chemical industry has responded amazingly fast to the ozone depletion emergency. It is encouraging that we can act rapidly when an environmental crisis occurs. Now we need to get better at keeping the environment at a higher priority as we plan for the future.

A modern refrigerator, one of many appliances that now use HFC-134a. This compound is replacing CFCs, which lead to the decomposition of the atmospheric ozone.



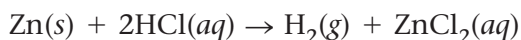
Notice that a chlorine atom is used up in the first reaction but a chlorine atom is formed again by the second reaction. Therefore, the amount of chlorine does not change as the overall process occurs. This means that the chlorine atom is a true catalyst: it participates in the process but is not consumed. Estimates show that *one chlorine atom can catalyze the destruction of about one million ozone molecules per second*.

The chlorine atoms that promote this damage to the ozone layer are present because of pollution. Specifically, they come from the decomposition of compounds called Freons, such as CF_2Cl_2 , which have been widely used in refrigerators and air conditioners. The Freons have leaked into the atmosphere, where they are decomposed by light to produce chlorine atoms and other substances. As a result, the manufacture of Freons was banned by agreement among the nations of the world as of the end of 1996. Substitute compounds are now being used in most countries in newly manufactured refrigerators and air conditioners.

C. Heterogeneous Reactions

Most of the reactions we have considered in this text occur in one phase. That is, the reactants and products are all gases or the reaction occurs with all reactants and products dissolved in a solution. Reactions involving only one phase are called **homogeneous reactions**.

Many other important reactions involve reactants in two phases. These reactions are **heterogeneous reactions**. An example of a heterogeneous reaction is zinc metal reacting with hydrochloric acid to produce hydrogen gas and aqueous zinc chloride:



This reaction is shown in **Figure 17.5**. The speed of this reaction depends on the surface area of the Zn available to react with H^+ ions from the solution. Large pieces of zinc are associated with a relatively slow reaction; if the zinc is ground up, however, the reaction is greatly accelerated by the increased surface area of the zinc.

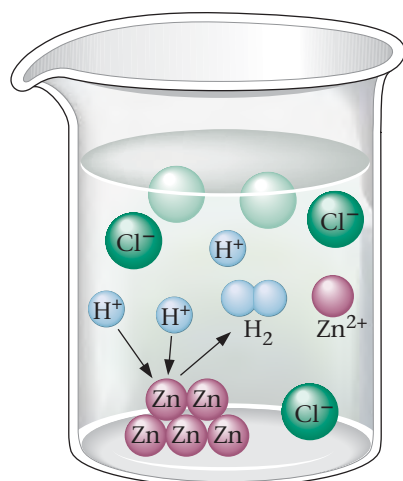


Figure 17.5

(a) The reaction between Zn and $\text{HCl}(aq)$.
(b) Molecular-level picture of the reaction

DID YOU KNOW

Although O atoms are too reactive to exist near the earth's surface, they do exist in the upper atmosphere.

Image not available for electronic use. Please refer to the image in the textbook.

(a)

(b)

Homogeneous reaction

Reaction involving reactants and products in the same state

Heterogeneous reaction

Reaction involving reactants and products in different states



Figure 17.6
A grain elevator near Haysville, Kansas, was severely damaged by a grain dust explosion.

Another example of the effect of surface area on the rate of a heterogeneous reaction is the explosive combustion of grain dust. Although grains of wheat will burn, the combustion reaction is relatively slow and controlled. In a storage silo, however, the tiny bits of dry plant materials (“grain dust”) that become airborne as grain is poured into the silo can combust explosively if a spark occurs. The force of these explosions is quite devastating (see **Figure 17.6**). The reaction of grain dust occurs explosively because the tiny sizes of the dust particles expose a tremendous amount of surface area to the oxygen in the air.

Let's Review

Factors That Affect Reaction Rates

Nature of Reactants: Substances vary greatly in their tendency to react depending on their bond strengths and structures.

Concentration (Pressure): The rate of a homogeneous reaction depends on the number of collisions that occur between reactants. Reaction rates typically increase as concentration (solution reactions) or pressure (gaseous reactions) increases.

Temperature: Because increased temperature accelerates reactant speeds, and thus increases the number of high-energy collisions, reaction rates increase with an increase in temperature.

Surface Area: For heterogeneous reactions, reaction rates increase with increased surface area.

CHEMISTRY IN YOUR WORLD

Consumer Connection

Some Like It Hot: The Top Ten Spices

Have you ever noticed that a direct relationship exists between the temperature of a place and the spiciness of its food? That is, the closer a country is to the equator, the spicier its food tends to be. Think about the difference between Scandinavian food, which uses relatively few spices, and Thai food, which often is extremely “hot.” An extensive study headed by biologist Paul Sherman of Cornell University has confirmed this relationship. Sherman and his students reviewed 4600 recipes from more than 90 cookbooks of traditional food from 36 countries and found that the amount of spicing is directly proportional to the warmth of the country’s climate.

What’s responsible for this interesting correlation? Kinetics. Food spoils by chemical reactions, and

Top Ten Antimicrobial Spices
Garlic
Onion
Allspice
Oregano
Thyme
Cinnamon
Tarragon
Cumin
Cloves
Lemon grass

that “hot” spices protect food against hot temperatures.

Garlic is the most popular spice.

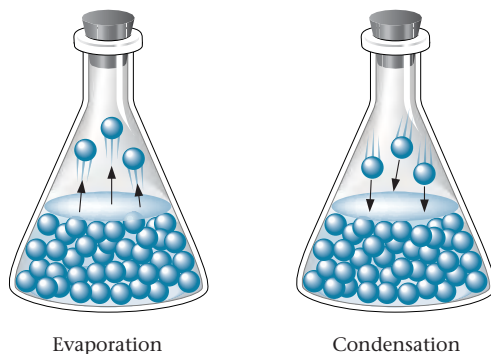


D. The Equilibrium Condition

Equilibrium is a word that implies balance or steadiness. When we say that someone is maintaining his or her equilibrium, we are describing a state of balance among various opposing forces. The term is used in a similar but more specific way in chemistry. Chemists define **equilibrium** as the *exact balancing of two processes, one of which is the opposite of the other*.

Equilibrium

The exact balance of two processes, one of which is opposite of the other



We first encountered the concept of equilibrium in Chapter 14, when we described the way vapor pressure develops over a liquid in a closed container. This equilibrium process is summarized in **Figure 17.7**. The *equilibrium state* occurs when the rate of evaporation exactly equals the rate of condensation.

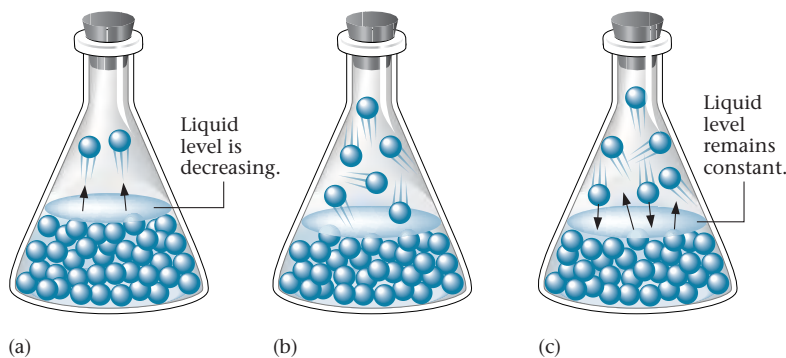
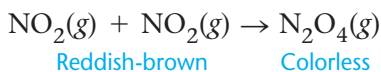


Figure 17.7

The establishment of the equilibrium vapor pressure over a liquid in a closed container.
 (a) At first there is a net transfer of molecules from the liquid state to the vapor state.
 (b) After a while, the amount of the substance in the vapor state becomes constant—both the pressure of the vapor and the level of the liquid stay the same. This is the equilibrium state.
 (c) The equilibrium state is very dynamic. The vapor pressure and liquid level remain constant because exactly the same number of molecules escape the liquid as return to it.

So far in this textbook we have usually assumed that reactions proceed to completion—that is, until one of the reactants “runs out.” Indeed, many reactions *do* proceed essentially to completion. For such reactions we can assume that the reactants are converted to products until the limiting reactant is completely consumed. On the other hand, there are many chemical reactions that “stop” far short of completion when they are allowed to take place in a closed container. An example is the reaction of nitrogen dioxide to form dinitrogen tetroxide.



The reactant NO_2 is a reddish-brown gas, and the product N_2O_4 is a colorless gas. Imagine an experiment where pure NO_2 is placed in an empty, sealed glass vessel at 25°C . The initial dark brown color will decrease in intensity as the NO_2 is converted to colorless N_2O_4 . However, even over a long period of time, the contents of the reaction vessel do not become colorless. Instead, the intensity of the brown color eventually becomes

Chemical equilibrium

A dynamic state where the concentrations of the reactants and the products remain constant over time, as long as the conditions are not changed

Information

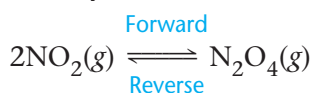
A double arrow (\rightleftharpoons) is used to show that a reaction is occurring in both directions.

constant, which means that the concentration of NO_2 is no longer changing. This simple observation is a clear indication that the reaction has “stopped” short of completion. In fact, the reaction has not stopped. Rather, the system has reached **chemical equilibrium**, a *dynamic state where the concentrations of all reactants and products remain constant*.

This situation is similar to the one where a liquid in a closed container develops a constant vapor pressure, except that in this case two opposite chemical reactions are involved. When pure NO_2 is first placed in the closed flask, there is no N_2O_4 present. As collisions between NO_2 molecules occur, N_2O_4 is formed and the concentration of N_2O_4 in the container increases. However, the reverse reaction can also occur. A given N_2O_4 molecule can decompose into two NO_2 molecules.



That is, chemical reactions are *reversible*; they can occur in either direction. We usually indicate this fact by using double arrows.



In this case the double arrows mean that either two NO_2 molecules can combine to form an N_2O_4 molecule (the *forward* reaction) or an N_2O_4 molecule can decompose to give two NO_2 molecules (the *reverse* reaction).

Equilibrium is reached whether pure NO_2 , pure N_2O_4 , or a mixture of NO_2 and N_2O_4 is initially placed in a closed container. In any of these cases, conditions will eventually be reached in the container such that N_2O_4 is being formed and is decomposing at exactly the same rate. This leads to chemical equilibrium, a dynamic situation where the concentrations of reactants and products remain the same indefinitely, as long as the conditions are not changed.

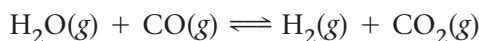
Active Reading Question

What do the double arrows in a chemical equation indicate?

E. Chemical Equilibrium: A Dynamic Condition

Because no changes occur in the concentrations of reactants or products in a reaction system at equilibrium, it may appear that everything has stopped. However, this is not the case. On the molecular level there is frantic activity. Equilibrium is not static but is a highly *dynamic* situation. Consider again the analogy between chemical equilibrium and two island cities connected by a single bridge. Suppose the traffic flow on the bridge is the same in both directions. It is obvious that there is motion (we can see the cars traveling across the bridge), but the number of cars in each city does not change because there is an equal flow of cars entering and leaving. The result is no *net* change in the number of cars in each of the two cities.

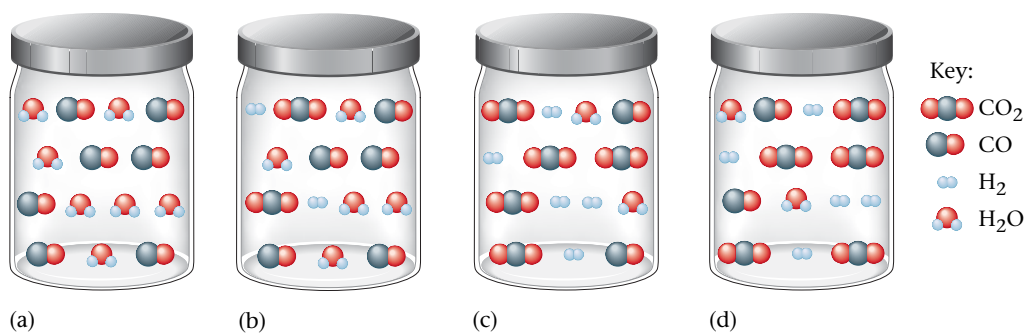
To see how this concept applies to chemical reactions, let's consider the reaction between steam and carbon monoxide in a closed vessel at a high temperature where the reaction takes place rapidly.



Assume that the same number of moles of gaseous CO and gaseous H_2O are placed in a closed vessel and allowed to react (see **Figure 17.8a**).

Information

Equilibrium is a dynamic situation.

**Figure 17.8**

The reaction of H₂O and CO to form CO₂ and H₂ as time passes.

(a) Equal numbers of moles of H₂O and CO are mixed in a closed container.

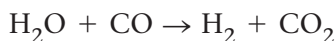
(b) The reaction begins to occur, and some products (H₂ and CO₂) are formed.

(c) The reaction continues as time passes and more reactants are changed to products.

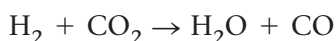
(d) Although time continues to pass, the numbers of reactant and product molecules are the same as in (c). No further changes are seen as time continues to pass. The system has reached equilibrium.

When CO and H₂O, the reactants, are mixed, they immediately begin to react to form the products, H₂ and CO₂. This leads to a decrease in the concentrations of the reactants, but the concentrations of the products, which were initially at zero, are increasing (see **Figure 17.8b**). After a certain period of time, the concentrations of reactants and products no longer change at all—equilibrium has been reached (see **Figure 17.8c and d**). Unless the system is somehow disturbed, no further changes in the concentrations will occur.

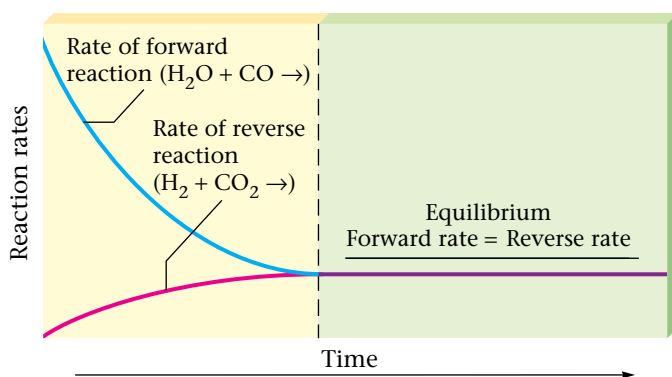
Why does equilibrium occur? We saw earlier in this chapter that molecules react by colliding with one another, and that the more collisions, the faster the reaction. This is why the speed of a reaction depends on concentrations. In this case the concentrations of H₂O and CO are lowered as the forward reaction occurs—that is, as products are formed.



As the concentrations of the reactants decrease, the forward reaction slows down (see **Figure 17.9**). But, there is also movement in the reverse direction.



Initially in this experiment, no H₂ and CO₂ are present, so this reverse reaction cannot occur. However, as the forward reaction proceeds, the concentrations of H₂ and CO₂ build up and the speed (or rate) of the reverse reaction increases (Figure 17.9) as the forward reaction slows down. Eventually the concentrations reach levels at which *the rate of the forward reaction equals the rate of the reverse reaction*. The system has reached equilibrium.

**Figure 17.9**

The changes with time in the rates of the forward and reverse reactions for $\text{H}_2\text{O}(g) + \text{CO}(g) \rightleftharpoons \text{H}_2(g) + \text{CO}_2(g)$ when equal numbers of moles of $\text{H}_2\text{O}(g)$ and $\text{CO}(g)$ are mixed. At first, the rate of the forward reaction decreases and the rate of the reverse reaction increases. Equilibrium is reached when the forward rate and the reverse rate become the same.

**Reaching Equilibrium: Are We There Yet?****Materials**

- two large containers
- water
- large cup, small cup
- permanent marker

Procedure

1. Fill two large containers with different amounts of water. (Plastic milk jugs cut in half horizontally work well.) You will also need a large cup, a small cup, a marker, and a partner.
2. Mark the level of water on the outside of each container.
3. Each person should do the following using his or her container:
 - a. Fill the cup as much as possible by dipping it into the container.
 - b. Pour the cup of water into your partner's container.

4. Repeat step 3 until the water level in each container is no longer changing. It is important that the same number of transfers occur in each direction.
5. Make five additional transfers as in step 3. Observe what happens to the water levels.

Results/Analysis

1. How many transfers were required until the water levels stopped changing?
2. Why does the level of water in each container stop changing? How is this situation similar to chemical equilibrium?
3. Even after the volume of water in each container stopped changing, water was still being transferred from one container to another. How is this situation similar to chemical equilibrium?

**SECTION 17.1****REVIEW QUESTIONS**

- 1 What is the main idea of the collision model?
- 2 How does a catalyst speed up a chemical reaction?
- 3 Use collision theory to explain why reactions should occur more slowly at lower temperatures.
- 4 Explain how grinding a solid in a heterogeneous reaction can speed up the reaction. Use collision theory in your answer.
- 5 Provide an example of a heterogeneous reaction and an example of a homogeneous reaction. Support your answer.
- 6 What is equal at equilibrium?
- 7 At the macroscopic level a system at equilibrium appears to be unchanging. Is it also unchanging at the molecular level? Explain.

**RESEARCH LINKS**
CLASSZONE.COM

Characteristics of Equilibrium

Objectives

- To understand the law of chemical equilibrium
- To learn to calculate values for the equilibrium constant
- To understand how the presence of solids or liquids affects the equilibrium expression

A. The Equilibrium Constant: An Introduction

Science is based on the results of experiments. The development of the equilibrium concept is typical. On the basis of their observations of many chemical reactions, two Norwegian chemists, Cato Maximilian Guldberg and Peter Waage, proposed in 1864 the **law of chemical equilibrium** (originally called the *law of mass action*) as a general description of the equilibrium condition. Guldberg and Waage postulated that for a reaction of the type



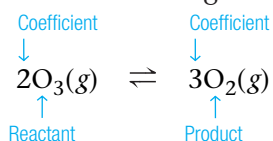
where A, B, C, and D represent chemical species and *a*, *b*, *c*, and *d* are their coefficients in the balanced equation, the law of chemical equilibrium is represented by the following **equilibrium expression**:

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

The square brackets indicate the concentrations of the chemical species *at equilibrium* (in units of mol/L), and *K* is a constant called the **equilibrium constant**. Note that the equilibrium expression is a special ratio of the concentrations of the products to the concentrations of the reactants. Each concentration is raised to a power corresponding to its coefficient in the balanced equation.

The law of chemical equilibrium as proposed by Guldberg and Waage is based on experimental observations. Experiments on many reactions showed that the equilibrium condition could always be described by this special ratio, called the equilibrium expression.

To see how to construct an equilibrium expression, consider the reaction where ozone changes to oxygen:



To obtain the equilibrium expression, we place the concentration of the product in the numerator and the concentration of the reactant in the denominator.

$$\frac{[\text{O}_2]}{[\text{O}_3]} \leftarrow \begin{array}{l} \text{Product} \\ \text{Reactant} \end{array}$$

Then we use the coefficients as powers.

$$K = \frac{[\text{O}_2]^3}{[\text{O}_3]^2} \leftarrow \begin{array}{l} \text{Coefficients become powers} \end{array}$$

Key Terms

- Law of chemical equilibrium
- Equilibrium expression
- Equilibrium constant
- Equilibrium position
- Homogeneous equilibria
- Heterogeneous equilibria

MATH



Square brackets, [], indicate concentration units of mol/L.

Active Reading Question

Why must you first balance a chemical equation before you can write the equilibrium expression?

EXAMPLE 17.1

Writing Equilibrium Expressions

Write the equilibrium expression for the following reactions.



Solution

Applying the law of chemical equilibrium, we place products over reactants (using square brackets to denote concentrations in units of moles per liter) and raise each concentration to the power that corresponds to the coefficient in the balanced chemical equation.

a. $K = \frac{[\text{HF}]^2}{[\text{H}_2][\text{F}_2]}$ ← Product (coefficient of 2 becomes power of 2)
← Reactants (coefficients of 1 become powers of 1)

Note that when a coefficient (power) of 1 occurs, it is not written but is understood.

b. $K = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$

MATH

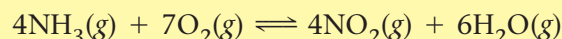


A number raised to a power of 1 shows no exponent:

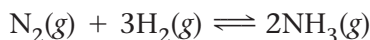
$$x^1 = x$$

Practice Problem • Exercise 17.1

Write the equilibrium expression for the following reaction.

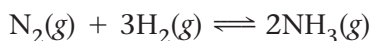


What does the equilibrium expression mean? It means that, for a given reaction at a given temperature, the special ratio of the concentrations of the products to reactants defined by the equilibrium expression will always be equal to the same number—namely, the equilibrium constant K . For example, consider a series of experiments on the ammonia synthesis reaction



carried out at 500 °C to measure the concentrations of N_2 , H_2 , and NH_3 present at equilibrium. The results of these experiments are shown in **Table 17.1**. In this table, subscript zeros next to square brackets are used to *indicate initial concentrations*: the concentrations of reactants and products originally mixed together before any reaction has occurred.

Consider the results of experiment I. One mole each of N_2 and H_2 were sealed into a 1-L vessel at 500 °C and allowed to reach chemical equilibrium. At equilibrium the concentrations in the flask were found to be $[\text{N}_2] = 0.921 \text{ M}$, $[\text{H}_2] = 0.763 \text{ M}$, and $[\text{NH}_3] = 0.157 \text{ M}$. The equilibrium expression for the reaction



is

$$K = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(0.157)^2}{(0.921)(0.763)^3} = 0.0602 = 6.02 \times 10^{-2}$$

Table 17.1

Results of Three Experiments for the Reaction $\text{N}_2(g) + 3\text{H}_2(g) \rightleftharpoons 2\text{NH}_3(g)$ at 500 °C

Experiment	Initial Concentrations			Equilibrium Concentrations			$\frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = K^*$
	$[\text{N}_2]_0$	$[\text{H}_2]_0$	$[\text{NH}_3]_0$	$[\text{N}_2]$	$[\text{H}_2]$	$[\text{NH}_3]$	
I	1.00 M	1.000 M	0	0.921 M	0.763 M	0.157 M	$\frac{(0.157)^2}{(0.921)(0.763)^3} = 0.0602$
II	0	0	1.000 M	0.399 M	1.197 M	0.203 M	$\frac{(0.203)^2}{(0.399)(1.197)^3} = 0.0602$
III	2.00 M	1.00 M	3.00 M	2.59 M	2.77 M	1.82 M	$\frac{(1.82)^2}{(2.59)(2.77)^3} = 0.0602$

*The units for K are customarily omitted.

Similarly, as shown in Table 17.1, we can calculate for experiments II and III that K , the equilibrium constant, has the value 6.02×10^{-2} . In fact, whenever N_2 , H_2 , and NH_3 are mixed together at this temperature, the system *always* comes to an equilibrium position such that

$$K = 6.02 \times 10^{-2}$$

regardless of the amounts of the reactants and products that are mixed together initially.

It is important to see from Table 17.1 that the *equilibrium concentrations are not always the same*. However, even though the individual sets of equilibrium concentrations are quite different for the different situations, the *equilibrium constant, which depends on the ratio of the concentrations, remains the same*.

Each set of *equilibrium concentrations* is called an **equilibrium position**. It is essential to distinguish between the equilibrium constant and the equilibrium positions for a given reaction system. There is only *one* equilibrium constant for a particular system at a particular temperature, but there is an *infinite* number of equilibrium positions. The specific equilibrium position adopted by a system depends on the initial concentrations; the equilibrium constant does not.

Note that in the preceding discussion, the equilibrium constant was given without units. In certain cases the units are included when the values of equilibrium constants are given, and in other cases they are omitted. We will not discuss the reasons for this. We will omit the units in this text.

Equilibrium position

A particular set of equilibrium concentrations of all reactants and products in a chemical system

Information

For a reaction at a given temperature, there are many equilibrium positions but only one value for K .

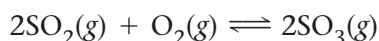
Active Reading Question

What is the difference between an “equilibrium constant” and an “equilibrium position”?

EXAMPLE 17.2

Calculating Equilibrium Constants

The reaction of sulfur dioxide with oxygen in the atmosphere to form sulfur trioxide has important environmental implications because SO_3 combines with moisture to form sulfuric acid droplets, an important component of acid rain. The following results were collected for two experiments involving the reaction at 600°C between gaseous sulfur dioxide and oxygen to form gaseous sulfur trioxide:



The law of chemical equilibrium predicts that the value of K should be the same for both experiments. Verify this by calculating the equilibrium constant observed for each experiment.

	Initial	Equilibrium
<i>Experiment I</i>	$[\text{SO}_2]_0 = 2.00\text{ M}$ $[\text{O}_2]_0 = 1.50\text{ M}$ $[\text{SO}_3]_0 = 3.00\text{ M}$	$[\text{SO}_2] = 1.50\text{ M}$ $[\text{O}_2] = 1.25\text{ M}$ $[\text{SO}_3] = 3.50\text{ M}$
<i>Experiment II</i>	$[\text{SO}_2]_0 = 0.500\text{ M}$ $[\text{O}_2]_0 = 0$ $[\text{SO}_3]_0 = 0.350\text{ M}$	$[\text{SO}_2] = 0.590\text{ M}$ $[\text{O}_2] = 0.045\text{ M}$ $[\text{SO}_3] = 0.260\text{ M}$

Solution

The balanced equation for the reaction is



From the law of chemical equilibrium, we can write the equilibrium expression:

$$K = \frac{[\text{SO}_3]^2}{[\text{SO}_2]^2[\text{O}_2]}$$

For Experiment I we calculate the value of K by substituting the observed *equilibrium* concentrations,

$$[\text{SO}_3] = 3.50\text{ M}; [\text{SO}_2] = 1.50\text{ M}; [\text{O}_2] = 1.25\text{ M}$$

into the equilibrium expression:

$$K_{\text{I}} = \frac{(3.50)^2}{(1.50)^2(1.25)} = 4.36$$

For Experiment II at equilibrium:

$$[\text{SO}_3] = 0.260\text{ M}; [\text{SO}_2] = 0.590\text{ M}; [\text{O}_2] = 0.045\text{ M}$$

and

$$K_{\text{II}} = \frac{(0.260)^2}{(0.590)^2(0.045)} = 4.32$$

Notice that the values calculated for K_{I} and K_{II} are nearly the same, as we expected. That is, the value of K is constant, within differences due to rounding off and due to experimental error. These experiments show *two different equilibrium positions* for this system, but K , the equilibrium constant, is, indeed, constant.

B. Heterogeneous Equilibria

So far we have discussed equilibria only for systems in the gaseous state, where all reactants and products are gases. These are examples of **homogeneous equilibria**, in which all substances are in the same state. However, many equilibria involve more than one state and are called **heterogeneous equilibria**. For example, the thermal decomposition of calcium carbonate in the commercial preparation of lime occurs by a reaction involving solids and gases.



Straightforward application of the law of equilibrium leads to the equilibrium expression

$$K = \frac{[\text{CO}_2][\text{CaO}]}{[\text{CaCO}_3]}$$

The position of a heterogeneous equilibrium does not depend on the amounts of pure solids or liquids present.

The fundamental reason for this behavior is that the concentrations of pure solids and liquids cannot change. In other words, we might say that the concentrations of pure solids and liquids are constants. Therefore, we can write the equilibrium expression for the decomposition of solid calcium carbonate as

$$K' = \frac{[\text{CO}_2]C_1}{C_2}$$

where C_1 and C_2 are constants representing the concentrations of the solids CaO and CaCO₃, respectively. This expression can be rearranged to give

$$\frac{C_2K'}{C_1} = K = [\text{CO}_2]$$

where the constants C_2 , K' , and C_1 are combined into a single constant K . This leads us to the following general statement:

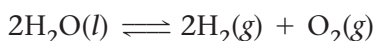
The concentrations of pure solids or pure liquids involved in a chemical reaction are not included in the equilibrium expression for the reaction.

This applies *only* to pure solids or liquids. It does not apply to solutions or gases, because their concentrations can vary.

Active Reading Question

Which species are not included in the equilibrium expression?

For example, consider the decomposition of liquid water to gaseous hydrogen and oxygen:



where

$$K = [\text{H}_2]^2[\text{O}_2]$$

Homogeneous equilibrium

An equilibrium system in which all reactants and products are in the same state

Heterogeneous equilibrium

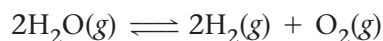
An equilibrium system in which all reactants and products are in different states

DID YOU KNOW

In terms of amount produced, lime is among the top ten chemicals manufactured in the United States.



Water is not included in the equilibrium expression because it is a pure liquid. However, when the reaction is carried out under conditions where the water is a gas rather than a liquid,



we have

$$K = \frac{[\text{H}_2]^2[\text{O}_2]}{[\text{H}_2\text{O}]^2}$$

because the concentration of water vapor can change.

EXAMPLE 17.3

Writing Equilibrium Expressions for Heterogeneous Equilibria

Write the expressions for K for the following processes.

- Solid phosphorus pentachloride is decomposed to liquid phosphorus trichloride and chlorine gas.
- Deep-blue solid copper(II) sulfate pentahydrate is heated to drive off water vapor to form white solid copper(II) sulfate.

Solution

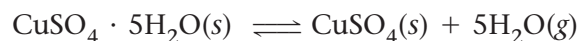
- a. The reaction is



In this case, neither the pure solid PCl_5 nor the pure liquid PCl_3 is included in the equilibrium expression. The equilibrium expression is

$$K = [\text{Cl}_2]$$

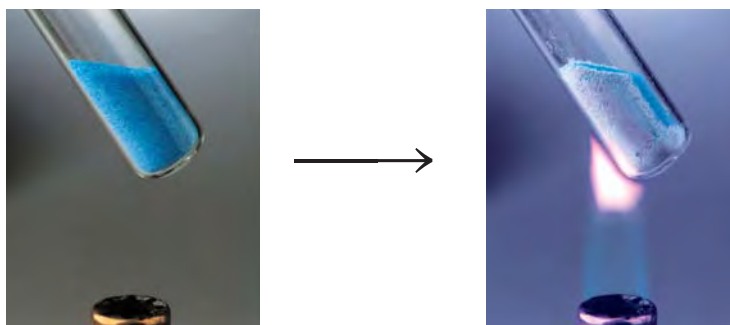
- b. The reaction is



The two solids are not included. The equilibrium expression is

$$K = [\text{H}_2\text{O}]^5$$

As solid copper(II) sulfate pentahydrate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, is heated, it loses H_2O , eventually forming white CuSO_4 .



Practice Problem • Exercise 17.3

Write the equilibrium expression for each of the following reactions.

- $2\text{KClO}_3(s) \rightleftharpoons 2\text{KCl}(s) + 3\text{O}_2(g)$
- $\text{NH}_4\text{NO}_3(s) \rightleftharpoons \text{N}_2\text{O}(g) + 2\text{H}_2\text{O}(g)$
- $\text{CO}_2(g) + \text{MgO}(s) \rightleftharpoons \text{MgCO}_3(s)$
- $\text{SO}_3(g) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{SO}_4(l)$

CELEBRITY CHEMICAL

Carbon Dioxide (CO₂)

Carbon dioxide is a very familiar molecule that is the product of human respiration and of the combustion of fossil fuels. This gas also puts the “fizz” in soda pop. It turns out that the simple, familiar CO₂ molecule has profound implications for our continued existence on this planet.

The concentration of carbon dioxide in the earth’s atmosphere has risen steadily since the beginning of the Industrial Revolution and is expected to double from today’s levels in the next 50 years. This change will have important effects on our environment. Consider the case of global warming. Many scientists feel that the increased CO₂ levels will trap more of the sun’s energy near the earth, significantly increasing the earth’s average temperature. Some evidence suggests that this is already occurring, although scientists disagree about the contribution of CO₂ to the changes. One thing that everyone does agree on is that the increased CO₂ levels will make plants grow faster.

On the surface the increase in plant growth due to increased CO₂ might seem like a good thing. Surprisingly, this situation could spell disaster for plant eaters, from caterpillars to antelopes, and for the animals that eat these herbivores. Faster plant

growth often leads to lower nutritional value. As the plants increase their rate of photosynthesis and use the carbon in CO₂ to build more fiber and starch, the amount of nitrogen—which indicates the amount of proteins present—declines. Studies show that new leaves on plants grown in an atmosphere rich in CO₂ are starchy, but protein-poor. This is bad news for caterpillars, which need to bulk up before they pupate. Studies have indicated that caterpillars eat 40% more of the starchy, protein-poor leaves but grow 10% slower and produce smaller than normal adult butterflies.

Studies on larger herbivores, such as cows and sheep, have been more difficult to carry out. Nevertheless, indications are that plants grown in a CO₂-enriched environment provide less protein and produce slower growth in these species as well.

Research is continuing to try to assess the effects of the increasing CO₂ levels on the food chain.

A Monarch butterfly caterpillar eating a leaf



SECTION 17.2

REVIEW QUESTIONS

- Write the equilibrium expression for $2\text{NO}(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}_2(g)$.
- Suppose that for the reaction $\text{N}_2(g) + 3\text{Cl}_2(g) \rightleftharpoons 2\text{NCl}_3(g)$ it is determined that, at a particular temperature, the equilibrium concentrations are $[\text{N}_2] = 0.000104\text{ M}$, $[\text{Cl}_2] = 0.000201\text{ M}$, and $[\text{NCl}_3] = 0.141\text{ M}$. Calculate the value of K for the reaction at this temperature.
- How can there be an infinite number of equilibrium *positions* for a particular system at a particular temperature, but only one equilibrium *constant* for this system at this temperature?
- Why are solids and liquids not included in the equilibrium expression?
- Write the equilibrium expression for the following chemical equations.
 - $\text{NH}_3(g) + \text{HCl}(g) \rightleftharpoons \text{NH}_4\text{Cl}(s)$
 - $2\text{H}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{H}_2\text{O}(g)$
 - $2\text{H}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{H}_2\text{O}(l)$

SECTION 17.3

Application of Equilibria

Key Terms

- Le Châtelier's principle
- Solubility product constant (K_{sp})

Objectives

- To learn to predict the changes that occur when a system at equilibrium is disturbed
- To learn to calculate equilibrium concentrations
- To learn to calculate the solubility product of a salt
- To learn to calculate solubility from the solubility product

A. Le Châtelier's Principle

It is important to understand the factors that control the *position* of a chemical equilibrium. For example, when a chemical is manufactured, the chemists and chemical engineers in charge of production want to choose conditions that favor the desired product as much as possible. That is, they want the equilibrium to lie far to the right (toward products). When the process for the synthesis of ammonia was being developed, extensive studies were carried out to determine how the equilibrium concentration of ammonia depended on the conditions of temperature and pressure.

In this section we will explore how various changes in conditions affect the equilibrium position of a reaction system. We can predict the effects of changes in concentration, pressure, and temperature on a system at equilibrium by using **Le Châtelier's principle**:

When a change is imposed on a system at equilibrium, the position of the equilibrium shifts in a direction that tends to reduce the effect of that change.

Le Châtelier's principle

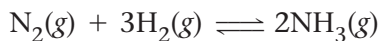
If a change is imposed on a system at equilibrium, the equilibrium position will shift to reduce the effect of that change.

The Effect of a Change in Concentration

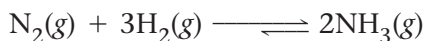
Let us consider the ammonia synthesis reaction. Suppose there is an equilibrium position described by these concentrations:

$$[\text{N}_2] = 0.399 \text{ M} \quad [\text{H}_2] = 1.197 \text{ M} \quad [\text{NH}_3] = 0.203 \text{ M}$$

Increasing $[\text{N}_2]$ What will happen if 1.000 mol/L of N_2 is suddenly injected into the system? We can begin to answer this question by remembering that for the system at equilibrium, the rates of the forward and reverse reactions exactly balance,



as indicated here by arrows of the same length. When the N_2 is added, there are suddenly more collisions between N_2 and H_2 molecules. This increases the rate of the forward reaction (shown here by the greater length of the arrow pointing in that direction),



and the reaction produces more NH_3 . As the concentration of NH_3 increases, the reverse reaction also speeds up (as more collisions between NH_3 molecules occur) and the system again comes to equilibrium. However, the new equilibrium position has more NH_3 than was present in the

original position. We say that the equilibrium has shifted to the *right*—toward the products. The original and new equilibrium positions are shown below.

<i>Equilibrium Position I</i>	→	<i>Equilibrium Position II</i>
$[\text{N}_2] = 0.399 \text{ M}$		$[\text{N}_2] = 1.348 \text{ M}$
$[\text{H}_2] = 1.197 \text{ M}$	1.000 mol/L of N_2 added	$[\text{H}_2] = 1.044 \text{ M}$
$[\text{NH}_3] = 0.203 \text{ M}$		$[\text{NH}_3] = 0.304 \text{ M}$

Note that the equilibrium does in fact shift to the right; the concentration of H_2 decreases (from 1.197 M to 1.044 M), the concentration of NH_3 increases (from 0.203 M to 0.304 M), and, of course, because nitrogen was added, the concentration of N_2 shows an increase relative to the original amount present.

It is important to note at this point that, although the equilibrium shifted to a new position, the *value of K did not change*. We can demonstrate this by inserting the equilibrium concentrations from positions I and II into the equilibrium expression.

- Position I: $K = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(0.203)^2}{(0.399)(1.197)^3} = 0.0602$
- Position II: $K = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(0.304)^2}{(1.348)(1.044)^3} = 0.0602$

These values of K are the same. Therefore, although the *equilibrium position* shifted when we added more N_2 , the *equilibrium constant K* remained the same.

Using Le Châtelier's Principle Could we have predicted this shift by using Le Châtelier's principle? Because the change in this case was to add nitrogen, Le Châtelier's principle predicts that the system will shift in a direction that *consumes* nitrogen. This tends to offset the original change—the addition of N_2 . Therefore, Le Châtelier's principle correctly predicts that adding nitrogen will cause the equilibrium to shift to the right (see **Figure 17.10**) as some of the added nitrogen is consumed.

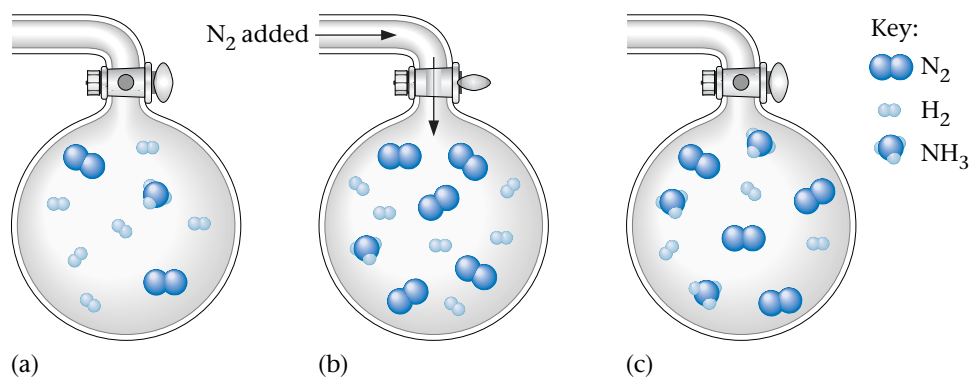


Figure 17.10

(a) The initial equilibrium mixture of N_2 , H_2 , and NH_3 .
 (b) Addition of N_2 .
 (c) The new equilibrium position for the system containing more N_2 (because of the addition of N_2), less H_2 , and more NH_3 than in (a).

Information

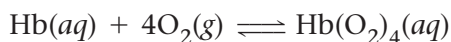
A system at equilibrium shifts in the direction that compensates for any imposed change.

Increasing $[\text{NH}_3]$ If ammonia had been added instead of nitrogen, the system would have shifted to the left, consuming ammonia. Another way of stating Le Châtelier's principle is as follows:

When a reactant or product is added to a system at equilibrium, the system shifts away from the added component. On the other hand, if a reactant or product is removed, the system shifts toward the removed component.

For example, if we had removed nitrogen, the system would have shifted to the left and the amount of ammonia present would have been reduced.

In Real Life An example that shows the importance of Le Châtelier's principle is the effect of high elevations on the oxygen supply to the body. If you have ever traveled to the mountains on vacation, you may have noticed that you felt lightheaded and especially tired during the first few days of your visit. These feelings resulted from a decreased supply of oxygen to your body because of the lower air pressure that exists at higher elevations. For example, the oxygen supply in Leadville, Colorado (elevation $\sim 10,000$ ft), is only about two-thirds that found at sea level. We can understand the effects of diminished oxygen supply in terms of the following equilibrium:



where Hb represents hemoglobin, the iron-containing protein that transports O_2 from your lungs to your tissues, where it is used to support metabolism. The coefficient 4 in the equation signifies that each hemoglobin molecule picks up four O_2 molecules in the lungs. Note by Le Châtelier's principle that a lower oxygen pressure will cause this equilibrium to shift to the left, away from oxygenated hemoglobin. This leads to an inadequate oxygen supply at the tissues, which in turn results in fatigue and a "woozy" feeling.

This problem can be solved in extreme cases, such as when climbing Mount Everest or flying in a plane at high altitudes, by supplying extra oxygen from a tank. This extra oxygen pushes the equilibrium to its normal position. However, lugging around an oxygen tank would not be very practical for people who live in the mountains. In fact, nature solves this problem in a very interesting way. The body adapts to living at high elevations by producing additional hemoglobin—the other way to shift this equilibrium to the right. Thus people who live at high elevations have significantly higher hemoglobin levels than those living at sea level. For example, the Sherpas who live in Nepal can function in the rarefied air at the top of Mt. Everest without an auxiliary oxygen supply.



Active Reading Question

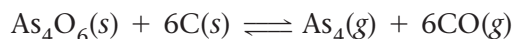
If a reactant is removed from a chemical system at equilibrium, what happens to the concentration of the products?



Sherpas in Nepal climb easily at high elevations.

EXAMPLE 17.4**Using Le Châtelier's Principle: Changes in Concentration**

Arsenic, As_4 , is obtained from nature by first reacting its ore with oxygen (called *roasting*) to form solid As_4O_6 . (As_4O_6 , a toxic compound fatal in doses of 0.1 g or more, is the “arsenic” made famous in detective stories.) The As_4O_6 is then reduced using carbon:



Predict the direction of the shift in the equilibrium position for this reaction that occurs in response to each of the following changes in conditions.

- Addition of carbon monoxide
- Addition or removal of $\text{C}(\text{s})$ or $\text{As}_4\text{O}_6(\text{s})$
- Removal of $\text{As}_4(\text{g})$

Solution

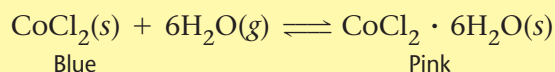
- Le Châtelier's principle predicts a shift away from the substance whose concentration is increased. The equilibrium position will shift to the left when carbon monoxide is added.
- Because the amount of a pure solid has no effect on the equilibrium position, changing the amount of carbon or tetraarsenic hexoxide will have no effect.
- When gaseous arsenic is removed, the equilibrium position will shift to the right to form more products. In industrial processes, the desired product is often continuously removed from the reaction system to increase the yield.



An ore sample containing arsenic

Practice Problem • Exercise 17.4

Novelty devices for predicting rain contain cobalt(II) chloride and are based on the following equilibrium:



What color will this indicator be when rain is likely due to increased water vapor in the air?



When blue anhydrous CoCl_2 reacts with water, pink $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}$ is formed.

The Effect of a Change in Volume

When the volume of a gas is decreased (when a gas is compressed), the pressure increases. This occurs because the molecules present are now contained in a smaller space and they hit the walls of their container more often, giving a greater pressure. Therefore, when the volume of a gaseous reaction system at equilibrium is suddenly reduced, leading to a sudden increase in pressure, by Le Châtelier's principle the system will shift in the direction that reduces the pressure.

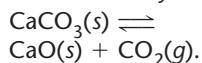
For example, consider the reaction



in a container with a movable piston (see **Figure 17.11**).

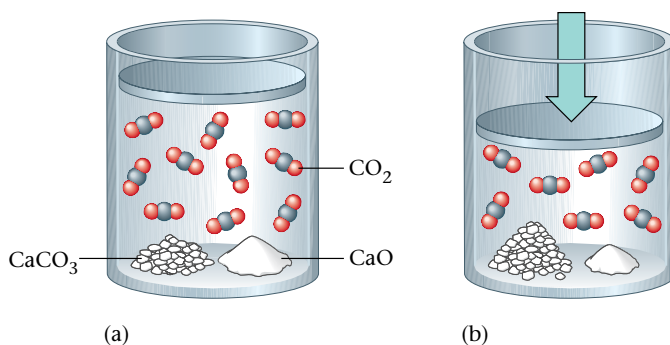
Figure 17.11

The reaction system



(a) The system is initially at equilibrium.

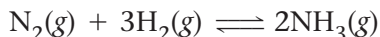
(b) The piston is pushed in, decreasing the volume and increasing the pressure. The system shifts in the direction that consumes CO_2 molecules, lowering the pressure again.



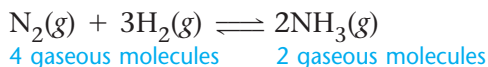
Decreasing the Volume If the volume is suddenly decreased by pushing in the piston, the pressure of the CO_2 gas initially increases. How can the system offset this pressure increase? By shifting to the left—the direction that reduces the amount of gas present. That is, a shift to the left will use up CO_2 molecules, thereby lowering the pressure. (There will be fewer molecules present to hit the walls, because more of the CO_2 molecules have combined with CaO and thus become part of the solid CaCO_3 .)

Therefore, when the volume of a gaseous reaction system at equilibrium is decreased (thus increasing the pressure), *the system shifts in the direction that gives the smaller number of gas molecules*. So a decrease in the system volume leads to a shift that decreases the total number of gaseous molecules in the system.

Suppose we are running the reaction



and we have a mixture of the gases nitrogen, hydrogen, and ammonia at equilibrium (see **Figure 17.12a**). If we suddenly reduce the volume, what will happen to the equilibrium position? Because the decrease in volume initially increases the pressure, the system moves in the direction that lowers its pressure. The reaction system can reduce its pressure by reducing the number of gas molecules present. This means that the reaction



shifts to the right, because in this direction four molecules (one of nitrogen and three of hydrogen) react to produce two molecules (of ammonia), thus *reducing the total number of gaseous molecules present*. The equilibrium position shifts to the right—toward the side of the reaction that involves the smaller number of gaseous molecules in the balanced equation.

i nformation

Shifts in equilibrium brought about by volume changes also produce changes in pressure.

Increasing the Volume When the container volume is increased (which lowers the pressure of the system), the system shifts so as to increase its pressure. An increase in volume in the ammonia synthesis system produces a shift to the left to increase the total number of gaseous molecules present (to increase the pressure).

Active Reading Question

How does the system shift when the volume of a gaseous reaction system at equilibrium is increased?

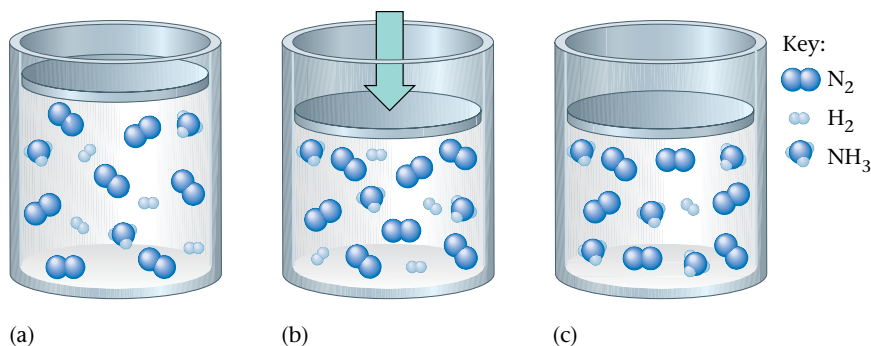


Figure 17.12

(a) A mixture of NH₃(g), N₂(g), and H₂(g) at equilibrium.

(b) The volume is suddenly decreased.

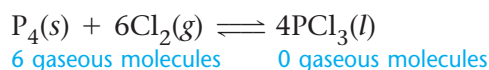
(c) The new equilibrium position for the system containing more NH₃ and less N₂ and H₂. The reaction N₂(g) + 3H₂(g) ⇌ 2NH₃(g) shifts to the right (toward the side with fewer molecules) when the container volume is decreased.

EXAMPLE 17.5

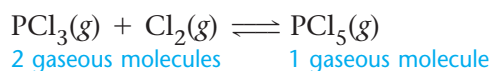
Using Le Châtelier's Principle: Changes in Volume

Predict the shift in equilibrium position that will occur for each of the following processes when the volume is reduced.

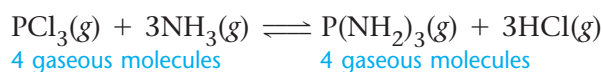
- a. The preparation of liquid phosphorus trichloride by the reaction



- b. The preparation of gaseous phosphorus pentachloride according to the equation



- c. The reaction of phosphorus trichloride with ammonia



Solution

- a. P₄ and PCl₃ are a pure solid and a pure liquid, respectively, so we need to consider only the effect on Cl₂. If the volume is decreased, the Cl₂ pressure will initially increase, so the position of the equilibrium will shift to the right, consuming gaseous Cl₂ and lowering the pressure (to counteract the original change).
- b. Decreasing the volume (increasing the pressure) will shift this equilibrium to the right, because the product side contains only one gaseous molecule while the reactant side has two. That is, the system will respond to the decreased volume (increased pressure) by lowering the number of molecules present.

- c. Both sides of the balanced reaction equation have four gaseous molecules. A change in volume will have no effect on the equilibrium position. There is no shift in this case, because the system cannot change the number of molecules present by shifting in either direction.

Practice Problem • Exercise 17.5

For each of the following reactions, predict the direction in which the equilibrium will shift when the volume of the container is increased.

- a. $\text{H}_2(\text{g}) + \text{F}_2(\text{g}) \rightleftharpoons 2\text{HF}(\text{g})$
b. $\text{CO}(\text{g}) + 2\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_3\text{OH}(\text{g})$
c. $2\text{SO}_3(\text{g}) \rightleftharpoons 2\text{SO}_2(\text{g}) + \text{O}_2(\text{g})$

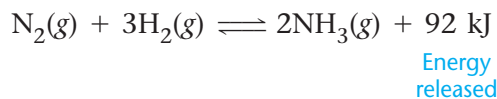
The Effect of a Change in Temperature

It is important to remember that although the changes we have just discussed may alter the equilibrium *position*, they do not alter the equilibrium *constant*. For example, the addition of a reactant shifts the equilibrium position to the right but has no effect on the value of the equilibrium constant; the new equilibrium concentrations satisfy the original equilibrium constant. This was demonstrated earlier in this section for the addition of N_2 to the ammonia synthesis reaction.

The effect of temperature on equilibrium is different, however, because *the value of K changes with temperature*. We can use Le Châtelier's principle to predict the direction of the change in K .

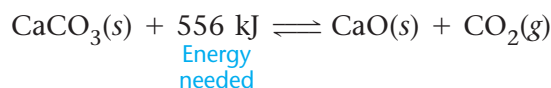
To do this we need to classify reactions according to whether they produce heat or absorb heat. A reaction that produces heat (heat is a "product") is said to be *exothermic*. A reaction that absorbs heat is called *endothermic*. Because heat is needed for an endothermic reaction, energy (heat) can be regarded as a "reactant" in this case.

Exothermic Reaction In an exothermic reaction, heat is treated as a product. For example, the synthesis of ammonia from nitrogen and hydrogen is exothermic (produces heat). We can represent this by treating energy as a product:



Le Châtelier's principle predicts that when we add energy to this system at equilibrium by heating it, the shift will be in the direction that consumes energy—that is, to the left.

Endothermic Reaction On the other hand, for an endothermic reaction (one that absorbs energy), such as the decomposition of calcium carbonate,



energy is treated as a reactant. In this case an increase in temperature causes the equilibrium to shift to the right.

Let's Review

Using Le Châtelier's Principle for Changes in Temperature

- Treat energy as a reactant in endothermic processes.
- Treat energy as a product in exothermic processes.
- Predict the direction of the shift in equilibrium position the same way as if a product or reactant were added or removed.

Active Reading Question

What property of a reaction allows you to predict the effect of a temperature change on the equilibrium?

EXAMPLE 17.6

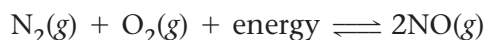
Using Le Châtelier's Principle: Changes in Temperature

For each of the following reactions, predict how the equilibrium will shift as the temperature is increased.

- $\text{N}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g})$ (endothermic)
- $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$ (exothermic)

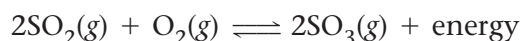
Solution

- This is an endothermic reaction, so energy can be viewed as a reactant.



Thus the equilibrium will shift to the right as the temperature is increased (energy added).

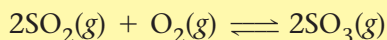
- This is an exothermic reaction, so energy can be regarded as a product.



As the temperature is increased, the equilibrium will shift to the left.

Practice Problem • Exercise 17.6

For the exothermic reaction



predict the equilibrium shift caused by each of the following changes.

- SO_2 is added.
- SO_3 is removed.
- The volume is decreased.
- The temperature is decreased.

We have seen how Le Châtelier's principle can be used to predict the effects of several types of changes on a system at equilibrium. To summarize these ideas, **Table 17.2** shows how various changes affect the equilibrium position of the endothermic reaction $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$. The effect of a temperature change on this system is shown in **Figure 17.13**.



Figure 17.13

Shifting the $\text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g)$ equilibrium by changing the temperature. (a) At 100 °C the flask is definitely reddish-brown due to a large amount of NO_2 present. (b) At 0 °C the equilibrium is shifted toward colorless $\text{N}_2\text{O}_4(g)$.

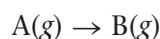
Table 17.2

**Shifts in the Equilibrium Position for the Reaction:
Energy + $\text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g)$**

Change	Shift
addition of $\text{N}_2\text{O}_4(g)$	right
addition of $\text{NO}_2(g)$	left
removal of $\text{N}_2\text{O}_4(g)$	left
removal of $\text{NO}_2(g)$	right
decrease in container volume	left
increase in container volume	right
increase in temperature	right
decrease in temperature	left

B. Applications Involving the Equilibrium Constant

Knowing the value of the equilibrium constant for a reaction allows us to do many things. For example, the size of K tells us the inherent tendency of the reaction to occur. A value of K much larger than 1 means that at equilibrium, the reaction system will consist of mostly products—the equilibrium lies to the right. For example, consider a general reaction of the type



where

$$K = \frac{[\text{B}]}{[\text{A}]}$$

If K for this reaction is 10,000 (10^4), then at equilibrium,

$$\frac{[\text{B}]}{[\text{A}]} = 10,000 \text{ or } = \frac{[\text{B}]}{[\text{A}]} = \frac{10,000}{1}$$

That is, at equilibrium $[\text{B}]$ is 10,000 times greater than $[\text{A}]$. This means that the reaction strongly favors the product B. Another way of saying this is that the reaction goes essentially to completion. That is, virtually all of A becomes B.

On the other hand, a small value of K means that the system at equilibrium consists largely of reactants—the equilibrium position is far to the left. The given reaction does not occur to any significant extent.

Let's Review

The Meaning of K

- $K > 1 \rightarrow$ the equilibrium position is far to the right
- $K < 1 \rightarrow$ the equilibrium position is far to the left

Another way we use the equilibrium constant is to calculate the equilibrium concentrations of reactants and products. For example, if we know the value of K and the concentrations of all the reactants and products except one, we can calculate the missing concentration. This is illustrated in Example 17.7 below.

Active Reading Question

What does the value of the equilibrium constant, K , tell us?

EXAMPLE 17.7

Calculating Equilibrium Concentration Using Equilibrium Expressions

Gaseous phosphorus pentachloride decomposes to chlorine gas and gaseous phosphorus trichloride. In a certain experiment, at a temperature where $K = 8.96 \times 10^{-2}$, the equilibrium concentrations of PCl_5 and PCl_3 were found to be $6.70 \times 10^{-3} \text{ M}$ and 0.300 M , respectively. Calculate the concentration of Cl_2 present at equilibrium.

Solution

Where do we want to go?

$$[\text{Cl}_2] = ? \text{ M}$$

What do we know?

- $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$
- $[\text{PCl}_5] = 6.70 \times 10^{-3} \text{ M}$ at equilibrium
- $K = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]} = 8.96 \times 10^{-2}$
- $[\text{PCl}_3] = 0.300 \text{ M}$ at equilibrium

How do we get there?

We want to calculate $[\text{Cl}_2]$. We will rearrange the equilibrium expression to solve for the concentration of Cl_2 . First we divide both sides of the expression

$$K = \frac{[\text{PCl}_3][\text{Cl}_2]}{[\text{PCl}_5]}$$

by $[\text{PCl}_3]$ to give

$$\frac{K}{[\text{PCl}_3]} = \frac{[\cancel{\text{PCl}_3}][\text{Cl}_2]}{[\cancel{\text{PCl}_3}][\text{PCl}_5]} = \frac{[\text{Cl}_2]}{[\text{PCl}_5]}$$

Next we multiply both sides by $[\text{PCl}_5]$.

$$\frac{K[\text{PCl}_5]}{[\text{PCl}_3]} = \frac{[\text{Cl}_2][\cancel{\text{PCl}_5}]}{[\cancel{\text{PCl}_5}]} = [\text{Cl}_2]$$

Then we can calculate $[\text{Cl}_2]$ by substituting the known information.

$$[\text{Cl}_2] = K \times \frac{[\text{PCl}_5]}{[\text{PCl}_3]} = (8.96 \times 10^{-2}) \frac{(6.70 \times 10^{-3})}{(0.300)}$$

$$[\text{Cl}_2] = 2.00 \times 10^{-3} \text{ M}$$

The equilibrium concentration of Cl_2 is $2.00 \times 10^{-3} \text{ M}$.



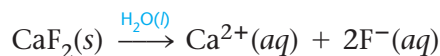
This X ray of the large intestine has been enhanced by the patient's consumption of barium sulfate.

C. Solubility Equilibria

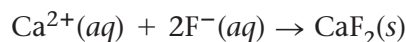
Solubility is a very important phenomenon. Consider the following examples.

- Because sugar and table salt dissolve readily in water, we can flavor foods easily.
- Because calcium sulfate is less soluble in hot water than in cold water, it coats tubes in boilers, reducing thermal efficiency.
- When food lodges between teeth, acids form that dissolve tooth enamel, which contains the mineral hydroxyapatite, $\text{Ca}_5(\text{PO}_4)_3\text{OH}$. Tooth decay can be reduced by adding fluoride to toothpaste. Fluoride replaces the hydroxide in hydroxyapatite to produce the corresponding fluorapatite, $\text{Ca}_5(\text{PO}_4)_3\text{F}$, and calcium fluoride, CaF_2 , both of which are less soluble in acids than the original enamel.
- The use of a suspension of barium sulfate improves the clarity of X rays of the digestive tract. Barium sulfate contains the toxic ion Ba^{2+} , but its very low solubility makes ingestion of solid BaSO_4 safe.

We will consider the equilibria associated with dissolving solids in water to form aqueous solutions. When a typical ionic solid dissolves in water, it dissociates completely into separate cations and anions. For example, calcium fluoride dissolves in water as follows:



When the solid salt is first added to the water, no Ca^{2+} and F^{-} ions are present. However, as dissolving occurs, the concentrations of Ca^{2+} and F^{-} increase, and it becomes more and more likely that these ions will collide and re-form the solid. Thus two opposite (competing) processes are occurring—the dissolving reaction shown above and the reverse reaction to re-form the solid:



Ultimately, equilibrium is reached. No more solid dissolves and the solution is said to be saturated.

We can write an equilibrium expression for this process according to the law of chemical equilibrium:

$$K_{\text{sp}} = [\text{Ca}^{2+}][\text{F}^{-}]^2$$

where $[\text{Ca}^{2+}]$ and $[\text{F}^{-}]$ are expressed in mol/L. The constant K_{sp} is called the **solubility product constant**, or simply the **solubility product**. Because CaF_2 is a pure solid, it is not included in the equilibrium expression.

Active Reading Question

Do very soluble salts have large or small K_{sp} values?

Information

Pure liquids and pure solids are never included in an equilibrium expression.

Solubility product constant (solubility constant; K_{sp})

The constant for the equilibrium expression representing the dissolving of an ionic solid in water

EXAMPLE 17.8**Writing Solubility Product Expressions**

Write the balanced equation describing the reaction for dissolving each of the following solids in water. Also write the K_{sp} expression for each solid.

- a. $\text{PbCl}_2(s)$ b. $\text{Ag}_2\text{CrO}_4(s)$ c. $\text{Bi}_2\text{S}_3(s)$

Solution

- a. $\text{PbCl}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2\text{Cl}^{-}(aq); K_{sp} = [\text{Pb}^{2+}][\text{Cl}^{-}]^2$
 b. $\text{Ag}_2\text{CrO}_4(s) \rightleftharpoons 2\text{Ag}^{+}(aq) + \text{CrO}_4^{2-}(aq); K_{sp} = [\text{Ag}^{+}]^2[\text{CrO}_4^{2-}]$
 c. $\text{Bi}_2\text{S}_3(s) \rightleftharpoons 2\text{Bi}^{3+}(aq) + 3\text{S}^{2-}(aq); K_{sp} = [\text{Bi}^{3+}]^2[\text{S}^{2-}]^3$

Practice Problem • Exercise 17.8

Write the balanced equation for the reaction describing the dissolving of each of the following solids in water. Also write the K_{sp} expression for each solid.

- a. $\text{BaSO}_4(s)$
 b. $\text{Fe}(\text{OH})_3(s)$
 c. $\text{Ag}_3\text{PO}_4(s)$

EXAMPLE 17.9**Calculating Solubility Products**

Copper(I) bromide, CuBr , has a measured solubility of 2.0×10^{-4} mol/L at 25°C . That is, when excess $\text{CuBr}(s)$ is placed in 1.0 L of water, we can determine that 2.0×10^{-4} mol of the solid dissolves to produce a saturated solution. Calculate the solid's K_{sp} value.

Solution**Where do we want to go?**

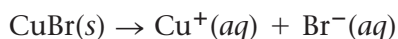
$$K_{sp} = ?$$

What do we know?

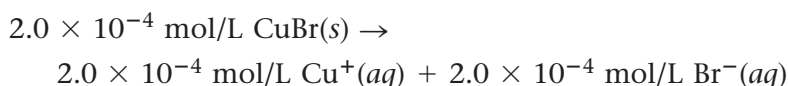
- Solubility $\text{CuBr} = 2.0 \times 10^{-4} \text{ M}$
- $\text{CuBr}(s) \rightleftharpoons \text{Cu}^{+}(aq) + \text{Br}^{-}(aq)$
- $K_{sp} = [\text{Cu}^{+}][\text{Br}^{-}]$

How do we get there?

We can calculate the value of K_{sp} if we know $[\text{Cu}^{+}]$ and $[\text{Br}^{-}]$, the equilibrium concentrations of the ions. We know that the measured solubility of CuBr is 2.0×10^{-4} mol/L. This means that 2.0×10^{-4} mol of solid CuBr dissolves per 1.0 L of solution to come to equilibrium. The reaction is



so

**i n f o r m a t i o n**

Solubilities must be expressed in mol/L in K_{sp} calculations.

We can now write the equilibrium concentrations

$$[\text{Cu}^+] = 2.0 \times 10^{-4} \text{ mol/L}$$

and

$$[\text{Br}^-] = 2.0 \times 10^{-4} \text{ mol/L}$$

These equilibrium concentrations allow us to calculate the value of K_{sp} for CuBr.

$$\begin{aligned} K_{\text{sp}} &= [\text{Cu}^+][\text{Br}^-] = (2.0 \times 10^{-4})(2.0 \times 10^{-4}) \\ &= 4.0 \times 10^{-8} \end{aligned}$$

The units for K_{sp} values are omitted.

Practice Problem • Exercise 17.9

Calculate the K_{sp} value for barium sulfate, BaSO_4 , which has a solubility of $3.9 \times 10^{-5} \text{ mol/L}$ at 25°C .

We have seen that the known solubility of an ionic solid can be used to calculate its K_{sp} value. The reverse is also possible: the solubility of an ionic solid can be calculated if its K_{sp} value is known.

EXAMPLE 17.10

Calculating Solubility from K_{sp} Values

The K_{sp} value for solid $\text{AgI}(s)$ is 1.5×10^{-16} at 25°C . Calculate the solubility of $\text{AgI}(s)$ in water at 25°C .

Solution

Where do we want to go?

Solubility for $\text{AgI}(s) = ? M$

What do we know?

- $K_{\text{sp}} = 1.5 \times 10^{-16}$
- $\text{AgI}(s) \rightleftharpoons \text{Ag}^+(aq) + \text{I}^-(aq)$
- $K_{\text{sp}} = [\text{Ag}^+][\text{I}^-]$

How do we get there?

Because we do not know the solubility of this solid, we will assume that x moles per liter dissolves to reach equilibrium. Therefore,



and at equilibrium,

$$[\text{Ag}^+] = x \frac{\text{mol}}{\text{L}}$$

$$[\text{I}^-] = x \frac{\text{mol}}{\text{L}}$$

Substituting these concentrations into the equilibrium expression gives

$$K_{\text{sp}} = 1.5 \times 10^{-16} = [\text{Ag}^+][\text{I}^-] = (x)(x) = x^2$$

Thus

$$x^2 = 1.5 \times 10^{-16}$$

$$x = \sqrt{1.5 \times 10^{-16}} = 1.2 \times 10^{-8} \text{ mol/L}$$

The solubility of $\text{AgI}(s)$ is $1.2 \times 10^{-8} \text{ mol/L}$.

Practice Problem • Exercise 17.10

The K_{sp} value for lead chromate, PbCrO_4 , is 2.0×10^{-16} at 25°C . Calculate its solubility at 25°C .



SECTION 17.3 REVIEW QUESTIONS

- When additional reactant is added to a system at equilibrium, what happens to the equilibrium position? What happens to the equilibrium constant?
- Suppose the reaction system $\text{UO}_2(s) + 4\text{HF}(g) \rightleftharpoons \text{UF}_4(g) + 2\text{H}_2\text{O}(g)$ has already reached equilibrium. Predict the effect of each of the following changes on the position of the equilibrium. Tell whether the equilibrium will shift to the right, shift to the left, or will not be affected.
 - Additional $\text{UO}_2(s)$ is added to the system.
 - 5.0 mol of $\text{Xe}(g)$ is added to the system at constant volume.
 - The reaction is performed in a glass reaction vessel; $\text{HF}(g)$ attacks and reacts with glass.
 - Water vapor is removed.
 - The size of the reaction vessel is increased.
- Why does changing the volume of a system at equilibrium that contains gas shift the equilibrium position? How can you predict this shift?
- When an exothermic reaction is at equilibrium what effect does increasing the temperature have on the equilibrium position? What effect does increasing the temperature have on the equilibrium constant?
- A reaction has a small value for K . At equilibrium what does this tell you about each of the following:
 - amount of reactants
 - amount of products
 - extent of reaction
- Write the equilibrium expression for $\text{AgCl}(s) \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq)$.
- The solubility of PbCrO_4 is $1.7 \times 10^{-4} \text{ g/L}$. Calculate the K_{sp} value for PbCrO_4 .



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Key Terms

17.1 Collision model

Activation energy (E_a)

Enzymes

Homogeneous reactions

Heterogeneous reactions

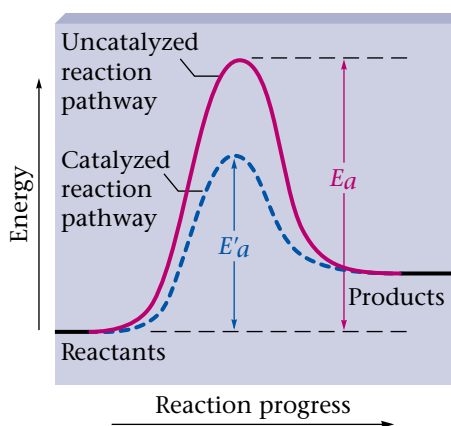
Equilibrium

Chemical equilibrium

Key Ideas

17.1 Reaction Rates and Equilibrium

- Collision model for chemical reactions
 - Reactants must collide to react.
 - A certain threshold energy (the activation energy, E_a) must be supplied by the collision for a reaction to occur.
- A catalyst
 - Speeds up a reaction without being consumed
 - Provides a new pathway for the reaction that has a smaller E_a
 - Enzymes are biological catalysts.



- Chemical equilibrium is established when a chemical reaction is carried out in a closed vessel.
 - The concentrations of both reactants and products remain constant over time.
 - Equilibrium is a highly dynamic state on the microscopic level.
 - ◆ Forward rate = reverse rate
 - In homogeneous reactions all reactants and products are in the same phase.
 - In heterogeneous reactions one or more reactants or products are in different phases.

17.2 Law of chemical equilibrium

Equilibrium expression

Equilibrium constant

Equilibrium position

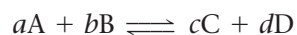
Homogeneous equilibria

Heterogeneous equilibria

17.2 Characteristics of Equilibrium

- The equilibrium expression is based on the law of chemical equilibrium.

- For the reaction



$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

- The equilibrium constant (K) is constant for a given chemical system at a given temperature.
- The equilibrium position is a set of equilibrium concentrations that satisfy K .
- There are an infinite number of equilibrium positions.
- Heterogeneous equilibria contain reactants or products in different phases.
 - A pure liquid or solid never appears in the equilibrium expression.

17.3 Le Châtelier's Principle

Solubility Product Constant (K_{sp})

17.3 Application of Equilibria

- Le Châtelier's principle states that when a change is imposed on a system at equilibrium the position of the equilibrium shifts in the direction which reduces the effect of that change.
- Applications of equilibria
 - The value of K for a system can be calculated from a known set of equilibrium concentrations.
 - Unknown equilibrium concentrations can be calculated if the value of K and the remaining equilibrium concentrations are known.
 - The equilibrium conditions also apply to a saturated solution containing excess solid, $MX(s)$.
 - ◆ $K_{sp} = [M^+][X^-]$ = solubility product constant
 - ◆ The value of the K_{sp} can be calculated from the measured solubility of $MX(s)$.

All exercises with **blue numbers** have answers in the back of this book.

17.1 Reaction Rates and Equilibrium

A. How Chemical Reactions Occur

1. For the reaction $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightarrow 2\text{HBr}(\text{g})$, list the types of bonds that must be broken and the type of bond that must form for the chemical reaction to take place.
2. For the reaction $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$, list the types of bonds that must be broken and the type of bond that must form for the chemical reaction to take place.

B. Conditions That Affect Reaction Rates

3. How do chemists envision reactions taking place in terms of the *collision model* for reactions? Give an example of a simple reaction and how you might envision the reaction taking place by means of a collision between the molecules.
4. What does the *activation energy* for a reaction represent? How is the activation energy related to whether a collision between molecules is successful?
5. What is a *catalyst*? How does a catalyst speed up a reaction?
6. What are the catalysts found in living cells called? Why are these biological catalysts necessary?

C. Heterogeneous Reactions

7. Explain the difference between a heterogeneous reaction and a homogeneous reaction.
8. Why does increasing the temperature speed up the rate of a chemical reaction?
9. Why is pouring grain into a silo a potentially explosive process? Use the collision theory in your explanation.

D. The Equilibrium Condition

10. How does *equilibrium* represent the balancing of opposing processes? Give an example of an “equilibrium” encountered in everyday life, showing how the processes involved oppose each other.
11. What do chemists mean by a state of *equilibrium*? Give an example of a *physical* equilibrium and of a *chemical* equilibrium.

12. What does it mean to say that chemical reactions are *reversible*? Are all chemical reactions, in principle, reversible? Are some reactions more likely to occur in one direction than in the other?

13. How do chemists recognize a system that has reached a state of chemical equilibrium? When writing chemical equations, how do we indicate reactions that come to a state of chemical equilibrium?

E. Chemical Equilibrium: A Dynamic Condition

14. When a reaction system has reached chemical equilibrium, the concentrations of the reactants and products no longer change with time. Why does the amount of product no longer increase, even though large concentrations of the reactants may still be present?
15. What does it mean to say that the condition of chemical equilibrium is a *dynamic* situation? Although the reaction overall may appear to have stopped, what is still going on in the system?

17.2 Characteristics of Equilibrium

A. The Equilibrium Constant: An Introduction

16. In general terms, what does the equilibrium constant for a reaction represent? What is the algebraic form of the equilibrium constant for a typical reaction? What do square brackets indicate when we write an equilibrium constant?
17. There is only one value of the equilibrium constant for a particular system at a particular temperature, but there is an infinite number of equilibrium positions. Explain.
18. Write the equilibrium expression for each of the following reactions.
 - a. $\text{N}_2(\text{g}) + 3\text{Cl}_2(\text{g}) \rightleftharpoons 2\text{NCl}_3(\text{g})$
 - b. $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$
 - c. $\text{N}_2(\text{g}) + 2\text{H}_2(\text{g}) \rightleftharpoons \text{N}_2\text{H}_4(\text{g})$
19. Write the equilibrium expression for each of the following reactions.
 - a. $\text{CO}(\text{g}) + 2\text{H}_2(\text{g}) \rightleftharpoons \text{CH}_3\text{OH}(\text{g})$
 - b. $2\text{NO}_2(\text{g}) \rightleftharpoons 2\text{NO}(\text{g}) + \text{O}_2(\text{g})$
 - c. $\text{P}_4(\text{g}) + 6\text{Br}_2(\text{g}) \rightleftharpoons 4\text{PBr}_3(\text{g})$

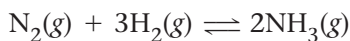
20. Suppose that for the reaction



it is determined that, at a particular temperature, the equilibrium concentrations are $[\text{CH}_3\text{OH}] = 0.00215 \text{ M}$, $[\text{CH}_2\text{O}] = 0.441 \text{ M}$, and $[\text{H}_2] = 0.0331 \text{ M}$.

Calculate the value of K for the reaction at this temperature.

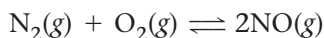
21. Ammonia, a very important industrial chemical, is produced by the direct combination of the following elements under carefully controlled conditions:



Suppose, in an experiment, that the reaction mixture is analyzed after equilibrium is reached and it is found, at the particular temperature, that $[\text{NH}_3] = 0.34 \text{ M}$, $[\text{H}_2] = 2.1 \times 10^{-3} \text{ M}$, and $[\text{N}_2] = 4.9 \times 10^{-4} \text{ M}$.

Calculate the value of K at this temperature.

22. At high temperatures, elemental nitrogen and oxygen react with each other to form nitrogen monoxide.

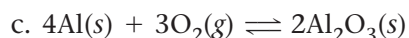
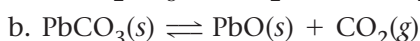
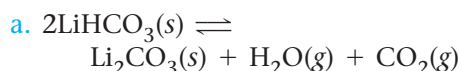


Suppose the system is analyzed at a particular temperature, and the equilibrium concentrations are found to be $[\text{N}_2] = 0.041 \text{ M}$, $[\text{O}_2] = 0.0078 \text{ M}$, and $[\text{NO}] = 4.7 \times 10^{-4} \text{ M}$.

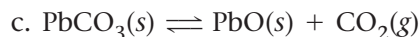
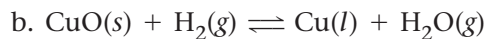
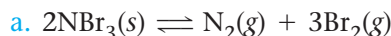
Calculate the value of K for the reaction.

B. Heterogeneous Equilibria

23. What is a *homogeneous* equilibrium system? Give an example of a homogeneous equilibrium reaction. What is a *heterogeneous* equilibrium system? Write two chemical equations that represent heterogeneous equilibria.
24. Explain why the position of a heterogeneous equilibrium does not depend on the amounts of pure solid or pure liquid reactants or products present.
25. Write the equilibrium expression for each of the following heterogeneous equilibria.



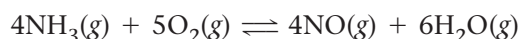
26. Write the equilibrium expression for each of the following heterogeneous equilibria.



17.3 Applications of Equilibria

A. Le Châtelier's Principle

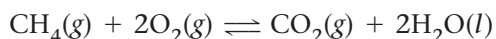
27. In your own words, describe what Le Châtelier's principle tells us about how we can change the position of a reaction system at equilibrium.
28. Discuss in general terms the effect on the net amount of product when an additional amount of one of the reactants is added to an equilibrium system. Does the value of the equilibrium constant change in this situation?
29. For an equilibrium involving gaseous substances, what effect, in general terms, is realized when the volume of the system is decreased?
30. What is the effect on the equilibrium position if an endothermic reaction is performed at a higher temperature? Does the net amount of product increase or decrease? Does the value of the equilibrium constant change if the temperature is increased?
31. For the reaction system



which has already reached a state of equilibrium, predict the effect that each of the following changes will have on the position of the equilibrium. Tell whether the equilibrium will shift to the right, will shift to the left, or will not be affected.

- a. The pressure of oxygen is increased by injecting one additional mole of oxygen into the reaction vessel.
- b. A desiccant (a material that absorbs water) is added to the system.
- c. The system is compressed and the ammonia liquefies.

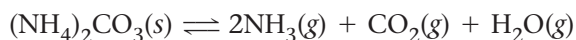
32. Suppose the reaction system



has already reached equilibrium. Predict the effect of each of the following changes on the position of the equilibrium. Tell whether the equilibrium will shift to the right, will shift to the left, or will not be affected.

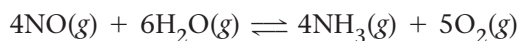
- Any liquid water present is removed from the system.
- CO_2 is added to the system by dropping a chunk of dry ice into the reaction vessel.
- The reaction is performed in a metal cylinder fitted with a piston, and the piston is compressed to decrease the total volume of the system.
- Additional $\text{O}_2(g)$ is added to the system from a cylinder of pure O_2 .

33. Old-fashioned “smelling salts” consist of ammonium carbonate, $(\text{NH}_4)_2\text{CO}_3$. The reaction for the decomposition of ammonium carbonate



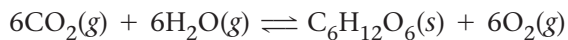
is endothermic. What would be the effect on the position of this equilibrium if the reaction were performed at a lower temperature?

34. The reaction



is strongly endothermic. Will an increase in temperature shift the equilibrium position toward the products or toward the reactants?

35. Plants synthesize the sugar dextrose according to the following reaction by absorbing radiant energy from the sun (photosynthesis).



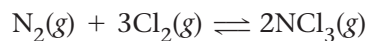
Will an increase in temperature tend to favor or discourage the production of $\text{C}_6\text{H}_{12}\text{O}_6(s)$?

B. Applications Involving the Equilibrium Constant

36. Suppose a reaction has the equilibrium constant $K = 1.3 \times 10^8$. What does the magnitude of this constant tell you about the relative concentrations of products and reactants that will be present once equilibrium is reached? Is this reaction likely to be a good source of the products?

37. Suppose a reaction has the equilibrium constant $K = 4.5 \times 10^{-6}$ at a particular temperature. If an experiment is set up with this reaction, will there be large relative concentrations of products present at equilibrium? Is this reaction useful as a means of producing the products? How might the reaction be made more useful?

38. For the reaction



an analysis of an equilibrium mixture is performed. It is found that

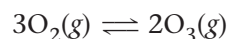
$$[\text{NCl}_3] = 1.9 \times 10^{-1} M,$$

$$[\text{N}_2] = 1.4 \times 10^{-3} M,$$

$$\text{and } [\text{Cl}_2] = 4.3 \times 10^{-4} M.$$

Calculate K for the reaction.

39. For the reaction



$K = 1.8 \times 10^{-7}$ at a particular temperature.

If the equilibrium system is analyzed and it is found that $[\text{O}_2] = 0.0012 M$, what is the concentration of O_3 in the system?

40. For the reaction

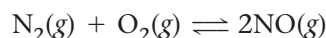


it is found that at equilibrium

$$[\text{CO}_2] = 2.1 \times 10^{-3} M$$

at a particular temperature. Calculate K for the reaction at this temperature.

41. For the reaction

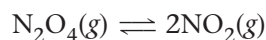


the equilibrium constant K has the value

$$1.71 \times 10^{-3}$$

at a particular temperature. If the concentrations of both $\text{N}_2(g)$ and $\text{O}_2(g)$ are $0.0342 M$ in an equilibrium mixture at this temperature, what is the concentration of $\text{NO}(g)$ under these conditions?

42. For the reaction



the equilibrium constant K has the value

$$8.1 \times 10^{-3}$$

at a particular temperature. If the concentration of $\text{NO}_2(g)$ is found to be $0.0021 M$ in the equilibrium system, what is the concentration of $\text{N}_2\text{O}_4(g)$ under these conditions?

C. Solubility Equilibria

43. Explain how the dissolving of an ionic solute in water represents an equilibrium process.
44. Why does the amount of excess solid solute present in a solution not affect the amount of solute that ultimately dissolves in a given amount of solvent?
45. Why does neither stirring nor grinding a solute affect the amount of solute that ultimately dissolves in a given amount of solvent?
46. Write the balanced chemical equations describing the dissolving of the following solids in water. Write the expression for K_{sp} for each process.
- | | |
|--------------------------|--|
| a. NiS(s) | c. BaCrO ₄ (s) |
| b. CuCO ₃ (s) | d. Ag ₃ PO ₄ (s) |
47. Write the balanced chemical equations describing the dissolving of the following solids in water. Write the expression for K_{sp} for each process.
- | | |
|--------------------------|--|
| a. PbBr ₂ (s) | c. PbCO ₃ (s) |
| b. Ag ₂ S(s) | d. Sr ₃ (PO ₄) ₂ (s) |
48. Zinc carbonate dissolves in water to the extent of 1.12×10^{-4} g/L at 25 °C. Calculate the solubility product K_{sp} for ZnCO₃ at 25 °C.
49. A saturated solution of nickel(II) sulfide contains approximately 3.6×10^{-4} g of dissolved NiS per liter at 20 °C. Calculate the solubility product K_{sp} for NiS at 20 °C.
50. The solubility product constant, K_{sp} , for calcium carbonate at room temperature is approximately 3.0×10^{-9} . Calculate the solubility of CaCO₃ in grams per liter under these conditions.
51. Chromate ion is used as a qualitative test for lead(II) ion, forming a bright yellow precipitate of lead chromate, PbCrO₄(s). The K_{sp} for PbCrO₄(s) is 2.8×10^{-13} at 25 °C. Calculate the solubility of PbCrO₄ in grams per liter at 25 °C.
52. Lead(II) chloride, PbCl₂(s), dissolves in water to the extent of approximately 3.6×10^{-2} M at 20 °C. Calculate the K_{sp} for PbCl₂(s), and calculate its solubility in grams per liter.
53. Mercury(I) chloride, Hg₂Cl₂, was formerly administered orally to induce vomiting. Although we usually think of mercury compounds as highly toxic, the K_{sp} of mercury(I) chloride is small enough (1.3×10^{-18}) that the amount of mercury that dissolves and enters the bloodstream is tiny. Calculate the concentration of mercury(I) ion present in a saturated solution of Hg₂Cl₂.

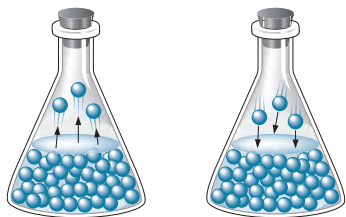
Critical Thinking

54. Before two molecules can react, chemists envision that the molecules must first *collide* with each other. Is collision among molecules the only consideration for the molecules to react with one another?

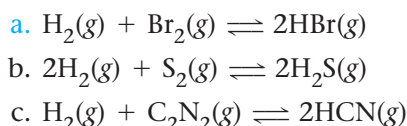


55. Why does an increase in temperature favor an increase in the speed of a reaction?
56. What does it mean to say that all chemical reactions are, to one extent or another, *reversible*?
57. What does it mean to say that chemical equilibrium is a *dynamic* process?
58. Why does increasing the temperature for an exothermic process tend to favor the conversion of products back to reactants?
59. Suppose $K = 4.5 \times 10^{-3}$ at a certain temperature for the reaction
- $$\text{PCl}_5(g) \rightleftharpoons \text{PCl}_3(g) + \text{Cl}_2(g)$$
- If it is found that the concentration of PCl₅ is twice the concentration of PCl₃, what must be the concentration of Cl₂ under these conditions?
60. How does the collision model account for the fact that a reaction proceeds faster when the concentrations of the reactants are increased?
61. How does an increase in temperature result in an increase in the number of successful collisions between reactant molecules? What does an increase in temperature mean on a molecular basis?

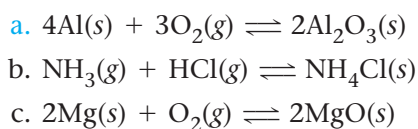
- 62.** Explain why the development of a vapor pressure above a liquid in a closed container represents an equilibrium. What are the opposing processes? How do we recognize when the system has reached a state of equilibrium?



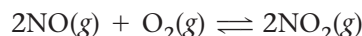
- 63.** Write the equilibrium expression for each of the following reactions.



- 64.** Write the equilibrium expression for each of the following heterogeneous equilibria.



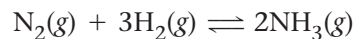
- 65.** Suppose the reaction system



has already reached equilibrium. Predict the effect of each of the following changes on the position of the equilibrium. Tell whether the equilibrium will shift to the right, will shift to the left, or will not be affected.

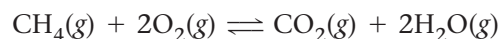
- Additional oxygen is injected into the system.
 - NO_2 is removed from the reaction vessel.
 - 1.0 mol of helium is injected into the system.
- 66.** Approximately 0.14 g of nickel(II) hydroxide, $\text{Ni}(\text{OH})_2(\text{s})$, dissolves per liter of water at 20°C . Calculate the K_{sp} for $\text{Ni}(\text{OH})_2(\text{s})$ at this temperature.

- 67.** Consider the reaction as represented by the following equation.



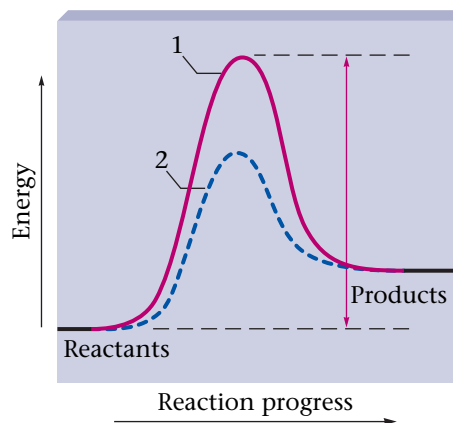
Initially, $[\text{N}_2] = 5.00\text{ M}$ and $[\text{H}_2] = 4.00\text{ M}$. At equilibrium, $[\text{H}_2] = 1.00\text{ M}$ at a certain temperature. Determine the value for K for this reaction at this temperature.

- 68.** Consider the combustion of methane (as represented by the following equation). This is the reaction that occurs for a Bunsen burner, which is a source of heat for chemical reactions in the laboratory.



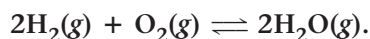
For the system at chemical equilibrium, which of the following explains what happens if the temperature is raised?

- The equilibrium position is shifted to the right and the value for K decreases.
 - The equilibrium position is shifted to the left and the value for K decreases.
 - The equilibrium position is shifted to the left and the value for K increases.
 - The equilibrium position is shifted but the value for K stays constant.
- 69.** Consider the figure below in answering the following questions.



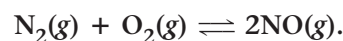
- What does a catalyst do to a chemical reaction?
- Which of the pathways in the figure is the catalyzed reaction pathway? How do you know?
- What is represented by the double-headed arrow?

- 1 Which of the following is true about chemical equilibrium?
- A It is microscopically and macroscopically static.
 - B It is microscopically and macroscopically dynamic.
 - C It is microscopically static and macroscopically dynamic.
 - D It is microscopically dynamic and macroscopically static.
- 2 Which of the following is true about a system at equilibrium?
- A The concentration(s) of the reactant(s) is equal to the concentration(s) of the product(s).
 - B No new product molecules are formed.
 - C The concentration of reactant(s) is constant over time.
 - D The rate of the reverse reaction is equal to the rate of the forward reaction and both rates are equal to zero.
- 3 Choose the correct equilibrium expression for the equation



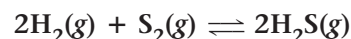
- A $\frac{[\text{H}_2\text{O}]^2}{[\text{H}_2]^2[\text{O}_2]}$
- B $\frac{[\text{H}_2\text{O}]^2}{[\text{H}_2][\text{O}_2]}$
- C $\frac{[2\text{H}_2\text{O}]^2}{[2\text{H}_2]^2[\text{O}_2]}$
- D $\frac{1}{[\text{H}_2]^2[\text{O}_2]}$

- 4 Consider the reaction represented by the equation



For the system at chemical equilibrium, which of the following explains what happens after the addition of oxygen gas (assume constant temperature)?

- A The amount of $\text{NO}(\text{g})$ increases and the value for K increases.
 - B The amount of $\text{NO}(\text{g})$ decreases and the value for K increases.
 - C The amount of $\text{NO}(\text{g})$ decreases and the value for K stays the same.
 - D The amount of $\text{NO}(\text{g})$ increases and the value for K stays the same.
- 5 Write a K_{sp} expression for the process of $\text{PbI}_2(\text{s})$ dissolving in water.
- A $[\text{Pb}^{2+}][\text{I}^-]$
 - B $[\text{Pb}^{2+}][\text{I}^-]^2$
 - C $\frac{[\text{Pb}^{2+}][\text{I}^-]^2}{[\text{PbI}_2]}$
 - D $\frac{[\text{Pb}^{2+}][\text{I}^-]}{[\text{PbI}_2]}$
- 6 At a given temperature, the solubility of CaF_2 is $2.15 \times 10^{-4} \text{ M}$. Calculate the K_{sp} value for CaF_2 at this temperature.
- A 4.64×10^{-8}
 - B 3.98×10^{-11}
 - C 9.94×10^{-12}
 - D 2.15×10^{-4}
- 7 At a certain temperature the equilibrium constant (K) for the following reaction is 9.00×10^2 .



It is found that $[\text{H}_2] = 0.010 \text{ M}$ and $[\text{H}_2\text{S}] = 0.15 \text{ M}$. Calculate $[\text{S}_2]$.

Oxidation–Reduction Reactions and Electrochemistry

LOOKING AHEAD

18.1 Electron Transfer Reactions

- A. Oxidation–Reduction Reactions
- B. Oxidation States

18.2 Balancing Oxidation–Reduction Reactions

- A. Oxidation–Reduction Reactions Between Nonmetals
- B. Balancing Oxidation–Reduction Reactions by the Half-Reaction Method

18.3 Electrochemistry and Its Applications

- A. Electrochemistry: An Introduction
- B. Batteries
 - Lead Storage Battery
 - Dry Cell Batteries
- C. Corrosion
- D. Electrolysis



A nickel electroplated with copper

IN YOUR LIFE

What do a forest fire, rusting steel, combustion in an automobile engine, and the metabolism of food in a human body have in common? All of these important processes involve oxidation–reduction reactions. In fact, virtually all of the processes that provide energy to heat buildings, power vehicles, and allow people to work and play depend on oxidation–reduction reactions. And every time you start your car, turn on your calculator, look at your digital watch, or listen to a radio at the beach, you are depending on an oxidation–reduction reaction to power the battery in each of these devices. In addition, because “pollution-free” vehicles have been promoted in California, hybrid cars are more common on U.S. roads. This will lead to increased reliance of our society on batteries and will spur the search for new, more efficient batteries.



The power generated by an alkaline AA battery and a mercury battery results from oxidation–reduction reactions.

WHAT DO YOU KNOW?

Prereading Questions

1. What is meant by the term *oxidation*? What is meant by the term *reduction*?
2. When forming compounds, do metals tend to gain or lose electrons? What about nonmetals?
3. What are the general trends for electronegativity across rows and down columns of the periodic table?

SECTION 18.1

Electron Transfer Reactions

Key Terms

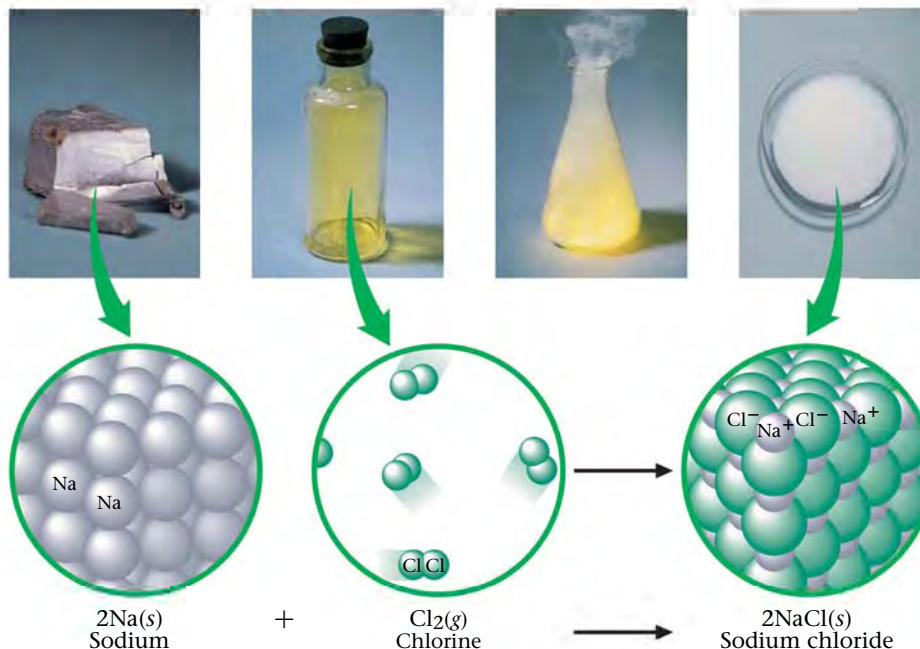
- Oxidation–reduction reactions (redox reactions)
- Oxidation
- Reduction
- Oxidation states

Objectives

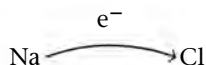
- To learn about metal–nonmetal oxidation–reduction reactions
- To learn to assign oxidation states

A. Oxidation–Reduction Reactions

In Chapter 8 we discussed the chemical reactions between metals and nonmetals. For example, sodium chloride is formed by the reaction of elemental sodium and chlorine.



Because elemental sodium and chlorine contain uncharged atoms and because sodium chloride is known to contain Na^+ and Cl^- ions, this reaction must involve a transfer of electrons from sodium atoms to chlorine atoms.



Reactions like this one, in which one or more electrons are transferred, are called **oxidation–reduction reactions**, or **redox reactions**. **Oxidation** is defined as a *loss of electrons*. **Reduction** is defined as a *gain of electrons*. In the reaction of elemental sodium and chlorine, each sodium atom loses one electron, forming a 1+ ion. Therefore, sodium is oxidized. Each chlorine atom gains one electron, forming a negative chloride ion, and is thus reduced. Whenever a metal reacts with a nonmetal to form an ionic compound, electrons are transferred from the metal to the nonmetal. So these reactions are always oxidation–reduction reactions where the metal is oxidized (loses electrons) and the nonmetal is reduced (gains electrons).

Oxidation–reduction reaction (redox reaction)

A chemical reaction involving the transfer of electrons

Oxidation

A loss of electrons; an increase in oxidation state

Reduction

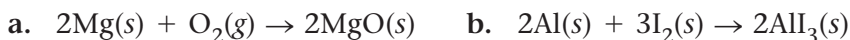
A gain of electrons; a decrease in oxidation state

Active Reading Question

Which element is oxidized and which element is reduced when a sodium metal reacts with chlorine gas?

EXAMPLE 18.1**Identifying Oxidation and Reduction in a Reaction**

In the following reactions, identify which element is oxidized and which element is reduced.

**Solution**

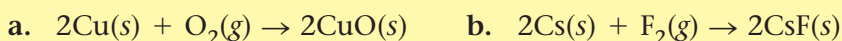
- a. We have learned that Group 2 metals form $2+$ cations and that Group 6 nonmetals form $2-$ anions, so we can predict that magnesium oxide contains Mg^{2+} and O^{2-} ions. This means that in the reaction given, each Mg loses two electrons to form Mg^{2+} and so is oxidized. Also each O gains two electrons to form O^{2-} and so is reduced.
- b. Aluminum iodide contains the Al^{3+} and I^- ions. Thus aluminum atoms lose electrons (are oxidized). Iodine atoms gain electrons (are reduced).



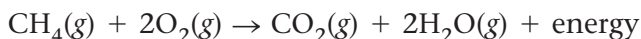
Magnesium burns in air to give a bright, white flame.

Practice Problem • Exercise 18.1

For the following reactions, identify the element oxidized and the element reduced.



Reactions Between Nonmetals Although we can identify reactions between metals and nonmetals as redox reactions, it is more difficult to decide whether a given reaction between nonmetals is a redox reaction. In fact, many of the most significant redox reactions involve only nonmetals. For example, combustion reactions such as methane burning in oxygen are oxidation–reduction reactions.



Even though none of the reactants or products in this reaction is ionic, the reaction does involve a transfer of electrons from carbon to oxygen. To explain this, we must introduce the concept of oxidation states.

B. Oxidation States

The concept of **oxidation states** (sometimes called *oxidation numbers*) lets us keep track of electrons in oxidation–reduction reactions by assigning charges to the various atoms in a compound. Sometimes these charges are quite apparent. For example, in a binary ionic compound the ions have easily identified charges: in sodium chloride, sodium is $+1$ and chlorine is -1 ; in magnesium oxide, magnesium is $+2$ and oxygen is -2 ; and so on.

In such binary ionic compounds the oxidation states are simply the charges of the ions.

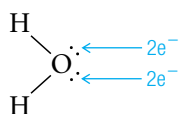
Ion	Oxidation State
Na^+	$+1$
Cl^-	-1
Mg^{2+}	$+2$
O^{2-}	-2

Oxidation state

A concept that allows us to keep track of electrons in oxidation–reduction reactions.

In an uncombined element, all of the atoms are uncharged (neutral). For example, sodium metal contains neutral sodium atoms, and chlorine gas is made up of Cl_2 molecules, each of which contains two neutral chlorine atoms. Therefore, an atom in a pure element has no charge and is assigned an oxidation state of zero.

Oxidation States in Covalent Compounds In a covalent compound such as water, although no ions are actually present, chemists find it useful to assign imaginary charges to the elements in the compound. The oxidation states of the elements in these compounds are equal to the imaginary charges we determine by assuming that the most electronegative atom in a bond controls or possesses *both* of the shared electrons (see Chapter 12). For example, in the O—H bonds in water, it is assumed for purposes of assigning oxidation states that the much more electronegative oxygen atom controls both of the shared electrons in each bond. This gives the oxygen eight valence electrons.



In effect we say that each hydrogen has lost its single electron to the oxygen. This gives each hydrogen an oxidation state of +1 and the oxygen an oxidation state of -2 (the oxygen atom has formally gained two electrons). In virtually all covalent compounds, oxygen is assigned an oxidation state of -2 and hydrogen is assigned an oxidation state of +1.

Because fluorine is so electronegative, it is always assumed to control any shared electrons. So fluorine is always assumed to have a complete octet of electrons and is assigned an oxidation state of -1 . That is, for purposes of assigning oxidation states, fluorine is always imagined to be F^- in its covalent compounds.

The most electronegative elements are F, O, N, and Cl. In general, we give each of these elements an oxidation state equal to its charge as an anion: fluorine is -1 , chlorine is -1 , oxygen is -2 , and nitrogen is -3 . When two of these elements are found in the same compound, we assign them in order of electronegativity, starting with the element that has the largest electronegativity.



For example, in the compound NO_2 , because oxygen has a greater electronegativity than nitrogen, we assign each oxygen an oxidation state of -2 . This gives a total “charge” of -4 (2×-2) on the two oxygen atoms. Because the NO_2 molecule has zero overall charge, the N must be $+4$ to exactly balance the -4 on the oxygens. In NO_2 , then, the oxidation state of *each* oxygen is -2 and the oxidation state of the nitrogen is $+4$.

Active Reading Question

Provide an example when sulfur has a negative oxidation state in a compound, and an example when sulfur has a positive oxidation state in a compound.

The rules for assigning oxidation states are given below and are illustrated in **Table 18.1**. Application of these rules allows us to assign oxidation states in most compounds. The principles are illustrated by Example 18.2.

Rules for Assigning Oxidation States

1. The oxidation state of an atom in an uncombined element is 0.
2. The oxidation state of a monatomic ion is the same as its charge.
3. Oxygen is assigned an oxidation state of -2 in most of its covalent compounds. Important exception: peroxides (compounds containing the O_2^{2-} group), in which each oxygen is assigned an oxidation state of -1 .
4. In its covalent compounds with nonmetals, hydrogen is assigned an oxidation state of $+1$.
5. In binary compounds, the element with the greater electronegativity is assigned a negative oxidation state equal to its charge as an anion in its ionic compounds.
6. For an electrically neutral compound, the sum of the oxidation states must be zero.
7. For an ionic species, the sum of the oxidation states must equal the overall charge.

Table 18.1

Examples of Oxidation States

Substance	Oxidation States	Comments
sodium, metal, Na	Na, 0	rule 1
magnesium sulfide, MgS	Mg, +2 S, -2	rule 2 rule 2
carbon monoxide, CO	C, +2 O, -2	rule 3
sulfur dioxide, SO_2	S, +4 O, -2	rule 3
hydrogen peroxide, H_2O_2	H, +1 O, -1	rule 3 (exception)
ammonia, NH_3	H, +1 N, -3	rule 4 rule 5
hydrogen sulfide, H_2S	H, +1 S, -2	rule 4 rule 5
sodium carbonate, Na_2CO_3	Na, +1 O, -2 C, +4	rule 2 rule 3 For CO_3^{2-} , the sum of the oxidation states is $+4 + 3(-2) = -2$. rule 7

EXAMPLE 18.2

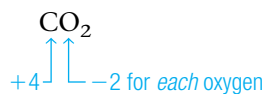
Assigning Oxidation States

Assign oxidation states to all atoms in the following molecules or ions.

- CO_2
- SF_6
- NO_3^-

Solution

- a. Rule 3 is most important here: oxygen is assigned an oxidation state of -2 . We determine the oxidation state for carbon by recognizing that because CO_2 has no charge, the sum of the oxidation states for oxygen and carbon must be 0 (rule 6). Each oxygen is -2 and there are two oxygen atoms, so the carbon atom must be assigned an oxidation state of $+4$.



Check: $+4 + 2(-2) = 0$

- b. Because fluorine has the greater electronegativity, we assign its oxidation state first. Its charge as an anion is always -1 , so we assign -1 as the oxidation state of each fluorine atom (rule 5). The sulfur must then be assigned an oxidation state of $+6$ to balance the total of -6 from the six fluorine atoms (rule 7).



Check: $+6 + 6(-1) = 0$

- c. Oxygen has a greater electronegativity than nitrogen, so we assign its oxidation state of -2 first (rule 5). Because the overall charge on NO_3^- is -1 and because the sum of the oxidation states of the three oxygens is -6 , the nitrogen must have an oxidation state of $+5$.



Check: $+5 + 3(-2) = -1$

This is correct; NO_3^- has a -1 charge.

Practice Problem • Exercise 18.2

Assign oxidation states to all atoms in the following molecules or ions.

- SO_3
- SO_4^{2-}
- N_2O_5
- PF_3
- C_2H_6

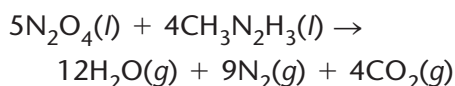
CELEBRITY CHEMICAL

Hydrazine (N₂H₄)

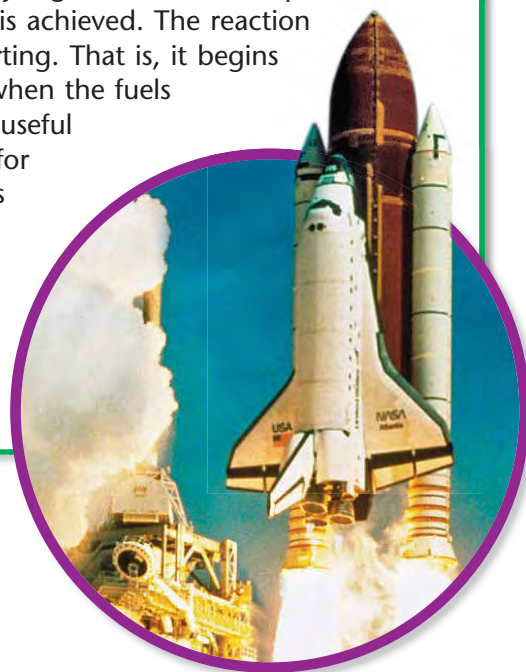
Hydrazine, N₂H₄, is a colorless liquid with an ammonia-like odor that freezes at 2 °C and boils at 114 °C. This powerful reducing agent reacts with oxygen in a highly exothermic reaction,



which produces 622 kJ of energy per mole of N₂H₄. Substituted hydrazines, where one or more of the hydrogen atoms are replaced by other groups, are useful as rocket fuels. For example, methylhydrazine (CH₃N₂H₃) is used with the oxidizing agent N₂O₄ (dinitrogen tetroxide) to power the U.S. space shuttle orbiter. The reaction is



Because of the large number of gaseous molecules produced and the highly exothermic nature of this reaction, a very high value of thrust per gram of fuels is achieved. The reaction is also self-starting. That is, it begins immediately when the fuels are mixed—a useful characteristic for rocket engines that must be started and stopped frequently.



SECTION 18.1

REVIEW QUESTIONS

- In a reaction between a metal and a nonmetal which substance tends to be oxidized? Reduced? Why?
- How do oxidation states help us to decide whether an oxidation–reduction reaction has occurred?
- Assign oxidation states to all atoms in each of the following. Explain the order in which you determine the oxidation states for each element, if applicable.
 - O₂
 - HSO₄[−]
 - Na₂HPO₄
 - CrCl₃
- For each of the following oxidation–reduction reactions of metals and nonmetals, identify the element that is oxidized and the element that is being reduced.
 - 4Fe(s) + 3O₂(g) → 2Fe₂O₃(s)
 - Zn(s) + 2AgNO₃(aq) → Zn(NO₃)₂(aq) + 2Ag(s)
 - 2K(s) + Cl₂(g) → 2KCl(s)
 - 2Ca(s) + O₂(g) → 2CaO(s)



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SECTION 18.2

Balancing Oxidation–Reduction Reactions

Key Terms

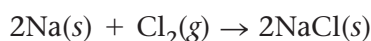
- Oxidizing agent (electron acceptor)
- Reducing agent (electron donor)
- Half-reactions

Objectives

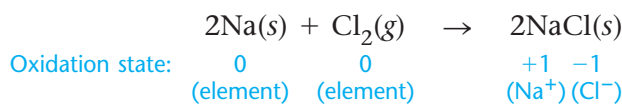
- To understand oxidation and reduction in terms of oxidation states
- To learn to identify oxidizing and reducing agents
- To learn to balance oxidation–reduction equations using half reactions

A. Oxidation–Reduction Reactions Between Nonmetals

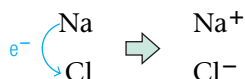
We have seen that oxidation–reduction reactions are characterized by a transfer of electrons. In some cases, the transfer literally occurs to form ions, such as in the reaction



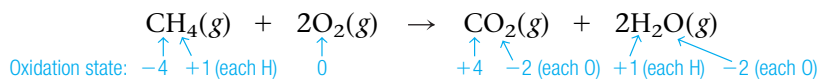
We can use oxidation states to verify that electron transfer has occurred.



Thus in this reaction, we represent the electron transfer as follows:



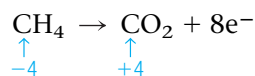
In other cases the electron transfer occurs in a different sense, such as in the combustion of methane (the oxidation state for each atom is given below each reactant and product).



Note that the oxidation state of oxygen in O_2 is 0 because the oxygen is in elemental form. In this reaction there are no ionic compounds, but we can still describe the process in terms of the transfer of electrons. Note that carbon undergoes a change in oxidation state from -4 in CH_4 to $+4$ in CO_2 . Such a change can be accounted for by a loss of eight electrons:



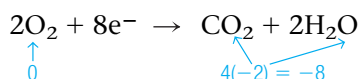
or, in equation form,



On the other hand, each oxygen changes from an oxidation state of 0 in O_2 to -2 in H_2O and CO_2 , signifying a gain of two electrons per atom. Four oxygen atoms are involved, so this is a gain of eight electrons:



or, in equation form,



Information



$$\text{H} = +1 \text{ (rule 4)}$$

$$\text{C} = -4$$

$$-4 + 4(+1) = 0$$



$$\text{O} = -2 \text{ (rule 3)}$$

$$\text{C} = +4$$

$$+4 + (2(-2)) = 0$$



$$\text{H} = +1 \text{ (rule 4)}$$

$$\text{O} = -2 \text{ (rule 3)}$$

$$2(+1) + 1(-2) = 0$$

CHEMISTRY IN YOUR WORLD

Consumer Connections

U.S. Coins

We all use coins every day without thinking much about them. It turns out that some interesting chemistry is associated with the making of the coins used in the United States. For example, the “copper” penny is really mostly zinc with a coating of copper. The “nickel” coin contains much more copper than nickel—it’s made of a copper/nickel alloy that is 75% copper and 25% nickel. All of the other coins used in the United States are clad—they have a pure copper core that is surrounded (sandwiched) by an alloy. The dime, the quarter, the half-dollar, and the Susan B. Anthony dollar all have a copper core sandwiched by a copper/nickel alloy. The Sacagawea dollar has a copper core surrounded by a special alloy containing copper, zinc, nickel, and manganese that gives it the same electromagnetic properties as the Anthony dollar so it will work in current vending machines.

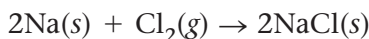
The composition and physical properties of the coins are listed in the accompanying table.



Coin	Composition	Mass (g)	Diameter (mm)
Penny	Copper-plated zinc (97.5% Zn, 2.5% Cu)	2.5	19.1
Nickel	Copper/nickel alloy (75% Cu, 25% Zn)	5.0	21.2
Dime	Copper-nickel clad (91.7% Cu, 8.3% Ni)	2.3	17.9
Quarter	Copper-nickel clad (91.7% Cu, 8.3% Ni)	5.7	24.3
Half-dollar	Copper-nickel clad (91.7% Cu, 8.3% Ni)	11.3	30.6
Susan B. Anthony dollar	Copper-nickel clad (87.5% Cu, 12.5% Ni)	8.1	26.5
Sacagawea dollar	Manganese-brass clad (88.5% Cu, 6.0% Zn, 3.5% Mn, 2.0% Ni)	8.1	26.5

Note that eight electrons are required because four oxygen atoms are going from an oxidation state of 0 to -2 . Each oxygen requires two electrons. No change occurs in the oxidation state of hydrogen, and it is not involved in the electron transfer process.

With this background, we can now define *oxidation* and *reduction* in terms of oxidation states. **Oxidation** is an *increase* in oxidation state (a loss of electrons). **Reduction** is a *decrease* in oxidation state (a gain of electrons). Thus in the reaction



sodium is oxidized and chlorine is reduced. Cl_2 is called the **oxidizing agent (electron acceptor)** and Na is called the **reducing agent (electron donor)**. We can also define the *oxidizing agent* as the reactant containing the element that is reduced (gains electrons). The *reducing agent* can be defined similarly as the reactant containing the element that is oxidized (loses electrons).

Concerning the reaction



we can say the following:

Oxidizing agent

A reactant that accepts electrons from another reactant; the oxidizing agent is reduced since it gains one or more electrons.

Reducing agent

A reactant that donates electrons to another reactant; the reducing agent is oxidized since it donates one or more electrons.

Information

In a redox reaction, an oxidizing agent is reduced (gains electrons) and a reducing agent is oxidized (loses electrons).

- Carbon is oxidized because there is an increase in its oxidation state (carbon has apparently lost electrons).
- The reactant CH_4 contains the carbon that is oxidized, so CH_4 is the reducing agent. It is the reactant that furnishes the electrons (those lost by carbon).
- Oxygen is reduced because there has been a decrease in its oxidation state (oxygen has apparently gained electrons).
- The reactant that contains the oxygen atoms is O_2 , so O_2 is the oxidizing agent. That is, O_2 accepts the electrons.

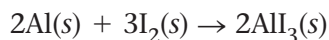
Note that when the oxidizing or reducing agent is named, the *whole compound* is specified, not just the element that undergoes the change in oxidation state.

EXAMPLE 18.3

Identifying Oxidizing and Reducing Agents, I



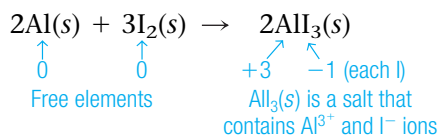
When powdered aluminum metal is mixed with pulverized iodine crystals and a drop of water is added, the resulting reaction produces a great deal of energy. The mixture bursts into flames, and a purple smoke of I_2 vapor is produced from the excess iodine. The equation for the reaction is



For this reaction, identify the atoms that are oxidized and those that are reduced, and specify the oxidizing and reducing agents.

Solution

The first step is to assign oxidation states.

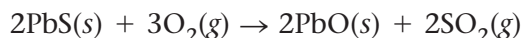


Because each aluminum atom changes its oxidation state from 0 to +3 (an increase in oxidation state), aluminum is *oxidized* (loses electrons). On the other hand, the oxidation state of each iodine atom decreases from 0 to -1 , and iodine is *reduced* (gains electrons). Because Al furnishes electrons for the reduction of iodine, it is the *reducing agent*. I_2 is the *oxidizing agent* (the reactant that accepts the electrons).

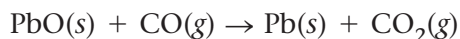
EXAMPLE 18.4

Identifying Oxidizing and Reducing Agents, II

Metallurgy, the process of producing a metal from its ore, always involves oxidation–reduction reactions. In the metallurgy of galena (PbS), the principal lead-containing ore, the first step is the conversion of lead sulfide to its oxide (a process called *roasting*).



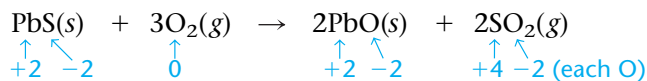
The oxide is then treated with carbon monoxide to produce the free metal.



For each reaction, identify the atoms that are oxidized and those that are reduced, and specify the oxidizing and reducing agents.

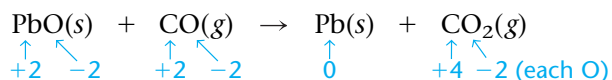
Solution

For the first reaction, we can assign the following oxidation states:



The oxidation state for the sulfur atom increases from -2 to $+4$, so sulfur is oxidized (loses electrons). The oxidation state for each oxygen atom decreases from 0 to -2 . Oxygen is reduced (gains electrons). The oxidizing agent (electron acceptor) is O_2 , and the reducing agent (electron donor) is PbS .

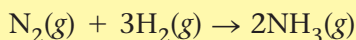
For the second reaction, we have



Lead is reduced (gains electrons; its oxidation state decreases from $+2$ to 0), and carbon is oxidized (loses electrons; its oxidation state increases from $+2$ to $+4$). PbO is the oxidizing agent (electron acceptor), and CO is the reducing agent (electron donor).

Practice Problem • Exercise 18.4

Ammonia, NH_3 , which is widely used as a fertilizer, is prepared by reaction of the elements:



Is this an oxidation–reduction reaction? If so, specify the oxidizing agent and the reducing agent.

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Real Gold

Recently, a woman wrote to Annie’s Mailbox complaining that the necklace her boyfriend had given her turned her neck green. Her concern was whether the necklace was “real gold.” This leads to the question: What is “real gold”? To chemists, real gold is pure gold, called 24-carat gold by jewelers. Thus, when you see 24K (for “carat,” which is also spelled “karat”) on a piece of gold, it means that it is 100% gold. However, the woman’s necklace was surely not 24K gold, because pure gold is too soft for making jewelry. Jewelry is usually made of 18K gold (which is $18/24 = 75\%$ gold by mass) or 14K gold ($14/24 = 58\%$ gold). The carat system was invented by the British in about 1300 to provide a standard for the use of gold as currency.

In the United States, the lowest carat designation for gold is 10K. Because a 0.5-carat error is allowed, 10K gold is actually 9.5K or 39.6% gold by mass. So if 9.5K gold is only about 40% gold, what makes

up the rest of the “gold” jewelry? The metals most commonly used to form gold alloys are copper, silver, zinc, and nickel, depending on the specific use and color desired. For example, “yellow gold” typically contains Au, Cu, Ag, and Zn, whereas “white gold” contains Au, Cu, Ni, and Zn.

Now back to the woman’s original question. Was her necklace “real gold”? Not likely. The necklace was probably mostly copper with a thin layer of gold plated on the surface. The woman’s necklace turned green because copper “leaked” through cracks in the gold plating. Alloys of 14K gold do not behave in this way. She was right to worry about her friend’s honesty.

Gold jewelry in a store in Thailand



B. Balancing Oxidation–Reduction Reactions by the Half-Reaction Method

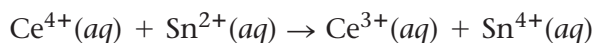
Many oxidation–reduction reactions can be balanced readily by trial and error. That is, we use the procedure described in Chapter 7 to find a set of coefficients that give the same number of each type of atom on both sides of the equation.

However, the oxidation–reduction reactions that occur in aqueous solution are often so complicated that it becomes very tedious to balance them by trial and error. In this section we will develop a systematic approach for balancing the equations for these reactions.

To balance the equations for oxidation–reduction reactions that occur in aqueous solution, we separate the reaction into two half-reactions. **Half-reactions** are equations that have electrons as reactants or products. One half-reaction represents a reduction process and the other half-reaction represents an oxidation process.

In a reduction half-reaction, electrons are shown on the reactant side (electrons are gained by a reactant in the equation). In an oxidation half-reaction, the electrons are shown on the product side (electrons are lost by a reactant in the equation).

For example, consider the unbalanced equation for the oxidation–reduction reaction between the cerium(IV) ion and the tin(II) ion.



This reaction can be separated into a half-reaction involving the substance being *reduced*:



and a half-reaction involving the substance being *oxidized*:



Notice that Ce^{4+} must gain one electron to become Ce^{3+} , so one electron is shown as a reactant along with Ce^{4+} in this half-reaction. On the other hand, for Sn^{2+} to become Sn^{4+} , it must lose two electrons. This means that two electrons must be shown as products in this half-reaction.

Information

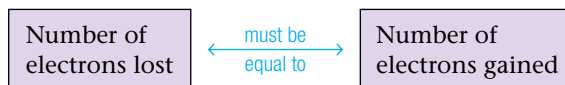
Ce^{4+} gains 1e^{-} to form Ce^{3+} and is thus reduced.

Sn^{2+} loses 2e^{-} to form Sn^{4+} and is thus oxidized.

Active Reading Question

In an oxidation half-reaction, is an electron a reactant or product?

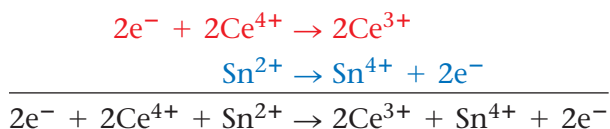
The key principle in balancing oxidation–reduction reactions is that the number of electrons lost (from the reactant that is oxidized) must equal the number of electrons gained (from the reactant that is reduced).



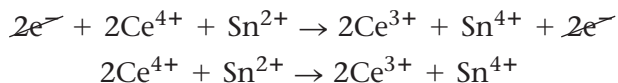
In the half-reaction shown above, one electron is gained by each Ce^{4+} while two electrons are lost by each Sn^{2+} . We must equalize the number of electrons gained and lost. To do this, we first multiply the reduction half-reaction by 2.



Then we add this half-reaction to the oxidation half-reaction.



Finally, we cancel the $2e^-$ on each side to give the overall balanced equation



Let's Review

Balancing Oxidation–Reduction Reactions In Aqueous Solution

1. Separate the reaction into an oxidation half-reaction and a reduction half-reaction.
2. Balance the half-reactions separately.
3. Equalize the number of electrons gained and lost.
4. Add the half-reactions together and cancel electrons to give the overall balanced equation.

It turns out that most oxidation–reduction reactions occur in solutions that are distinctly basic or distinctly acidic. We will cover only the acidic case in this text, because it is the most common. The detailed procedure for balancing the equations for oxidation–reduction reactions that occur in acidic solution is given below, and Example 18.5 illustrates the use of these steps.

The Half-Reaction Method for Balancing Equations for Oxidation–Reduction Reactions Occurring in Acidic Solution

- Step 1** Identify and write the equations for the oxidation and reduction half-reactions.
- Step 2** For each half-reaction:
- a. Balance all of the elements except hydrogen and oxygen.
 - b. Balance oxygen using H_2O .
 - c. Balance hydrogen using H^+ .
 - d. Balance the charge using electrons.
- Step 3** If necessary, multiply one or both balanced half-reactions by an integer to equalize the number of electrons transferred in the two half-reactions.
- Step 4** Add the half-reactions, and cancel identical species that appear on both sides.
- Step 5** Check to be sure the elements and charges balance.

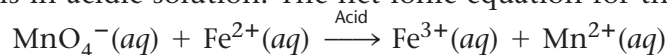
EXAMPLE 18.5

Balancing Oxidation–Reduction Reactions Using the Half-Reaction Method, I



A solution containing MnO_4^- ions (top) and a solution containing Fe^{2+} ions (bottom).

Balance the equation for the reaction between permanganate and iron(II) ions in acidic solution. The net ionic equation for this reaction is

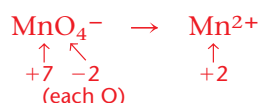


This reaction is used to analyze iron ore for its iron content.

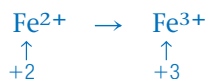
Solution

Step 1 Identify and write equations for the half-reactions.

The oxidation states for the half-reaction involving the permanganate ion show that manganese is reduced.



Because manganese changes from an oxidation state of +7 to +2, it is reduced. So this is the *reduction half-reaction*. It will have electrons as reactants, although we will not write them yet. The other half-reaction involves the oxidation of iron(II) to the iron(III) ion and is the *oxidation half-reaction*.



This reaction will have electrons as products, although we will not write them yet.

Step 2 Balance each half-reaction.

For the reduction reaction, we have



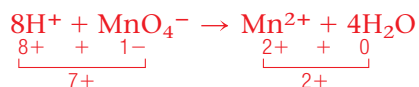
- The manganese is already balanced.
- We balance oxygen by adding $4\text{H}_2\text{O}$ to the right side of the equation.



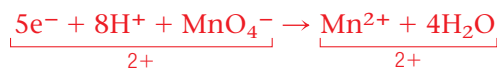
- Next we balance hydrogen by adding 8H^+ to the left side.



- All of the elements have been balanced, but we need to balance the charge using electrons. At this point we have the following charges for reactants and products in the reduction half-reaction.



We can equalize the charges by adding five electrons to the left side.



Information

The H^+ comes from the acidic solution in which the reaction is taking place.

Information

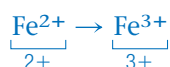
Always add electrons to the side of the half-reaction with excess positive charge.

Both the *elements* and the *charges* are now balanced, so this represents the balanced reduction half-reaction. The fact that five electrons appear on the reactant side of the equation makes sense, because five electrons are required to reduce MnO_4^- (in which Mn has an oxidation state of +7) to Mn^{2+} (in which Mn has an oxidation state of +2).

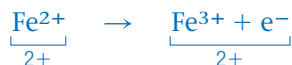
For the oxidation reaction,



the elements are balanced, so all we have to do is balance the charge.



One electron is needed on the right side to give a net 2+ charge on both sides.

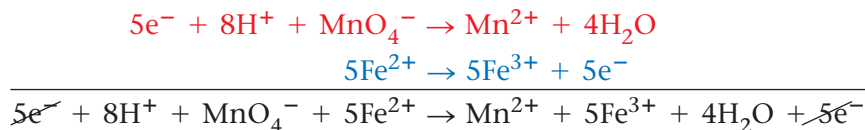


Step 3 Equalize the number of electrons transferred in the two half-reactions.

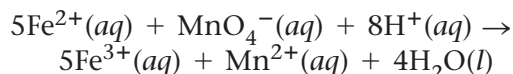
Because the reduction half-reaction involves a transfer of five electrons and the oxidation half-reaction involves a transfer of only one electron, the oxidation half-reaction must be multiplied by 5.



Step 4 Add the half-reactions and cancel identical species.

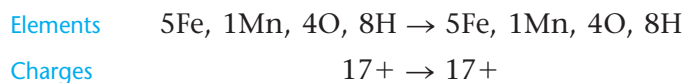


Note that the electrons cancel (as they must) to give the final balanced equation



Note that we show the physical states of the reactants and products—(aq) and (l) in this case—only in the final balanced equation.

Step 5 Check to be sure that elements and charges balance.



The equation is balanced.

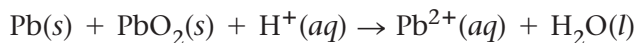
Information

The number of electrons gained in the reduction half-reaction must equal the number of electrons lost in the oxidation half-reaction.

EXAMPLE 18.6

Balancing Oxidation–Reduction Reactions Using the Half-Reaction Method, II

When an automobile engine is started, it uses the energy supplied by a lead storage battery. This battery uses an oxidation–reduction reaction between elemental lead (lead metal) and lead(IV) oxide to provide the power to start the engine. The unbalanced equation for a simplified version of the reaction is



Balance this equation using the half-reaction method.

Solution

Information

Because Pb^{2+} is the only lead-containing product, it must be the product in both half-reactions.

Step 1 First we identify and write the two half-reactions. One half-reaction must be $\text{Pb} \rightarrow \text{Pb}^{2+}$ and the other is $\text{PbO}_2 \rightarrow \text{Pb}^{2+}$.

The first reaction involves the oxidation of Pb to Pb^{2+} . The second reaction involves the reduction of Pb^{4+} (in PbO_2) to Pb^{2+} .

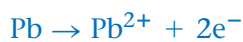
Step 2 Now we will balance each half-reaction separately.

The oxidation half-reaction



a–c. All the elements are balanced.

d. The charge on the left is zero and that on the right is 2+, so we must add $2e^-$ to the right to give zero overall charge.



This half-reaction is balanced.

The reduction half-reaction



a. All elements are balanced except O.

b. The left side has two oxygen atoms and the right side has none, so we add $2\text{H}_2\text{O}$ to the right side.



c. Now we balance hydrogen by adding 4H^+ to the left.

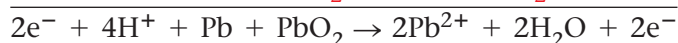


d. Because the left side has a 4+ overall charge and the right side has a 2+ charge, we must add $2e^-$ to the left side.

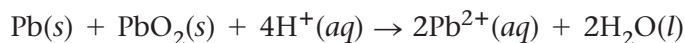


The half-reaction is balanced.

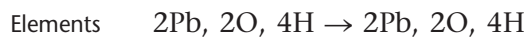
Step 3 Because each half-reaction involves $2e^-$, we can simply add the half-reactions as they are.

Step 4

Canceling electrons gives the balanced overall equation



where the appropriate states are also indicated.

Step 5 Both the elements and the charges balance.

The equation is correctly balanced.



Copper metal reacting with concentrated nitric acid. The solution is colored by the presence of Cu^{2+} ions. The brown gas is NO_2 .

Practice Problem • Exercise 18.6

Copper metal reacts with dilute nitric acid, $\text{HNO}_3(aq)$, to give aqueous copper(II) nitrate, water, and nitrogen monoxide gas as products. Write and balance the equation for this reaction.

SECTION 18.2

REVIEW QUESTIONS

- Explain how you can determine the oxidizing agent and reducing agent in an oxidation–reduction reaction.
- For each of the following oxidation–reduction reactions of metals and nonmetals, identify the oxidizing agent and reducing agent.
 - $\text{Cu}(s) + 2\text{AgNO}_3(aq) \rightarrow \text{Cu}(\text{NO}_3)_2(aq) + 2\text{Ag}(s)$
 - $4\text{Al}(s) + 3\text{O}_2(g) \rightarrow 2\text{Al}_2\text{O}_3(s)$
 - $\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g)$
- For each of the following reactions, is oxygen gas an oxidizing agent or a reducing agent? Does that mean oxygen is oxidized or reduced?
 - $4\text{Al}(s) + 3\text{O}_2(g) \rightarrow 2\text{Al}_2\text{O}_3(s)$
 - $\text{C}(s) + \text{O}_2(g) \rightarrow \text{CO}_2(g)$
- Write two definitions of oxidation. In the first definition, use the phrase *oxidation state*. In the second definition, use the term *electrons*.
- Write two definitions of reduction. In the first definition, use the phrase *oxidation state*. In the second definition, use the term *electrons*.
- Write the balanced half-reactions corresponding to each of the following:
 - The reduction of $\text{Br}_2(l)$ to $\text{Br}^{-}(aq)$.
 - The oxidation of $\text{Zn}(s)$ to $\text{Zn}^{2+}(aq)$.
- Each of the oxidation–reduction reactions place in acidic solution. Balance each reaction by using the half-reaction method.
 - $\text{NO}_3^{-}(aq) + \text{Br}^{-}(aq) \rightarrow \text{NO}(g) + \text{Br}_2(l)$
 - $\text{Ni}(s) + \text{NO}_3^{-}(aq) \rightarrow \text{Ni}^{2+}(aq) + \text{NO}_2(g)$
 - $\text{ClO}_4^{-}(aq) + \text{Cl}^{-}(aq) \rightarrow \text{ClO}_3^{-}(aq) + \text{Cl}_2(g)$

SECTION 18.3

Electrochemistry and Its Applications

Key Terms

- Electrochemistry
- Electrochemical battery (galvanic cell)
- Anode
- Cathode
- Electrolysis
- Lead storage battery
- Potential
- Dry cell batteries
- Corrosion
- Cathodic protection

Objectives

- To understand the concept of electrochemistry
- To learn to identify the components of an electrochemical (galvanic) cell
- To learn about commonly used batteries
- To understand corrosion and ways of preventing it
- To understand electrolysis
- To learn about the commercial preparation of aluminum

A. Electrochemistry: An Introduction

Our lives would be very different without batteries. We would have to crank the engines in our cars by hand, wind our watches, and buy very long extension cords if we wanted to listen to an MP3 player on a picnic. Indeed, our society sometimes seems to run on batteries. In this section and the next, we will find out how these important devices produce electrical energy.

A battery uses the energy from an oxidation–reduction reaction to produce an electric current. This is an important illustration of **electrochemistry**, *the study of the interchange of chemical and electrical energy*.

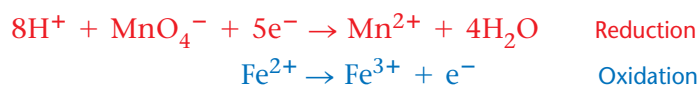
Electrochemistry

The study of the interchange of chemical and electrical energy

Electrochemistry involves two types of processes:

1. The production of an electric current from a chemical (oxidation–reduction) reaction
2. The use of an electric current to produce a chemical change

To understand how a redox reaction can be used to generate a current, let's reconsider the aqueous reaction between MnO_4^- and Fe^{2+} that we worked with in Example 18.5. We can break this redox reaction into the following half-reactions:



When the reaction between MnO_4^- and Fe^{2+} occurs in solution, electrons are transferred directly as the reactants collide. No useful work is obtained from the chemical energy involved in the reaction. How can we harness this energy? The key is to *separate the oxidizing agent (electron acceptor) from the reducing agent (electron donor)*, thus requiring the electron transfer to occur through a wire. That is, to get from the reducing agent to the oxidizing agent, the electrons must travel through a wire. The current produced in the wire by this electron flow can be directed through a device, such as an electric motor, to do useful work.

DID YOU KNOW

The energy involved in a chemical reaction is customarily not shown in the balanced equation. In the reaction of MnO_4^- with Fe^{2+} , energy is released that can be used to do useful work.



Active Reading Question

What is the key to getting useful energy from an oxidation–reduction reaction?

For example, consider the system illustrated in **Figure 18.1**. If our reasoning has been correct, electrons should flow through the wire from Fe^{2+} to MnO_4^- . However, when we construct the apparatus as shown, no flow of electrons occurs. Why? The problem is that if the electrons flowed from the right to the left compartment, the left compartment would become negatively charged and the right compartment would experience a build-up of positive charge. Creating a charge separation of this type would require large amounts of energy. Therefore, electron flow does not occur under these conditions.

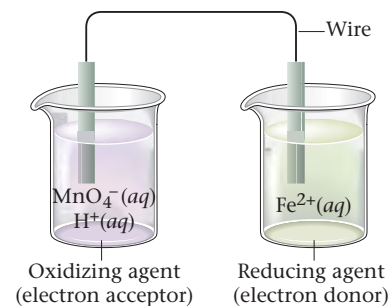
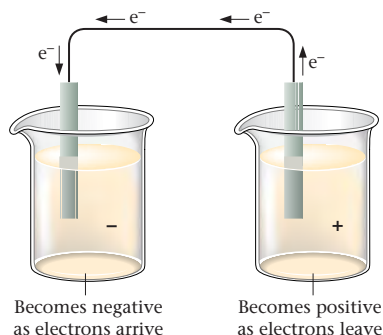
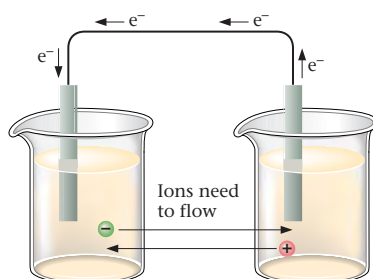


Figure 18.1

Schematic of a method for separating the oxidizing and reducing agents in a redox reaction. (The solutions also contain other ions to balance the charge.) This cell is incomplete at this point.

We can, however, solve this problem very simply. The solutions must be connected (without allowing them to mix extensively) so that *ions* can also flow to keep the net charge in each compartment zero. This can be accomplished by using a salt bridge (a U-shaped tube filled with a strong electrolyte) or a porous disk in a tube connecting the two solutions (see **Figure 18.2**). Either of these devices allows ion flow but prevents extensive mixing



of the solutions. When we make a provision for ion flow, the circuit is complete. Electrons flow through the wire from reducing agent to oxidizing agent, and ions in the two aqueous solutions flow from one compartment to the other to keep the net charge zero.



An electric car being charged

Active Reading Question

What is the purpose of a salt bridge in an electrochemical cell?

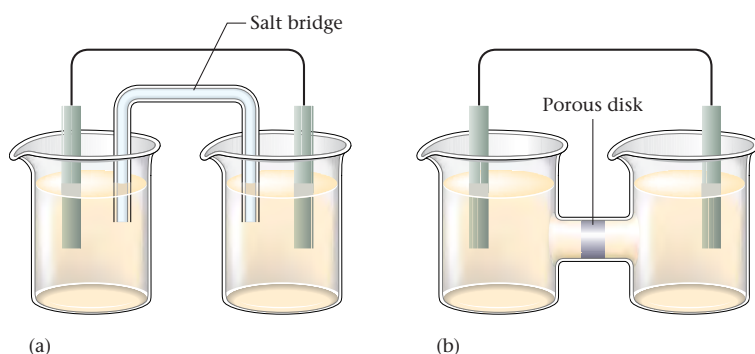


Figure 18.2

A salt bridge or a porous-disk connection allows ions to flow, completing the electric circuit. (a) The salt bridge contains a strong electrolyte either as a gel or as a solution; both ends are covered with a membrane that allows only ions to pass. (b) The porous disk allows ion flow but does not permit overall mixing of the solutions in the two compartments.

B. Batteries

In the previous section we saw that a galvanic cell is a device that uses an oxidation–reduction reaction to generate an electric current by separating the oxidizing agent from the reducing agent. In this section we will consider several specific galvanic cells and their applications.

Lead Storage Battery

Since about 1915, when self-starters were first used in automobiles, the **lead storage battery** has been a major factor in making the automobile a practical means of transportation. This type of battery can function for several years under temperature extremes ranging from $-30\text{ }^{\circ}\text{F}$ to $100\text{ }^{\circ}\text{F}$ and under incessant punishment from rough roads. The fact that this same type of battery has been in use for so many years in the face of all of the changes in science and technology over that span of time attests to how well it does its job.

In the lead storage battery, the reducing agent is lead metal, Pb , and the oxidizing agent is lead(IV) oxide, PbO_2 . We have already considered a simplified version of this reaction in Example 18.6. In an actual lead storage battery, sulfuric acid, H_2SO_4 , furnishes the H^+ needed in the reaction; it also furnishes SO_4^{2-} ions that react with the Pb^{2+} ions to form solid PbSO_4 . A schematic of one cell of the lead storage battery is shown in **Figure 18.3**.

In this cell the anode is constructed of lead metal, which is oxidized. In the cell reaction, lead atoms lose two electrons each to form Pb^{2+} ions, which combine with SO_4^{2-} ions present in the solution to give solid PbSO_4 .

The cathode of this battery has lead(IV) oxide coated onto lead grids. Lead atoms in the +4 oxidation state in PbO_2 accept two electrons each (are reduced) to give Pb^{2+} ions that also form solid PbSO_4 .

In the cell the anode and cathode are separated (so that the electrons must flow through an external wire) and bathed in sulfuric acid. The half-reactions that occur at the two electrodes and the overall cell reaction are shown below.

Anode reaction:



Cathode reaction:



Overall reaction:

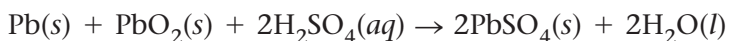


Figure 18.3

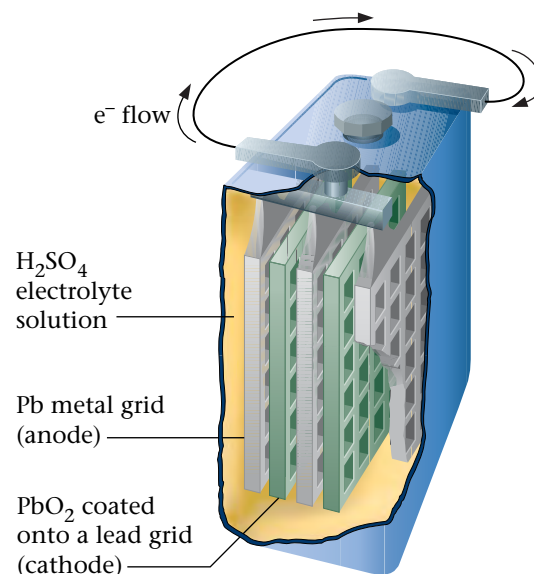
In a lead storage battery each cell consists of several lead grids that are connected by a metal bar. These lead grids furnish electrons (the lead atoms lose electrons to form Pb^{2+} ions, which combine with SO_4^{2-} ions to give solid PbSO_4). Because the lead is oxidized, it functions as the anode of the cell. The substance that gains electrons is PbO_2 ; it is coated onto lead grids, several of which are hooked together by a metal bar. The PbO_2 formally contains Pb^{4+} , which is reduced to Pb^{2+} , which in turn combines with SO_4^{2-} to form solid PbSO_4 . The PbO_2 accepts electrons, so it functions as the cathode.

Lead storage battery

A battery used in cars in which the anode is lead, the cathode is lead coated with lead(IV) oxide, and the electrolyte is sulfuric acid

information

The oxidizing agent accepts electrons and the reducing agent furnishes electrons.



Active Reading Question

What is the reducing agent in the lead storage battery? The oxidizing agent?

Electrical Potential The tendency for electrons to flow from the anode to the cathode in a battery depends on the ability of the reducing agent to release electrons and on the ability of the oxidizing agent to capture electrons. If a battery consists of a reducing agent that releases electrons readily and an oxidizing agent with a high affinity for electrons, the electrons are driven through the connecting wire with great force and can provide much electrical energy.

To understand this think of the analogy of water flowing through a pipe. The greater the pressure on the water, the more vigorously the water flows. The “pressure” on electrons to flow from one electrode to the other in a battery is called the **potential** of the battery and is measured in volts. For example, each cell in a lead storage battery produces about 2 volts of potential. In an actual automobile battery, six of these cells are connected to produce about 12 volts of potential.

Potential

The “pressure” on electrons to flow from the anode to the cathode in a battery

Dry cell battery

A battery that does not contain a liquid electrolyte; commonly used in calculators and watches

Dry Cell Batteries

The calculators, electronic watches, personal music players that are so familiar to us are all powered by small, efficient **dry cell batteries**. They are called dry cells because they do not contain a liquid electrolyte. The common dry cell battery was invented more than 100 years ago by George Leclanché (1839–1882), a French chemist. In its *acid version*, the dry cell battery contains a zinc inner case that acts as the anode and a carbon (graphite) rod in contact with a moist paste of solid MnO_2 , solid NH_4Cl , and carbon that acts as the cathode (see **Figure 18.4**). The half-cell reactions are complex but can be approximated as follows:

Anode reaction:



Cathode reaction:



This cell produces a potential of about 1.5 volts.

In the *alkaline version* of the dry cell battery, the NH_4Cl is replaced with KOH or NaOH . In this case the half-reactions can be approximated as follows:

Anode reaction:



Cathode reaction:



The alkaline dry cell lasts longer, mainly because the zinc anode corrodes less rapidly under basic conditions than under acidic conditions.

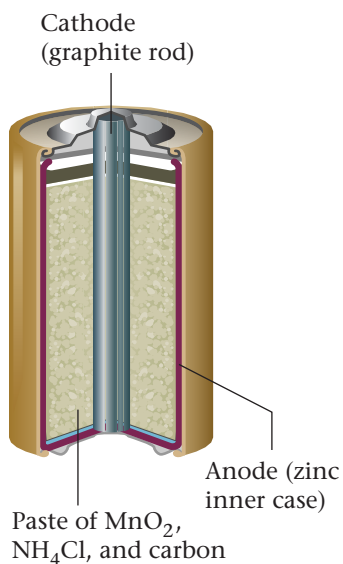


Figure 18.4
A common dry cell battery

Other types of dry cell batteries include the *silver cell*, which has a Zn anode and a cathode that employs Ag_2O as the oxidizing agent in a basic environment. *Mercury cells*, often used in calculators, have a Zn anode and a cathode involving HgO as the oxidizing agent in a basic medium (see **Figure 18.5**).

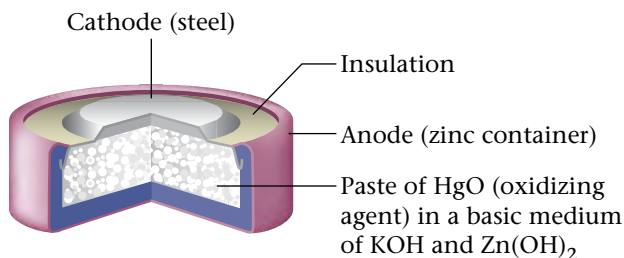


Figure 18.5
A mercury battery of the type used in small calculators

An especially important type of dry cell is the *nickel-cadmium battery*, in which the electrode reactions are

Anode reaction:



Cathode reaction:



In this cell, as in the lead storage battery, the products adhere to the electrodes. Therefore, a nickel-cadmium battery can be recharged an indefinite number of times, because the products can be turned back into reactants by the use of an external source of current.

C. Corrosion

Most metals are found in nature in compounds with nonmetals such as oxygen and sulfur. For example, iron exists as iron ore (which contains Fe_2O_3 and other oxides of iron).

Corrosion can be viewed as the process of returning metals to their natural state—the ores from which they were originally obtained. Corrosion involves oxidation of the metal. Because corroded metal often loses its strength and attractiveness, this process causes great economic loss. For example, approximately one-fifth of the iron and steel produced annually is used to replace rusted metal.

Self-Protecting Metals Because most metals react with O_2 , we might expect them to corrode so fast in air that they wouldn't be useful. It is surprising, therefore, that the problem of corrosion does not virtually prevent the use of metals in air. Part of the explanation is that most metals develop a thin oxide coating, which tends to protect their internal atoms against further oxidation. The best example of this is aluminum. Aluminum readily loses electrons, so it should be very easily oxidized by O_2 . Given this fact, why is aluminum so useful for building airplanes, bicycle frames, and so on? Aluminum is such a valuable structural material because it forms a thin adherent layer of aluminum oxide, Al_2O_3 , which greatly inhibits further corrosion. Aluminum protects itself with this tough oxide coat. Many other metals, such as chromium, nickel, and tin, do the same thing.

Top Ten Most Active Metals (Best Reducing Agents in Water)

Li
K
Ba
Ca
Na
La
Mg
Al
Mn
Zn

Corrosion

The process by which metals are oxidized in the atmosphere



Lemon Power**Materials**

- three lemons
- three galvanized nails
- three pennies
- alligator clips
- wire
- voltmeter

Procedure

1. Stick one nail and one penny into each lemon, making sure that they do not touch. Connect the penny from the first lemon to the nail of the second lemon using the alligator clips and a piece of wire. Connect the third lemon to the second lemon in the same way.

2. Connect the nail of the first lemon to a terminal of a voltmeter using an alligator clip and a piece of wire. Connect the penny of the third lemon to a terminal of a voltmeter in the same way. Determine the potential of the lemon battery.

Results/Analysis

1. Explain why this battery works.

Active Reading Question

How do some metals protect themselves from corrosion?

Information

An alloy is a mixture of elements with metallic properties.

Iron can also form a protective oxide coating. However, this oxide is not a very effective shield against corrosion, because it scales off easily, exposing new metal surfaces to oxidation. Under normal atmospheric conditions, copper forms an external layer of greenish copper sulfate or carbonate called *patina*. *Silver tarnish* is silver sulfide, Ag_2S , which in thin layers gives the silver surface a richer appearance. Gold shows no appreciable corrosion in air.

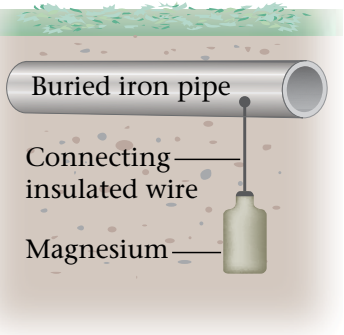
Cathodic protection

The connection of an active metal, such as magnesium, to steel to protect the steel from corrosion

Preventing Corrosion Preventing corrosion is an important way of conserving our natural supplies of metals and energy. The primary means of protection is the application of a coating—most often paint or metal plating—to protect the metal from oxygen and moisture. Chromium and tin are often used to plate steel because they oxidize to form a durable, effective oxide coating.

Alloying is also used to prevent corrosion. *Stainless steel* contains chromium and nickel, both of which form oxide coatings that protect the steel.

Cathodic protection is the method most often employed to protect steel in buried fuel tanks and pipelines. A metal that furnishes electrons more easily than iron, such as magnesium, is connected by a wire to the pipeline or tank that is to be protected (see **Figure 18.6**). Because the magnesium is a better reducing agent than iron, electrons flow through the wire from the magnesium to the iron pipe. Thus the electrons are furnished by the magnesium rather than by the iron, keeping the iron from being oxidized. As oxidation of the magnesium occurs, the magnesium dissolves, so it must be replaced periodically.



Buried iron pipe

Connecting
insulated wire

Magnesium

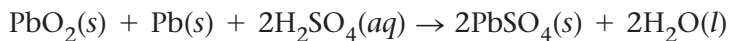
Figure 18.6

Cathodic protection of an underground pipe

D. Electrolysis

Unless it is recharged, a battery “runs down” because the substances in it that furnish and accept electrons (to produce the electron flow) are

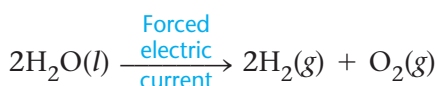
consumed. For example, in the lead storage battery (see Section 18.3B), PbO_2 and Pb are consumed to form PbSO_4 as the battery runs.



However, one of the most useful characteristics of the lead storage battery is that it can be recharged. *Forcing* current through the battery in the direction opposite to the normal direction reverses the oxidation–reduction reaction. That is, PbSO_4 is consumed and PbO_2 and Pb are formed in the charging process. This recharging is done continuously by the automobile’s alternator, which is powered by the engine.

The process of **electrolysis** involves *forcing a current through a cell to produce a chemical change that would not otherwise occur*.

One important example of this type of process is the electrolysis of water. Water is a very stable substance that can be broken down into its elements by using an electric current.



The electrolysis of water produces hydrogen gas at the cathode (on the left) and oxygen gas at the anode (on the right). A nonreacting strong electrolyte such as Na_2SO_4 is needed to furnish ions to allow the flow of current.

The electrolysis of water to produce hydrogen and oxygen occurs when a current is forced through an aqueous solution. Thus, when the lead storage battery is charged, or “jumped,” potentially explosive mixtures of H_2 and O_2 are produced by the current flow through the solution in the battery. This is why it is very important not to produce a spark near the battery during these operations.



Electrolysis

A process involving forcing a current through a cell to produce a chemical change that would not otherwise occur

CRITICAL THINKING ?

The air we breathe is approximately 80% nitrogen and 20% oxygen.

What if our air was 80% oxygen and 20% nitrogen? How would it affect our lives?

Active Reading Question

Why is it important to avoid a spark near a car battery when you are jump-starting it?

CHEMISTRY EXPLORERS

Charles Martin Hall • 1863–1914

Hall was a student at Oberlin College in Ohio when he first became interested in aluminum. One of his professors commented that anyone who could find a way to manufacture aluminum cheaply would make a fortune, and Hall decided to give it a try. The 21-year-old worked in a wooden shed near his house with an iron frying pan as a container, a blacksmith’s forge as a heat source, and galvanic cells constructed from fruit jars. Using these crude galvanic cells, Hall found that he could produce aluminum by passing a current through a molten mixture of Al_2O_3 and Na_3AlF_6 . By a strange coincidence, Paul Heroult, a French chemist who was born and died in the same years as Hall, made the same discovery at about the same time.



Production of Aluminum Another important use of electrolysis is in the production of metals from their ores. The metal produced in the greatest quantities by electrolysis is aluminum.

Aluminum is one of the most abundant elements on earth, ranking third behind oxygen and silicon. Because aluminum is a very reactive metal, it is found in nature as its oxide in an ore called *bauxite* (named after Les Baux, France, where it was discovered in 1821). Production of aluminum metal from its ore proved to be more difficult than the production of most other metals. In 1782 Lavoisier, the pioneering French chemist, recognized aluminum as a metal “whose affinity for oxygen is so strong that it cannot be overcome by any known reducing agent.” As a result, pure aluminum metal remained unknown. Finally, in 1854, a process was found for producing metallic aluminum by using sodium, but aluminum remained a very expensive rarity. In fact, it is said that Napoleon III served his most honored guests with aluminum forks and spoons, while the others had to settle for gold and silver utensils!

The breakthrough came in 1886 when two men, Charles M. Hall in the United States and Paul Heroult in France, almost simultaneously discovered a practical electrolytic process for producing aluminum, which greatly increased the availability of aluminum for many purposes. **Table 18.2** shows how dramatically the price of aluminum dropped after this discovery. The effect of the electrolysis process is to reduce Al^{3+} ions to neutral Al atoms that form aluminum metal. The aluminum produced in this electrolytic process is 99.5% pure. To be useful as a structural material, aluminum is alloyed with metals such as zinc (for trailer and aircraft construction) and manganese (for cooking utensils, storage tanks, and highway signs). The production of aluminum consumes about 4.5% of all electricity used in the United States.

Table 18.2

The Price of Aluminum

Date	Price of Aluminum (\$/lb)*
1855	\$100,000
1885	100
1890	2
1895	0.50
1970	0.30
1980	0.80
1990	0.74
2005	0.92

*Note the huge drop in price after the discovery of the Hall-Heroult process in 1886.

SECTION 18.3 REVIEW QUESTIONS

- 1 What is the difference between a galvanic cell and an electrolytic cell?
- 2 Draw the galvanic cell shown in Figure 18.2b.
 - a. Label the electrons in the wire as flowing from right to left.
 - b. Label the anode and the cathode.
 - c. Show where oxidation occurs. Show where reduction occurs.
 - d. What would happen to the cell if you removed the salt bridge?
- 3 What is electrochemical potential?
- 4 Why does an alkaline battery last longer than an acid dry cell?
- 5 What is corrosion? Why do we want to prevent it?
- 6 What are three ways to protect against corrosion?
- 7 What happens to a battery when it is “recharged”?



Key Terms

18.1 Oxidation–reduction reactions (redox reactions)

Oxidation

Reduction

Oxidation states

18.2 Oxidizing agent (electron acceptor)

Reducing agent

(electron donor)

Half-reactions

18.3 Electrochemistry

Electrochemical battery (galvanic cell)

Anode

Cathode

Electrolysis

Lead storage battery

Potential

Dry cell batteries

Corrosion

Cathodic protection

Key Ideas

18.1 Electron Transfer Reactions

- Oxidation–reduction reactions involve electron transfer.
 - Oxidation states are used to keep track of electrons.
 - Oxidation—an increase in oxidation state (a loss of electrons)
 - Reduction—a decrease in oxidation state (a gain of electrons)
 - The oxidizing agent accepts electrons.
 - The reducing agent furnishes electrons.

18.2 Balancing Oxidation–Reduction Reactions

- Balancing oxidation–reduction reactions can be accomplished by two methods:
 - Inspection (trial and error)
 - Half-reactions

18.3 Electrochemistry and Its Applications

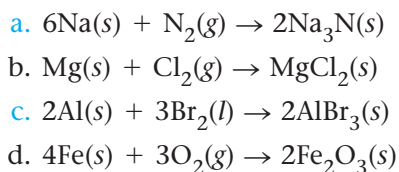
- Electrochemistry is the study of the interchange of chemical and electrical energy.
- A galvanic cell is a device for converting chemical energy to electrical energy.
 - Consists of separate compartments for the oxidizing and reducing agents which are connected by a salt bridge and a wire
 - Anode—electrode at which oxidation occurs
 - Cathode—electrode at which reduction occurs
 - Battery—a galvanic cell or a group of cells connected together
- Corrosion is the oxidation of metals to form mainly oxides and sulfides.
 - Some metals, such as aluminum, protect themselves with their oxide coating.
 - Corrosion of iron can be prevented by coatings, by alloying, and cathodic protection.
- Electrolysis is the use of electrical energy to produce a chemical change.

All exercises with **blue numbers** have answers in the back of this book.

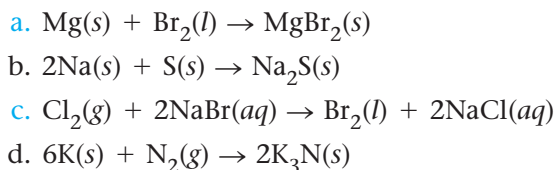
18.1 Electron Transfer Reactions

A. Oxidation–Reduction Reactions

- How is *oxidation* defined? Write an equation showing an atom being *oxidized*.
- How is *reduction* defined? Write an equation showing an atom being *reduced*.
- For each of the following oxidation–reduction reactions of metals with nonmetals, identify which element is oxidized and which is reduced.



- For each of the following oxidation–reduction reactions, identify which element is oxidized and which is reduced.



B. Oxidation States

- What is an oxidation state? Why is this concept useful?
- Explain why, although it is not an ionic compound, we assign oxygen an oxidation state of -2 in water, H_2O . Give an example of a compound in which oxygen is *not* in the -2 oxidation state.
- Why must the sum of all the oxidation states of the atoms in a neutral molecule be zero?
- What must be the sum of the oxidation states of all the atoms in a polyatomic ion?
- Assign oxidation states to all of the atoms in each of the following:
 - NCl_3
 - SF_6
 - PCl_5
 - SiH_4
- Assign oxidation states to all of the atoms in each of the following:
 - HBr
 - HOBr
 - Br_2
 - HBrO_4

- Assign oxidation states to all of the atoms in each of the following:



- Assign oxidation states to all of the atoms in each of the following:



- Assign oxidation states to all of the atoms in each of the following:



18.2 Balancing Oxidation–Reduction Reactions

A. Oxidation–Reduction Reactions Between Nonmetals

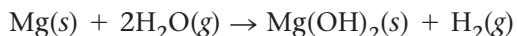
- Oxidation can be defined as a loss of electrons or as an increase in oxidation state. Explain why the two definitions mean the same thing, and give an example to support your explanation.
- Reduction can be defined as a gain of electrons or as a decrease in oxidation state. Explain why the two definitions mean the same thing, and give an example to support your explanation.
- Is an oxidizing agent itself oxidized or reduced? Is a reducing agent itself oxidized or reduced?
- Does an oxidizing agent increase or decrease its own oxidation state when it acts on another atom? Does a reducing agent increase or decrease its own oxidation state when it acts on another substance?
- In each of the following reactions, identify which element is oxidized and which is reduced by assigning oxidation numbers.
 - $\text{Zn}(s) + 2\text{HNO}_3(aq) \rightarrow \text{Zn}(\text{NO}_3)_2(aq) + \text{H}_2(g)$
 - $\text{H}_2(g) + \text{CuSO}_4(aq) \rightarrow \text{Cu}(s) + \text{H}_2\text{SO}_4(aq)$
 - $\text{N}_2(g) + 3\text{Br}_2(l) \rightarrow 2\text{NBr}_3(g)$
 - $2\text{KBr}(aq) + \text{Cl}_2(g) \rightarrow 2\text{KCl}(aq) + \text{Br}_2(l)$
- In each of the following reactions, identify which element is oxidized and which is reduced by assigning oxidation states.
 - $\text{Cu}(s) + 2\text{AgNO}_3(aq) \rightarrow 2\text{Ag}(s) + \text{Cu}(\text{NO}_3)_2(aq)$
 - $\text{N}_2(g) + 3\text{F}_2(g) \rightarrow 2\text{NF}_3(g)$
 - $2\text{Fe}_2\text{O}_3(s) + 3\text{S}(s) \rightarrow 4\text{Fe}(s) + 3\text{SO}_2(g)$
 - $2\text{H}_2\text{O}_2(l) \rightarrow 2\text{H}_2\text{O}(l) + \text{O}_2(g)$

- 20.** In ordinary photography, the light-sensitive portion of the film is a thin coating of a silver halide (usually silver bromide). When the film is exposed to light, the reaction



occurs. Identify which element is reduced and which is oxidized.

- 21.** Although magnesium metal does not react with water at room temperature, it does react vigorously with steam at higher temperatures, releasing elemental hydrogen gas from the water.

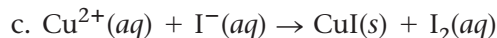
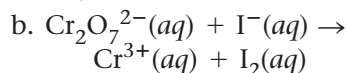


Identify which element is being oxidized and which is being reduced.

B. Balancing Oxidation–Reduction Reactions by the Half-Reaction Method

- 22.** In what *two* respects must oxidation–reduction reactions be balanced?
- 23.** Why must the number of electrons lost in the oxidation equal the number of electrons gained in the reduction? Is it possible to have “leftover” electrons in a reaction?
- 24.** Balance each of the following half-reactions.
- $\text{N}_2(g) \rightarrow \text{N}_3^-(aq)$
 - $\text{O}_2^{2-}(aq) \rightarrow \text{O}_2(g)$
 - $\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq)$
 - $\text{F}_2(g) \rightarrow \text{F}^-(aq)$
- 25.** Balance each of the following half-reactions, which take place in acidic solution.
- $\text{O}_2(g) \rightarrow \text{H}_2\text{O}(l)$
 - $\text{IO}_3^-(aq) \rightarrow \text{I}_2(s)$
 - $\text{VO}^{2+}(aq) \rightarrow \text{V}^{3+}(aq)$
 - $\text{BiO}^+(aq) \rightarrow \text{Bi}(s)$
- 26.** Balance each of the following oxidation–reduction reactions, which take place in acidic solution, by using the half-reaction method.
- $\text{MnO}_4^-(aq) + \text{Zn}(s) \rightarrow \text{Mn}^{2+}(aq) + \text{Zn}^{2+}(aq)$
 - $\text{Sn}^{4+}(aq) + \text{H}_2(g) \rightarrow \text{Sn}^{2+}(aq) + \text{H}^+(aq)$
 - $\text{Zn}(s) + \text{NO}_3^-(aq) \rightarrow \text{Zn}^{2+}(aq) + \text{NO}_2(g)$
 - $\text{H}_2\text{S}(g) + \text{Br}_2(l) \rightarrow \text{S}(s) + \text{Br}^-(aq)$

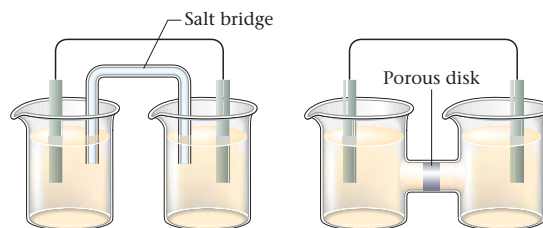
- 27.** Iodide ion, I^- , is one of the most easily oxidized species. Balance each of the following oxidation–reduction reactions, which take place in acidic solution, by using the half-reaction method.



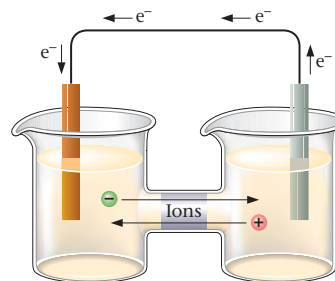
18.3 Electrochemistry and Its Applications

A. Electrochemistry: An Introduction

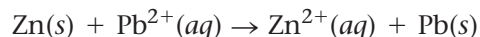
- 28.** How is an oxidation–reduction reaction set up as a galvanic cell (battery)? How is the transfer of electrons between reducing agent and oxidizing agent made useful?
- 29.** What is a salt bridge? Why is a salt bridge necessary in a galvanic cell? Can a different method be used in place of the salt bridge?



- 30.** In which direction do electrons flow in a galvanic cell, anode to cathode, or vice versa?
- 31.** Consider the oxidation–reduction reaction
- $$\text{Al}(s) + \text{Ni}^{2+}(aq) \rightarrow \text{Al}^{3+}(aq) + \text{Ni}(s)$$
- Sketch a galvanic cell that makes use of this reaction. Which metal ion is reduced? Which metal is oxidized? What half-reaction takes place at the anode in the cell? What half-reaction takes place at the cathode?



- 32.** Consider the oxidation–reduction reaction



Sketch a galvanic cell that uses this reaction. Which metal ion is reduced? Which metal is oxidized? What half-reaction takes place at the anode in the cell? What half-reaction takes place at the cathode?

B. Batteries

- 33.** Write the chemical equation for the overall cell reaction that occurs in a lead storage automobile battery. What species is oxidized in such a battery? What species is reduced? Why can such a battery be “recharged”?
- 34.** Why does an alkaline dry cell battery typically last longer than a normal dry cell? Write the chemical equation for the overall cell reaction in an alkaline dry cell.

C. Corrosion

- 35.** What process is represented by the *corrosion* of a metal? Why is corrosion undesirable?
- 36.** Explain how some metals, notably aluminum, naturally resist complete oxidation by the atmosphere.

- 37.** Pure iron ordinarily rusts quickly, but steel does not corrode nearly as fast. How does steel resist corrosion?
- 38.** What is *cathodic protection*, and how is it applied to prevent oxidation of steel tanks and pipes?

D. Electrolysis

- 39.** What is *electrolysis*? What types of reactions is electrolysis capable of causing?
- 40.** What reactions go on during the recharging of an automobile battery?
- 41.** How is aluminum metal produced by electrolysis? Why can't simpler chemical reduction methods be used?
- 42.** Jewelry is often manufactured by plating an expensive metal such as gold over a cheaper metal. How might such a process be set up as an electrolysis reaction?

Critical Thinking

- 43.** An oxidizing agent causes the (oxidation/reduction) of another species, and the oxidizing agent itself is (oxidized/reduced).
- 44.** To function as a good reducing agent, a species must _____ electrons easily.
- 45.** Jump-starting a dead automobile battery can be dangerous if precautions are not taken, because of the production of an explosive mixture of _____ and _____ gases in the battery.
- 46.** For each of the following unbalanced oxidation–reduction chemical equations, balance the equation by inspection, and identify which species is undergoing oxidation and which is undergoing reduction.
- $\text{Fe}(s) + \text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s)$
 - $\text{Al}(s) + \text{Cl}_2(g) \rightarrow \text{AlCl}_3(s)$
 - $\text{Mg}(s) + \text{P}_4(s) \rightarrow \text{Mg}_3\text{P}_2(s)$
- 47.** Balance each of the following oxidation–reduction reactions, which take place in acidic solution.
- $\text{MnO}_4^-(aq) + \text{H}_2\text{O}_2(aq) \rightarrow \text{Mn}^{2+}(aq) + \text{O}_2(g)$
 - $\text{BrO}_3^-(aq) + \text{Cu}^+(aq) \rightarrow \text{Br}^-(aq) + \text{Cu}^{2+}(aq)$
 - $\text{HNO}_2(aq) + \text{I}^-(aq) \rightarrow \text{NO}(g) + \text{I}_2(aq)$

- 48.** For each of the following oxidation–reduction reactions of metals with nonmetals, identify which element is oxidized and which is reduced.
- $3\text{Zn}(s) + \text{N}_2(g) \rightarrow \text{Zn}_3\text{N}_2(s)$
 - $\text{Co}(s) + \text{S}(s) \rightarrow \text{CoS}(s)$
 - $4\text{K}(s) + \text{O}_2(g) \rightarrow 2\text{K}_2\text{O}(s)$
 - $4\text{Ag}(s) + \text{O}_2(g) \rightarrow 2\text{Ag}_2\text{O}(s)$
- 49.** Assign oxidation states to all of the atoms in each of the following:
- MnO_2
 - BaCrO_4
 - H_2SO_3
 - $\text{Ca}_3(\text{PO}_4)_2$
- 50.** In each of the following reactions, identify which element is oxidized and which is reduced by assigning oxidation states.
- $2\text{B}_2\text{O}_3(s) + 6\text{Cl}_2(g) \rightarrow 4\text{BCl}_3(l) + 3\text{O}_2(g)$
 - $\text{GeH}_4(g) + \text{O}_2(g) \rightarrow \text{Ge}(s) + 2\text{H}_2\text{O}(g)$
 - $\text{C}_2\text{H}_4(g) + \text{Cl}_2(g) \rightarrow \text{C}_2\text{H}_4\text{Cl}_2(l)$
 - $\text{O}_2(g) + 2\text{F}_2(g) \rightarrow 2\text{OF}_2(g)$
- 51.** Consider the oxidation–reduction reaction
- $$\text{Mg}(s) + \text{Cu}^{2+}(aq) \rightarrow \text{Mg}^{2+}(aq) + \text{Cu}(s)$$
- Sketch a galvanic cell that uses this reaction. Which metal ion is reduced? Which metal is oxidized? What half-reaction takes place at the anode in the cell? What half-reaction takes place at the cathode?

- 1 In the following reaction, identify which element is oxidized, and which element is reduced.



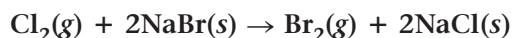
- A Calcium is oxidized and oxygen is reduced.
 B Oxygen is oxidized and calcium is reduced.
 C One of the oxygen atoms is oxidized, and the other oxygen atom is reduced.
 D The reaction above is not an oxidation–reduction reaction.
- 2 In which of the following compounds or ions is the oxidation state for nitrogen -3 ?

- A N_2
 B NI_3
 C NO_2
 D NH_3

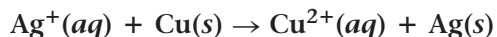
- 3 Which of the following is not an oxidation reduction reaction?

- A $2\text{Al}(s) + 3\text{I}_2(s) \rightarrow 2\text{AlI}_3(s)$
 B $\text{Mg}(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g)$
 C $\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)$
 D $\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l)$

- 4 In the following reaction, identify the oxidizing agent and the reducing agent.



- A Cl_2 = oxidizing agent, NaBr = reducing agent
 B NaBr = oxidizing agent, Cl_2 = reducing agent
 C Cl_2 = oxidizing agent, Br_2 = reducing agent
 D NaCl = oxidizing agent, NaBr = reducing agent
- 5 Which of the following is true concerning the following equation?



- A When the equation is balanced in standard form, the coefficient in front of the Cu and Cu^{2+} is 2.
 B When the equation is balanced in standard form, the coefficient in front of the Ag and Ag^+ is 2.
 C When the equation above is balanced in standard form, electrons need to appear in the equation.
 D The equation as written above is balanced in standard form.

- 6 The oxidizing agent in the lead storage battery is

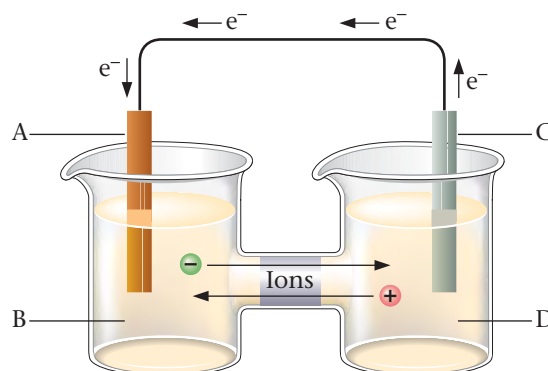
- A Pb
 B PbO_2
 C H_2SO_4
 D PbSO_4

- 7 The process in which electrical energy is used to produce chemical energy is called

- A cathodic protection.
 B electrolysis.
 C corrosion.
 D electrochemistry.

- 8 Label the following parts of the galvanic cell.

- anode
 cathode
 reducing agent
 oxidizing agent



- 9 The reaction $\text{Cu}(s) + \text{NO}_3^-(aq) \rightarrow \text{Cu}^{2+}(aq) + \text{NO}(g)$ takes place in acidic solution.

- a Write the balanced oxidation half-reaction.
 b Write the balanced reduction half-reaction.
 c Write the overall balanced reaction.
 d Label the oxidizing agent and the reducing agent.

Radioactivity and Nuclear Energy

LOOKING AHEAD

19.1 Radioactivity


- A. Radioactive Decay
 - Types of Radioactive Decay
- B. Nuclear Transformations
- C. Detection of Radioactivity and the Concept of Half-life

19.2 Application of Radioactivity

- A. Dating by Radioactivity
- B. Medical Applications of Radioactivity

19.3 Using the Nucleus as a Source of Energy

- A. Nuclear Energy
- B. Nuclear Fission
- C. Nuclear Reactors
 - Breeder Reactors
- D. Nuclear Fusion
- E. Effects of Radiation



The sun's energy comes from nuclear reactions.

IN YOUR LIFE

Although the electrons are the part of the atom that mostly determines its chemical properties, the nucleus is very important in all of our lives. The high energy radiation produced by radioactive nuclides can be very harmful to us, but it is also used in many types of medical tests. Called radiotracers, certain radioactive nuclides are introduced into the body to analyze biological functions and to provide images of organs such as the thyroid and the heart.

The nucleus is also a very important, though controversial, source of energy in our society. For example, France uses nuclear energy as a source for about 80% of its electrical power. In the United States nuclear energy produces over 20% of our electrical power. Although the production of energy in nuclear reactors has some safety issues, nuclear energy production is likely to become more important in the future because it does not result in the release of “greenhouse gases” such as carbon dioxide.

Nuclear processes are also important to geologists and archeologists who use radioactivity to find the age of rocks and other artifacts.

Wooden artifacts from Easter Island can be dated by carbon-14 analysis.



WHAT DO YOU KNOW?



Prereading Questions

1. What is meant by the term *atomic number*? What is meant by the term *mass number*?
2. What are isotopes?
3. What comes to mind when you hear the term *radioactive*?
4. Have you heard of carbon dating? What do you know about it?

SECTION 19.1

Radioactivity

Key Terms

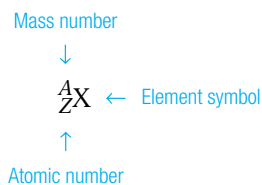
- Nucleons (protons and neutrons)
- Atomic number (Z)
- Mass number (A)
- Isotopes
- Nuclide
- Radioactive
- Beta particle
- Nuclear equation
- Alpha particle
- Alpha-particle production
- Beta-particle production
- Gamma ray
- Positron
- Positron production
- Electron capture
- Decay series
- Nuclear transformation
- Transuranium elements
- Geiger–Müller counter
- Scintillation counter
- Half-life

Objectives

- To learn the types of radioactive decay
- To learn to write nuclear equations for radioactive decay
- To learn how one element may be changed to another by particle bombardment
- To learn about radiation detection instruments
- To understand half-life

Several facts about the nucleus are immediately impressive: its very small size, its very large density, and the energy that holds it together. The radius of a typical nucleus is about 10^{-13} cm, only a hundred-thousandth the radius of a typical atom. In fact, if the nucleus of the hydrogen atom were the size of a ping-pong ball, the electron in the 1s orbital would be, on the average, 0.5 km (0.3 mile) away. The density of the nucleus is equally impressive; it is approximately 1.6×10^{14} g/cm³. A sphere of nuclear material the size of a ping-pong ball would have a mass of 2.5 *billion tons*! Finally, the energies involved in nuclear processes are typically millions of times larger than those associated with normal chemical reactions, a fact that makes nuclear processes potentially attractive for generating energy.

The nucleus is believed to be made of particles called **nucleons** (**neutrons** and **protons**). Recall from Chapter 3 that the number of protons in a nucleus is called the **atomic number (Z)** and that the sum of the numbers of neutrons and protons is the **mass number (A)**. Atoms that have identical atomic numbers but different mass numbers are called **isotopes**. The general term **nuclide** is applied to each unique atom, and we represent it as follows:



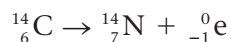
where X represents the symbol for a particular element. For example, the following nuclides constitute the common isotopes of carbon: carbon-12, $^{12}_6\text{C}$; carbon-13, $^{13}_6\text{C}$; and carbon-14, $^{14}_6\text{C}$. Notice that all of the carbon nuclides have six protons ($Z = 6$) and that they have six, seven, and eight neutrons, respectively.

information

The atomic number (Z) represents the number of protons in a nucleus; the mass number (A) represents the sum of the numbers of protons and neutrons in a nucleus.

A. Radioactive Decay

Many nuclei are **radioactive**; that is, they spontaneously decompose, forming a different nucleus and producing one or more particles. An example is carbon-14, which decays as shown in the equation



where $^0_{-1}\text{e}$ represents an electron, which in nuclear terminology is called a **beta particle**, or **β particle**. This **nuclear equation**, which is typical of those representing radioactive decay, is quite different from the chemical equations we have written before. Recall that in a balanced chemical equation the atoms must be conserved.

DID YOU KNOW

Over 85% of all known nuclides are radioactive.

In a nuclear equation both the atomic number (Z) and the mass number (A) must be conserved.

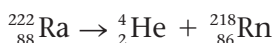
That is, the sums of the Z values on both sides of the arrow must be equal, and the same restriction applies to the A values. For example, in the equation on the previous page, the sum of the Z values is 6 on both sides of the arrow (6 and $7 - 1$), and the sum of the A values is 14 on both sides of the arrow (14 and $14 + 0$). Notice that the mass number for the β particle is zero; the mass of the electron is so small that it can be neglected here. Of the approximately 2000 known nuclides, only 279 do not undergo radioactive decay. Tin has the largest number of nonradioactive isotopes—ten.

Active Reading Question

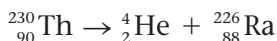
What is the difference between a chemical equation and a nuclear equation?

Types of Radioactive Decay

There are several different types of radioactive decay. One frequently observed decay process involves production of an **alpha (α) particle**, which is a helium nucleus (${}^4_2\text{He}$). **Alpha-particle production** is a very common mode of decay for heavy radioactive nuclides. For example, ${}^{222}_{88}\text{Ra}$, radium-222, decays by α -particle production to give radon-218.

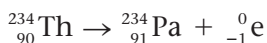


Notice in this equation that the mass number is conserved ($222 = 4 + 218$) and the atomic number is conserved ($88 = 2 + 86$). Another α -particle producer is ${}^{230}_{90}\text{Th}$:

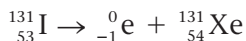


Notice that the production of an α particle results in a loss of 4 in mass number (A) and a loss of 2 in atomic number (Z).

Beta-particle production is another common decay process. For example, the thorium-234 nuclide produces a β particle as it changes to protactinium-234.



Iodine-131 is also a β -particle producer.



Recall that the β particle is assigned mass number 0, because its mass is tiny compared to that of a proton or neutron. The value of Z is -1 for the β particle, so the atomic number for the new nuclide is greater by 1 than the atomic number for the original nuclide. Therefore, *the net effect of β -particle production is to change a neutron to a proton.*

Production of a β particle results in no change in mass number (A), and an increase of 1 in atomic number (Z).

Alpha particle (α particle)

A helium nucleus produced in radioactive decay

Alpha-particle production

A common mode of decay for heavy radioactive nuclides resulting in a loss of 4 in mass number and a loss of 2 in atomic number

Beta-particle production

A decay process for radioactive nuclides in which the mass number does not change and the atomic number increases by 1

Information

α -particle production

$$\Delta A = -4$$

$$\Delta Z = -2$$

Information

Notice that both Z and A balance in each of these nuclear equations.

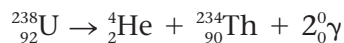
Information

β -particle production

$$\Delta A = 0$$

$$\Delta Z = +1$$

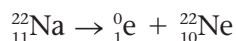
A **gamma ray**, or **γ ray**, is a high-energy photon of light. A nuclide in an excited nuclear energy state can release excess energy by producing a gamma ray, and γ -ray production often accompanies nuclear decays of various types. For example, in the α -particle decay of ${}_{92}^{238}\text{U}$,



two γ rays of different energies are produced in addition to the α particle (${}^4_2\text{He}$). Gamma rays are photons of light and so have zero charge and zero mass number.

Production of a γ ray results in no change in mass number (A), and no change in atomic number (Z).

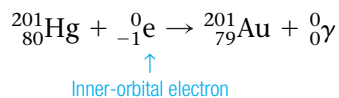
The **positron** is a particle with the same mass as the electron but opposite charge. An example of a nuclide that decays by **positron production** is sodium-22.



Note that the production of a positron has the effect of changing a proton to a neutron.

Production of a positron results in no change in mass number (A), and a decrease of 1 in atomic number (Z).

Electron capture is a process in which one of the inner-orbital electrons is captured by the nucleus, as illustrated by the process



This reaction would have been of great interest to the alchemists, but unfortunately it does not occur often enough to make it a practical means of changing mercury to gold. Gamma rays are always produced along with electron capture.

Information

The ${}_0^0\gamma$ notation indicates $Z = 0$ and $A = 0$ for a γ ray. A gamma ray is often simply indicated by γ .

Information

Positron production

$$\Delta A = 0$$

$$\Delta Z = -1$$

Gamma ray (γ ray)

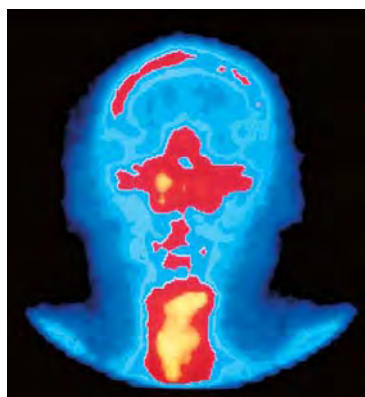
A high-energy photon produced in radioactive decay

Positron

A particle with the same mass as an electron but opposite in charge

Active Reading Question

For which radioactive processes does the reactant have a greater atomic number than the product?



Bone scintigraph of a patient's cranium following administration of the radiopharmaceutical technetium-99

Table 19.1 lists the common types of radioactive decay, with examples.

Table 19.1

Various Types of Radioactive Processes

Process	Example
β -particle (electron) production	${}^{227}_{89}\text{Ac} \rightarrow {}^{227}_{90}\text{Th} + {}^0_{-1}\text{e}$
positron production	${}^{13}_7\text{N} \rightarrow {}^{13}_6\text{C} + {}^0_1\text{e}$
electron capture	${}^{73}_{33}\text{As} + {}^0_{-1}\text{e} \rightarrow {}^{73}_{32}\text{Ge}$
α -particle production	${}^{210}_{84}\text{Po} \rightarrow {}^{206}_{82}\text{Pb} + {}^4_2\text{He}$
γ -ray production	excited nucleus \rightarrow ground-state nucleus + ${}^0_0\gamma$ excess energy lower energy

Often a radioactive nucleus cannot achieve a stable (nonradioactive) state through a single decay process. In such a case, a **decay series** occurs until a stable nuclide is formed. A well-known example is the decay series that starts with ${}^{238}_{92}\text{U}$ and ends with ${}^{206}_{82}\text{Pb}$, as shown in **Figure 19.1**. Similar series exist for ${}^{235}_{92}\text{U}$:



and for ${}^{232}_{90}\text{Th}$:

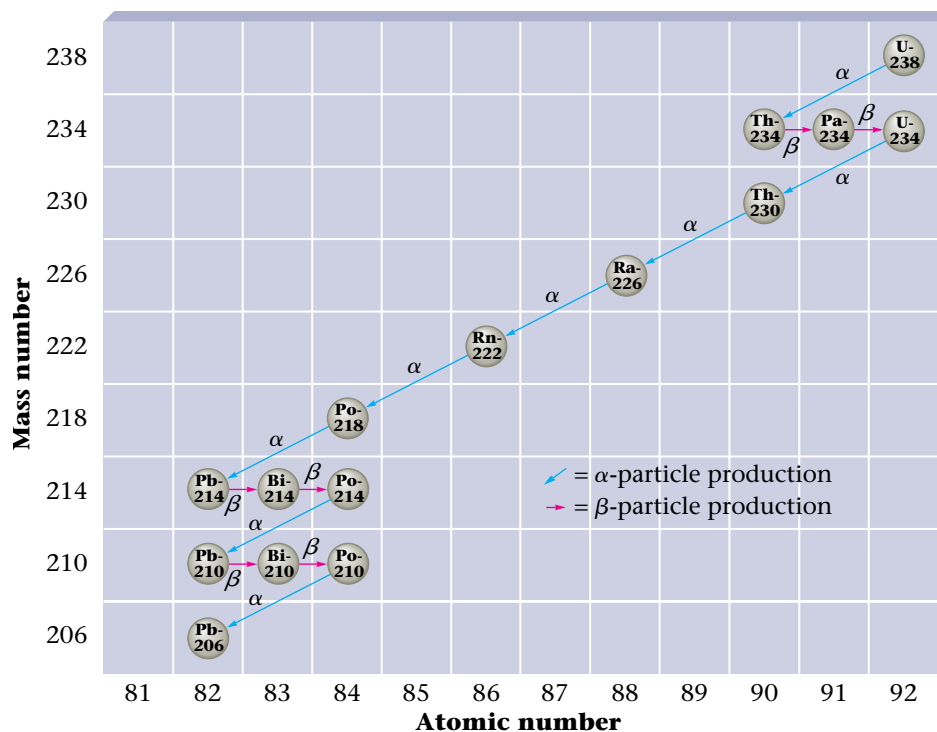


Figure 19.1

The decay series from ${}^{238}_{92}\text{U}$ to ${}^{206}_{82}\text{Pb}$. Each nuclide in the series except ${}^{206}_{82}\text{Pb}$ is radioactive, and the successive transformations (shown by the arrows) continue until ${}^{238}_{82}\text{Pb}$ is finally formed. The horizontal red arrows indicate β -particle production (Z increases by 1 and A is unchanged). The diagonal blue arrows signify α -particle production (both A and Z decrease).

CHEMISTRY EXPLORERS

Marie Curie • 1867–1934

Marie Sklodowska Curie, one of the truly monumental figures of modern science, was born in Warsaw, Poland, on November 7, 1867.

Marie emigrated to Paris in 1891 at the age of 24, where she decided to pursue a degree in science at the Sorbonne Institute. While studying there, Marie met Pierre Curie, a well-respected physicist, and they were married in 1895, after which Marie decided to pursue a doctorate in physics. As the subject of her doctoral thesis, she studied the strange radiation emitted by uranium ore, which had been accidentally discovered by Henri Becquerel. In her studies, Madame Curie noticed that pitchblende produced more radiation than uranium, and she became convinced that an as-yet-unknown element in pitchblende was responsible for this “radioactivity”—a term that she coined.

The next step was to identify and isolate the radioactive element or elements in pitchblende. Pierre interrupted his own research—he thought it would be for just a few weeks—to collaborate with his wife on the project.

They convinced the Austrian government to send them one ton of pitchblende from the mines at Joachimsthal. After receiving this 5-cubic-foot pile of “sand” from Austria, the Curies worked to chemically digest the ore. In this process they worked with batches as large

as 40 pounds in an improvised laboratory with a leaky roof. Working through the bitter winter of 1896 and all through 1897 (in which they had their first daughter, Irène, who also became a prominent scientist), in July 1898 the Curies finally isolated a previously unknown element they named polonium after Marie’s homeland. Although most people would be satisfied by discovering a new element 400 times more radioactive than uranium, the Curies kept working. By this time, the one ton of pitchblende had been concentrated to an amount that would fit into an ordinary flask. Marie continued to extract and crystallize increasingly smaller amounts of material. Finally, in November 1898 she obtained crystals of the salt of another new element that the Curies named radium, which turned out to be 900 times more radioactive than uranium.

For their work the Curies were awarded the Nobel Prize in Physics in 1903, sharing the award with Henri Becquerel.

In 1911, Marie was awarded her second Nobel Prize—this time in chemistry, for the chemical processes discovered in the identification of radium and polonium and for the subsequent characterization of these elements.



EXAMPLE 19.1

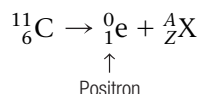
Writing Nuclear Equations, I

Write balanced nuclear equations for each of the following processes.

- $^{11}_6\text{C}$ produces a positron.
- $^{214}_{83}\text{Bi}$ produces a β particle.
- $^{237}_{93}\text{Np}$ produces an α particle.

Solution

- We must find the product nuclide represented by ^A_ZX in the following equation:



The key to solving this problem is to recognize that both A and Z must be conserved. That is, we can find the identity of A_ZX by recognizing that the sums of the Z and the A values must be the same on both sides of the equation. Thus for X , Z must be 5 because $Z + 1 = 6$. A must be 11 because $11 + 0 = 11$. Therefore, A_ZX is ${}^{11}_5\text{B}$. (The fact that Z is 5 tells us that the nuclide is boron. See the periodic table on the inside back cover of the book.) So the balanced equation is ${}^{11}_6\text{C} \rightarrow {}^0_1\text{e} + {}^{11}_5\text{B}$.

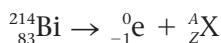
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$$Z = 6 \qquad Z = 5 + 1 = 6$$

→

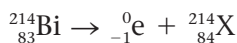
$$A = 11 \qquad A = 11 + 0 = 11$$

- b. Knowing that a β particle is represented by ${}^0_{-1}\text{e}$, we can write



where $Z - 1 = 83$ and $A + 0 = 214$. This means $Z = 84$ and $A = 214$.

We can now write



Using the periodic table, we find that $Z = 84$ for the element polonium, so ${}^{214}_84\text{X}$ must be ${}^{214}_{84}\text{Po}$.

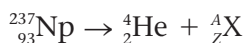
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$$Z = 83 \qquad Z = 84 - 1 = 83$$

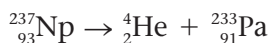
→

$$A = 214 \qquad A = 214 + 0 = 214$$

- c. Because an α particle is represented by ${}^4_2\text{He}$, we can write



where $A + 4 = 237$ or $A = 237 - 4 = 233$, and $Z + 2 = 93$ or $Z = 93 - 2 = 91$. Thus $A = 233$, $Z = 91$, and the balanced equation must be



Check: Left Side **Right Side**

$$Z = 93 \qquad Z = 91 + 2 = 93$$

→

$$A = 237 \qquad A = 133 + 4 = 237$$

Practice Problem • Exercise 19.1

The decay series for ${}^{238}_{92}\text{U}$ is represented in Figure 19.1. Write the balanced nuclear equation for each of the following radioactive decays.

- Alpha-particle production by ${}^{226}_{88}\text{Ra}$
- Beta-particle production by ${}^{214}_{82}\text{Pb}$

EXAMPLE 19.2

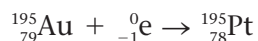
Writing Nuclear Equations, II

In each of the following nuclear reactions, supply the missing particle.

- a. ${}_{79}^{195}\text{Au} + ? \rightarrow {}_{78}^{195}\text{Pt}$
b. ${}_{19}^{38}\text{K} \rightarrow {}_{18}^{38}\text{Ar} + ?$

Solution

- a. A does not change and Z for Pt is 1 lower than Z for Au, so the missing particle must be an electron.



Check: Left Side **Right Side**

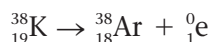
$$Z = 79 - 1 = 78 \qquad Z = 78$$

→

$$A = 195 + 0 = 195 \qquad A = 195$$

This is an example of electron capture.

- b. For Z and A to be conserved, the missing particle must be a positron.



Check: Left Side **Right Side**

$$Z = 19 \qquad Z = 18 + 1 = 19$$

→

$$A = 38 \qquad A = 38 + 0 = 38$$

Potassium-38 decays by positron production.

Practice Problem • Exercise 19.2

Supply the missing species in each of the following nuclear equations.

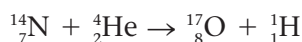
- a. ${}_{86}^{222}\text{Rn} \rightarrow {}_{84}^{218}\text{Po} + ?$ b. ${}_{8}^{15}\text{O} \rightarrow ? + {}_{+1}^0\text{e}$

Nuclear transformation

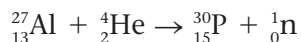
The change of one element into another

B. Nuclear Transformations

In 1919 Lord Rutherford observed the first **nuclear transformation**, *the change of one element into another*. He found that bombarding ${}_{7}^{14}\text{N}$ with α particles produced the nuclide ${}_{8}^{17}\text{O}$,



with a proton (${}_{1}^1\text{H}$) as another product. Fourteen years later, Irene Curie and her husband Frederick Joliot observed a similar transformation from aluminum to phosphorus:



where ${}_{0}^1\text{n}$ represents a neutron that is produced in the process.

Notice that in both of these cases the bombarding particle is a helium nucleus (an α particle). Other small nuclei, such as $^{12}_6\text{C}$ and $^{15}_7\text{N}$, can also be used to bombard heavier nuclei and cause transformations. However, because these positive bombarding ions are repelled by the positive charge of the target nucleus, the bombarding particle must be moving at a very high speed to penetrate the target. These high speeds are achieved in various types of *particle accelerators*.

Neutrons are also employed as bombarding particles to effect nuclear transformations. However, because neutrons are uncharged (and thus not repelled by a target nucleus), they are readily absorbed by many nuclei, producing new nuclides. The most common source of neutrons for this purpose is a fission reactor (see Section 19.3).

By using neutron and positive-ion bombardment, scientists have been able to extend the periodic table—that is, to produce chemical elements that are not present naturally. Prior to 1940, the heaviest known element was uranium ($Z = 92$), but in 1940, neptunium ($Z = 93$) was produced by neutron bombardment of $^{238}_{92}\text{U}$. The process initially gives $^{239}_{92}\text{U}$, which decays to $^{239}_{93}\text{Np}$ by β -particle production



In the years since 1940, the elements with atomic numbers 93 through 112, called the **transuranium elements**, have been synthesized. **Table 19.2** gives some examples of these processes.

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Irene Curie and Frederick Joliot

Table 19.2

Syntheses of Some of the Transuranium Elements

Neutron Bombardment	neptunium ($Z = 93$)	$^{238}_{92}\text{U} + {}^1_0\text{n} \rightarrow ^{239}_{92}\text{U} \rightarrow ^{239}_{93}\text{Np} + {}^0_{-1}\text{e}$
	americium ($Z = 95$)	$^{239}_{94}\text{Pu} + 2 {}^1_0\text{n} \rightarrow ^{241}_{94}\text{Pu} \rightarrow ^{241}_{95}\text{Am} + {}^0_{-1}\text{e}$
Positive-Ion Bombardment	curium ($Z = 96$)	$^{239}_{94}\text{Pu} + {}^4_2\text{He} \rightarrow ^{242}_{96}\text{Cm} + {}^1_0\text{n}$
	californium ($Z = 98$)	$^{242}_{96}\text{Cm} + {}^4_2\text{He} \rightarrow ^{245}_{98}\text{Cf} + {}^1_0\text{n}$ or $^{238}_{92}\text{U} + {}^{12}_6\text{C} \rightarrow ^{246}_{98}\text{Cf} + 4 {}^1_0\text{n}$
	rutherfordium ($Z = 104$)	$^{249}_{98}\text{Cf} + {}^{12}_6\text{C} \rightarrow ^{257}_{104}\text{Rf} + 4 {}^1_0\text{n}$
	dubnium ($Z = 105$)	$^{249}_{98}\text{Cf} + {}^{15}_7\text{N} \rightarrow ^{260}_{105}\text{Db} + 4 {}^1_0\text{n}$
	seaborgium ($Z = 106$)	$^{249}_{98}\text{Cf} + {}^{18}_8\text{O} \rightarrow ^{263}_{106}\text{Sg} + 4 {}^1_0\text{n}$

C. Detection of Radioactivity and the Concept of Half-life

Geiger-Müller counter (Geiger counter)

An instrument that measures the rate of radioactive decay by registering the ions and electrons produced as a radioactive particle passes through a gas-filled chamber

DID YOU KNOW

Geiger counters are commonly called survey meters.

The most familiar instrument for measuring radioactivity levels is the **Geiger-Müller counter**, or **Geiger counter** (see **Figure 19.2**). High-energy particles from radioactive decay produce ions when they travel through matter. The probe of the Geiger counter contains argon gas. The argon atoms have no charge, but they can be ionized by a rapidly moving particle.



That is, the fast-moving particle “knocks” electrons off some of the argon atoms. Although a sample of uncharged argon atoms does not conduct a current, the ions and electrons formed by the high-energy particle allow a current to flow momentarily, so a “pulse” of current flows every time a particle enters the probe. The Geiger counter detects each pulse of current, and these events are counted.

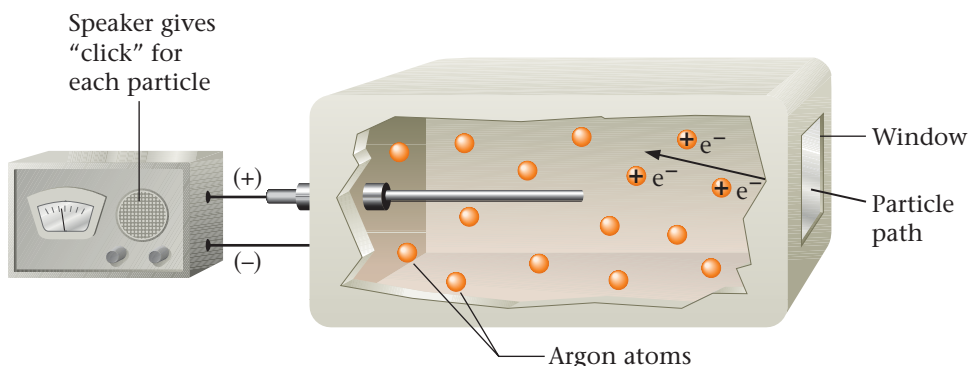
Active Reading Question

How does a Geiger counter work?

Figure 19.2

A schematic representation of a Geiger-Müller counter. The high-speed particle knocks electrons off argon atoms to form ions,

$\text{Ar} \xrightarrow{\text{Particle}} \text{Ar}^+ + e^-$
and a pulse of current flows.



Scintillation counter

An instrument that measures the rate of radioactive decay by sensing flashes of light that the radiation produces in a detector

Half-life

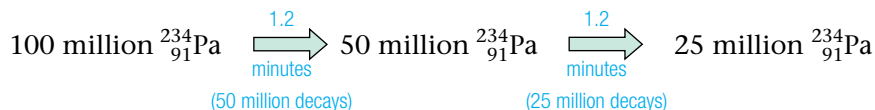
The time required for half of the original sample of radioactive nuclides to decay

A **scintillation counter** is another instrument often employed to detect radioactivity. This device uses a substance, such as sodium iodide, that gives off light when it is struck by a high-energy particle. A detector senses the flashes of light and counts the decay events.

Half-Life One important characteristic of a given type of radioactive nuclide is its half-life. The **half-life** is the *time required for half of the original sample of nuclei to decay*. For example, if a certain radioactive sample contains 1000 nuclei at a given time and 500 nuclei (half of the original number) 7.5 days later, this radioactive nuclide has a half-life of 7.5 days.

A given type of radioactive nuclide always has the same half-life. However, the various radioactive nuclides have half-lives that cover a tremendous range. For example, ${}_{91}^{234}\text{Pa}$, protactinium-234, has a half-life of 1.2 minutes, and ${}_{92}^{238}\text{U}$, uranium-238, has a half-life of 4.5×10^9 (4.5 billion) years. This means that a sample containing 100 million ${}_{91}^{234}\text{Pa}$ nuclei will have only 50 million ${}_{91}^{234}\text{Pa}$ nuclei in it (half of 100 million) after 1.2 minutes have passed. In another

1.2 minutes, the number of nuclei will decrease to half of 50 million, or 25 million nuclei.



This means that a sample of ${}_{91}^{234}\text{Pa}$ with 100 million nuclei will show 50 million decay events (50 million ${}_{91}^{234}\text{Pa}$ nuclei will decay) over a time of 1.2 minutes. By contrast, a sample containing 100 million ${}_{92}^{238}\text{U}$ nuclei will undergo 50 million decay events over 4.5 billion years. Therefore, ${}_{91}^{234}\text{Pa}$ shows much greater activity than ${}_{92}^{238}\text{U}$. We sometimes say that ${}_{91}^{234}\text{Pa}$ is “hotter” than ${}_{92}^{238}\text{U}$.

Thus, at a given moment, a radioactive nucleus with a short half-life is much more likely to decay than one with a long half-life.

EXAMPLE 19.3

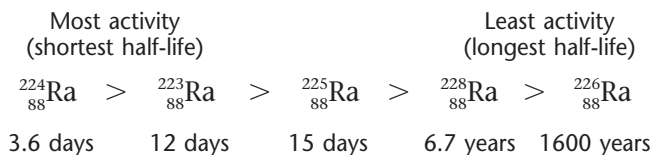
Understanding Half-life

Table 19.3 lists various radioactive nuclides of radium.

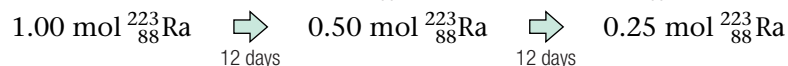
- Order these nuclides in terms of activity (from most to least decays per day).
- How long will it take for a sample containing 1.00 mol of ${}_{88}^{223}\text{Ra}$ to reach a point where it contains only 0.25 mol of ${}_{88}^{223}\text{Ra}$?

Solution

- The shortest half-life indicates the greatest activity (the most decays over a given period of time). Therefore, the order is



- In one half-life (12 days), the sample will decay from 1.00 mol of ${}_{88}^{223}\text{Ra}$ to 0.50 mol of ${}_{88}^{223}\text{Ra}$. In the next half-life (another 12 days), it will decay from 0.50 mol of ${}_{88}^{223}\text{Ra}$ to 0.25 mol of ${}_{88}^{223}\text{Ra}$.



Therefore, it will take 24 days (two half-lives) for the sample to change from 1.00 mol of ${}_{88}^{223}\text{Ra}$ to 0.25 mol of ${}_{88}^{223}\text{Ra}$.

Practice Problem • Exercise 19.3

Watches with numerals that “glow in the dark” were formerly made by including radioactive radium in the paint used to letter the watch faces. Assume that to make the numeral 3 on a given watch, a sample of paint containing 8.0×10^{-7} mol of ${}_{88}^{228}\text{Ra}$ was used. This watch was then put in a drawer and forgotten. Many years later someone finds the watch and wishes to know when it was made. Analyzing the paint, this person finds 1.0×10^{-7} mol of ${}_{88}^{228}\text{Ra}$ in the numeral 3. How much time elapsed between the making of the watch and the finding of the watch?

Hint: Use the half-life of ${}_{88}^{228}\text{Ra}$ from Table 19.3.

Table 19.3

The Half-lives for Some of the Radioactive Nuclides of Radium

Nuclide	Half-life
${}_{88}^{223}\text{Ra}$	12 days
${}_{88}^{224}\text{Ra}$	3.6 days
${}_{88}^{225}\text{Ra}$	15 days
${}_{88}^{226}\text{Ra}$	1600 years
${}_{88}^{228}\text{Ra}$	6.7 years



A watch dial with radium paint

CHEMISTRY EXPLORERS

Darleane Hoffman • 1926–

Darleane Hoffman was born on November 8, 1926, in Terril, Iowa. She began her college career at Iowa State College (now University) as an art major. She decided to switch her major from art to chemistry in part because of the influence of her first chemistry professor, also a woman.

Dr. Hoffman worked at Los Alamos National Laboratory for 31 years where she was part of the team that discovered plutonium-244 in nature, discovered element 100 (fermium), and confirmed the discovery of element 106 (which is named seaborgium after Hoffman's colleague Glenn Seaborg). She was awarded the prestigious Priestly Medal by the American Chemical Society in March 2000, the highest honor of the society, and is the only woman to receive the ACS Award for Nuclear Chemistry.

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SECTION 19.1

REVIEW QUESTIONS

- When balancing a nuclear equation, what two quantities must be conserved? How is this different from balancing other chemistry equations?
- Determine the net result to mass number and atomic number for each of the following processes.
 - alpha-particle production
 - beta-particle production
 - electron capture
 - positron production
- Complete each of the following nuclear equations by supplying the missing particle.
 - ${}_{13}^{23}\text{Al} \rightarrow ? + {}_{14}^{28}\text{Si}$
 - ${}_{24}^{56}\text{Cr} \rightarrow {}_{-1}^0\text{e} + ?$
 - ${}_{30}^{72}\text{Zn} \rightarrow {}_{-1}^0\text{e} + ?$
- What is the difference between nuclear decay and nuclear transformation reactions?
- Explain why one fourth of a given radioactive sample still remains at the end of two half-lives.
- The half-life of radium-255 is 15 days. Approximately what percentage of radium-255 is left after 3 months?

Application of Radioactivity

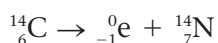
Objectives

- To learn how objects can be dated by radioactivity
- To understand the use of radiotracers in medicine

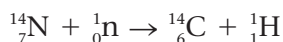
A. Dating by Radioactivity

Archaeologists, geologists, and others involved in reconstructing the ancient history of the earth rely heavily on the half-lives of radioactive nuclei to provide accurate dates for artifacts and rocks. A method for dating ancient articles made from wood or cloth is **radiocarbon dating**, or **carbon-14 dating**, a technique originated in the 1940s by Willard Libby, an American chemist who received the Nobel Prize for his efforts.

Radiocarbon dating is based on the radioactivity of $^{14}_6\text{C}$, which decays by β -particle production.



Carbon-14 is continuously produced in the atmosphere when high-energy neutrons from space collide with nitrogen-14.



Just as carbon-14 is continuously produced by this process, it continuously decomposes through β -particle production. Over the years, these two opposing processes have come into balance, causing the amount of $^{14}_6\text{C}$ present in the atmosphere to remain approximately constant.

Carbon-14 can be used to date wood and cloth artifacts because the $^{14}_6\text{C}$, along with the other carbon isotopes in the atmosphere, reacts with oxygen to form carbon dioxide. A living plant consumes this carbon dioxide in the photosynthesis process and incorporates the carbon, including $^{14}_6\text{C}$, into its molecules. As long as the plant lives, the $^{14}_6\text{C}$ content in its molecules remains the same as in the atmosphere because of the plant's continuous uptake of carbon. However, as soon as a tree is cut to make a wooden bowl or a flax plant is harvested to make linen, it stops taking in carbon. There is no longer a source of $^{14}_6\text{C}$ to replace that lost to radioactive decay, so the material's $^{14}_6\text{C}$ content begins to decrease.

Because the half-life of $^{14}_6\text{C}$ is known to be 5730 years, a wooden bowl found in an archaeological dig that shows a $^{14}_6\text{C}$ content of half that found in currently living trees is approximately 5730 years old. That is, because half of the $^{14}_6\text{C}$ present when the tree was cut has disappeared, the tree must have been cut one half-life of $^{14}_6\text{C}$ ago.

Active Reading Question

What must we assume about the atmospheric content of carbon-14 when using carbon dating?

Key Terms

- Radiocarbon dating (Carbon-14 dating)
- Radiotracers

Radiocarbon dating (carbon-14 dating)

A method for dating ancient wood or cloth on the basis of the radioactive decay of the carbon-14 nuclide

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Brigham Young researcher Scott Woodward takes a bone sample for carbon-14 dating at an archeological site in Egypt.

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Dating Diamonds

While connoisseurs of gems value the clearest possible diamonds, geologists learn the most from impure diamonds. Diamonds are formed in the earth's crust at depths of about 200 kilometers, where the high pressures and temperatures favor the most dense form of carbon. As the diamond is formed, impurities are sometimes trapped and these can be used to determine the diamond's date of "birth." One valuable dating impurity is $^{238}_{92}\text{U}$, which is radioactive and decays in a series of steps to $^{206}_{82}\text{Pb}$, which is stable (nonradioactive). Because the rate at which $^{238}_{92}\text{U}$ decays is known, determining how much $^{238}_{92}\text{U}$ has been converted to $^{206}_{82}\text{Pb}$ tells scientists the amount of time that has elapsed since the $^{238}_{92}\text{U}$ was trapped in the diamond as it was formed.

Using these dating techniques, Peter D. Kinney of Curtin University of Technology in Perth, Australia, and Henry O. A. Meyer of Purdue University in West Lafayette, Indiana, have recently identified the youngest diamond ever found. Discovered in Mbuji Mayi, Zaire, the diamond is 628 million years old,

far younger than all previously dated diamonds, which range from 2.4 to 3.2 billion years old.

The great age of all previously dated diamonds had caused some geologists to speculate that all diamond formation occurred billions of years ago. However, this "youngster" suggests that diamonds have formed throughout geologic time and are probably being formed right now in the earth's crust. These diamonds won't be seen for a long time, because diamonds typically remain deeply buried in the earth's crust for millions of years until they are brought to the surface by volcanic blasts called kimberlite eruptions.

It's good to know that eons from now there will be plenty of diamonds to mark the engagements of future couples.

The Hope Diamond



Radiotracer

A radioactive nuclide introduced into an organism and traced (followed) for diagnostic purposes

B. Medical Applications of Radioactivity

Although we owe the rapid advances of the medical sciences in recent decades to many causes, one of the most important has been the discovery and use of **radiotracers**—radioactive nuclides that can be introduced into organisms in food or drugs and subsequently *traced* by monitoring their radioactivity. For example, the incorporation of nuclides such as $^{14}_6\text{C}$ and $^{32}_{15}\text{P}$ into nutrients has yielded important information about how these nutrients are used to provide energy for the body.

Iodine-131 has proved very useful in the diagnosis and treatment of illnesses of the thyroid gland. Patients drink a solution containing a small amount of NaI that includes ^{131}I , and the uptake of the iodine by the thyroid gland is monitored with a scanner (**Figure 19.3**).

DID YOU KNOW

Nuclides used as radiotracers have short half-lives so that they disappear rapidly from the body.

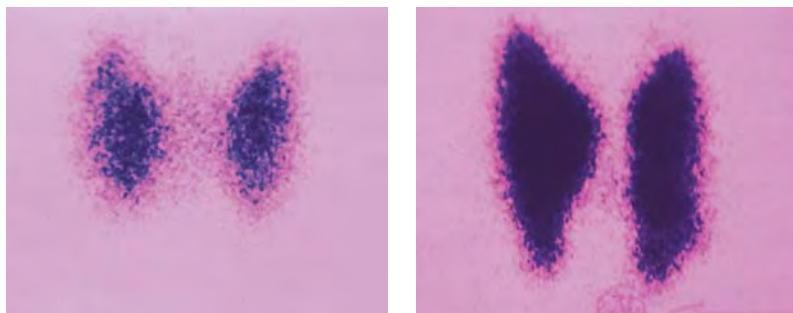


Figure 19.3

After consumption of Na^{131}I , the patient's thyroid is scanned for radioactivity levels to determine the efficiency of iodine absorption. (a) Scan of radioactive iodine in a normal thyroid (b) Scan of an enlarged thyroid

Thallium-201 can be used to assess the damage to the heart muscle in a person who has suffered a heart attack, because thallium becomes concentrated in healthy muscle tissue. Technetium-99, which is also taken up by normal heart tissue, is used for damage assessment in a similar way.

Radiotracers provide sensitive and nonsurgical methods for learning about biological systems, for detecting disease, and for monitoring the action and effectiveness of drugs. Some useful radiotracers are listed in **Table 19.4**.

Table 19.4

Some Radioactive Nuclides, Their Half-lives, and Their Medical Applications as Radiotracers*

Nuclide	Half-Life	Area of the Body Studied
^{131}I	8.1 days	thyroid
^{59}Fe	45.1 days	red blood cells
^{99}Mo	67 hours	metabolism
^{32}P	14.3 days	eyes, liver, tumors
^{51}Cr	27.8 days	red blood cells
^{87}Sr	2.8 hours	bones
^{99}Tc	6.0 hours	heart, bones, liver, lungs
^{133}Xe	5.3 days	lungs
^{24}Na	14.8 hours	circulatory system

*Z is sometimes not written when listing nuclides.

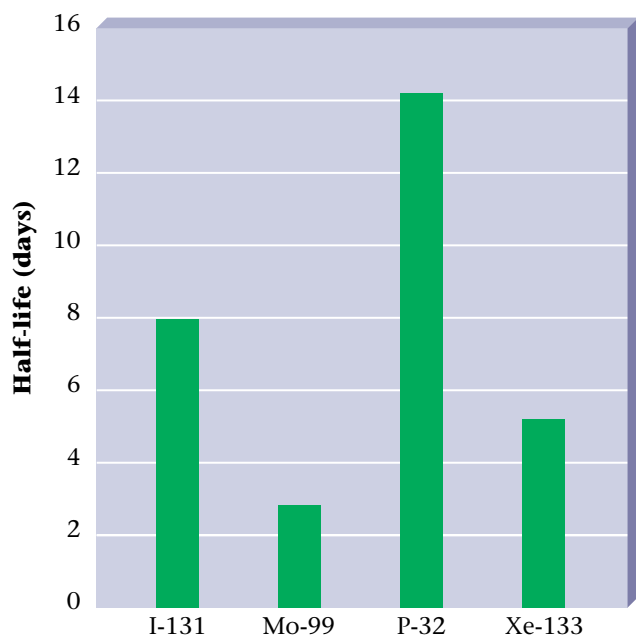
Active Reading Question

In general, how do the half-lives of radiotracers compare to the half-lives of some of the radioactive nuclides of radium?

SECTION 19.2 REVIEW QUESTIONS

- 1 Explain the basic ideas of radiocarbon dating in your own words.
- 2 Diamonds are made of carbon. In “Chemistry In Your World: Dating Diamonds,” however, it is stated that uranium and not carbon is used to determine when a diamond was formed. Why don’t scientists use carbon dating to determine this date?
- 3 How do half-lives of tracers compare to the half-life of carbon-14? Why would these nuclides be a better choice for medical applications than carbon-14?
- 4 Use the graph to answer the following questions.
 - a. Which radiotracer has the shortest half-life?
 - b. Which radiotracer has a half-life equal to the sum of the half-lives of Xe-133 and Mo-99?
 - c. If we started with equal amounts of each radiotracer, which would be present in the least amount after 15 days? Explain your answer.
 - d. If we started with equal amounts of each radiotracer, which would be present in the greatest amount after 15 days? Explain your answer.

Half-lives of Some Radiotracers



Using the Nucleus as a Source of Energy

Objectives

- To introduce fusion and fission as sources of energy
- To learn about nuclear fission
- To understand how a nuclear reactor works
- To learn about nuclear fusion
- To see how radiation damages human tissue

A. Nuclear Energy

The protons and the neutrons in atomic nuclei are bound together with forces that are much greater than the forces that bind atoms together to form molecules. In fact, the energies associated with nuclear processes are more than a million times those associated with chemical reactions. This potentially makes the nucleus a very attractive source of energy.

Because medium-sized nuclei contain the strongest binding forces ($^{56}_{26}\text{Fe}$ has the strongest binding forces of all), there are two types of nuclear processes that produce energy:

1. Combining two light nuclei to form a heavier nucleus. This process is called **fusion**.
2. Splitting a heavy nucleus into two nuclei with smaller mass numbers. This process is called **fission**.

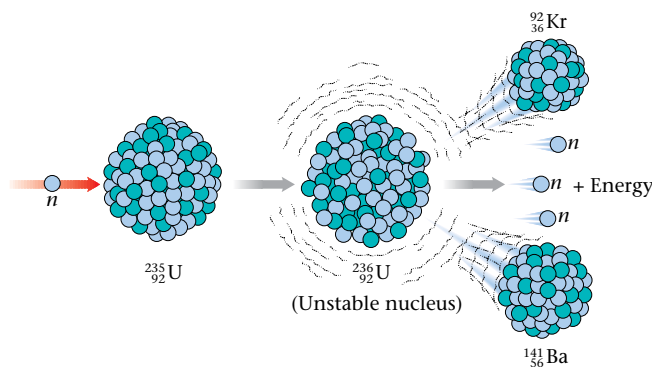
As we will see, these two processes can supply amazing quantities of energy with relatively small masses of materials consumed.

B. Nuclear Fission

Nuclear fission was discovered in the late 1930s when $^{235}_{92}\text{U}$ nuclides bombarded with neutrons were observed to split into two lighter elements.



This process, shown below, releases 2.1×10^{13} joules of energy per mole of $^{235}_{92}\text{U}$. Compared with what we get from typical fuels, this is a huge amount of energy. For example, the fission of 1 mol of $^{235}_{92}\text{U}$ produces about *26 million times* as much energy as the combustion of 1 mol of methane.



Key Terms

- Fusion
- Fission
- Chain reaction
- Critical mass
- Breeder reactors

Fusion

The process of combining two light nuclei to form a heavier, more stable nucleus

Fission

The process of splitting a heavy nucleus into two more stable nuclei with smaller mass numbers

CRITICAL THINKING ?

Nuclear fission processes can provide useful energy but can also be dangerous.

What if Congress decided to outlaw all processes that involve fission? How would it change our society?

Upon capturing a neutron, the ${}^{235}_{92}\text{U}$ nucleus undergoes fission to produce two lighter nuclides, more neutrons (typically three), and a large amount of energy.

The process shown on the previous page is only one of the many fission reactions that ${}^{235}_{92}\text{U}$ can undergo. In fact, over 200 different isotopes of 35 different elements have been observed among the fission products of ${}^{235}_{92}\text{U}$.

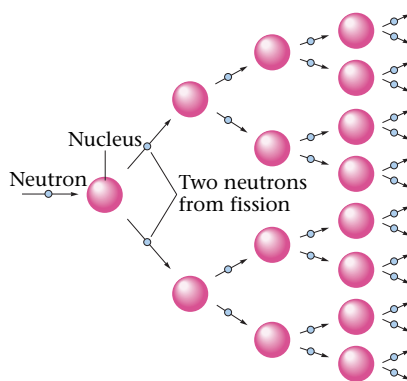
In addition to the product nuclides, neutrons are produced in the fission reactions of ${}^{235}_{92}\text{U}$. As these neutrons fly through the solid sample of uranium, they may collide with other ${}^{235}_{92}\text{U}$ nuclei, producing additional fission events. Each of these fission events produces more neutrons that can, in turn, produce the fission of more ${}^{235}_{92}\text{U}$ nuclei. Because each fission event produces neutrons, the process can be self-sustaining. We call it a **chain reaction**.

Chain reaction

A self-sustaining fission process caused by the production of neutrons that proceed to split other nuclei

Critical mass

The mass of fissionable material required to produce a chain reaction



For the fission process to be self-sustaining, at least one neutron from each fission event must go on to split another nucleus. If, on the average, *less than one* neutron causes another fission event, the process dies out. If *exactly one* neutron from each fission event causes another fission event, the process sustains itself at the same level and is said to be *critical*. If *more than one* neutron from each fission event causes another fission event, the process rapidly escalates and the heat build-up causes a violent explosion.

To achieve the critical state, a certain mass of fissionable material, called the **critical mass**, is needed. If the sample is too small, too many neutrons escape before they have a chance to cause a fission event, and the process stops.

Active Reading Question

What is the relationship between *critical mass* and *chain reaction*?

During World War II, the United States carried out an intense research effort called the Manhattan Project to build a bomb based on the principles of nuclear fission. This program produced the fission bomb, which was used with devastating effect on the cities of Hiroshima and Nagasaki in 1945. Basically, a fission bomb operates by suddenly combining subcritical masses, which results in rapidly escalating fission events that produce an explosion of incredible intensity.

Top Ten Countries Producing Electricity by Nuclear Power (in order of total nuclear output)

Country	Percentage of Country's Total Power Production
United States	21.9
France	77.4
Japan	34.0
Germany	30.3
Russia	13.1
Canada	16.0
Ukraine	43.8
United Kingdom	26.0
Sweden	52.4
South Korea	35.8

C. Nuclear Reactors

Because of the tremendous energies involved, fission has been developed as an energy source to produce electricity in reactors where controlled fission can occur. The resulting energy is used to heat water to produce steam that runs turbine generators, in much the same way that a coal-burning power plant generates energy by heating water to produce steam. A schematic diagram of a nuclear power plant is shown in **Figure 19.4**.

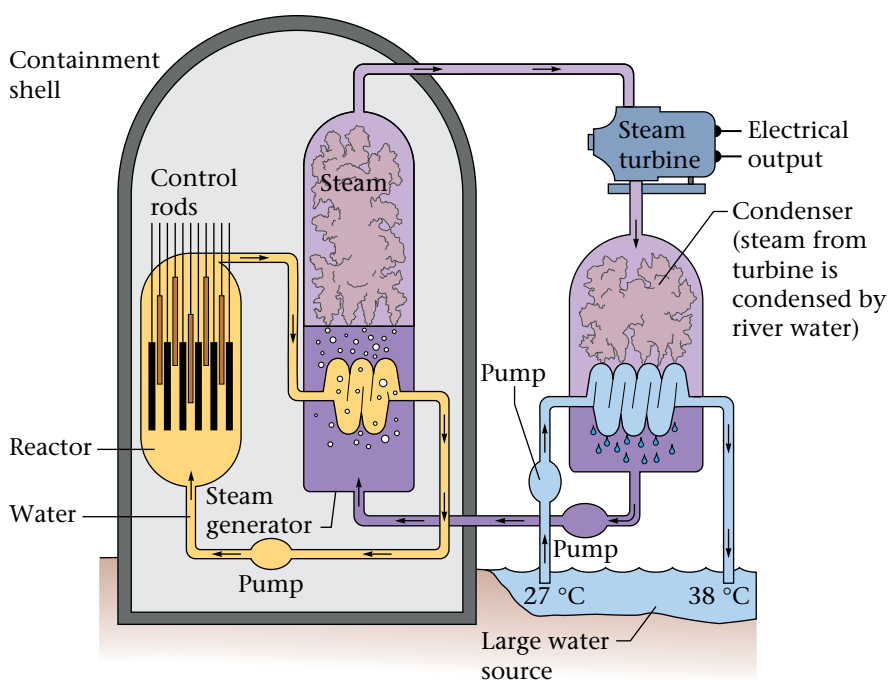


Figure 19.4

A schematic diagram of a nuclear power plant. The energy from the fission process is used to boil water, producing steam for use in a turbine-driven generator. Cooling water from a lake or river is used to condense the steam after it leaves the turbine.

In the reactor core (see **Figure 19.5**), uranium that has been enriched to approximately 3% $^{235}_{92}\text{U}$ (natural uranium contains only 0.7% $^{235}_{92}\text{U}$) is housed in metal cylinders. A *moderator* surrounding the cylinders slows down the neutrons so that the uranium fuel can capture them more efficiently. *Control rods*, composed of substances (such as cadmium) that absorb neutrons, are used to regulate the power level of the reactor. The reactor is designed so that if a malfunction occurs, the control rods are automatically inserted into the core to absorb neutrons and stop the reaction. A liquid (usually water) is circulated through the core to extract the heat generated by the energy of fission. This heat energy is then used to change water to steam, which runs turbines that in turn run electrical generators.

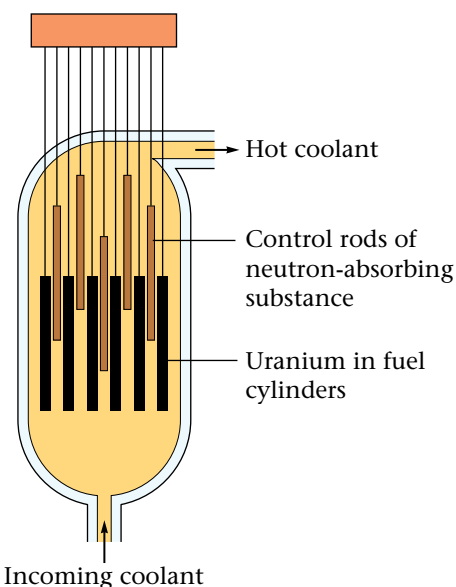


Figure 19.5

A schematic of the reactor core

CELEBRITY CHEMICAL Uranium (U)

Uranium was first discovered in 1789 as a component of pitchblende. Its name came from the planet Uranus, which had been discovered only shortly before. Uranium, which is not a rare element in the earth's crust (it is more abundant than tin), is widely scattered on earth. The most easily accessible deposits of uranium are found in the United States, Canada, South Africa, and Australia. Although uranium is used in small amounts as a coloring agent for glass and ceramics, its major use is as a nuclear fuel in fission reactors.

Uranium consists principally of $^{238}_{92}\text{U}$ mixed with smaller amounts of $^{235}_{92}\text{U}$. Both isotopes are radioactive, although ^{235}U decays more rapidly than ^{238}U by about a factor of six. The ^{235}U isotope splits into two smaller nuclei (fission) when it is bombarded by neutrons, which makes it useful for producing energy in nuclear reactions.

Because fission of ^{235}U requires carefully controlled conditions, scientists were shocked to find that a natural nuclear reactor existed millions of years ago

in Oklo, Gabon. Apparently, conditions were just right in a uranium deposit at Oklo to allow natural fission to occur. Water that was trapped in clay mixed with the ore and acted as the moderator, slowing down the neutrons to promote efficient fission. As the "reactor" heated up, it apparently drove off some of the water. This loss caused the fission to slow down, thus preventing an explosion. As the area cooled down, the water returned, leading to an increase in fission again. This speeding up and slowing down of the natural reactor apparently occurred repeatedly over many years until conditions changed enough to cause the reactor to stop. It seems clear that nature invented the nuclear reactor millions of years before humans thought of the idea.

Uranium-238 kernels for use in a reactor



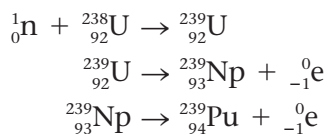
Although the concentration of $^{235}_{92}\text{U}$ in the fuel elements is not great enough to allow an explosion such as that which occurs in a fission bomb, a failure of the cooling system can lead to temperatures high enough to melt the reactor core. This means that the building housing the core must be designed to contain the core even in the event of such a "melt-down." A great deal of controversy now exists about the efficiency of the safety systems in nuclear power plants. Accidents such as the one at the Three Mile Island facility in Pennsylvania in 1979 and the one at Chernobyl in the Soviet Union in 1986 have led many people to question the wisdom of continuing to build fission-based power plants.

Breeder Reactors

One potential problem facing the nuclear power industry is the limited supply of $^{235}_{92}\text{U}$. Some scientists believe we have nearly depleted the uranium deposits that are rich enough in $^{235}_{92}\text{U}$ to make the production of fissionable fuel economically feasible. Because of this possibility, reactors have been developed in which fissionable fuel is actually produced while the reactor runs. In these **breeder reactors**, the major component of natural uranium, nonfissionable $^{238}_{92}\text{U}$, is changed to fissionable $^{239}_{94}\text{Pu}$. The reaction involves absorption of a neutron, followed by production of two β particles.

Breeder reactor

A nuclear reactor in which fissionable fuel is produced while the reactor runs



As the reactor runs and ${}_{92}^{235}\text{U}$ is split, some of the excess neutrons are absorbed by ${}_{92}^{238}\text{U}$ to produce ${}_{94}^{239}\text{Pu}$. The ${}_{94}^{239}\text{Pu}$ is then separated out and used to fuel another reactor. Such a reactor thus “breeds” nuclear fuel as it operates.

Although breeder reactors are now used in Europe, the United States is proceeding slowly with their development because much controversy surrounds their use. One problem involves the hazards that arise in handling plutonium, which is very toxic and flames on contact with air.

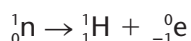
CHEMISTRY IN YOUR WORLD

Connection to Astronomy

Making Elements in Stars

How did all the matter around us originate? One scientific answer to this question is a theory called *stellar nucleosynthesis*—the formation of nuclei in stars.

Many scientists believe that our universe originated as a cloud of neutrons that became unstable and produced an immense explosion, giving this model its name: *the big bang theory*. The model postulates that after the initial explosion, neutrons decomposed into protons and electrons,



which eventually combined to form clouds of hydrogen. Over the eons, gravitational forces caused many of these hydrogen clouds to contract and heat up sufficiently to reach temperatures at which proton fusion began to occur, releasing large quantities of energy. When the tendency to expand in response to the heat from fusion and the tendency to contract in response to the forces of gravity are balanced, a stable young star such as our sun can form.

Eventually, when the star’s supply of hydrogen is exhausted, the core of the star again contracts, with further heating, until it reaches temperatures at which fusion of helium nuclei can occur, leading to the formation of ${}_{6}^{12}\text{C}$ and ${}_{8}^{16}\text{O}$ nuclei. In turn, when the supply of helium nuclei runs out, further contraction and heating occur, until the fusion of heavier nuclei takes place. This process occurs

repeatedly, forming heavier and heavier nuclei, until iron nuclei are formed. Because the iron nucleus has the greatest binding forces of all, iron nuclei do not release excess energy when they fuse to form heavier nuclei. This means that there is no further fusion energy to sustain the star, and it cools to a small, dense *white dwarf*.

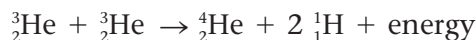
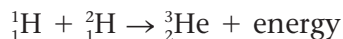
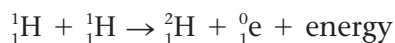
The evolution just described is characteristic of small and medium-sized stars. Much larger stars, however, become unstable at some time during their evolution and undergo a *supernova explosion*. In this explosion, nuclei with medium mass numbers are fused to produce heavy elements. Also in the explosion, light nuclei capture neutrons. These neutron-rich nuclei then produce β particles, increasing their atomic number with each β decay. This eventually leads to nuclei with large atomic numbers. In fact, almost all nuclei beyond iron are thought to originate in supernova explosions. The debris of such an explosion contains a wide variety of elements, and it is thought that this debris might eventually form a solar system such as our own.

A view of a portion of the Cygnus Loop supernova remnant taken by the Hubble Space Telescope

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D. Nuclear Fusion

The process of combining two light nuclei—called **nuclear fusion**—produces even more energy per mole than does nuclear fission. In fact, stars produce their energy through nuclear fusion. Our sun, which presently consists of 73% hydrogen, 26% helium, and 1% other elements, gives off vast quantities of energy from the fusion of protons to form helium. One possible scheme for this process is



Intense efforts are under way to develop a feasible fusion process because of the ready availability of many light nuclides (deuterium, ${}^2_1\text{H}$, in seawater, for example) that can serve as fuel in fusion reactors. However, initiating the fusion process is much more difficult than initiating fission. The forces that bind nucleons together to form a nucleus become effective only at *very small* distances (approximately 10^{-13} cm), so for two protons to bind together and thereby release energy, they must get very close together. But protons, because they are identically charged, repel each other. This suggests that to get two protons (or two deuterons) close enough to bind together (the strong nuclear binding force is *not* related to charge), they must be “shot” at each other at speeds high enough to overcome their repulsion from each other. The repulsive forces between two ${}^1_1\text{H}$ nuclei are so great that temperatures of about 40 million K are thought to be necessary. Only at these temperatures are the nuclei moving fast enough to overcome the repulsions.

Currently, scientists are studying two types of systems to produce the extremely high temperatures required: high-powered lasers and heating by electric currents. At present many technical problems remain to be solved, and it is not clear whether either method will prove useful.

E. Effects of Radiation

Everyone knows that being hit by a train is a catastrophic event. The energy transferred in such a collision is very large. In fact, any source of energy is potentially harmful to organisms. Energy transferred to cells can break chemical bonds and cause malfunctioning of the cell systems. This fact is behind our present concern about maintaining the ozone layer in the earth’s upper atmosphere, which screens out high-energy ultraviolet radiation arriving from the sun.

Radioactive elements, which are sources of high-energy particles, are also potentially hazardous. However, the effects are usually quite subtle, because even though high-energy particles are involved, the quantity of energy actually deposited in tissues *per decay event* is quite small. The resulting damage is no less real, but the effects may not be apparent for years.

Information

${}^2_1\text{H}$ particles are called deuterons.

Top Ten Elements in the Sun (by mass)

Hydrogen
Helium
Oxygen
Carbon
Iron
Neon
Nitrogen
Silicon
Magnesium
Sulfur

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A solar flare erupts from the surface of the sun.

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Nuclear Waste Disposal

Our society does not have an impressive record for safe disposal of industrial wastes. We have polluted our water and air, and some land areas have become virtually uninhabitable because of the improper burial of chemical wastes. As a result, many people are wary about the radioactive wastes from nuclear reactors. The potential threats of cancer and genetic mutations make these materials especially frightening.

Because of its controversial nature, most of the nuclear waste generated over the past 50 years has been placed in temporary storage. However, in 1982 the U.S. Congress passed the Nuclear Waste Policy Act, which established a timetable for choosing and preparing sites for the deep underground disposal of radioactive materials.

The tentative disposal plan calls for incorporation of the spent nuclear fuel into blocks of glass that will be packed in corrosion-resistant metal containers and then buried in a deep, stable rock formation indicated by the rock layers in **Figure 19.6**.

There are indications that this method will isolate the waste until the radioactivity decays to safe levels. Some reassuring evidence comes from the natural fission “reactor” that was discovered at Oklo in Gabon, Africa (see “Celebrity Chemical: Uranium” on page 686). Spawned about 2 billion years ago when uranium in ore deposits there formed a critical mass, this “reactor” produced fission and fusion products for several thousand years. Although some of these products have migrated away from the site in the intervening 2 billion years, most have stayed in place.

In 1999 the Waste Isolation Pilot Plant (WIPP) in New Mexico began receiving nuclear waste. This facility employs tunnels carved into the salt beds of an ancient ocean. Once a repository room becomes full, the salt will collapse around the waste, encapsulating it forever.

Another waste depository, under Yucca Mountain in Nevada, is being contemplated. For nearly two decades, this area has been studied to determine its suitability for storage of high-level radioactive wastes. At present it looks to be a long time before this issue is settled.

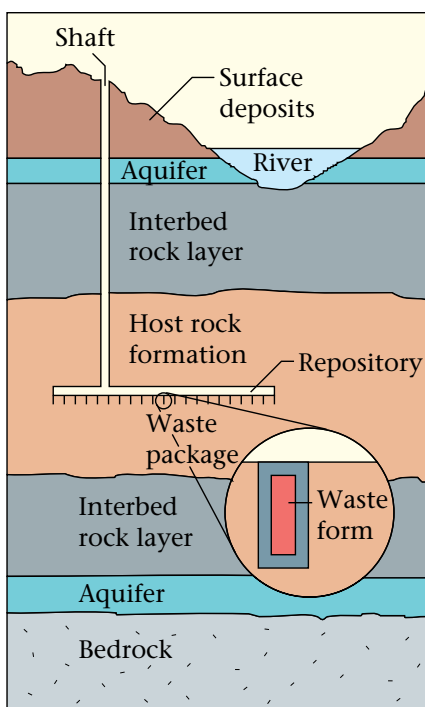


Figure 19.6

A schematic diagram for the tentative plan for deep underground isolation of nuclear waste. The disposal system would consist of a waste package buried in tunnels like the one shown above.

Radiation damage to organisms can be classified as somatic or genetic damage. *Somatic damage* is damage to the organism itself, resulting in sickness or death. The effects may appear almost immediately if a massive dose of radiation is received; for smaller doses, damage may appear years later, usually in the form of cancer. *Genetic damage* is damage to the genetic machinery of reproductive cells, creating problems that often afflict the offspring of the organism.

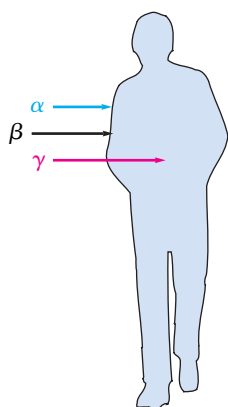


Figure 19.7

Radioactive particles and rays vary greatly in penetrating power. Gamma rays are by far the most penetrating.

Rem

A unit of radiation dosage that accounts for both the energy of the dose and its effectiveness in causing biological damage

The Factors Determining Biological Effects of Radiation

1. *The energy of the radiation.* The higher the energy content of the radiation, the more damage it can cause.
2. *The penetrating ability of the radiation.* The particles and rays produced in radioactive processes vary in their ability to penetrate human tissue: γ rays are highly penetrating, β particles can penetrate approximately 1 cm, and α particles are stopped by the skin (see **Figure 19.7**).
3. *The ionizing ability of the radiation.* Because ions behave quite differently from neutral molecules, radiation that removes electrons from molecules in living tissues seriously disturbs their functions. The ionizing ability of radiation varies dramatically. For example, γ rays penetrate very deeply but cause only occasional ionization. On the other hand, α particles, although they are not very penetrating, are very effective at causing ionization and produce serious damage. Therefore, the ingestion of a producer of α particles, such as plutonium, is particularly damaging.
4. *The chemical properties of the radiation source.* When a radioactive nuclide is ingested, its capacity to cause damage depends on how long it remains in the body. For example, both $^{85}_{36}\text{Kr}$ and $^{90}_{38}\text{Sr}$ are β -particle producers. Because krypton, being a noble gas, is chemically inert, it passes through the body quickly and does not have much time to do damage. Strontium, on the other hand, is chemically similar to calcium. It can collect in bones, where it may cause leukemia and bone cancer.

Because of the differences in the behavior of the particles and rays produced by radioactive decay, we have invented a unit called the **rem** that indicates the danger the radiation poses to humans.

Table 19.5 shows the physical effects of short-term exposure to various doses of radiation, and **Table 19.6** gives the sources and amounts of the radiation to which a typical person in the United States is exposed each year. Note that natural sources contribute about twice as much as human activities do to the total exposure. However, although the nuclear industry contributes only a small percentage of the total exposure, controversy surrounds nuclear power plants because of their *potential* for creating radiation hazards. These hazards arise mainly from two sources: accidents allowing the release of radioactive materials, and improper disposal of the radioactive products in spent fuel elements.

Active Reading Question

What human activity typically exposes people to the greatest amount of radiation per year?

Table 19.5

Effects of Short-Term Exposures to Radiation

Dose (rem)	Clinical Effect
0–25	nondetectable
25–50	temporary decrease in white blood cell counts
100–200	strong decrease in white blood cell counts
500	death of half the exposed population within 30 days after exposure

Table 19.6

Typical Radiation Exposures for a Person Living in the United States (1 millirem = 10^{-3} rem)

Source	Exposure (millirems/year)
cosmic	50
from the earth	47
from building materials	3
in human tissues	21
inhalation of air	5
<i>Total from natural sources</i>	126
X-ray diagnosis	50
radiotherapy X rays, radioisotopes	10
internal diagnosis and therapy	1
nuclear power industry	0.2
luminous watch dials, TV tubes, industrial wastes	2
radioactive fallout	4
<i>Total from human activities</i>	97
<i>Total</i>	193 = 0.193 rems



SECTION 19.3

REVIEW QUESTIONS

- 1 What is the difference between fission and fusion?
- 2 If there is no danger of a nuclear power plant having a nuclear explosion (if the reaction got out of control), why do we need to build containment buildings for the core?
- 3 Explain the role of *critical mass* for a chain reaction.
- 4 How is a typical nuclear power plant in the United States different from a breeder reactor commonly used in Europe?
- 5 Why is initiating fusion much more difficult than initiating fission?
- 6 What are the factors that determine the biological effects of radiation?



RESEARCH LINKS
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Key Terms

19.1 Nucleons (protons and neutrons)
 Atomic number (Z)
 Mass number (A)
 Isotopes
 Nuclide
 Radioactive
 Beta particle
 Nuclear equation
 Alpha particle
 Alpha-particle production
 Beta-particle production
 Gamma ray
 Positron
 Positron production
 Electron capture
 Decay series
 Nuclear transformation
 Transuranium elements
 Geiger–Müller counter
 Scintillation counter
 Half-life

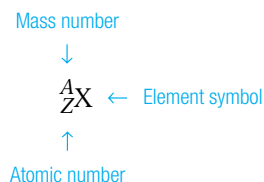
19.2 Radiocarbon dating
 (Carbon-14 dating)
 Radiotracers

Key Ideas

19.1 Radioactivity

■ Radioactivity is the spontaneous decomposition of a nucleus to form another nucleus.

- Atomic number (Z) and atomic mass number (A) are both conserved in a radioactive decay.



- Types of radioactive decay
 - ◆ Alpha (α) particle production
 - ◆ Beta (β) particle production
 - ◆ Gamma (γ) ray production
 - ◆ Electron capture
 - ◆ Positron (β^+) production
 - Often a series of decays occurs until a stable nucleus is formed.
 - Radioactivity is detected by instruments such as the Geiger–Müller counter.
 - Half-life for a radioactive nucleus is the time required for half of the sample to decay.
- Nuclear transformations change one atom into another by particle bombardment.

19.2 Application of Radioactivity

- Radiocarbon or carbon-14 dating—objects containing carbon can be dated by measuring their ^{14}C content.
- Radiotracers are radioactive nuclides that can be introduced into an organism and traced by observing their radioactive decay.

19.3 Fusion

Fission

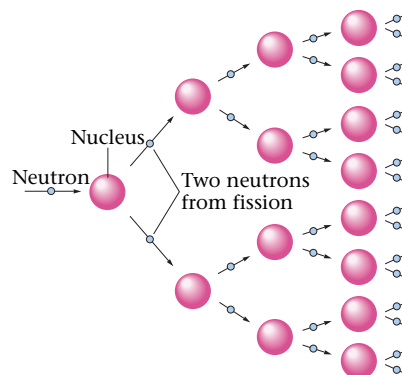
Chain reaction

Critical mass

Breeder reactors

19.3 Using the Nucleus as a Source of Energy

- Energy can be obtained from the nucleus in two ways.
 - Fusion is the combining of two light nuclei to form a heavier nucleus.
 - Fission is the splitting of a heavy nucleus into two lighter nuclei.
 - ◆ Current nuclear reactors use fission to produce electrical energy.



- Nuclear radiation can cause damage.
 - Somatic—damage to an organism
 - Genetic—damage to an organism's offspring
 - Biological effects of radiation depend on the following:
 - ◆ Its energy
 - ◆ Its penetrating ability
 - ◆ The chemical properties of the source

All exercises with **blue numbers** have answers in the back of this book.

19.1 Radioactivity

A. Radioactive Decay

- Does the nucleus of an atom have much of an effect on the atom's chemical properties? Explain.
- How large is a typical atomic nucleus, and how does the size of the nucleus of an atom compare with the overall size of the atom?
- What do the *atomic number* and the *mass number* of a nucleus represent? Use a specific element as an example to illustrate your answer.
- What are *isotopes*? Do the isotopes of an element have the same atomic number? Do they have the same mass number? Explain.
- Using Z to represent the atomic number and A to represent the mass number, give the general symbol for a nuclide of element X . Give also a specific example of the use of such symbolism.
- When an unstable nucleus produces an alpha particle, by how many units does the atomic number of the nucleus change? Does the atomic number increase or decrease?
- When an unstable nucleus produces a beta particle, by how many units does the atomic number of the nucleus change? Does the atomic number increase or decrease?
- What does a *gamma ray* represent? Is a gamma ray a particle? Is there a change in mass or atomic number when a nucleus produces only a gamma ray?
- What is a *positron*? What are the mass number and charge of a positron? How do the mass number and atomic number of a nucleus change when the nucleus produces a positron?
- What do we mean when we say a nucleus has undergone an *electron capture* process? What type of electron is captured by the nucleus in this process?
- Naturally occurring magnesium consists primarily of three isotopes, of mass numbers 24, 25, and 26. How many protons does each of these nuclides contain? How many neutrons does each of these nuclides contain? Write nuclear symbols for each of these isotopes.

- Complete each of the following nuclear equations by supplying the missing particle.
 - ${}_{85}^{196}\text{At} \rightarrow {}_2^4\text{He} + ?$
 - ${}_{84}^{208}\text{Po} \rightarrow {}_2^4\text{He} + ?$
 - ${}_{86}^{210}\text{Rn} \rightarrow {}_2^4\text{He} + ?$
- Complete each of the following nuclear equations by supplying the missing particle.
 - ${}_{80}^{201}\text{Hg} + ? \rightarrow {}_{79}^{201}\text{Au}$
 - ${}_{82}^{210}\text{Pb} \rightarrow {}_{83}^{210}\text{Bi} + ?$
 - $? \rightarrow {}_{84}^{210}\text{Po} + {}_{-1}^0\text{e}$
- Write a balanced nuclear equation for the decay of each of the following nuclides to produce a beta particle.
 - ${}_{53}^{136}\text{I}$
 - ${}_{51}^{133}\text{Sb}$
 - ${}_{49}^{117}\text{In}$
- Write a balanced nuclear equation for the decay of each of the following nuclides to produce an alpha particle.

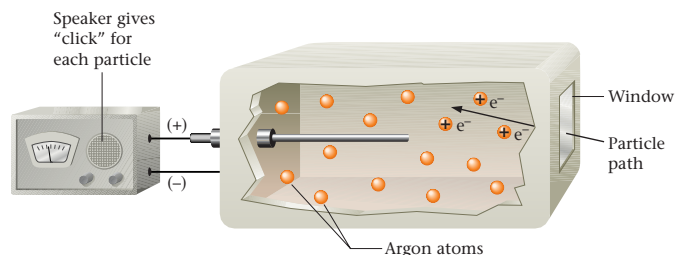
a. ${}_{88}^{226}\text{Ra}$	c. ${}_{94}^{239}\text{Pu}$
b. ${}_{86}^{222}\text{Rn}$	d. ${}_{4}^8\text{Be}$

B. Nuclear Transformations

- What does a *nuclear transformation* represent? How is a nuclear transformation performed?
- Why are particle accelerators needed for bombardment processes?
- What are the elements with atomic numbers greater than 92 called? How have these elements been prepared?
- Write a balanced nuclear equation for the bombardment of ${}_{13}^{27}\text{Al}$ with alpha particles to produce ${}_{15}^{30}\text{P}$ and a neutron.

C. Detection of Radioactivity and the Concept of Half-life

- Describe the operation of a Geiger counter. How does a Geiger counter detect radioactive particles? How does a scintillation counter differ from a Geiger counter?



- 21.** What is the *half-life* of a radioactive nucleus? Does a given type of nucleus always have the same half-life? Do nuclei of different elements have the same half-life?
- 22.** What do we mean when we say that one radioactive nucleus is “hotter” than another? Which element would have more decay events over a given period of time?
- 23.** The following isotopes (listed with their half-lives) have been used in the medical and biological sciences. Arrange these isotopes in order of their relative decay activities: ^3H (12.2 years), ^{24}Na (15 hours), ^{131}I (8 days), ^{60}Co (5.3 years), ^{14}C (5730 years).
- 24.** A list of several important radioactive nuclides is given in Table 19.4. Arrange these nuclides in order of their relative decay activities.
- 25.** For the first three isotopes of radium listed in Table 19.3, given a starting amount of 1000 mg, *estimate* the approximate amount of each isotope remaining after 1 month.
- 26.** Technetium-99 has been used as a radiographic agent in bone scans (^{99}Tc is absorbed by bones). If ^{99}Tc has a half-life of 6.0 hours, what fraction of an administered dose of 100 μg of ^{99}Tc remains in a patient’s body after 2.0 days?

19.2 Applications of Radioactivity

A. Dating by Radioactivity

- 27.** What nuclide is commonly used in the dating of artifacts?
- 28.** How is ^{14}C produced in the atmosphere? Write a balanced equation for this process.
- 29.** Why is it assumed that the amount of ^{14}C in the atmosphere remains constant?
- 30.** Why does an ancient wood or cloth artifact contain less ^{14}C than contemporary or more recently fabricated articles made of similar materials?

B. Medical Applications of Radioactivity

- 31.** What is a *radiotracer*? Which tracers have been used to study the conversion of nutrients into energy in living cells?
- 32.** List four radioisotopes used in medical diagnosis or treatment, and give their nuclear symbols. Discuss why the isotopes you choose are particularly well suited to their uses.

19.3 Using the Nucleus as a Source of Energy

A. Nuclear Energy

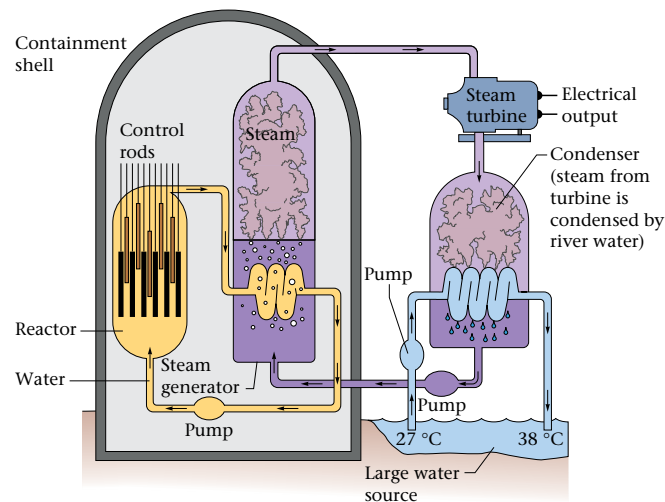
- 33.** How do the forces that hold an atomic nucleus together compare in strength with the forces between atoms in a molecule?
- 34.** Define the terms *nuclear fission* and *nuclear fusion*. Which process results in the production of a heavier nucleus? Which results in the production of smaller nuclei?

B. Nuclear Fission

- 35.** How do the energies released by nuclear processes compare in magnitude with the energies of ordinary chemical processes?
- 36.** Write an equation for the fission of ^{235}U by bombardment with neutrons.
- 37.** How is it possible for the fission of ^{235}U , once started, to lead to a chain reaction?
- 38.** What does it mean to say that fissionable material possesses a *critical mass*? Can a chain reaction occur when a sample has less than the critical mass?

C. Nuclear Reactors

- 39.** Describe the purpose of each of the major components of a nuclear reactor (moderator, control rods, containment, cooling liquid, and so on).



- 40.** Can a nuclear explosion take place in a reactor? Is the concentration of fissionable material used in reactors large enough for this?
- 41.** What is a *melt-down* and how can it occur? Most nuclear reactors use water as the cooling liquid. Is there any danger of a steam explosion if the reactor core becomes overheated?

42. In a *breeder* nuclear reactor, nonfissionable $^{238}_{92}\text{U}$ is converted to fissionable $^{239}_{94}\text{Pu}$. Write nuclear equations showing this transformation.

D. Nuclear Fusion

43. What is the nuclear *fusion* of small nuclei? How does the energy released by fusion compare in magnitude with that released by fission?

44. What are some reasons why a practical fusion reactor has not yet been developed?

45. Why is the development of nuclear fusion reactors desirable?

E. Effects of Radiation

46. Although the energy transferred per event when a living creature is exposed to radiation is small, why is such exposure dangerous?

47. Explain the difference between *somatic* damage from radiation and *genetic* damage. Which type causes immediate damage to the exposed individual?

48. Describe the relative penetrating powers of alpha, beta, and gamma radiation.

49. Explain why, although gamma rays are far more penetrating than alpha particles, the latter are actually more likely to cause damage to an organism. Which radiation is more effective at causing ionization of biomolecules?

50. How do the *chemical properties* of radioactive nuclei (as opposed to the nuclear decay they undergo) influence the degree of damage they do to an organism?

51. Although radiation has been measured in many terms, the *rem* is most commonly used in discussions of the exposure of human beings to radiation. Using Table 19.6, discuss the sources of radiation to which humans are exposed naturally, and indicate how the total exposure compares with the “acceptable” (nondetectable clinical effect) level of exposure given in Table 19.5.

Critical Thinking

52. The decay series from uranium-238 to lead-206 is indicated in Figure 19.1. For each step of the process indicated in the figure, specify what type of particle is produced by the particular nucleus involved at that point in the series.

53. Each of the following isotopes has been used medically for the purpose indicated. Suggest reasons why the particular element might have been chosen for this purpose.

- cobalt-57, for study of the body's use of vitamin B_{12}
- calcium-47, for study of bone metabolism
- iron-59, for study of red blood cell function
- mercury-197, for brain scans before the CAT scan became available

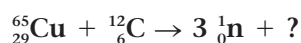
54. Discuss some of the problems associated with the storage and disposal of the waste products of the nuclear industry.

55. Aluminum exists in several isotopic forms, including $^{27}_{13}\text{Al}$, $^{28}_{13}\text{Al}$, and $^{29}_{13}\text{Al}$. Indicate the number of protons and the number of neutrons in each of these isotopes.

56. How have $^{131}_{53}\text{I}$ and $^{201}_{81}\text{Tl}$ been used in medical diagnosis? Why are these particular nuclides especially well suited for this purpose?

57. What is a *breeder* nuclear reactor? What difficulties with such reactors have led to their not yet being used in the United States for the generation of electricity?

- 1 Atoms with identical atomic numbers but different mass numbers are termed
- neutrons.
 - nucleons.
 - isotopes.
 - alpha particles.
- 2 Complete the nuclear equation by supplying the missing particle.

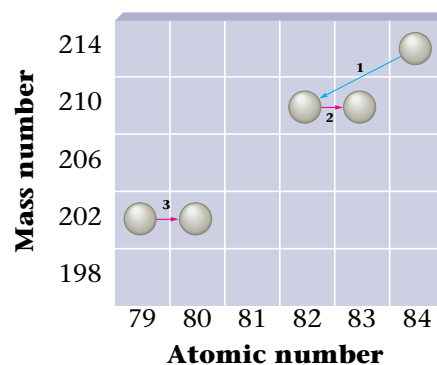


- ${}_{32}^{74}\text{Ge}$
 - ${}_{32}^{76}\text{Ge}$
 - ${}_{35}^{76}\text{Br}$
 - ${}_{35}^{74}\text{Br}$
- 3 The nuclide ${}^{99}\text{Tc}$ has a half-life of 6.0 hours. How much of a 200.0 g sample remains after 1 day?
- 100.0 g
 - 50.00 g
 - 25.00 g
 - 12.50 g
- 4 Nuclides used in medicine that can be useful in diagnosis and treatment of areas of the body are called
- gamma rays.
 - radiotracers.
 - positrons.
 - rems.
- 5 Splitting a heavy nuclide into two nuclei with smaller mass numbers is called
- critical mass.
 - fusion.
 - fission.
 - electron capture.

Write balanced nuclear equations for each of the processes in Questions 6–8.

- 6 ${}_{90}^{232}\text{Th}$ produces an alpha particle.
- 7 ${}_{91}^{234}\text{Pa}$ produces a beta particle.
- 8 ${}_{15}^{30}\text{P}$ produces a positron.

For questions 9–11, use the following figure.



- 9 Which of the following must occur for transformation 2 in the figure?
- alpha-particle production
 - beta-particle production
 - electron capture
- 10 Which of the following must occur for transformation 1 in the figure?
- alpha-particle production
 - beta-particle production
 - electron capture
- 11 Which of the following must occur for transformation 3 in the figure?
- alpha-particle production
 - beta-particle production
 - electron-capture
- 12 Carbon-14 dating is a technique which can provide accurate dates for artifacts and rocks. Answer the following questions concerning carbon-14 dating.
- Carbon-14 decays by beta particle production. Write the nuclear equation for this process.
 - Carbon-14 is produced in the atmosphere when high-energy neutrons collide with nitrogen-14 (in a 1:1 mol ratio), also producing a hydrogen nucleus. Write the nuclear equation for this process.
 - Explain how carbon-14 is used to determine the dates of wooden artifacts.

LOOKING AHEAD

20.1 Saturated Hydrocarbons

- A. Carbon Bonding
- B. Alkanes
- C. Structural Formulas and Isomerism
- D. Naming Alkanes
- E. Petroleum
- F. Reactions of Alkanes

20.2 Unsaturated Hydrocarbons

- A. Alkenes and Alkynes
 - Reactions of Alkenes
- B. Aromatic Hydrocarbons
- C. Naming Aromatic Compounds
 - Monosubstituted Benzenes
 - Disubstituted Benzenes

20.3 Introduction to Functional Groups and Alcohols

- A. Functional Groups
- B. Alcohols
- C. Properties and Uses of Alcohols

20.4 Additional Organic Compounds

- A. Aldehydes and Ketones
- B. Naming Aldehydes and Ketones
- C. Carboxylic Acids and Esters
- D. Polymers

The sails of these paragliders are made of nylon, an organic substance.



IN YOUR LIFE

The study of carbon-containing compounds and their properties is called **organic chemistry**. The industries based on organic substances have truly revolutionized our lives. In particular, the development of polymers, such as nylon for fabrics; Velcro for fasteners; Kevlar for the composites used in exotic cars, airplanes, and bicycles; and polyvinyl chloride (PVC) for pipes, siding, and toys has produced a wondrous new world.

Organic chemistry plays a vital role in our quest to understand living systems. Beyond that, the synthetic fibers, plastics, artificial sweeteners, and medicines that we take for granted are products of industrial organic chemistry. Finally, the energy which we rely on so heavily to power our civilization is based mostly on the combustion of organic materials found in coal and petroleum.

Image not available for electronic use.

Please refer to the image in the textbook.

A close-up photo of Velcro, a synthetic organic material used for fasteners

WHAT DO YOU KNOW?

 Prereading Questions

1. What do you know about methane, propane, and butane?
2. What are some uses for petroleum?
3. What first comes to mind when you hear the term *polymer*?

SECTION 20.1

Saturated Hydrocarbons

Key Terms

- Organic chemistry
- Biomolecule
- Hydrocarbons
- Saturated
- Unsaturated
- Alkanes
- Unbranched hydrocarbons (normal or straight chain)
- Structural isomerism
- Petroleum
- Natural gas
- Combustion reactions
- Substitution reactions
- Dehydrogenation reactions

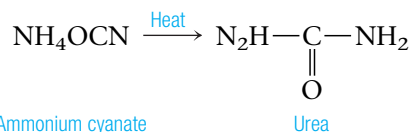
Objectives

- To understand the types of bonds formed by the carbon atom
- To learn about the alkanes
- To learn about structural isomers
- To learn to draw structural formulas
- To learn to name alkanes and substituted alkanes
- To learn about the composition and uses of petroleum
- To learn about the chemical reactions of alkanes

Two Group 4 elements, carbon and silicon, form the basis of most natural substances. Silicon, with its great affinity for oxygen, forms chains and rings containing Si—O—Si bonds to produce the silica and silicates that form the basic structures for most rocks, sands, and soils. Therefore, silicon compounds are the fundamental inorganic materials of the earth. What silicon is to the geological world, carbon is to the organic or biological world.

Carbon has the unusual ability to bond strongly to itself, forming long chains or rings of carbon atoms. In addition, carbon forms strong bonds to other nonmetals such as hydrogen, nitrogen, oxygen, sulfur, and the halogens. Because of these bonding properties, an extraordinary number of carbon compounds exists; several million are now known, and the number continues to grow rapidly. Among these many compounds are the **biomolecules**, those molecules that make possible the maintenance and reproduction of life.

Although a few compounds of carbon, such as the oxides of carbon and carbonates, are considered to be inorganic substances, the vast majority of carbon compounds are designated as organic compounds; compounds that typically contain chains or rings of carbon atoms. Originally, the distinction between inorganic and organic substances was based on whether they were produced by living systems. For example, until the early nineteenth century, it was believed that organic compounds had some kind of “life force” and could be synthesized only by living organisms. This misconception was dispelled in 1828 when the German chemist Friedrich Wöhler (1800–1882) prepared urea from the inorganic salt ammonium cyanate by simple heating.



Urea is a component of urine, so it is clearly an organic material formed by living things, yet here was clear evidence that it could be produced in the laboratory as well.

Because organic chemistry is such a vast subject, we can introduce it only briefly in this text. We will begin with the simplest class of organic compounds, the hydrocarbons, and then show how most other organic compounds can be considered to be derived from hydrocarbons.

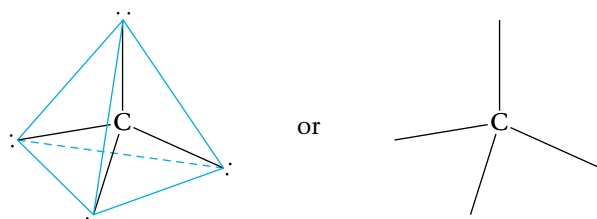
Biomolecule

A molecule that functions in maintaining and/or reproducing life

A. Carbon Bonding

The reason so many carbon-containing compounds exist is because carbon forms strong bonds to itself and to many other elements. A carbon atom can form bonds to a maximum of four other atoms; these can be either carbon atoms or atoms of other elements. One of the hardest, toughest materials known is diamond, a form of pure carbon in which each carbon atom is bound to four other carbon atoms.

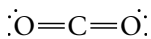
One of the most familiar compounds of carbon is methane, CH_4 , the main component of natural gas. The methane molecule consists of a carbon atom with four hydrogen atoms bound to it in a tetrahedral fashion. That is, as predicted by the VSEPR model (see Chapter 12), the four pairs of bonding electrons around the carbon have minimum repulsions when they are located at the corners of a tetrahedron.



This leads to the structure for CH_4 shown in **Figure 20.1**.

When carbon has four atoms bound to it, these atoms will always have a tetrahedral arrangement about the carbon.

Carbon can bond to fewer than four elements by forming one or more multiple bonds. Recall that a multiple bond involves the sharing of more than one pair of electrons. For example, a *double bond* involves sharing two pairs of electrons, as in carbon dioxide:

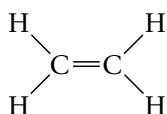


A *triple bond* involves sharing three pairs of electrons, as in carbon monoxide:



Note that carbon is bound to two other atoms in CO_2 and to only one other atom in CO .

Multiple bonding also occurs in organic molecules. Ethylene, C_2H_4 , has a double bond:



In this case each carbon is bound to three other atoms (one C atom and two H atoms). A molecule with a triple bond is acetylene, C_2H_2 :



Here each carbon is bound to two other atoms (one carbon atom and one hydrogen atom).

More than any other element, carbon has the ability to form chains of atoms, as illustrated by the structures of propane and butane shown in **Figure 20.3**. In these compounds each carbon atom is bound to four atoms in a tetrahedral fashion.

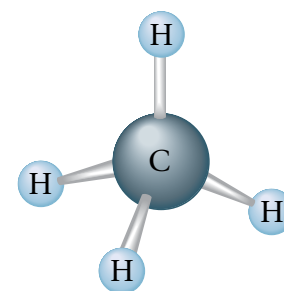


Figure 20.1
Methane is a tetrahedral molecule.

i n f o r m a t i o n

CO_2 and CO are classed as inorganic substances.

B. Alkanes

Hydrocarbon

A compound consisting only of hydrogen and carbon

Saturated

Describes a hydrocarbon in which all carbon-carbon bonds are single bonds

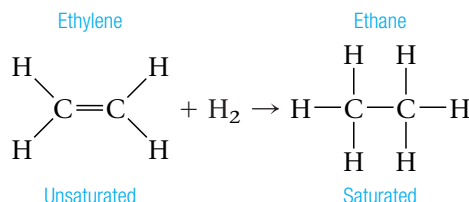
Unsaturated

Describes a hydrocarbon containing carbon-carbon multiple bonds

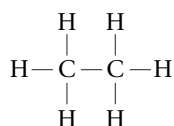
Alkane

Another name for saturated hydrocarbons; alkanes have the general formula C_nH_{2n+2} .

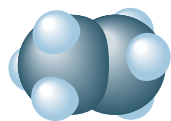
Hydrocarbons, as the name indicates, are compounds composed of carbon and hydrogen. Those whose carbon-carbon bonds are all single bonds are said to be **saturated**, because each carbon is bound to four atoms, the maximum number. Hydrocarbons containing carbon-carbon multiple bonds are described as being **unsaturated**, because the carbon atoms involved in a multiple bond can bond to one or more additional atoms. This is shown by the *addition* of hydrogen to ethylene.



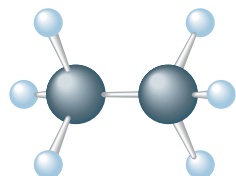
Note that each carbon in ethylene is bonded to three atoms (one carbon and two hydrogens) but can bond to one additional atom after one bond of the carbon-carbon double bond is broken. This leads to ethane, a saturated hydrocarbon (each carbon atom is bonded to four atoms).



(a)



(b)



(c)

Figure 20.2

(a) The Lewis structure of ethane, C_2H_6 . The molecular structure of ethane represented by (b) a space-filling model and (c) a ball-and-stick model.

Active Reading Question

What is the fewest number of carbons an unsaturated hydrocarbon may have?

Saturated hydrocarbons are called **alkanes**. The simplest alkane is *methane*, CH_4 , which has a tetrahedral structure (see Figure 20.1). The next alkane, the one containing two carbon atoms, is *ethane*, C_2H_6 , shown in **Figure 20.2**. Note that each carbon atom in ethane is bonded to four atoms.

The next two members of the series are *propane*, with three carbon atoms and the formula C_3H_8 , and *butane*, with four carbon atoms and the formula C_4H_{10} . These molecules are shown in **Figure 20.3**. Again, these are saturated hydrocarbons (alkanes); each carbon is bonded to four atoms.

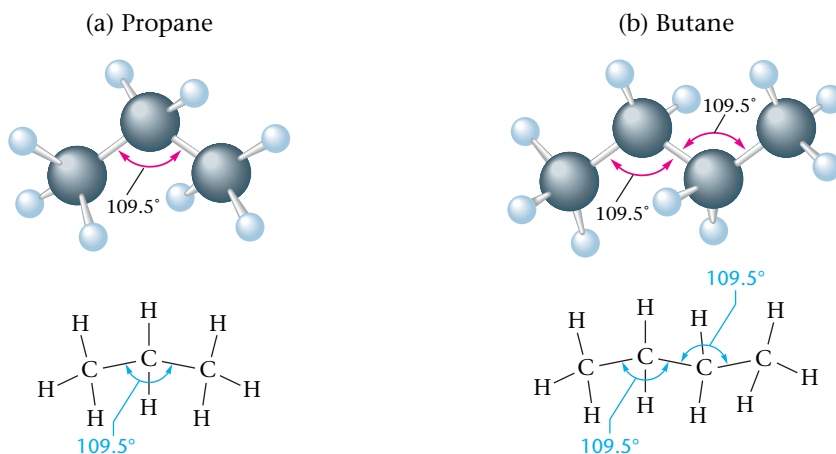
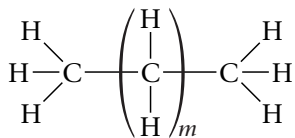


Figure 20.3

The structures of (a) propane and (b) butane

Alkanes in which the carbon atoms form long “strings” or chains are called **normal**, **straight-chain**, or **unbranched hydrocarbons**. As Figure 20.3 illustrates, the chains in normal alkanes are not really straight but zig-zag, because the tetrahedral C—C—C angle is 109.5° . The normal alkanes can be represented by the structure:



where m is a whole number. Note that each member is obtained from the previous one by insertion of a *methylene*, CH_2 , group. We can condense the structural formulas by omitting some of the C—H bonds. For example, the general formula for normal alkanes shown above can be condensed to:



EXAMPLE 20.1

Writing Formulas for Alkanes

Give the formulas for the normal (or straight-chain) alkanes with six and eight carbon atoms.

Solution

Six Atoms: The alkane with six carbon atoms can be written as

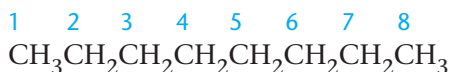


which can be condensed to



Notice that the molecule contains fourteen hydrogen atoms in addition to the six carbon atoms. Therefore, the formula is C_6H_{14} .

Eight Atoms: The alkane with eight carbons is



which can be written in condensed form as



This molecule has eighteen hydrogens. The formula is C_8H_{18} .

Practice Problem • Exercise 20.1

Give the molecular formulas for the alkanes with 10 and 15 carbon atoms.

The first ten straight-chain alkanes are shown in **Table 20.1**. Note that all alkanes can be represented by the general formula $\text{C}_n\text{H}_{2n+2}$, where n represents the number of carbon atoms. For example, nonane, which has nine carbon atoms, is represented by $\text{C}_9\text{H}_{(2 \times 9)+2}$, or C_9H_{20} . The formula $\text{C}_n\text{H}_{2n+2}$ reflects the fact that each carbon in the chain has two hydrogen atoms except the two end carbons, which have three each. Thus the number of hydrogen atoms present is twice the number of carbon atoms plus two (for the extra two hydrogen atoms on the ends).

Table 20.1

Formulas of the First Ten Straight-Chain Alkanes

Name	Condensed Formula (C_nH_{2n+2})	Extended Formula
methane	CH ₄	CH ₄
ethane	C ₂ H ₆	CH ₃ CH ₃
propane	C ₃ H ₈	CH ₃ CH ₂ CH ₃
<i>n</i> -butane	C ₄ H ₁₀	CH ₃ CH ₂ CH ₂ CH ₃
<i>n</i> -pentane	C ₅ H ₁₂	CH ₃ CH ₂ CH ₂ CH ₂ CH ₃
<i>n</i> -hexane	C ₆ H ₁₄	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃
<i>n</i> -heptane	C ₇ H ₁₆	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃
<i>n</i> -octane	C ₈ H ₁₈	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃
<i>n</i> -nonane	C ₉ H ₂₀	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃
<i>n</i> -decane	C ₁₀ H ₂₂	CH ₃ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₂ CH ₃

EXAMPLE 20.2

Using the General Formula for Alkanes

Show that the alkane with 15 carbon atoms can be represented in terms of the general formula C_nH_{2n+2} .

Solution

In this case $n = 15$. The formula is $C_{15}H_{2(15)+2}$, or $C_{15}H_{32}$.

Structural isomerism

Describes what occurs when two molecules have the same atoms but different bonds

C. Structural Formulas and Isomerism

Butane and all larger alkanes exhibit structural isomerism. **Structural isomerism** occurs when two molecules have the same atoms but different bonds. That is, the molecules have the same formulas but different arrangements of the atoms. For example, butane can exist as a straight-chain molecule (normal butane, or *n*-butane) or with a branched-chain structure (called isobutane), as shown in **Figure 20.4**. Because of their different structures, these structural isomers have different properties.

Active Reading Question

Which alkanes cannot have structural isomers? Why?

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Plastic Tasters

The scene: A lab near Brussels, Belgium; room temperature 21 °C; humidity 60%. A woman in a white lab coat is sitting in a private cubicle in front of a row of glasses each filled with a liquid. The woman rinses out her mouth with pure water and then begins the tasting. Exotic coffees? Premium teas? No. The woman is tasting water that has had various plastics soaking in it.

What is going on here? This scene takes place every Tuesday through Friday at Exxon Chemicals' Machelen Research Center in Belgium, where trained tasters (called *organolepticians*) sniff and taste plastic-marinated water. Their job is to make sure that you never detect the taste or smell of plastic in your favorite cheese, meat, pizza, or other packaged foods. Today's shoppers expect clear plastic wrapping and plastic containers that keep food fresh, safe, and ready to use right out of the package. They also expect it to taste just like it came from the farm or from mom's kitchen.

Ensuring that plastic packaging does not impart a taste to the food stored inside is quite a challenge. That's where organoleptics—the art and science of tasting and smelling—comes in. To test a new type of plastic, the material is sealed into an aluminum bag with mineral water. After about a week the water is tasted by a panel of highly trained organolepticians. The sample is also checked for odors. Although highly sophisticated instruments are used to test the samples as well, the human nose and tongue have been found to be the most sensitive and reliable detectors of odors and tastes. For the plastic manufacturers, no taste means good business.

Produce wrapped in plastic

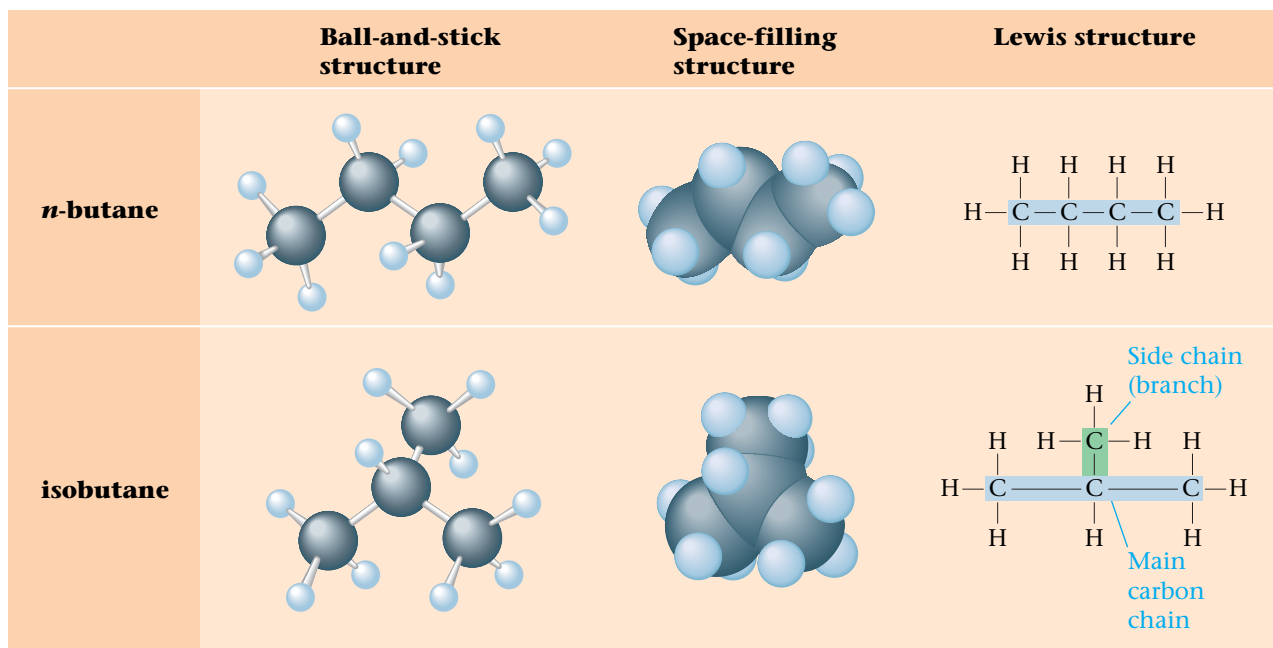


Figure 20.4
The structural isomers of C_4H_{10}

EXAMPLE 20.3

Drawing Structural Isomers of Alkanes

Draw the structural isomers of pentane, C_5H_{12} .

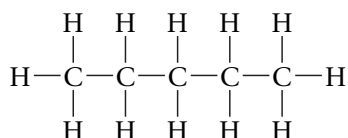
Solution

To find the isomeric structures for pentane, C_5H_{12} , we first write the straight carbon chain and then add the hydrogen atoms.

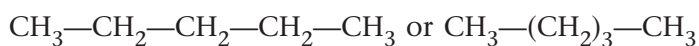
1. The straight-chain structure has the five carbon atoms in a row.



We can now add the H atoms.

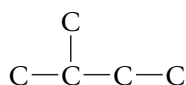


This can be written in shorthand form as

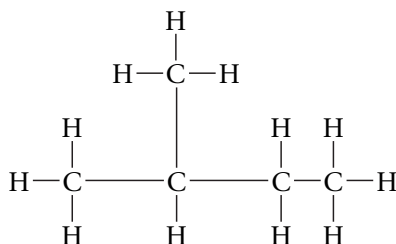


and is called *n*-pentane.

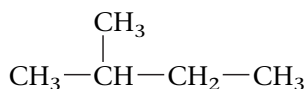
2. Next we remove one C atom from the main chain and bond it to the second carbon in the chain.



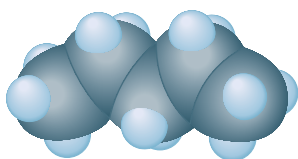
Then we put on the H atoms, so that each carbon has four bonds.



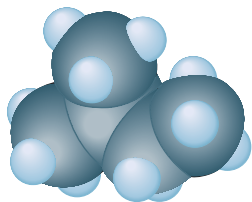
This structure can be represented as



and is called isopentane.

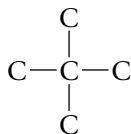


n-Pentane

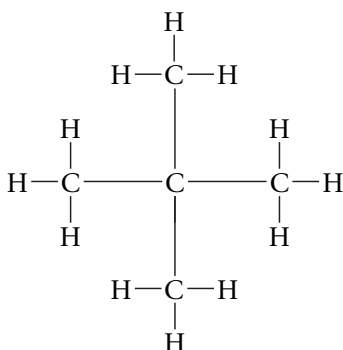


Isopentane

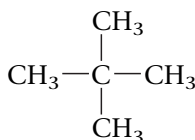
3. Finally, we take two carbons out of the chain to give the arrangement



Adding the H atoms gives



which can be written in shorthand form as

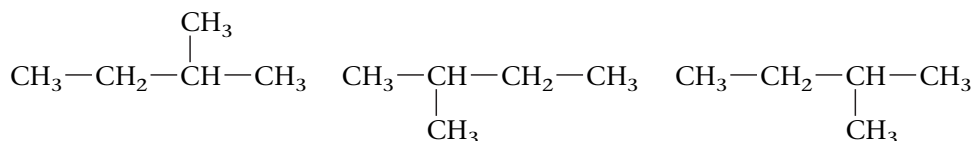


This molecule is called neopentane.

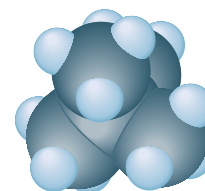
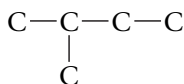
The space-filling models for these molecules are shown in the margin.

Note that all of these molecules have the formula C_5H_{12} , as required.

Also note that the structures



which might at first appear to be additional isomers, are actually identical to structure 2. All three of these structures have exactly the same skeleton of carbons as the structure shown in part 2. All of these structures have four carbons in the chain with one carbon on the side:



Neopentane

D. Naming Alkanes

Because there are literally millions of organic compounds, it would be impossible to remember common names for all of them. Just as we did in Chapter 4 for inorganic compounds, we must learn a systematic method for naming organic compounds. We will first consider the principles applied in naming alkanes and then summarize them as a set of rules.

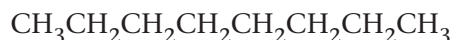
1. The first four members of the alkane series are called methane, ethane, propane, and butane. The names of the alkanes beyond butane are obtained by adding the suffix *-ane* to the Greek root for the number of carbon atoms.

Number	Greek Root
5	<i>pent</i>
6	<i>hex</i>
7	<i>hept</i>
8	<i>oct</i>
9	<i>non</i>
10	<i>dec</i>

Information

The complete name for this alkane is *n*-octane, the *n* indicating a straight-chain alkane.

Therefore, the alkane



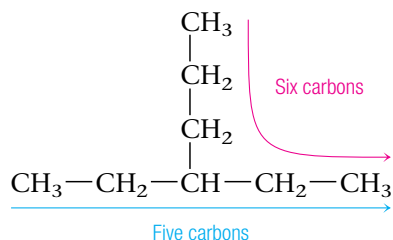
which has eight carbons in the chain, is called octane.

oct - ane

Tells us
there are
eight carbons

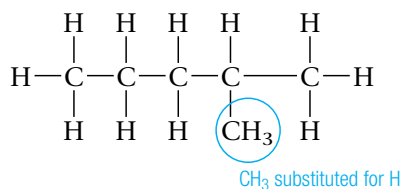
Tells us
it is an
alkane

2. For a branched hydrocarbon, the longest continuous chain of carbon atoms gives the root name for the hydrocarbon. For example, in the alkane



the longest continuous chain contains six carbon atoms. The specific name of this compound is not important at this point, but it will be named as a hexane (indicating a six-carbon chain).

3. Alkanes lacking one hydrogen atom can be attached to a hydrocarbon chain in place of one hydrogen atom. For example, the molecule



can be viewed as a pentane (five-carbon chain) in which one hydrogen atom has been replaced by a —CH₃ group, which is a methane, CH₄, molecule with a hydrogen removed.

Active Reading Question

What does the suffix *-yl* mean when naming an organic compound?

Let's Review

Rules for Naming Alkanes

1. Find the longest continuous chain of carbon atoms. This chain (called the parent chain) determines the base alkane name.
2. Number the carbons in the parent chain, starting at the end closest to any branching (the first alkyl substituent). When a substituent occurs the same number of carbons from each end, use the next substituent (if any) to determine from which end to start numbering.
3. Using the appropriate name for each alkyl group, specify its position on the parent chain with a number.
4. When a given type of alkyl group occurs more than once, attach the appropriate prefix (*di-* for two, *tri-* for three, and so on) to the alkyl name.
5. The alkyl groups are listed in alphabetical order, *disregarding any prefix*.

CHEMISTRY IN YOUR WORLD

Consumer Connection

www.Dot.Safety

Elbert Dysart Botts may have saved your life, even though you have probably never heard of him. Botts, who was born in rural Missouri in 1893, was a chemist who specialized in paints and who taught chemistry at San Jose State University in California. Eventually he went to work for the California Department of Transportation (Caltrans), where he tried unsuccessfully to find a paint for road stripes that would be both durable and visible even on rainy, gloomy nights. Discouraged by his failure, Botts decided that a reflective pavement marker might be a better solution. He and his team came up with a glass-and-ceramic marker that was visible at night for over 100 meters. The biggest problem was how to attach the markers to the pavement. Eventually Botts settled on a fast-setting epoxy

developed by one of his former chemistry students, which would bond the reflectors to virtually any solid surface.

The first "Botts Dots" were put in service in California in 1966. Today, hundreds of millions of them all over the world protect drivers from straying out of their lanes.

"Botts Dots" highlight traffic lanes.

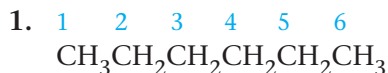


EXAMPLE 20.4**Naming Isomers of Alkanes**

Draw the structural isomers for the alkane C_6H_{14} , and give the systematic name for each one.

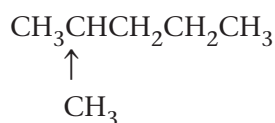
Solution

We proceed systematically, starting with the longest chain and then rearranging the carbons to form the shorter, branched chains.

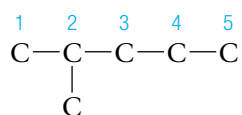


This alkane has six carbons all in the same continuous chain, so we call it hexane or, more properly, *n*-hexane, indicating that all of the carbon atoms are in the same chain.

2. We now take one carbon out of the main chain and make it a methyl substituent. This gives the molecule

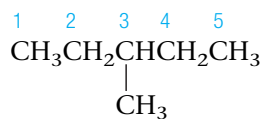


The carbon skeleton is as follows:



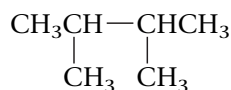
Because the longest chain has five carbons, the base name is pentane. We have numbered the chain from the left, starting closest to the substituent, a methyl group. We indicate the position of the methyl group on the chain by the number 2, the number of the carbon to which it is attached. So the name is 2-methylpentane. Note that if we numbered the chain from the right end, the methyl group would be on carbon 4. We want the smallest possible number, so the numbering shown is correct.

3. The methyl substituent can also be on the number-3 carbon:

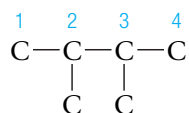


The name is 3-methylpentane. We have now exhausted all possibilities for placing a single methyl group on pentane.

4. Next we take two carbons out of the original six-member chain.

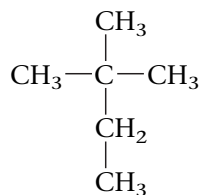


The carbon skeleton is

**Information**

Note that placing the $-CH_3$ group on carbon 4 gives the same molecule as placing it on carbon 2.

In searching for more isomers, we might try a structure such as

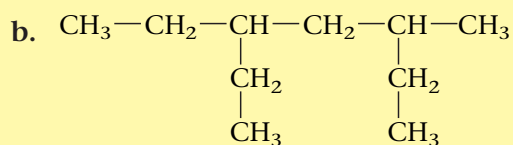
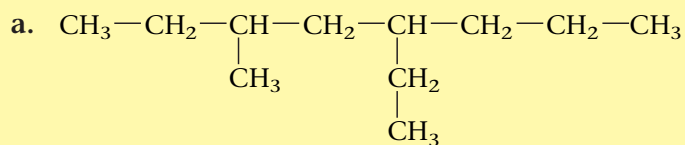


As we have drawn it, this molecule might appear to be a propane. However, the molecule has a longest chain of four atoms (look vertically), so the correct name is 2,2-dimethylbutane.

Thus there are five distinct structural isomers of C_6H_{14} : *n*-hexane, 2-methylpentane, 3-methylpentane, 2,3-dimethylbutane, and 2,2-dimethylbutane.

 **Practice Problem • Exercise 20.4**

Name the following molecules.



So far we have learned how to name a compound by examining its structural formula. We must also be able to do the reverse: to write the structural formula from the name.

EXAMPLE 20.5

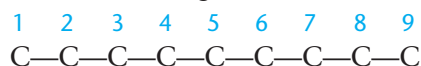
Writing Structural Isomers from Names

Write the structural formula for each of the following compounds.

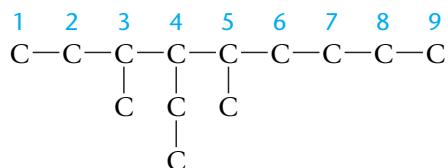
- 4-ethyl-3,5-dimethylnonane
- 4-*tert*-butylheptane

Solution

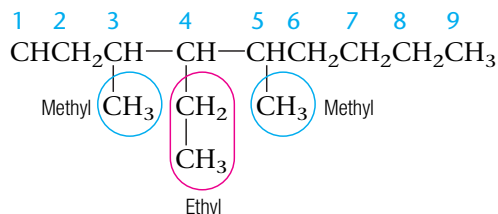
- The root name nonane signifies a nine-carbon chain. Therefore, we have the following main chain of carbons:



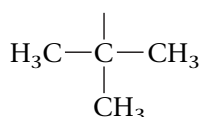
The name indicates an ethyl group attached to carbon 4 and two methyl groups, one on carbon 3 and one on carbon 5. This gives the following carbon skeleton:



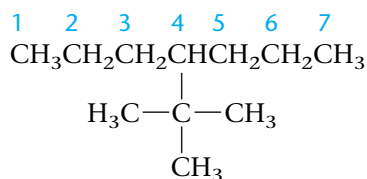
When we add the hydrogen atoms, we get the final structure



- Heptane signifies a seven-carbon chain, and the *tert*-butyl group (see Table 20.2) is



Thus we have the molecule



Practice Problem • Exercise 20.5

Write the structural formula for 5-isopropyl-4-methyldecane.

E. Petroleum

Woody plants, coal, petroleum, and natural gas provide a vast resource of energy that originally came from the sun. By the process of photosynthesis, plants store energy that we can claim by burning the plants themselves or, more commonly, burning the decay products that have been converted to fossil fuels. Although the United States currently depends heavily on petroleum for energy, this dependency is a relatively recent phenomenon (see **Figure 20.5**).

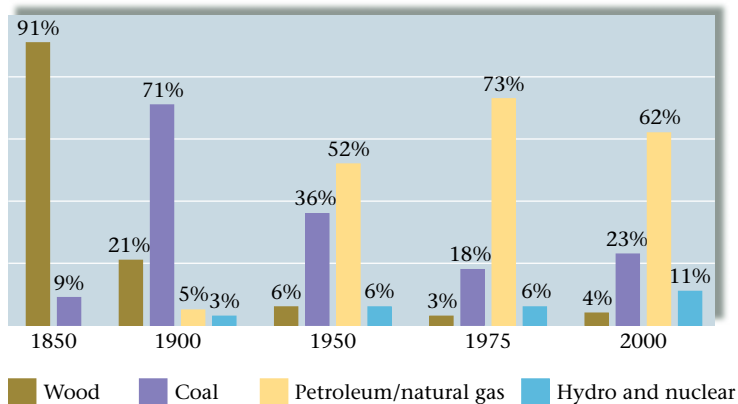


Figure 20.5
Energy sources used in the United States

Petroleum and natural gas deposits probably formed from the remains of marine organisms that lived approximately 500 million years ago. **Petroleum** is a thick, dark liquid composed largely of hydrocarbons containing from 5 to more than 25 carbon atoms. **Natural gas**, which is usually associated with petroleum deposits, consists mostly of methane but also contains significant amounts of ethane, propane, and butane.

To be used efficiently, petroleum must be separated by boiling or distilled into portions called fractions. The smaller hydrocarbons can be boiled off at relatively low temperatures; the larger molecules require successively higher temperatures. The major uses of various petroleum fractions are shown in **Table 20.3**.

Table 20.3

Uses of the Various Petroleum Fractions

Petroleum Fraction*	Major Uses
$C_5 \rightarrow C_{12}$	gasoline
$C_{10} \rightarrow C_{18}$	kerosene jet fuel
$C_{15} \rightarrow C_{25}$	diesel fuel heating oil lubricating oil
$> C_{25}$	asphalt

*Shows the chain lengths present in each fraction.

CRITICAL THINKING ?

Petroleum is a very valuable raw material for the synthesis of polymers.

What if Congress decided that petroleum must be conserved as a raw material and could not be used as fuel? What might our society use as alternative sources of energy?

Petroleum

A thick, dark liquid composed mostly of hydrocarbon compounds

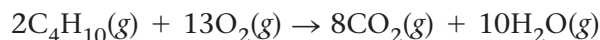
Natural gas

A gaseous fossil fuel, mostly consisting of methane (with some ethane, propane, and butane) and usually associated with petroleum deposits

F. Reactions of Alkanes

At low temperatures alkanes are not very reactive because the C—C and C—H bonds in alkanes are relatively strong. For example, at 25 °C alkanes do not react with acids, bases, or strong oxidizing agents. This chemical inertness makes alkanes valuable as lubricating materials and as the backbone for structural materials such as plastics.

At sufficiently high temperatures, however, alkanes *react vigorously with oxygen*. These **combustion reactions** are the basis for the alkanes' widespread use as fuels. For example, the combustion reaction of butane with oxygen is



Combustion reaction

A chemical reaction involving oxygen as one of the reactants that produces enough heat so that a flame results

Substitution reaction

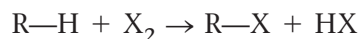
A reaction in which an atom (such as a halogen) replaces a hydrogen atom in a hydrocarbon



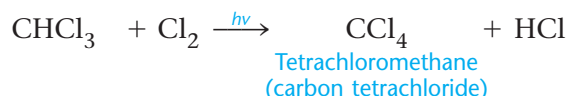
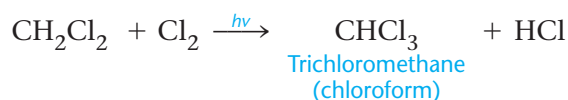
Active Reading Question

When hydrocarbons undergo combustion, what is the other reactant? What are the products?

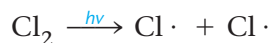
The alkanes can also undergo **substitution reactions**—reactions in which *one or more hydrogen atoms of the alkane are replaced* (substituted) *by different atoms*. We can represent the substitution reaction of an alkane with a halogen molecule as follows:



where R represents an alkyl group and X represents a halogen atom. For example, methane can react successively with chlorine as follows:



The $h\nu$ above each arrow signifies that ultraviolet light is needed to furnish the energy to break the Cl—Cl bond to produce chlorine atoms:



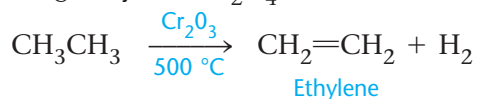
A chlorine atom has an unpaired electron, indicated by the dot, which makes it very reactive and able to disrupt the C—H bond.

Notice that each step in the process involves replacement of a C—H bond by a C—Cl bond. That is, a chlorine atom *substitutes* for a hydrogen atom. The names of the products of these reactions use the term *chloro* for the chlorine substituents with a prefix that gives the number of chlorine atoms present: *di-* for two, *tri-* for three, and *tetra-* for four. No number is used to describe the chlorine positions in this case, because methane has only one carbon atom. Note that the products of the last two reactions have two names, the systematic name and the common name in parentheses.

Information

The symbol $h\nu$ signifies ultraviolet light used to furnish energy for the reaction.

Besides substitution reactions, alkanes can undergo **dehydrogenation reactions** in which *hydrogen atoms are removed* and the product is an unsaturated hydrocarbon. For example, in the presence of a catalyst, such as chromium(III) oxide, at high temperatures, ethane can be dehydrogenated, yielding ethylene, C_2H_4 .



Dehydrogenation reaction

A reaction in which hydrogen atoms are removed from an alkane and the result is an unsaturated hydrocarbon



SECTION 20.1 REVIEW QUESTIONS

- How is carbon and its bonding unusual compared to other elements?
- Determine the shape of each of the following molecules (assume C is the central atom):
 - CH_4
 - H_2CO
 - C_2H_2
- How is a saturated carbon compound different from an unsaturated one?
- Draw the structural formulas for at least five isomeric alkanes having the formula C_8H_{18} .
- What are some of the uses of petroleum?
- Write a balanced equation for the combustion of propane.
- How is substitution reaction in alkanes different from a dehydrogenation reaction since both involve removing a hydrogen atom from the molecule?
- Give the systematic name for each of the following alkanes.
 - $$CH_3-CH-CH_2-CH_2-CH_2-CH_3$$

$$|$$

$$CH_2-CH_3$$
 - $$CH_3-CH-CH-CH_3$$

$$| \quad |$$

$$CH_3 \quad CH_3$$
 - $$CH_3-CH-CH-CH_3$$

$$| \quad |$$

$$CH_3 \quad CH_3$$
 - $$CH_3-CH_2-CH_2-CH-CH-CH-CH_3$$

$$| \quad | \quad |$$

$$CH_3 \quad CH_3 \quad CH_3$$



RESEARCH LINKS
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SECTION 20.2

Unsaturated Hydrocarbons

Key Terms

- Alkenes
- Alkynes
- Addition reactions
- Hydrogenation reactions
- Halogenation
- Polymerization
- Aromatic hydrocarbons
- Benzene
- Phenyl group

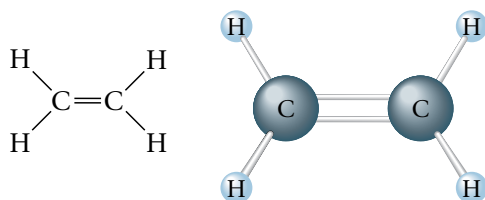
Objectives

- To learn to name hydrocarbons with double and triple bonds
- To understand addition reaction
- To learn about the aromatic hydrocarbons
- To learn to name aromatic compounds

A. Alkenes and Alkynes

We have seen that alkanes are saturated hydrocarbons—each of the carbon atoms is bound to four atoms by single bonds. Hydrocarbons that contain carbon–carbon *double bonds* are called **alkenes**. Hydrocarbons with carbon–carbon *triple bonds* are called **alkynes**. Alkenes and alkynes are unsaturated hydrocarbons.

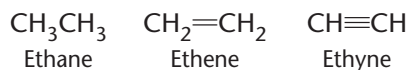
Multiple carbon–carbon bonds result when hydrogen atoms are removed from alkanes. Alkenes that contain one carbon–carbon double bond have the general formula C_nH_{2n} . The simplest alkene, C_2H_4 , commonly known as *ethylene*, has the structure



The system for naming alkenes and alkynes is similar to the one we have used for alkanes. The following rules are useful.

Rules for Naming Alkenes and Alkynes

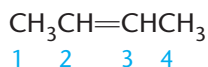
- Select the longest continuous chain of carbon atoms that contains the double or triple bond.
- For an alkene, the root name of the carbon chain is the same as for the alkane, except that the *-ane* ending is replaced by *-ene*. For an alkyne, the *-ane* is replaced by *-yne*. For example, for a two-carbon chain we have



- Number the parent chain, starting at the end closest to the double or triple bond. The location of the multiple bond is given by the lowest-numbered carbon involved in the bond. For example,



is called 1-butene and



is called 2-butene.

- Substituents on the parent chain are treated the same way as in naming alkanes. For example, the molecule $\text{ClCH}=\text{CHCH}_2\text{CH}_3$ is called 1-chloro-1-butene.

Alkene

An unsaturated hydrocarbon containing a carbon–carbon double bond; alkenes have the general formula C_nH_{2n} .

Alkyne

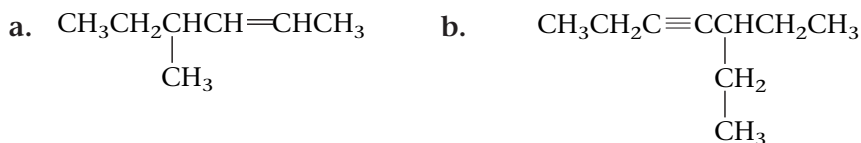
An unsaturated hydrocarbon containing a carbon–carbon triple bond; alkynes have the general formula C_nH_{2n-2} .

i n f o r m a t i o n

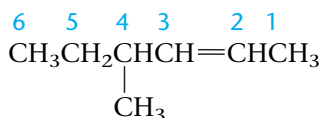
In writing shorthand formulas, the hydrogen atoms are often written just after the carbon to which they are attached. For example, the formula for $\text{H}-\text{C}\equiv\text{C}-\text{H}$ is often written as $\text{CH}\equiv\text{CH}$.

EXAMPLE 20.6**Naming Alkenes and Alkynes**

Name each of the following molecules.

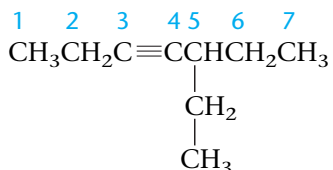
**Solution**

- a. The longest chain contains six carbon atoms, and we number the carbons starting from the end closest to the double bond.



Thus the root name for the hydrocarbon is 2-hexene. Remember to use the lower number of the two carbon atoms involved in the double bond. There is a methyl group attached to the number-4 carbon. Therefore, the name of the compound is 4-methyl-2-hexene.

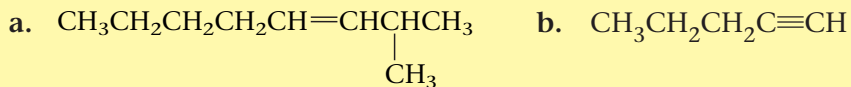
- b. The longest chain of carbon atoms is seven carbons long, and the chain is numbered as shown (starting from the end closest to the triple bond).



The hydrocarbon is a 3-heptyne (we use the lowest-numbered carbon in the triple bond). Because there is an ethyl group on carbon number 5, the full name is 5-ethyl-3-heptyne.

Practice Problem • Exercise 20.6

Name the following molecules.



Addition reaction

A reaction in which new atoms form single bonds to the carbon atoms in unsaturated hydrocarbons that were involved in double or triple bonds

Hydrogenation reaction

An addition reaction in which H_2 is a reactant

Halogenation

An addition reaction in which a halogen is a reactant

Polymerization

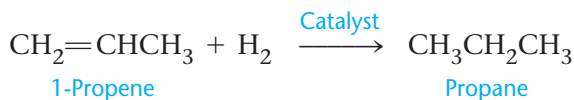
A process in which many small molecules (monomers) are joined together to form a large molecule

Reactions of Alkenes

Because alkenes and alkynes are unsaturated, their most important reactions are **addition reactions**.

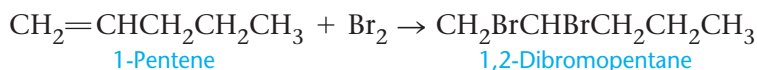
In addition reactions, new atoms form single bonds to the carbons formerly involved in the double or triple bonds.

An addition reaction for an alkene changes the carbon–carbon double bond to a single bond, giving a saturated hydrocarbon (each carbon bonded to four atoms). For example, **hydrogenation reactions**, which use H_2 as a reactant, lead to the addition of a hydrogen atom to each carbon formerly involved in the double bond.



Hydrogenation of molecules with double bonds is an important industrial process, particularly in the manufacture of solid shortenings. Unsaturated fats (fats containing double bonds) are generally liquids at room temperatures, whereas saturated fats (those containing C—C single bonds) are solids. The liquid unsaturated fats are converted to solid saturated fats by hydrogenation.

Halogenation of unsaturated hydrocarbons involves the addition of halogen atoms. Here is an example:



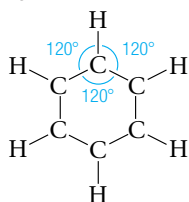
Another important reaction of certain unsaturated hydrocarbons is **polymerization**, a process in which many small molecules are joined together to form a large molecule. Polymerization will be discussed later in the chapter.

Active Reading Question

What is the difference between a substitution reaction of an alkane that involves a halogen, and a halogenation reaction of an alkene?

B. Aromatic Hydrocarbons

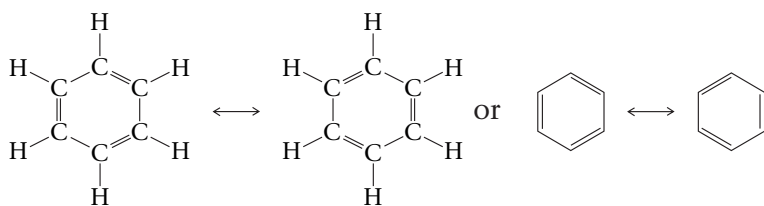
When mixtures of hydrocarbons from natural sources, such as petroleum or coal, are separated, certain of the compounds that emerge have pleasant odors and are thus known as **aromatic hydrocarbons**. When these substances, which include wintergreen, cinnamon, and vanillin, are examined, they are all found to contain a common feature: a six-membered ring of carbon atoms called the benzene ring. **Benzene** has the formula C_6H_6 and a planar (flat) structure in which all of the bond angles are 120° .



Aromatic hydrocarbon

One of a special class of cyclic unsaturated hydrocarbons, the simplest of which is benzene

When we examine the bonding in the benzene ring, we find that more than one Lewis structure can be drawn. That is, the double bonds can be located in different positions, as shown below.



Because the actual bonding is a combination of the structures represented above, the benzene ring is usually shown with a circle.

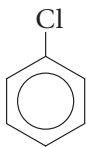


C. Naming Aromatic Compounds

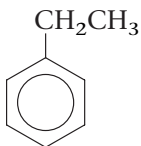
Substituted benzene molecules are formed by replacing one or more of the H atoms on the benzene ring with other atoms or groups of atoms. We will consider benzene rings with one substituent, called monosubstituted benzenes, first.

Monosubstituted Benzenes

The systematic method for naming monosubstituted benzenes uses the substituent name as a prefix of benzene. For example, the molecule

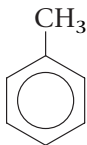


is called chlorobenzene, and the molecule

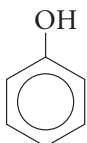


is called ethylbenzene.

Sometimes monosubstituted benzene compounds have special names. For example, the molecule



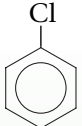
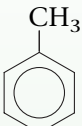
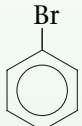
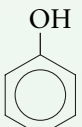
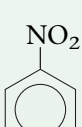
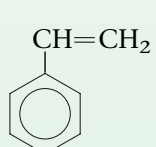
has the systematic name methylbenzene. However, for convenience it is given the name toluene. Likewise, the molecule



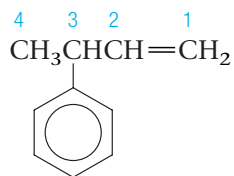
which might be called hydroxybenzene, has the special name phenol. Several examples of monosubstituted benzenes are shown in **Table 20.4**.

Table 20.4

Common Monosubstituted Benzenes

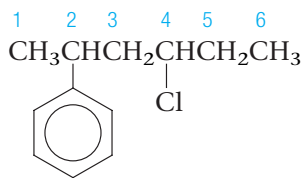
Name	Structure
Chlorobenzene	
Toluene	
Bromobenzene	
Phenol	
Nitrobenzene	
Styrene	

Sometimes it is more convenient to name compounds if we view the benzene ring itself as a substituent. For example, the compound



is most easily named as a 1-butene with a benzene ring as a substituent on the number-3 carbon. When the benzene ring is used as a substituent, it is

called the **phenyl** (pronounced fen'-ill) **group**. So the name of this compound is 3-phenyl-1-butene. As another example, the compound



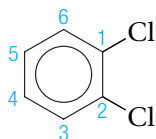
is named 4-chloro-2-phenylhexane. Remember, we start to number the chain from the end closest to the first substituent and name the substituents in alphabetical order (chloro before phenyl).

Active Reading Question

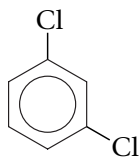
What are the two ways of naming monosubstituted benzenes?

Disubstituted Benzenes

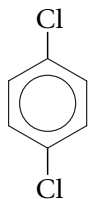
When there is more than one substituent on the benzene ring, numbers are used to indicate substituent position. For example, the compound



is named 1,2-dichlorobenzene. Another naming system uses the prefix *ortho-* (*o-*) for two adjacent substituents, *meta-* (*m-*) for two substituents with one carbon between them, and *para-* (*p-*) for two substituents opposite each other. This means that 1,2-dichlorobenzene can also be called *ortho*-dichlorobenzene or *o*-dichlorobenzene. Likewise, the compound

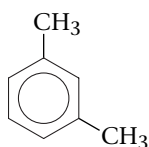


can be called 1,3-dichlorobenzene or *m*-dichlorobenzene. The compound



is named 1,4-dichlorobenzene or *p*-dichlorobenzene.

Benzenes that have two methyl substituents have the special name xylene. So the compound



which might be called 1,3-dimethylbenzene, is instead called *m*-xylene (*meta*-xylene).

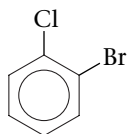
Information

Ortho- (*o-*) means two adjacent substituents.

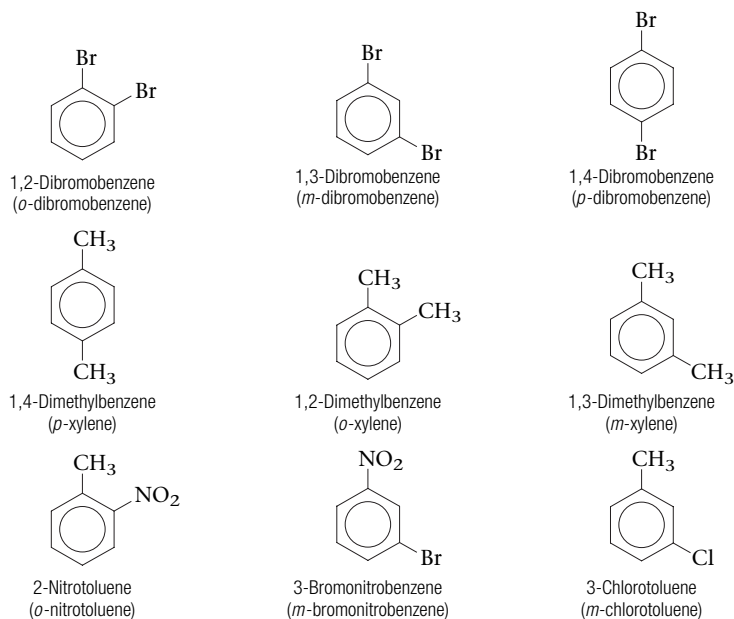
Para- (*p-*) means two substituents directly across the ring from each other.

Meta- (*m-*) means two substituents with one carbon between them.

When two different substituents are present on the benzene ring, one is always assumed to be at carbon number 1, and this number is not often specified in the name. For example, the compound



is named 2-bromochlorobenzene, not 2-bromo-1-chlorobenzene. Various examples of disubstituted benzenes are shown below.



Benzene is the simplest aromatic molecule. More complex aromatic systems can be viewed as consisting of a number of “fused” benzene rings. Some examples are given in **Table 20.5**.

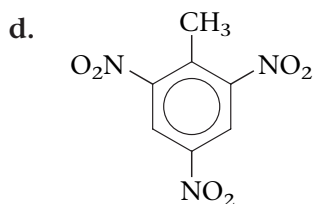
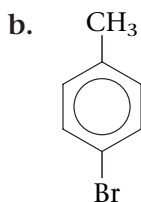
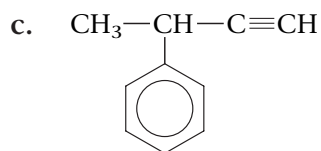
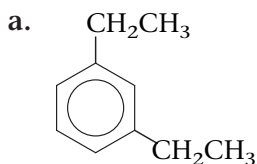
Table 20.5

More Complex Aromatic Molecules



Structural Formula	Name	Use
	naphthalene	
	anthracene	dyes
	phenanthrene	dyes, explosives, and synthesis of drugs

EXAMPLE 20.7**Naming Aromatic Compounds**

Name the following compounds.

**Solution**

a. There are ethyl groups in the 1 and 3 (or *meta*-) positions, so the name is 1,3-diethylbenzene or *m*-diethylbenzene.

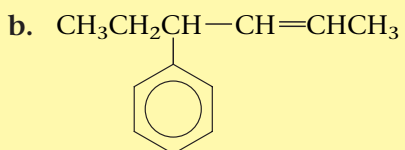
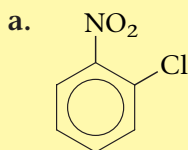
b. 
 The  group is called toluene. The bromine is in the 4 (or *para*-) position. The name is 4-bromotoluene or *p*-bromotoluene.

c. In this case we name the compound as a butyne with a phenyl substituent. The name is 3-phenyl-1-butyne.

d. We name this compound as a substituted toluene (with the —CH_3 group assumed to be on carbon number 1). So the name is 2,4,6-trinitrotoluene. This compound is more commonly known as TNT, a component of high explosives.

Practice Problem • Exercise 20.7

Name the following compounds.



CHEMISTRY IN YOUR WORLD

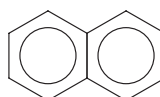
Science, Technology, and Society

Termite Mothballing

Termites typically do not get much respect. They are regarded as lowly, destructive insects. However, termites are the first insects known to fumigate their nests with naphthalene, a chemical long used by humans to prevent moths from damaging wool garments. Although termites are not concerned about holes in their sweaters, they may use naphthalene to ward off microbes and predatory ants, among other pests.

Gregg Henderson and Jian Chen, of the Louisiana State University Agricultural Center in Baton Rouge, have observed that Formosan termites are unusually resistant to naphthalene. In fact these insects build their underground galleries from chewed wood glued together with saliva and excrement. This

“glue” (called *carton*) contains significant amounts of naphthalene, which evaporates and permeates the air in the underground tunnels. The source of the naphthalene is unknown—it might be a metabolite from a food source of the termites or it might be produced from the carton by organisms present in the nest. Whatever the source of the naphthalene, this interesting example shows how organisms use chemistry to protect themselves.

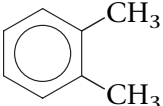
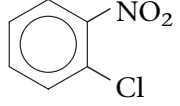
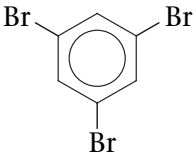
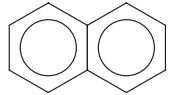


Naphthalene

Formosan subterranean termites

Image not available for electronic use. Please refer to the image in the textbook.

SECTION 20.2 REVIEW QUESTIONS

- Draw structural formulas for the following:
 - 2,5-dimethyl-3-hexene
 - 4-methyl-2-pentene
 - iodoethyne
 - 3-methylpentene
- Why do alkenes and alkynes react by addition reactions whereas alkanes do not?
- How is a halogenation reaction similar to a hydrogenation reaction?
- What is the common feature of an aromatic compound?
- Draw the structure for the ring compounds with formulas C_6H_{12} and C_6H_6 .
 - Name each compound.
 - How are they similar?
 - How are they different?
- Would C_6H_{12} react by addition or substitution reactions? Why?
 - 
 - 
 - 
 - 

Introduction to Functional Groups and Alcohols

Objectives

- To learn the common functional groups in organic molecules
- To learn about simple alcohols and how to name them
- To learn about how some alcohols are made and used

A. Functional Groups

The vast majority of organic molecules contain elements in addition to carbon and hydrogen. Most of these substances are fundamentally hydrocarbons that have additional atoms or groups of atoms called **functional groups**. The common functional groups are listed in **Table 20.6**; one example of a compound that contains that functional group is given for each.

Table 20.6

The Common Functional Groups

Class	Functional Group	General Formula*	Example
halohydrocarbons [†]	—X (F, Cl, Br, I)	R—X	CH ₃ I
alcohols	—OH	R—OH	CH ₃ OH
ethers	—O—	R—O—R'	CH ₃ —O—CH ₃
aldehydes	$\begin{array}{c} \text{O} \\ \parallel \\ \text{—C—H} \end{array}$	$\begin{array}{c} \text{O} \\ \parallel \\ \text{R—C—H} \end{array}$	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H—C—H} \end{array}$
ketones	$\begin{array}{c} \text{O} \\ \parallel \\ \text{—C—} \end{array}$	$\begin{array}{c} \text{O} \\ \parallel \\ \text{R—C—R}' \end{array}$	$\begin{array}{c} \text{O} \\ \parallel \\ \text{CH}_3\text{—C—CH}_3 \end{array}$
carboxylic acids	$\begin{array}{c} \text{O} \\ \parallel \\ \text{—C—OH} \end{array}$	$\begin{array}{c} \text{O} \\ \parallel \\ \text{R—C—OH} \end{array}$	$\begin{array}{c} \text{O} \\ \parallel \\ \text{CH}_3\text{—C—OH} \end{array}$
esters	$\begin{array}{c} \text{O} \\ \parallel \\ \text{—C—O—} \end{array}$	$\begin{array}{c} \text{O} \\ \parallel \\ \text{R—C—O—R}' \end{array}$	$\begin{array}{c} \text{O} \\ \parallel \\ \text{CH}_3\text{—C—OCH}_2\text{CH}_3 \end{array}$
amines	—NH ₂	R—NH ₂	CH ₃ NH ₂

*R and R' represent hydrocarbon fragments, which may be the same or different.

[†]These substances are also called alkyl halides.

Key Terms

- Functional groups
- Alcohols

Functional group

An atom or group of atoms in a hydrocarbon derivative that contains elements in addition to carbon and hydrogen

Alcohol

A hydrocarbon derivative in which the hydroxyl group ($-\text{OH}$) is the functional group

B. Alcohols

Alcohols are characterized by the presence of the $-\text{OH}$ group. Some common alcohols are listed in **Table 20.7**. The systematic name for an alcohol is obtained by replacing the final *-e* of the parent hydrocarbon name with *-ol*. The position of the $-\text{OH}$ group is specified by a number (where necessary) chosen such that it is the smallest of the substituent numbers.

Table 20.7

Some Common Alcohols

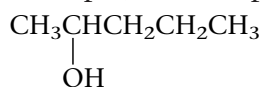
Formula	Systematic Name	Common Name
CH_3OH	methanol	methyl alcohol
$\text{CH}_3\text{CH}_2\text{OH}$	ethanol	ethyl alcohol
$\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$	1-propanol	<i>n</i> -propyl alcohol
$\begin{array}{c} \text{CH}_3\text{CHCH}_3 \\ \\ \text{OH} \end{array}$	2-propanol	isopropyl alcohol

The rules for naming alcohols follow.

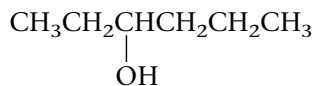
Rules for Naming Alcohols

1. Select the longest chain of carbon atoms containing the $-\text{OH}$ group.
2. Number the chain such that the carbon with the $-\text{OH}$ group gets the lowest possible number.
3. Obtain the root name from the name of the parent hydrocarbon chain by replacing the final *-e* with *-ol*.
4. Name any other substituents as usual.

For example, the compound

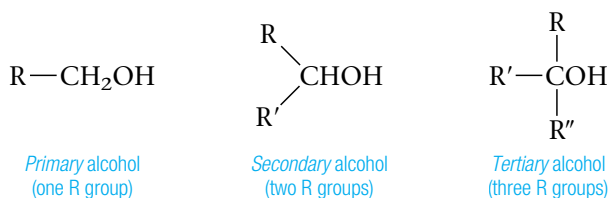


is called 2-pentanol because the parent carbon chain is pentane. The compound



is called 3-hexanol.

Alcohols are classified according to the number of hydrocarbon fragments (alkyl groups) bonded to the carbon where the —OH group is attached. Thus we have



where R, R', and R'' (which may be the same or different) represent hydrocarbon fragments (alkyl groups).

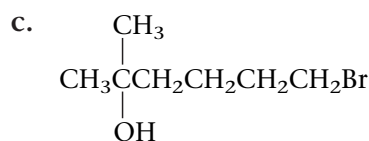
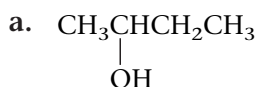
Active Reading Question

What is the difference among primary, secondary, and tertiary alcohols?

EXAMPLE 20.8

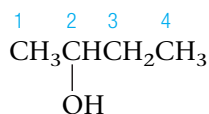
Naming Alcohols

Give the systematic name for each of the following alcohols, and specify whether the alcohol is primary, secondary, or tertiary.

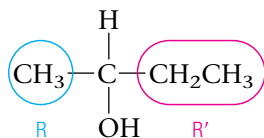


Solution

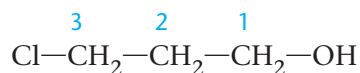
a. The chain is numbered as follows:



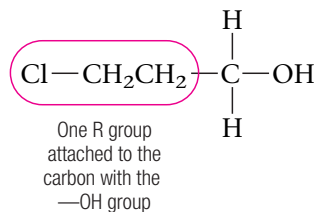
The compound is called 2-butanol, because the —OH group is located at the number-2 position of a four-carbon chain. Note that the carbon to which the —OH is attached also has two R groups (—CH₃ and —CH₂CH₃) attached. Therefore, this is a *secondary* alcohol.



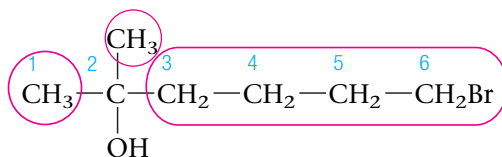
b. The chain is numbered as follows:



Remember that in naming an alcohol, we give the carbon with the —OH attached the lowest possible number. The name is 3-chloro-1-propanol. This is a *primary* alcohol.



c. The chain is numbered as follows:



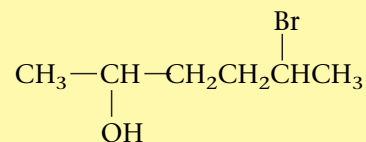
The name is 6-bromo-2-methyl-2-hexanol. This is a tertiary alcohol, because the carbon where the —OH is attached also has three R groups attached.

Practice Problem • Exercise 20.8

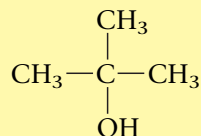
Name each of the following alcohols, and specify whether it is primary, secondary, or tertiary.



c.

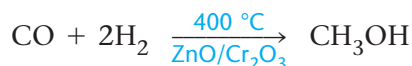


b.



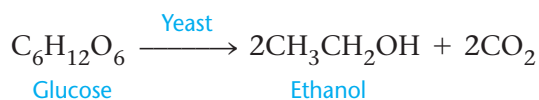
C. Properties and Uses of Alcohols

Although there are many important alcohols, the simplest ones, methanol and ethanol, have the greatest commercial value. Methanol, also known as *wood alcohol* because it was formerly obtained by heating wood in the absence of air, is prepared industrially (over 20 million tons per year in the United States) by the hydrogenation of carbon monoxide (catalyzed by a ZnO/Cr₂O₃ mixture).



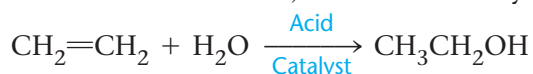
Methanol Methanol is used as a starting material for the synthesis of acetic acid and many types of adhesives, fibers, and plastics. It also can be used as a motor fuel. In fact, pure methanol has been used for many years in the engines of the cars that are driven in the Indianapolis 500 and similar races. Methanol is especially useful in racing engines because of its resistance to knocking. It is advantageous for regular cars because it produces less carbon monoxide (a toxic gas) in the exhaust than does gasoline. Methanol is highly toxic to humans, and swallowing it can lead to blindness and death.

Ethanol Ethanol, which is used in the production of some beverages, is produced by the fermentation of the sugar glucose in corn, barley, grapes, and so on.

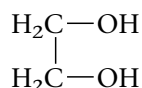


This reaction is catalyzed by the enzymes (biological catalysts) found in yeast, and it can proceed only until the alcohol content reaches approximately 13%, at which point the yeast can no longer survive. Beverages with higher alcohol content are made by distilling the fermentation mixture.

Ethanol, like methanol, can be burned in the internal combustion engines of automobiles and is now commonly added to gasoline to form gasohol. It is also used in industry as a solvent and for the preparation of acetic acid. The commercial production of ethanol (half a million tons per year in the United States) is carried out by reaction of water with ethylene.

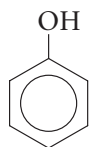


Other Alcohols Many alcohols are known that contain more than one —OH group. The one that is the most important commercially is ethylene glycol,

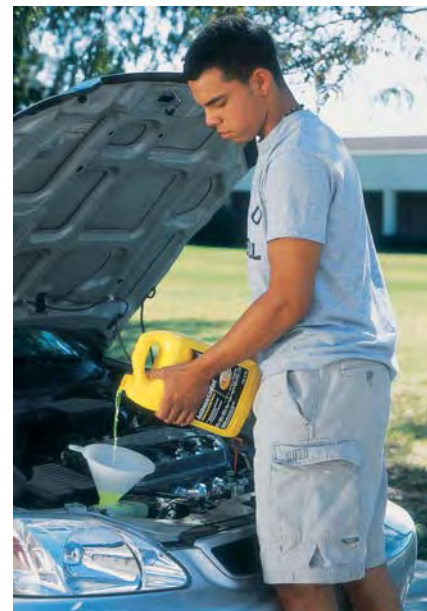


a toxic substance that is the major constituent of most automobile antifreezes.

The simplest aromatic alcohol is



commonly called **phenol**. Most of the 1 million tons of phenol produced annually in the United States is used to produce polymers for adhesives and plastics.



Ethylene glycol is a component of antifreeze, which is used to protect the cooling systems of automobiles.

SECTION 20.3

REVIEW QUESTIONS

- How are aldehydes and ketones similar? How can you distinguish an aldehyde from a ketone? (See Table 20.6.)
- How are carboxylic acids and esters similar? How can you distinguish a carboxylic acid from an ester? (See Table 20.6.)
- Name the following compounds:
 - $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{OH}$
 - $\begin{array}{c} \text{CHCHCH}_3 \\ | \\ \text{OH} \end{array}$
 - $\begin{array}{c} \text{CH}_3 \\ | \\ \text{HOCH}_2\text{CHCHCH}_2\text{CH}_2\text{OH} \\ | \\ \text{CH}_3 \end{array}$
- Draw a structural formula for each of the following alcohols. Indicate whether the alcohol is primary, secondary, or tertiary.
 - 2-propanol
 - 2-methyl-2-propanol
 - 4-isopropyl-2-heptanol
 - 2,3-dichloro-1-pentanol
- Why is methanol important commercially?



Additional Organic Compounds

Key Terms

- Carbonyl group
- Ketones
- Aldehydes
- Carboxylic acids
- Carboxyl group
- Ester
- Polymer
- Addition polymerization
- Condensation polymerization
- Copolymer
- Homopolymer
- Dimer
- Polyester

Objectives

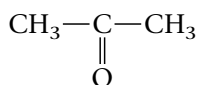
- To learn about aldehydes and ketones
- To learn to name aldehydes and ketones
- To learn about some common carboxylic acids and esters
- To learn about some common polymers

A. Aldehydes and Ketones

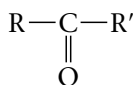
Aldehydes and ketones contain the **carbonyl group**



In **ketones** this group is bonded to two carbon atoms; an example is acetone:

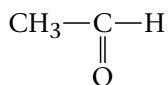


The general formula for a ketone is

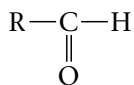


where R and R' are alkyl groups that may or may not be the same. In a ketone the carbonyl group is never at the end of the hydrocarbon chain. (If it were, it would be an aldehyde.)

In **aldehydes** the carbonyl group always appears at the end of the hydrocarbon chain. There is always at least one hydrogen bonded to the carbonyl carbon atom. An example of an aldehyde is acetaldehyde:



The general formula for an aldehyde is



We often use compact formulas for aldehydes and ketones. For example, formaldehyde (where R = H) and acetaldehyde (where R = CH₃) are usually represented as HCHO and CH₃CHO, respectively. Acetone is often written as CH₃COCH₃ or (CH₃)₂CO.

Many ketones have useful solvent properties (acetone is often found in nail polish remover, for example) and are frequently used in industry for this purpose. Aldehydes typically have strong odors. Vanillin is responsible for the pleasant odor of vanilla beans; cinnamaldehyde produces the characteristic odor of cinnamon. On the other hand, the unpleasant odor of rancid butter arises from the presence of butyraldehyde and butyric acid. (See **Table 20.8** for the structures of these compounds.)

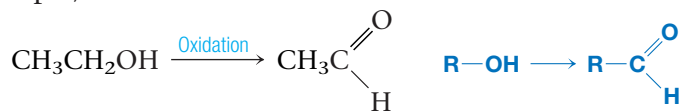
Ketone

A hydrocarbon derivative containing the carbonyl group bonded to two carbon atoms

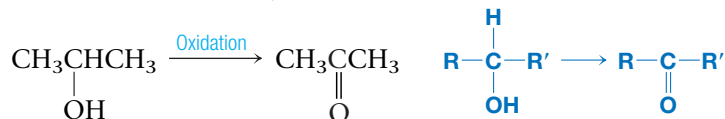
Aldehyde

A hydrocarbon derivative containing the carbonyl group at the end of the hydrocarbon chain

Aldehydes and ketones are most often produced commercially by the oxidation of alcohols. Oxidation of a *primary* alcohol gives the corresponding aldehyde, for example,



Oxidation of a *secondary* alcohol results in a ketone:

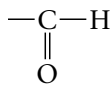


B. Naming Aldehydes and Ketones

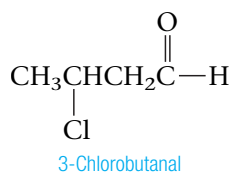
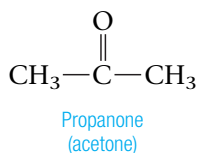
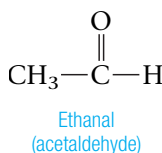
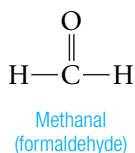
We obtain the systematic name for an aldehyde from the parent alkane by removing the final *-e* and adding *-al*. For ketones the final *-e* is replaced by *-one*, and a number indicates the position of the carbonyl group where necessary. The carbon chain in ketones is numbered such that the



carbon gets the lowest possible number. In aldehydes the



is always at the end of the chain and is always assumed to be carbon number 1. The positions of other substituents are specified by numbers, as usual. The following examples illustrate these principles.



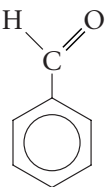
The names in parentheses are common names that are used much more often than the systematic names.

Table 20.8

Common Aldehydes

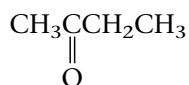
Name	Structure
Vanillin	
Cinnamaldehyde	
Butyraldehyde	$\text{CH}_3\text{CH}_2\text{CH}_2 \begin{array}{c} \text{O} \\ \parallel \\ \text{---C---} \\ \parallel \\ \text{O} \end{array} \text{H}$

Another common aldehyde is benzaldehyde (an aromatic aldehyde), which has the structure

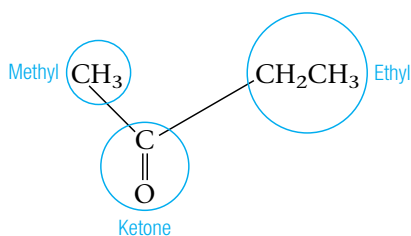


Benzaldehyde

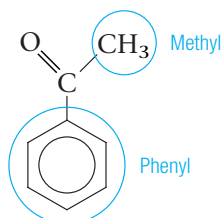
An alternative system for naming ketones specifies the substituents attached to the C=O group. For example, the compound



is called 2-butanone when we use the system just described. However, this molecule can also be named methyl ethyl ketone and is commonly referred to in industry as MEK (*methyl ethyl ketone*):



Another example is the use of the name ketone for the compound



which is commonly called methyl phenyl ketone.

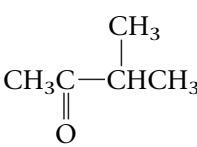
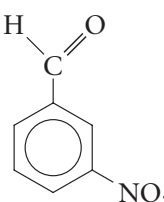
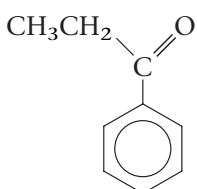
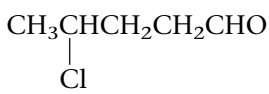
Active Reading Question

What are two ways of naming ketones?

EXAMPLE 20.9

Naming Aldehydes and Ketones

Name the following molecules.

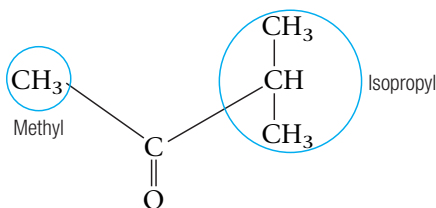
- a.  (Give two names.)
- b. 
- c. 
- d. 

Solution

- a. We can name this molecule as a 2-butanone because the longest chain has four carbon atoms (butane root) with the



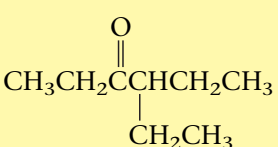
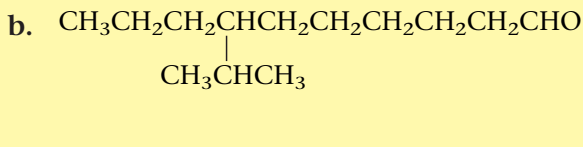
group in the number 2 position (the lowest possible number). Because the methyl group is in the number 3 position, the name is 3-methyl-2-butanone. We can also name this compound methyl isopropyl ketone.



- b. We name this molecule as a substituted benzaldehyde (the nitro group is in the number 3 position): 3-nitrobenzaldehyde. It might also be named *m*-nitrobenzaldehyde.
- c. We name this molecule as a ketone: ethyl phenyl ketone.
- d. The name is 4-chloropentanal. Note that an aldehyde group is always at the end of the chain and is automatically assigned as the number 1 carbon.

Practice Problem • Exercise 20.9

Name the following molecules.

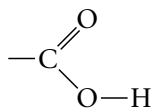
- a. 
- b. 

Carboxylic acid

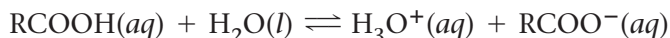
A hydrocarbon derivative containing the carboxyl group (—COOH)

C. Carboxylic Acids and Esters

Carboxylic acids are characterized by the presence of the **carboxyl group**, —COOH, which has the structure

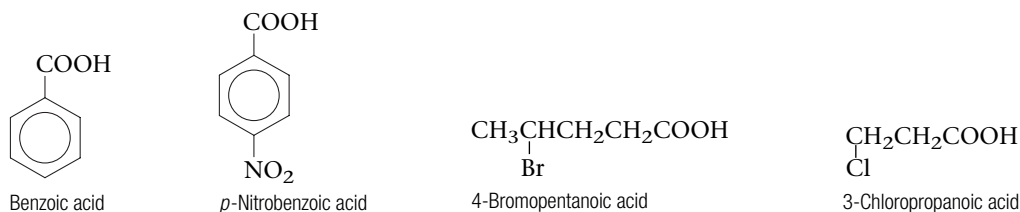


The general formula of a carboxylic acid is RCOOH, where R represents the hydrocarbon fragment. These molecules typically are weak acids in aqueous solution. That is, the dissociation (ionization) equilibrium



lies far to the left—only a small percentage of the RCOOH molecules are ionized.

We name carboxylic acids by dropping the final *-e* from the parent alkane (the longest chain containing the —COOH group) and adding *-oic*. Carboxylic acids are frequently known by their common names. For example, CH₃COOH, often written HC₂H₃O₂ and commonly called acetic acid, has the systematic name ethanoic acid because the parent alkane is ethane. Several carboxylic acids, their systematic names, and their common names are given in **Table 20.9**. Other examples of carboxylic acids are shown below.



Note that the —COOH group is always assigned the number 1 position in the chain.

Carboxylic acids can be produced by oxidizing primary alcohols with a strong oxidizing agent. For example, we can oxidize ethanol to acetic acid by using potassium permanganate.

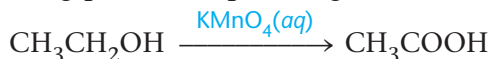
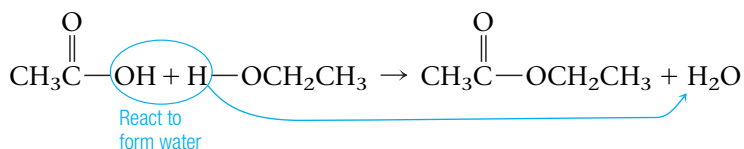


Table 20.9

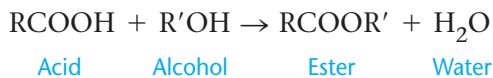
Several Carboxylic Acids, Their Systematic Names, and Their Common Names

Formula	Systematic Name	Common Name
HCOOH	methanoic acid	formic acid
CH ₃ COOH	ethanoic acid	acetic acid
CH ₃ CH ₂ COOH	propanoic acid	propionic acid
CH ₃ CH ₂ CH ₂ COOH	butanoic acid	butyric acid
CH ₃ CH ₂ CH ₂ CH ₂ COOH	pentanoic acid	valeric acid

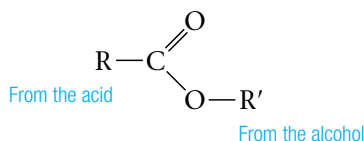
Esters A carboxylic acid reacts with an alcohol to form an ester and a water molecule. For example, the reaction of acetic acid and ethanol produces the ester ethyl acetate and water.



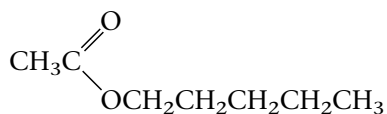
This reaction can be represented in general as follows:



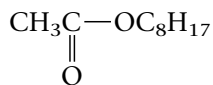
An **ester** has the following general formula:



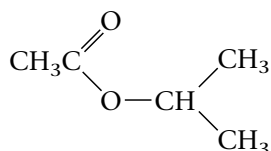
Esters often have a sweet, fruity odor that contrasts markedly with the often-pungent odors of the parent carboxylic acids. For example, the odor of bananas derives from *n*-amyl acetate:



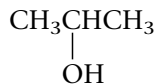
and that of oranges from *n*-octyl acetate:



Like carboxylic acids, esters are often referred to by their common names. The name consists of the alkyl name from the alcohol followed by the acid name, where the *-ic* ending is replaced by *-ate*. For example, the ester



is made from acetic acid, CH_3COOH , and isopropyl alcohol,



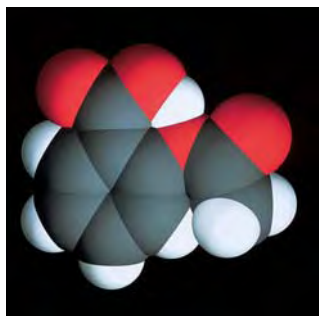
and is called isopropyl acetate. The systematic name for this ester is isopropylethanoate (from ethanoic acid, the systematic name for acetic acid).



The smell of bananas results from an aromatic ester.

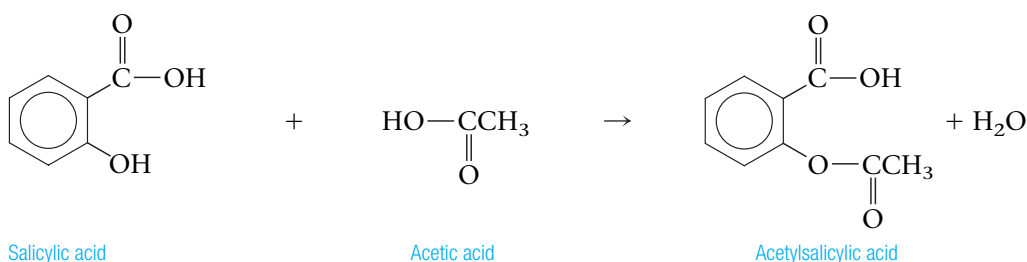
Information

Amyl is a common name for $\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2-$.



Computer-generated space-filled model of acetylsalicylic acid (*aspirin*)

A very important ester is formed from the reaction of salicylic acid and acetic acid.

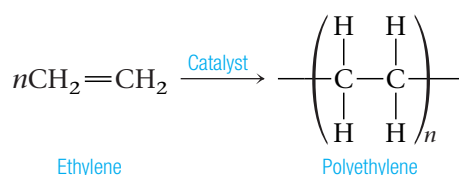


The product is acetylsalicylic acid, commonly known as *aspirin*, which is manufactured in huge quantities and is widely used as an analgesic (painkiller).

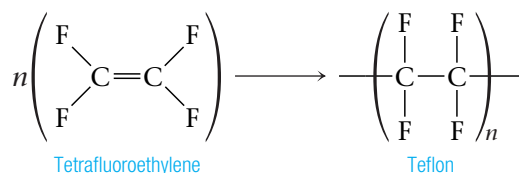
D. Polymers

Polymers are large, usually chainlike molecules that are built from small molecules called *monomers*. Polymers form the basis for synthetic fibers, rubbers, and plastics and have played a leading role in the revolution brought about in our lives by chemistry during the past 50 years. (Many important biomolecules are also polymers.)

The simplest and one of the best-known synthetic polymers is *polyethylene*, which is constructed from ethylene monomers. Its structure is



where n represents a large number (usually several thousand). Polyethylene is a tough, flexible plastic used for piping, bottles, electrical insulation, film for packaging, garbage bags, and many other purposes. Its properties can be varied by using substituted ethylene monomers. For example, when tetra-fluoroethylene is the monomer, the polymer Teflon is obtained.



Because of the resistance of the strong C—F bonds to chemical attack, Teflon is an inert, tough, and nonflammable material that is widely used for electrical insulation, nonstick coatings for cooking utensils, and bearings for low-temperature applications.

Other similar polyethylene-type polymers are made from monomers containing chloro, methyl, cyano, and phenyl substituents (see **Table 20.10**). In each case, the carbon–carbon double bond in the substituted ethylene monomer becomes a single bond in the polymer. The different substituents lead to a wide variety of properties.

The polyethylene polymers illustrate one of the major types of polymerization reactions, called **addition polymerization**, in which the monomers simply “add together” to form the polymer and there are no other products.

Polymers

A large, usually chainlike molecule made from small molecules called monomers

Addition polymerization

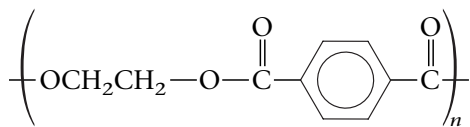
The process in which monomers simply “add together” to form polymers

Table 20.10

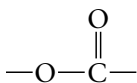
Some Common Synthetic Polymers, Their Monomers, and Applications

Monomer Name and Formula	Polymer Name and Formula	Applications
ethylene $\text{H}_2\text{C}=\text{CH}_2$	polyethylene $-(\text{CH}_2-\text{CH}_2)_n-$	plastic piping, bottles, electrical insulation, toys
propylene $\begin{array}{c} \text{H} \\ \\ \text{H}_2\text{C}=\text{C} \\ \\ \text{CH}_3 \end{array}$	polypropylene $\begin{array}{c} -(\text{CH}-\text{CH}_2-\text{CH}-\text{CH}_2)_n- \\ \quad \quad \\ \text{CH}_3 \quad \quad \text{CH}_3 \end{array}$	film for packaging, carpets, lab wares, toys
vinyl chloride $\begin{array}{c} \text{H} \\ \\ \text{H}_2\text{C}=\text{C} \\ \\ \text{Cl} \end{array}$	polyvinyl chloride (PVC) $\begin{array}{c} -(\text{CH}_2-\text{CH})_n- \\ \\ \text{Cl} \end{array}$	piping, siding, floor tile, clothing, toys
acrylonitrile $\begin{array}{c} \text{H} \\ \\ \text{H}_2\text{C}=\text{C} \\ \\ \text{CN} \end{array}$	polyacrylonitrile (PAN) $\begin{array}{c} -(\text{CH}_2-\text{CH})_n- \\ \\ \text{CN} \end{array}$	carpets, fabrics
tetrafluoroethylene $\text{F}_2\text{C}=\text{CF}_2$	Teflon $-(\text{CF}_2-\text{CF}_2)_n-$	coating for cooking utensils, electrical insulation, bearings
styrene $\begin{array}{c} \text{H} \\ \\ \text{H}_2\text{C}=\text{C} \\ \\ \text{C}_6\text{H}_5 \end{array}$	polystyrene $\begin{array}{c} -(\text{CH}_2\text{CH})_n- \\ \\ \text{C}_6\text{H}_5 \end{array}$	containers, thermal insulation, toys
butadiene $\begin{array}{c} \text{H} \quad \text{H} \\ \quad \\ \text{H}_2\text{C}=\text{C}-\text{C}=\text{CH}_2 \end{array}$	polybutadiene $-(\text{CH}_2\text{CH}=\text{CHCH}_2)_n-$	tire tread, coating resin
butadiene and styrene (see above)	styrene-butadiene rubber $\begin{array}{c} (\text{CH}-\text{CH}_2-\text{CH}_2-\text{CH}=\text{CH}-\text{CH}_2)_n- \\ \\ \text{C}_6\text{H}_5 \end{array}$	synthetic rubber

The repeating unit of Dacron is



Note that this polymerization involves a carboxylic acid and an alcohol to form an ester group,



Thus Dacron is called a **polyester**. By itself or blended with cotton, Dacron is widely used in fibers for the manufacture of clothing.

Active Reading Question

What is meant by the term *polyester*?

HANDS-ON CHEMISTRY

• MINI-LAB •

Guar Gum Slime

Materials

- 5-ounce paper cup
- graduated cylinder
- food coloring
- water
- guar gum
- 4% borax solution
- spoon
- plastic bag

Procedure

1. Measure 100 mL of water and pour it into a 5-ounce paper cup.
2. Add 2–3 drops of food coloring to the water.
3. Add 0.5 g ($\frac{1}{8}$ teaspoon) guar gum to the water and stir until dissolved. The mixture will thicken slightly as the guar gum dissolves.
4. Add 5 mL 4% borax solution to the cup and stir. The mixture should change in 1–2 minutes.
5. Store the slime in a plastic bag to prevent it from drying out.

SECTION 20.4

REVIEW QUESTIONS

- 1 Complete: Oxidation of a _____ alcohol results in a ketone. Oxidation of a _____ alcohol results in an aldehyde.
- 2 Draw a structural formula for each of the following:
 - a. 3-methylpentanal
 - b. 3-methyl-2-pentanone
 - c. methyl phenyl ketone
 - d. 2-hydroxybutanal
 - e. propanal
- 3 Why is it more common for organic chemists to write the formula for acetic acid as CH_3COOH instead of $\text{HC}_2\text{H}_3\text{O}_2$?
- 4 a. To make an ester, a carboxylic acid reacts with an alcohol. Water is removed to form the new ester. Circle the water that is removed in the reaction below and draw the ester which results:

$$\text{CH}_3\text{CH}_2\text{C—OH} + \text{H—OCH}_3$$
 - b. Name each of the compounds in your reaction.
- 5 Do all of the monomers in a polymer have to be the same? Give examples to support your answer.



Key Terms

20.1 Organic chemistry

Biomolecule
Hydrocarbons
Saturated
Unsaturated
Alkanes
Unbranched hydrocarbons
(normal or straight chain)
Structural isomerism
Petroleum
Natural gas
Combustion reactions
Substitution reactions
Dehydrogenation reactions

20.2 Alkenes

Alkynes
Addition reactions
Hydrogenation reaction
Halogenation
Polymerization
Aromatic hydrocarbons
Benzene
Phenyl group

20.3 Functional groups

Alcohols

Key Ideas

20.1 Saturated Hydrocarbons

- Organic chemistry is the study of carbon-containing compounds and their properties.
- Most organic compounds contain chains or rings of carbon atoms.
- Hydrocarbons are organic compounds composed of carbon and hydrogen.
 - Alkanes are called saturated hydrocarbons because they have four single bonds (the maximum possible) to each carbon.
 - ◆ General formula C_nH_{2n+2}
 - ◆ Exhibit structural isomerism: molecules with the same atoms but different arrangements of atoms
 - ◆ Undergo substitution reactions
 - Petroleum is a mixture of hydrocarbons formed by decay of ancient marine organisms which is used as a source of energy.

20.2 Unsaturated Hydrocarbons

- Alkenes are unsaturated hydrocarbons which contain one or more double carbon-carbon bonds.
 - Named as alkanes with the final *-ane* replaced by *-ene*
- Alkynes are unsaturated hydrocarbons which contain one or more triple carbon-carbon bonds.
 - Named as alkanes with the final *-ane* replaced by *-yne*
- Undergo addition reactions
- Some alkenes form special aromatic rings that are unusually stable.

20.3 Introduction to Functional Groups and Alcohols

- Alcohols contain one or more —OH groups
- Alcohols are named by using the root name of the parent alkane and replacing the final *-e* with *-ol*.

20.4 Carbonyl group

Ketones

Aldehydes

Carboxylic acids

Carboxyl group

Ester

Polymer

Addition polymerization

Condensation
polymerization

Copolymer

Homopolymer

Dimer

Polyester

20.4 Additional Organic Compounds

Aldehydes and ketones

- Contain the carbonyl —C— groups
 ||
 O

- Aldehyde R—C—H
 ||
 O

- ◆ Are named by using the parent alkane root and *-al* ending

- Ketone R—C—R'
 ||
 O

- ◆ Are named by using the parent alkane root and *-al* ending

Carboxylic acids

- Contain the carbonyl —C=O group
 ||
 O—H

- Named by dropping the *-e* from the parent alkane and adding *-oic* and the word *acid*

- React with alcohols to form esters

- Esters are sweet smelling compounds named by using the alkyl name from the alcohol followed by the acid name in which the *-ic* ending is replaced by *-ate*

Polymers

- Large molecules made by hooking small molecules (monomers) together

- Types of polymerization reactions

- ◆ Addition: where monomers simply add together

- ◆ Condensation: in which a small molecule (often water) is formed for each extension of the polymer chain

All exercises with **blue numbers** have answers in the back of this book.

20.1 Saturated Hydrocarbons

A. Carbon Bonding

1. What makes carbon able to form so many different compounds?
2. What is the maximum number of other atoms to which a given carbon atom can be attached? Why?
3. What does a double bond represent? Give two examples of molecules in which double bonds occur, and draw the Lewis structures of those molecules.
4. Structural representations of the hydrocarbon ethane are shown in Figure 20.2. To how many other atoms is each carbon atom bonded in ethane? What is the geometry around each carbon atom in ethane?

B. Alkanes

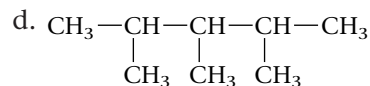
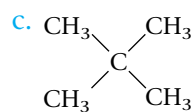
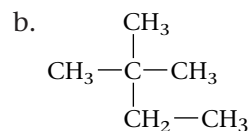
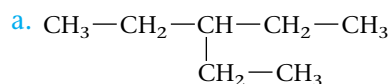
5. What does it mean to say that a hydrocarbon is *saturated*? To how many other atoms is each carbon atom in a saturated hydrocarbon bonded? What name is given to the family of saturated hydrocarbons?
6. Figure 20.3 shows the structures of the hydrocarbons propane and butane. Alkanes that contain an unbranched chain of carbon atoms are often called straight-chain alkanes. Why is such an unbranched chain not really “straight”? What are the bond angles between the carbon atoms in propane and butane?
7. For each of the following straight-chain alkanes, draw the structural formula.
 - a. heptane
 - b. nonane
 - c. propane
 - d. decane
8. Draw the structural formula for each of the following straight-chain alkanes.
 - a. butane
 - b. octane
 - c. pentane
 - d. hexane

C. Structural Formulas and Isomerism

9. What are structural *isomers*? What is the smallest alkane that has a structural isomer? Draw structures to illustrate the isomers.
10. What is a small chain of carbon atoms called when it is attached to a longer chain of carbon atoms?
11. Draw structural formulas for at least *five* isomeric alkanes having the formula C_6H_{14} .

D. Naming Alkanes

12. Give the *root names* for the alkanes with five through ten carbon atoms.
13. To what does the root name for a *branched* hydrocarbon correspond?
14. How is the *position* of substituents along the longest chain of a hydrocarbon indicated?
15. In giving the name of a hydrocarbon with several substituents, in what order do we list the substituents?
16. Give the systematic name for each of the following alkanes.



17. Draw structural formulas for each of the following compounds.
 - a. 2-methylhexane
 - b. 3-methylhexane
 - c. 2,2-dimethylhexane
 - d. 2,3-dimethylhexane
 - e. 3,3-dimethylhexane

E. Petroleum

18. What are the major constituents of crude petroleum? What is the major constituent of natural gas? How were these mixtures formed?
19. List several petroleum fractions and tell how each fraction is primarily used.
20. How is petroleum separated into its fractions?
21. What source of energy is the United States most dependent on today? How does this compare to 100 years ago?

F. Reactions of Alkanes

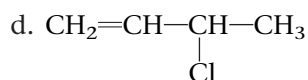
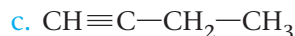
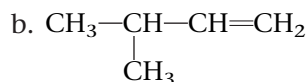
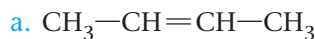
22. Explain why alkanes are relatively unreactive.
23. What is a *combustion* reaction? How have we made use of the combustion reactions of alkanes?
24. What is a *substitution* reaction? Give an example of a reactive molecule that is able to replace the hydrogen atoms of an alkane.
25. What is a *dehydrogenation* reaction? What results when an alkane is dehydrogenated?
26. Complete and balance each of the following chemical equations.
- $\text{C}_6\text{H}_{14}(l) + \text{O}_2(g) \rightarrow$
 - $\text{CH}_4(g) + \text{Cl}_2(g) \rightarrow$
 - $\text{CHCl}_3(l) + \text{Cl}_2(g) \rightarrow$

20.2 Unsaturated Hydrocarbons

A. Alkenes and Alkynes

27. What is an alkene? What structural feature characterizes alkenes? Give the general formula for alkenes.
28. What is an alkyne? What structural feature characterizes alkynes? Give the general formula for alkynes.
29. Complete each of the following chemical equations:
- $\text{CH}\equiv\text{C}-\text{CH}_3(g) + \text{H}_2(g) \rightarrow?$
 - $\text{CH}_3-\text{CH}=\text{CH}-\text{CH}_3(l) + \text{Br}_2(l) \rightarrow?$
 - $\text{CH}_3-\text{C}\equiv\text{C}-\text{CH}_3(l) + \text{O}_2(g) \rightarrow?$

30. Give the systematic name for each of the following unsaturated hydrocarbons and substituted unsaturated compounds.



31. Draw structural formulas, and give the systematic names, for at least *four* isomeric hydrocarbons containing seven carbon atoms and one double bond.

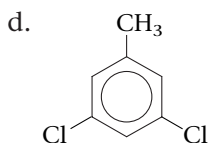
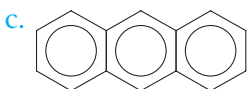
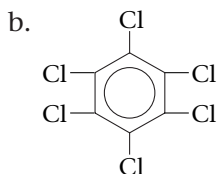
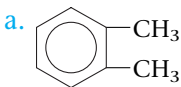
B. Aromatic Hydrocarbons

32. What structure do all aromatic hydrocarbons have in common?
33. Benzene exhibits resonance. Explain this statement in terms of the different Lewis structures that can be drawn for benzene.

C. Naming Aromatic Compounds

34. How is a monosubstituted benzene named? Give the structures and names of two examples. Also give two examples of monosubstituted benzenes that have special names.
35. What do the prefixes *ortho-*, *meta-*, and *para-* refer to in terms of the relative location of substituents in a disubstituted benzene?
36. Draw a structural formula for each of the following aromatic or substituted aromatic compounds.
- naphthalene
 - 2-bromophenol
 - 3-methylstyrene
 - 4-nitrochlorobenzene
 - 1,3-dinitrobenzene

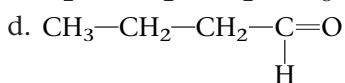
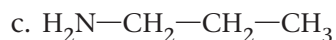
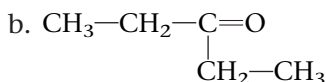
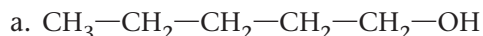
37. Name each of the following aromatic or substituted aromatic compounds.



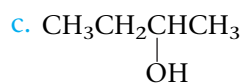
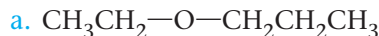
20.3 Introduction to Functional Groups and Alcohols

A. Functional Groups

38. On the basis of the functional groups listed in Table 20.6, identify the family of organic compounds to which each of the following belongs.



39. On the basis of the functional groups listed in Table 20.5, identify the family of organic compounds to which each of the following belongs.

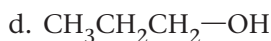
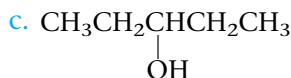
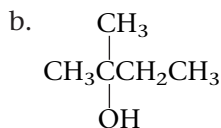


B. Alcohols

40. What functional group characterizes an alcohol? What ending is added to the name of the parent hydrocarbon to show that a molecule is an alcohol?

41. Distinguish among primary, secondary, and tertiary alcohols. Give a structural formula for an example of each type.

42. Give the systematic name for each of the following alcohols. Indicate whether the alcohol is primary, secondary, or tertiary.



43. Draw structural formulas for each of the following alcohols. Indicate whether the alcohol is primary, secondary, or tertiary.

a. 1-pentanol

b. 2-pentanol

c. 3-pentanol

d. 3-methyl-3-pentanol

C. Properties and Uses of Alcohols

44. Why is methanol sometimes called wood alcohol? Describe the modern synthesis of methanol. What are some uses of methanol?

45. Write the equation for the fermentation of glucose to form ethanol. Why can't ethanol solutions of greater than about 13% concentration be made directly by fermentation? How can the ethanol content be increased beyond this level in beverages?

46. Write the equation for the synthesis of ethanol from ethylene. What are some commercial uses of ethanol made by this process?

47. Give the names and structural formulas of two other commercially important alcohols. Cite the major use of each.

20.4 Additional Organic Compounds

A. Aldehydes and Ketones

48. What functional group is *common* to both aldehydes and ketones?

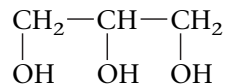
49. What structural feature *distinguishes* aldehydes from ketones?

Critical Thinking

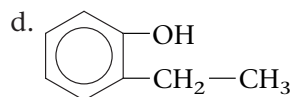
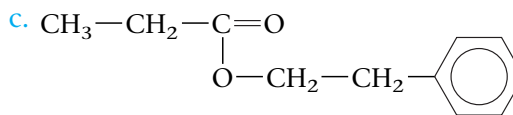
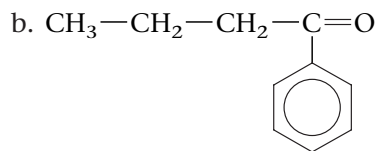
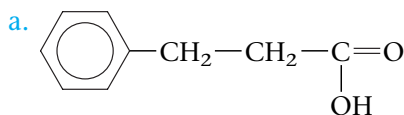
- 60.** Write the *general* formula for a normal alkane. How does a given alkane differ from the previous or the following member of the series?
- 61.** Draw structural formulas for all isomeric alkanes having the general formula C_5H_{12} .
- 62.** Give the systematic name of each of the following substituted alkanes.
- $CH_3-\underset{\substack{| \\ Cl}}{CH}-CH_2-CH_3$
 - $\underset{\substack{| \\ Br}}{CH_2}-\underset{\substack{| \\ Br}}{CH_2}$
 - CHI_3
 - $CH_3-\underset{\substack{| \\ Cl}}{CH}-\underset{\substack{| \\ Cl}}{CH}-\underset{\substack{| \\ Cl}}{CH}-CH_3$
 - $\begin{array}{ccccccc} & & Cl & & & & \\ & & | & & & & \\ CH_3 & - & C & - & CH_2 & - & CH & - & CH_2 & - & CH_2 & - & CH_3 \\ & & | & & & & | & & & & & & \\ & & Cl & & & & CH_3 & - & CH & - & CH_3 & & \end{array}$

- 63.** Write a structural formula for each of the following compounds.
- 2,3-dimethylheptane
 - 2,2-dimethyl-3-chloro-1-octanol
 - 2-chloro-1-hexene
 - 1-chloro-2-hexene
 - 2-methylphenol
- 64.** Draw the structural formula(s) for, and give the name(s) of, the organic product(s) of each of the following reactions. If a mixture of several, similar products is expected, indicate the type of product expected.
- $CH_3-CH_3 + Cl_2 \xrightarrow{\text{Light}}$
 - $CH_3-CH=CH-CH_3 + H_2 \xrightarrow{Pt}$
 - $CH_3-CH_2-\underset{\substack{| \\ CH_3-CH_2}}{C=CH}-CH_3 + Br_2 \longrightarrow$

- 65.** The alcohol glycerol (glycerine), which is produced in the human body by the digestion of fats, has the following structure. Give the systematic name of glycerol.



- 66.** The sugar glucose could conceivably be given the name 2,3,4,5,6-pentahydroxyhexanal (though this name is never used, because the actual sugar glucose is only one possible isomer having this name). Draw the structure implied by this name.
- 67.** Draw a structure corresponding to each of the following names.
- 2-methylpentanal
 - 3-hydroxybutanoic acid
 - 2-aminopropanal
 - 2,4-hexanedione
 - 3-methylbenzaldehyde
- 68.** On the basis of the functional groups listed in Table 20.6, identify the family of organic compounds to which each of the following belongs.

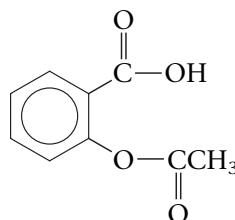


- 1 Which of the following is a saturated hydrocarbon?
- A C_3H_6
 B C_2H_4
 C C_4H_{10}
 D C_5H_{10}
- 2 While studying together, your friend names a molecule 3-methyl-4-isopropylpentane. You point out that although the correct structure could be drawn from this name, the name did not follow systematic rules. What is the correct systematic name for the molecule?
- A 3,4,5-trimethylhexane
 B 2,3,4-trimethylhexane
 C 2-isopropyl-3-methylpentane
 D 1,1,2,3-tetramethylpentane
- 3 Which of the following names is a correct one?
- A 3,4-dichloropentane
 B 1,1-dimethyl-2,2-diethylbutane
 C 4-pentene
 D 2-bromo-1-chloro-4,4-diethyloctane
- 4 The name for the alkene with the skeletal structure $C-C-C=C$ is
- A 3-butene, because the double bond is after the third carbon.
 B 2-butene because in this case you count from the right, and the double bond is before the second carbon.
 C 1-butene because in this case you count from the right, and the double bond is after the first carbon.
 D 4-butene, because the double bond is on a molecule with 4 carbon atoms.
- 5 Which of the following has a $C=O$ bond?
- A ester, aldehyde, secondary alcohol, ketone
 B secondary alcohol, ketone, aldehyde
 C ester, aldehyde, ketone
 D carboxylic acid, ether, tertiary alcohol

- 6 How many isomers does pentane have?

A 4
 B 3
 C 2
 D 1

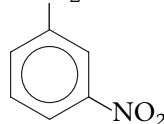
- 7 What functional groups are present in aspirin (acetylsalicylic acid) as shown below?



A carboxylic acid, alcohol
 B ether, alcohol
 C carboxylic acid, ester
 D ketone, ether

- 8 Name each of the following molecules.

a. NO_2



b. CH_3CHCH_2OH
 $\quad |$
 $\quad CH_3$

c. CH_3CH_2C-H
 $\quad \quad \quad ||$
 $\quad \quad \quad O$

d. $CH_3CCH_2CH_3$
 $\quad \quad ||$
 $\quad \quad O$

- 9 There are two isomers with the formula C_2H_6O .

a. Draw Lewis structures for each of these molecules.
 b. Circle and label the functional group for each of the molecules.
 c. Which of the molecules would exhibit the stronger intermolecular forces? Explain your answer.


LOOKING AHEAD

21.1 Introduction to Proteins

- A. Proteins
- B. Primary Structure of Proteins
- C. Secondary Structure of Proteins
- D. Tertiary Structure of Proteins
- E. Functions of Proteins
- F. Enzymes

21.2 Carbohydrates, Nucleic Acids, and Lipids

- A. Carbohydrates
- B. Nucleic Acids
 - DNA and Protein Synthesis
- C. Lipids

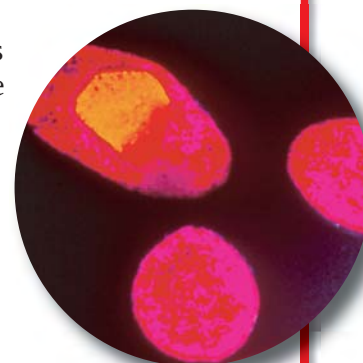
A photograph of a swimmer in a blue cap and goggles swimming in a pool. The swimmer is wearing a green and blue swimsuit. The water is clear blue with ripples. A white lane line is visible in the background.

The energy to exercise is furnished by chemical reactions in the body.

IN YOUR LIFE

Biochemistry, the study of the chemistry of living systems, is a vast and exciting field in which important discoveries about how life is maintained and how diseases occur are being made every day. In particular, there has been rapid growth in the understanding of how living cells manufacture and use the molecules necessary for life. This not only has been beneficial for detection and treatment of diseases but also has spawned a new field—**biotechnology**, which uses nature’s “machinery” to synthesize desired substances. For example, insulin is a complex biomolecule that is used in the body to regulate the metabolism of sugars. People who have diabetes are deficient in natural insulin and must take insulin by injection or other means. In the past this insulin was obtained from animal tissues (particularly that of cows). However, our increased understanding of the biochemical processes of the cell has allowed us to “farm” insulin. We have learned to insert the “instructions” for making insulin into the cells of bacteria such as *E. coli* so that as the bacteria grow they produce insulin that can then be harvested for use by diabetics. Many other products, including natural pesticides, are also being produced by the techniques of biotechnology.

An understanding of biochemistry also allows our society to produce healthier processed foods. For example, the food industry is making great efforts to reduce the fat levels in food without destroying the good taste that fats bring to food.



*This sample of *E. coli* bacteria has been genetically altered to produce human insulin. The insulin production sites are colored orange. The bacteria are cultured and pure human insulin is harvested from them.*

WHAT DO YOU KNOW?



Prereading Questions

1. What is meant by the phrase “like dissolves like”?
2. What do catalysts do to a chemical reaction?
3. What is a carbohydrate?
4. What do you know about DNA?
5. What do you know about the term “gene”?

SECTION 21.1

Introduction to Proteins

Key Terms

- Essential elements
- Trace elements
- Cell
- Proteins
- Fibrous proteins
- Globular proteins
- Alpha-amino acids
- Side chains
- Dipeptide
- Peptide linkage
- Polypeptide
- Primary structure
- Secondary structure
- Alpha-helix
- Pleated sheet
- Tertiary structure
- Disulfide linkage
- Denaturation
- Enzymes
- Lock-and-key model
- Substrate
- Active site

Objectives

- To learn about proteins
- To understand the primary structure of proteins
- To understand the secondary structure of proteins
- To understand the tertiary structure of proteins
- To learn the functions of proteins
- To understand how enzymes work

At present, 30 elements are definitely known to be essential to human life. These **essential elements** are shown in color in **Figure 21.1**. The most abundant elements are hydrogen, carbon, nitrogen, and oxygen; sodium, magnesium, potassium, calcium, phosphorus, sulfur, and chlorine are also present in relatively large amounts. Although present only in trace amounts, the first-row transition metals are essential for the action of many enzymes (biological catalysts). For example, zinc, one of the **trace elements**, is found in nearly 200 biologically important molecules. The functions of the essential elements are summarized in **Table 21.1**. In time, other elements will probably be found to be essential.

Life is organized around the functions of the **cell**, the smallest unit in living things that exhibits the properties normally associated with life, such as reproduction, metabolism, mutation, and sensitivity to external stimuli.

As the fundamental building blocks of all living systems, aggregates of cells form tissues, which in turn are assembled into the organs that make up complex living systems. Therefore, to understand how life is maintained and reproduced, we must learn how cells operate on the molecular level. This is the main thrust of biochemistry.

Essential elements

The 30 elements currently known to be essential to human life

Trace elements

First-row transition metals that are present in very small (trace) amounts in the human body

1																	8
1																	2
H																	He
3	4											5	6	7	8	9	10
Li	Be											B	C	N	O	F	Ne
11	12											13	14	15	16	17	18
Na	Mg											Al	Si	P	S	Cl	Ar
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
87	88	89	104	105	106	107	108	109	110	111	112	113	114	115			
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Uub	Uut	Uuq	Uup			
58	59	60	61	62	63	64	65	66	67	68	69	70	71				
Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu				
90	91	92	93	94	95	96	97	98	99	100	101	102	103				
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lw				

Figure 21.1

The chemical elements essential for life. Those most abundant in living systems are shown as purple. Nineteen elements, called the trace elements, are shown as green.

Table 21.1

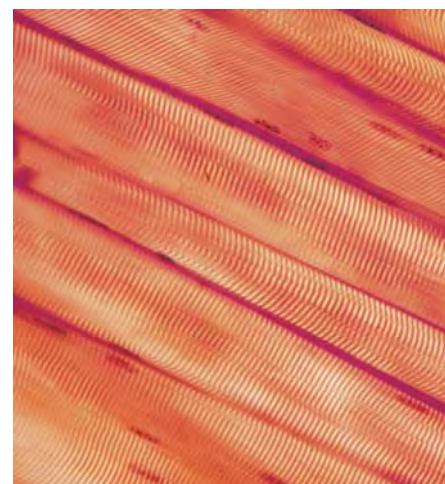
Some Essential Elements and Their Major Functions

Element	Percent by Mass in the Human Body	Function
oxygen	65	in water and many organic compounds
carbon	18	in all organic compounds
hydrogen	10	in water and many inorganic and organic compounds
nitrogen	3	in both inorganic and organic compounds
calcium	1.5	in bone; essential to some enzymes and to muscle action
phosphorus	1.2	essential in cell membranes and to energy transfer in cells
potassium	0.2	cation in cell fluid
chlorine	0.2	anion inside and outside the cells
sulfur	0.2	in proteins
sodium	0.1	cation in cell fluid
magnesium	0.05	essential to some enzymes
iron	<0.05	in molecules that transport and store oxygen
zinc	<0.05	essential to many enzymes
cobalt	<0.05	found in vitamin B ₁₂
iodine	<0.05	essential to thyroid hormones
fluorine	<0.01	in teeth and bones

A. Proteins

In Chapter 20 we saw that many useful synthetic materials are polymers. A great many natural materials are also polymers: starch, hair, silk and cotton fibers, and the cellulose in woody plants, to name only a few.

In this section we introduce a class of natural polymers, the **proteins**, which make up about 15% of our bodies and have molar masses that range from approximately 6,000 to over 1,000,000 grams. Proteins have many functions in the human body. **Fibrous proteins** provide structural integrity and strength for many types of tissue and are the main components of muscle, hair, and cartilage. Other proteins, usually called **globular proteins** because of their roughly spherical shape, are the “worker” molecules of the body. These proteins transport and store oxygen and nutrients, act as catalysts for the thousands of reactions that make life possible, fight invasion of the body by foreign objects, participate in the body’s many regulatory systems, and transport electrons in the complex process of metabolizing nutrients.



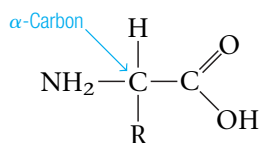
Striated muscle tissue is composed of fibrous proteins.

B. Primary Structure of Proteins

α -amino acid

An organic acid in which an amino group is attached to the carbon atom next to the carboxyl group

The building blocks of all proteins are the α -amino acids:



The R in this structure may represent H, CH₃, or more complex substituents. These molecules are called α -amino acids because the amino group (—NH₂) is always attached to the α -carbon, the one next to the carboxyl group (—COOH). The 20 amino acids most commonly found in proteins are shown in **Figure 21.2**.

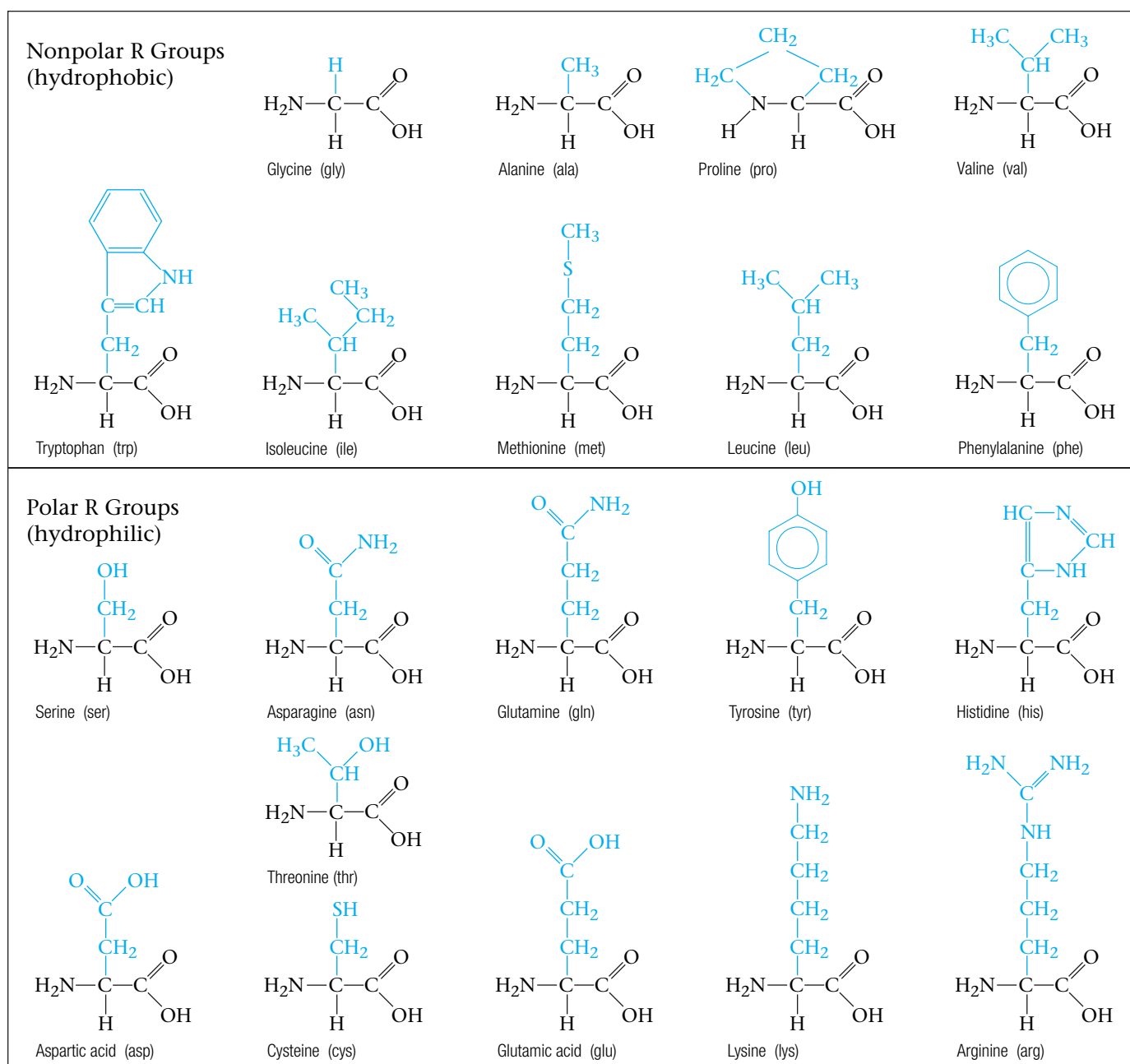


Figure 21.2

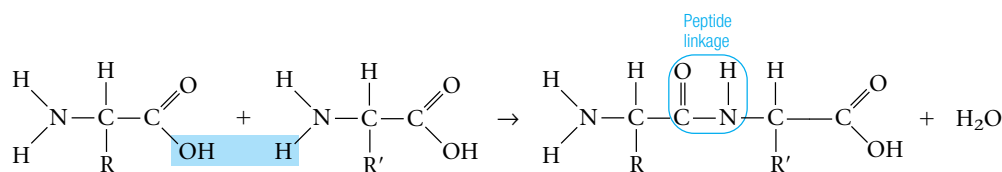
The 20 α -amino acids found in most proteins. The R group is shown in color.

Note from Figure 21.2 that the amino acids are grouped into polar and nonpolar classes on the basis of the composition of the R groups, also called the **side chains**. Nonpolar side chains contain mostly carbon and hydrogen atoms; polar side chains contain nitrogen and oxygen atoms. This difference is important because polar side chains are *hydrophilic* (water-loving), but nonpolar side chains are *hydrophobic* (water-fearing). This greatly affects the three-dimensional structure of the resulting proteins, because they exist in aqueous media in living things.

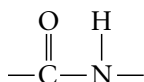
Active Reading Question

How does the nature of the side chain affect the three dimensional structure of a protein?

The protein polymer is built by reactions between amino acids. For example, two amino acids can react as follows, forming a C—N bond with the elimination of water.



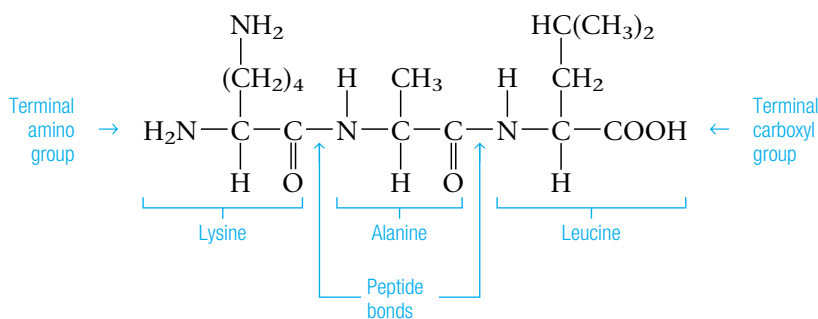
The product shown is called a **dipeptide**. The term *peptide* comes from the structure



which chemists call a **peptide linkage** or a peptide bond. The prefix *di-* indicates that two amino acids have been joined. Additional reactions lengthen the chain to produce a **polypeptide** and eventually a protein.

The 20 amino acids can be assembled in any order, which makes possible an enormous number of different proteins. This variety allows for the many types of proteins needed for the functions that organisms carry out.

The *order* or *sequence* of amino acids in the protein chain is called the **primary structure**. We indicate the primary structure by using three-letter codes for the amino acids (see Figure 21.2), where it is understood that the terminal carboxyl group is on the right and the terminal amino group is on the left. For example, one possible sequence for a tripeptide containing the amino acids lysine, alanine, and leucine is



which is represented in the shorthand notation (three-letter codes) as lys-ala-leu

Information

At the pH in biological fluids, the amino acids shown in Figure 21.2 exist in a different form, with the proton of the —COOH group transferred to the —NH₂ group. For example, glycine would be in the form ⁺H₃NCH₂COO⁻.

Primary structure

The order or sequence of amino acids in a protein chain

EXAMPLE 21.1

Understanding Primary Structure

Write the sequences of all possible tripeptides composed of the amino acids tyrosine (tyr), histidine (his), and cysteine (cys).

Solution

There are six possible sequences:

tyr-his-cys his-tyr-cys cys-tyr-his
tyr-cys-his his-cys-tyr cys-his-tyr

Secondary structure

The arrangement in space of the chain of the protein chain

A Closer Look The importance of the primary structure of polypeptides can be seen in the differences between *oxytocin* and *vasopressin*. Oxytocin is a hormone that triggers contraction of the uterus and milk secretion. Vasopressin raises blood pressure and regulates kidney function. Both of these molecules are nine-unit polypeptides, and they differ by only two amino acids.

cys-tyr-**ile**-gln-asn-cys-pro-**leu**-gly
oxytocin

cys-tyr-**phe**-gln-asn-cys-pro-**arg**-gly
vasopressin

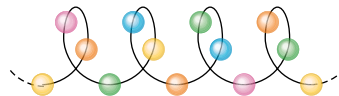
These molecules have completely different functions in the human body.

C. Secondary Structure of Proteins

So far we have considered the primary structure of proteins—the order of amino acids in the chain as shown below.



A second level of structure in proteins is the arrangement in space of the chain of the long molecule. This is called the **secondary structure** of the protein.



One common type of secondary structure resembles a spiral staircase. This spiral structure is called an **α -helix** (see **Figure 21.3**). A spiral-like secondary structure gives the protein elasticity (springiness) and is found in the fibrous proteins in wool, hair, and tendons. Another type of secondary structure involves joining several different protein chains in an arrangement called a **pleated sheet**, as shown below.

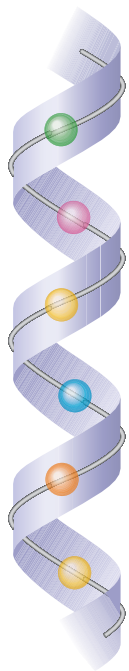
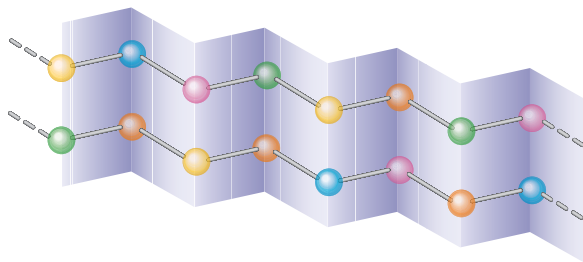


Figure 21.3

One type of secondary protein structure is like a spiral staircase and is called the α -helix. The spheres represent individual amino acids.

Silk has this arrangement of proteins, making its fibers flexible but very strong and resistant to stretching. The pleated sheet is also found in muscle fibers.

Active Reading Question

What is it about a pleated sheet arrangement that makes it suitable for muscle fibers?

As you might imagine, a molecule as large as a protein has a great deal of flexibility and can assume a variety of overall shapes. The function the protein is to serve influences the specific shape that it will have. For long, thin structures such as hair, wool and silk fibers, and tendons, an elongated shape is required. This may involve an α -helical secondary structure, as found in the protein α -keratin in hair and wool or in the collagen that occurs in tendons (see **Figure 21.4a**), or it may involve a pleated-sheet secondary structure, as found in silk (see **Figure 21.4b**).

Many of the proteins in the body that have nonstructural functions (such as serving as enzymes) are globular. One is myoglobin (see **Figure 21.5**), which absorbs an O_2 molecule and stores it for use by the cells as it is needed. Note that the secondary structure of myoglobin is basically α -helical. However, in the areas where the chain bends to give the protein its compact globular structure, the α -helix breaks down so that the protein can “turn the corner.”

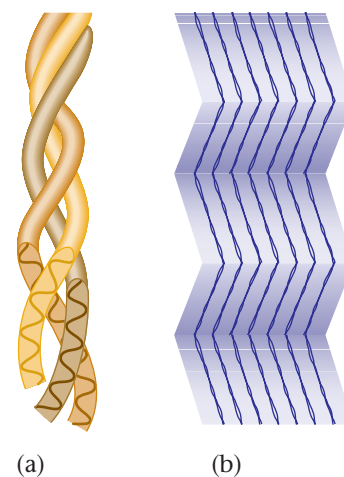


Figure 21.4

(a) Collagen, a protein found in tendons, consists of three protein chains (each with an α -helical structure) twisted together to form a super-helix. The result is a long, relatively narrow protein. (b) Many proteins are bound together in the pleated-sheet arrangement to form the elongated protein found in silk fibers.

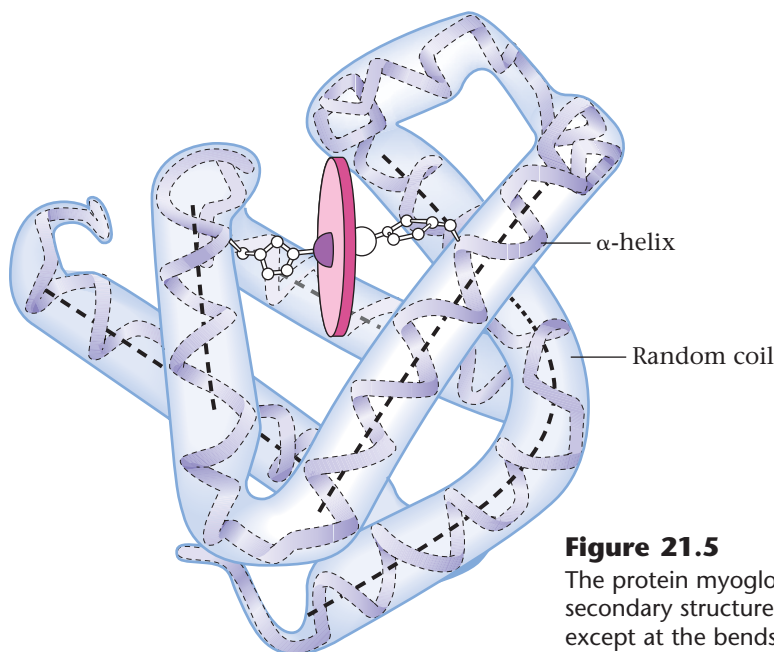


Figure 21.5

The protein myoglobin. Note that the secondary structure is basically α -helical except at the bends. The tertiary structure is globular.

D. Tertiary Structure of Proteins

Tertiary structure

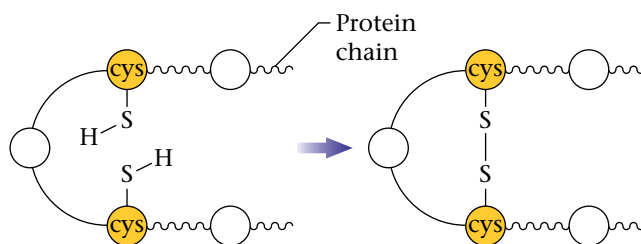
The overall shape of a protein

The overall shape of the protein, long and narrow or globular, is called its **tertiary structure**. To make sure the difference between secondary and tertiary structure is clear, examine Figure 21.5 again. The secondary structure of the myoglobin protein is helical except in the regions where it bends back on itself to give the overall compact (globular) tertiary structure. If the bends did not occur, the tertiary structure would be like a long tube (tubular). In both tertiary arrangements (globular and tubular), the secondary structure of the protein is basically helical.

The amino acid *cysteine* (cys) plays a special role in stabilizing the tertiary structure of many proteins, because the —SH groups on two cysteines can react to form an S—S bond called a **disulfide linkage**.



Chemicals used in permanent wave solutions create curls in hair.



The formation of a disulfide linkage can fasten together two parts of a protein chain to form and hold a bend in the chain, for example. A practical application of the chemistry of disulfide bonds is the permanent waving of hair (see **Figure 21.6**). The S—S linkages in the protein of hair are broken by treatment with a reducing agent. Next the hair is set in curlers to change the tertiary protein structure to the desired shape. Then treatment with an oxidizing agent causes new S—S bonds to form, which make the hair protein retain the new structure.

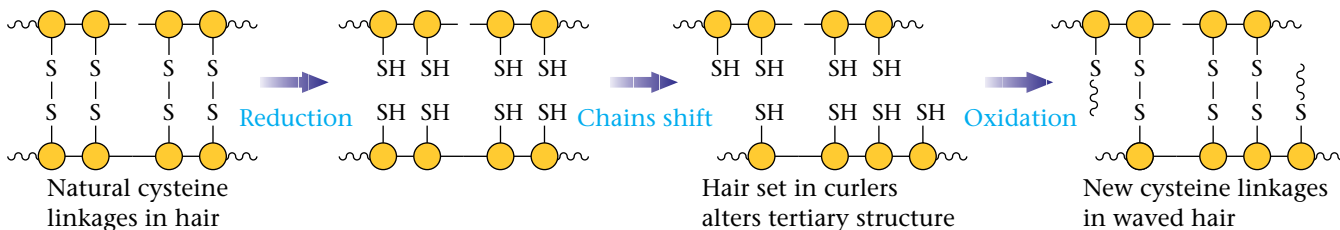


Figure 21.6

The permanent waving of hair

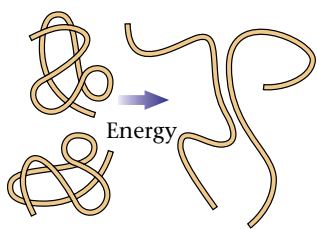


Figure 21.7

A schematic representation of the thermal denaturation of a protein

E. Functions of Proteins

The three-dimensional structure of a protein is crucial to its function. The process of breaking down this structure is called **denaturation** (see **Figure 21.7**). For example, heat causes the denaturation of egg proteins when an egg is cooked. Any source of energy can cause denaturation of proteins and is thus potentially dangerous to living organisms. For example, ultraviolet radiation, X-ray radiation, or nuclear radioactivity can disrupt protein structure, which may lead to cancer or genetic damage. The metals lead and mercury, which have a very high affinity for sulfur, cause protein denaturation by disrupting disulfide bonds.

The tremendous variability in the several levels of protein structure allows proteins to be tailored specifically to serve a wide range of functions, some of which are given in **Table 21.2**.

F. Enzymes

Enzymes are proteins that catalyze specific biological reactions. Without the several hundred enzymes now known, life would be impossible. Almost all of the critical biochemical reactions would occur far too slowly at the temperatures at which life exists. Enzymes are impressive for their tremendous efficiency (they are typically 1 to 10 million times as efficient as inorganic catalysts) and their incredible selectivity (an enzyme “ignores” the thousands of molecules in body fluids that are not involved in the reaction it catalyzes).

Enzyme

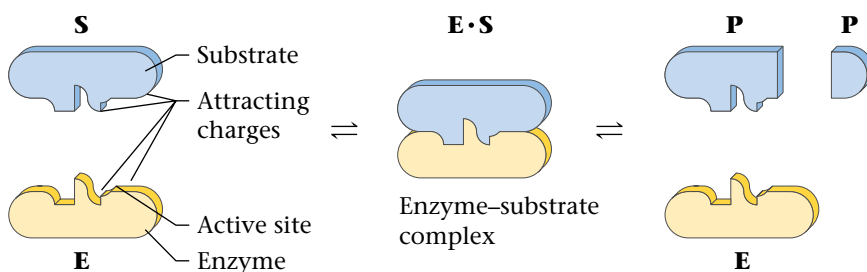
A protein that catalyzes a specific biological reaction

Table 21.2

Common Functions of Proteins

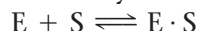
Function	Comment/Example
Structure	Proteins provide the strength of tendons, bones, and skin. Cartilage, hair, wool, fingernails, and claws are mainly protein. Viruses have an outer layer of protein.
Movement	Proteins are the major components of muscles and are directly responsible for the ability of muscles to contract. The swimming of sperm results from the contraction of protein filaments in their tails.
Catalysis	Nearly all chemical reactions in living organisms are catalyzed by enzymes, which are almost always proteins.
Transport	Oxygen is carried from the lungs to tissues by the protein hemoglobin in red blood cells.
Storage	The protein ferritin stores iron in the liver, spleen, and bone marrow.
Energy transformation	Cytochromes are proteins found in all cells. They extract energy from food molecules by transferring electrons in a series of oxidation–reduction reactions.
Protection	Antibodies are special proteins that are synthesized in response to foreign substances and cells, such as bacterial cells. They then bind to those substances or cells and provide us with immunity to various diseases. Interferon, a small protein made and released by cells when they are exposed to a virus, protects other cells against viral infection. Blood-clotting proteins protect against bleeding (hemorrhage).
Control	Many hormones are proteins produced in the body that have specific effects on the activity of certain organs.
Buffering	Because proteins contain both acidic and basic groups on their side chains, they can neutralize both acids and bases and thus provide buffering for blood and tissues.

Figure 21.8
Schematic diagram of the lock-and-key model of the functioning of an enzyme

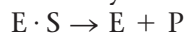


Although the mechanisms of enzyme activity are complex and not fully understood in most cases, a simple theory called the **lock-and-key model** (see **Figure 21.8**) seems to fit many enzymes. This model postulates that the shapes of the reacting molecule (the **substrate**) and the enzyme are such that they fit together much as a key fits a specific lock. The substrate and enzyme attach to each other in such a way that the part of the substrate where the reaction is to occur occupies the **active site** of the enzyme. After the reaction occurs, the products are liberated and the enzyme is ready for a new substrate. We can represent enzyme catalysis by the following steps:

Step 1 The enzyme E and the substrate S come together.



Step 2 The reaction occurs to give the product P, which is released from the enzyme.



After the product is released, the enzyme is free to engage another substrate.

Because this process occurs so rapidly, only a tiny amount of enzyme is required.

i nformation

An enzyme often changes shape slightly as the substrate is bound. This “clamps” the substrate into place.

Active Reading Question

What is the difference between a substrate and an enzyme? How are they related?

In some cases a substance other than the substrate can bind to the enzyme's active site. When this occurs, the enzyme is said to be *inhibited*. If the inhibition is permanent, the enzyme is said to be inactivated. Some of the most powerful toxins act by inhibiting, or inactivating, key enzymes.

Because enzymes are crucial to life and because we hope to learn how to mimic their efficiency in our industrial catalysts, the study of enzymes occupies a prominent place in chemical research.

CHEMISTRY IN YOUR WORLD

Science, Technology, and Society

Urine Farming

Nature is a very good chemist—organisms make hundreds of complex chemicals every second to survive. The rapidly expanding science of biotechnology is learning more every day about how to use natural chemical pathways to make valuable molecules. For example, insulin for diabetics is now made almost exclusively by “farming” genetically altered *E. coli* or yeast.

Turning livestock into four-footed pharmaceutical factories has seemed an attractive option ever since we learned to carry out “genetic engineering.” In fact, a great deal of effort has been expended on creating cows, sheep, and goats that are genetically modified to produce useful human proteins that are secreted into the animals’ milk, from which they can then be harvested. None of these proteins has yet reached the market—tests are now under way to check their efficacy and safety—but a blood-clotting agent, antithrombin III, developed by Genzyme Transgenics Corporation of Framingham, Massachusetts, is close to commercial application. Creating a transgenic (genetically modified) animal is expensive (costing approximately \$60,000), but then standard breeding techniques can lead to an entire herd of animals programmed to produce commercially interesting molecules.

As a medium, milk has some disadvantages. It is produced only by mature females and contains a wide variety of proteins that complicate the purification process to obtain the desired product. Urine may be a better alternative. It contains fewer natural proteins and is produced at an early age by both male and female animals. Robert J. Wall and his colleagues at the U.S. Department of Agriculture in Beltsville, Maryland, have developed transgenic mice that produce human growth hormone in the walls of their bladders. For obvious reasons, mice are not ideal for large-scale production of chemicals, but Wall’s experiments nevertheless show that the concept works. At present it is too early to know whether urine farming will prove feasible. Yields from the bladder are about 10,000 times lower than those from the mammary glands.

In addition, collecting urine from farm animals could prove to be a very tricky business.

A New Zealand goat, one of the animals used for urine farming

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SECTION 21.1

REVIEW QUESTIONS

- 1 What are the major types and functions of proteins?
- 2 What property of the side chain on an amino acid affects the three-dimensional structure of the protein?
- 3 Draw the general structure for an amino acid.
 - a. Circle the side chain group.
 - b. Draw a second amino acid next to the first and circle the water removed to form a peptide bond to link the two amino acids.
- 4 What are the differences among primary, secondary, and tertiary structure in proteins?
- 5 What level of protein structure is changed by denaturation?
- 6 What is the role of enzymes in biochemical reactions?
- 7 Explain the lock-and-key model of enzymes.

SECTION 21.2

Carbohydrates, Nucleic Acids, and Lipids

Key Terms

- Carbohydrates
- Monosaccharides (simple sugars)
- Disaccharide
- Sucrose
- Glycoside linkage
- Polysaccharides
- Starch
- Cellulose
- Glycogen
- DNA (deoxyribonucleic acid)
- RNA (ribonucleic acid)
- Nucleotide
- Protein synthesis
- Gene
- mRNA (messenger RNA)
- tRNA (transfer RNA)
- Lipids
- Fats
- Fatty acids
- Saponification
- Micelles
- Surfactant
- Phospholipids
- Waxes
- Steroids
- Cholesterol
- Adrenocorticoid hormones
- Sex hormones
- Bile acids

Objectives

- To understand the fundamental properties of carbohydrates
- To understand nucleic acid structures
- To learn the four classes of lipids

A. Carbohydrates

The **carbohydrates** are another class of biologically important molecules. Carbohydrates serve as food sources for most organisms and as structural materials for plants. Because many carbohydrates have the empirical formula CH_2O , it was originally believed that these substances were hydrates of carbon ($\text{C} \cdot \text{H}_2\text{O}$), which accounts for their name.

Like proteins, carbohydrates occur in almost bewildering varieties. Many of the most important carbohydrates are polymers—large molecules constructed by hooking together many smaller molecules. We have seen that proteins are polymers constructed from amino acids. The polymeric carbohydrates are constructed from molecules called **simple sugars** or, more precisely, **monosaccharides**. Monosaccharides are aldehydes or ketones that contain several hydroxyl ($-\text{OH}$) substituents. An example of a monosaccharide is fructose, with the structure

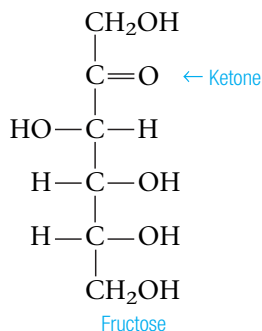


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Breads contain high levels of carbohydrates and fruits are a dietary source of fructose.

Fructose is a sugar found in honey and fruits. Monosaccharides can have various numbers of carbon atoms, and we name them according to the number of carbon atoms they contain by adding prefixes to the root *-ose*. The general names of monosaccharides are shown in **Table 21.3**. Notice that fructose is a ketone with six carbon atoms and five —OH substituents. Fructose is a member of the hexose family, where the prefix *hex-* means six.

Table 21.3**The General Names of Monosaccharides**

Number of Carbon Atoms	Prefix	General Name of Sugar
3	<i>tri-</i>	triose
4	<i>tetr-</i>	tetrose
5	<i>pent-</i>	pentose
6	<i>hex-</i>	hexose
7	<i>hept-</i>	heptose
8	<i>oct-</i>	octose
9	<i>non-</i>	nonose

Top Ten Sugar-Consuming Countries

Country	kg per Person per Year
Swaziland	320
Belize	280
Israel	276
Réunion	264
Costa Rica	251
New Zealand	242
Barbados	238
Brazil	231
Trinidad and Tobago	225
Malaysia	223

The most important monosaccharides found in living organisms are pentoses and hexoses. The most common pentoses and hexoses are shown in **Table 21.4**.

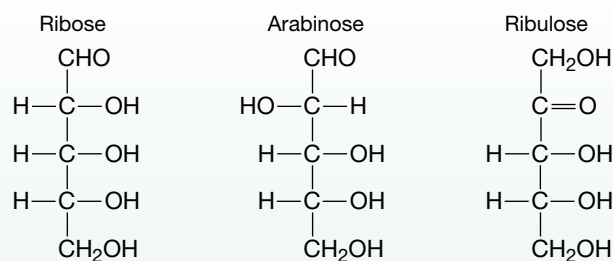
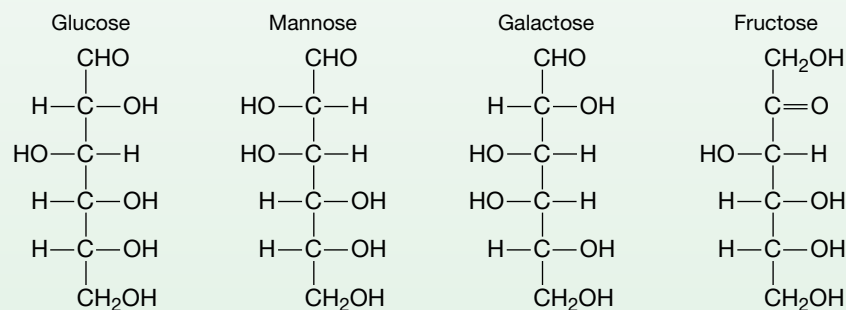
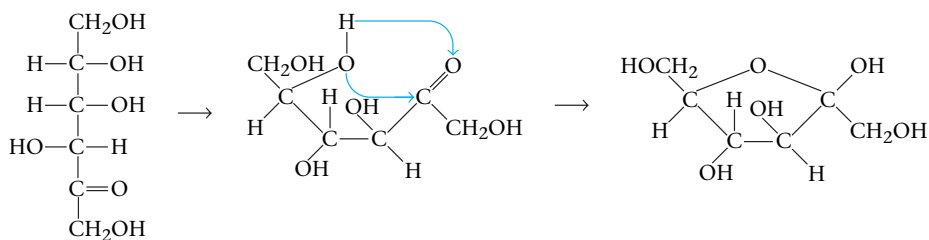
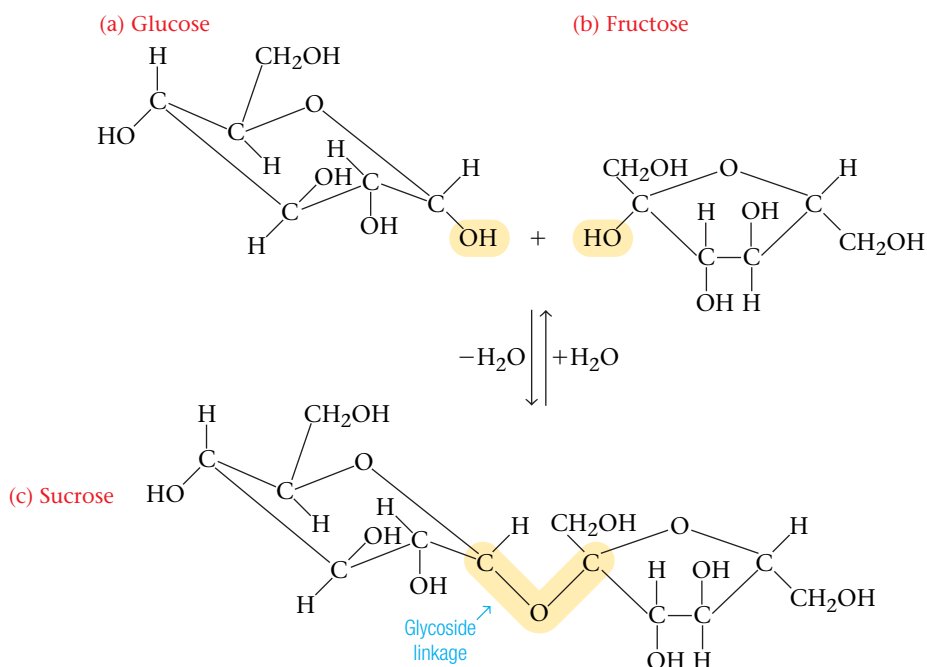
Table 21.4**Some Important Monosaccharides****Pentoses (five carbon atoms)****Hexoses (six carbon atoms)**

Figure 21.9
The formation of a ring structure for fructose



Although we have so far represented the monosaccharides as straight-chain molecules, they usually form ring structures in aqueous solution. **Figure 21.9** shows how the ring forms for fructose. Note that a new bond is formed between the oxygen of a hydroxyl group and the carbon of the ketone group. In the cyclic form, fructose is a five-membered ring containing a C—O—C bond. Glucose, a hexose that is an aldehyde, forms a six-membered ring structure. The six-membered glucose ring is shown in **Figure 21.10a**. Note that the ring is nonplanar.

Figure 21.10
Sucrose (c) is a disaccharide formed from glucose (a) and fructose (b)



Disaccharide

A sugar formed from two monosaccharides formed by a glycoside linkage

Sucrose

Table sugar; a disaccharide formed between glucose and fructose

Glycoside linkage

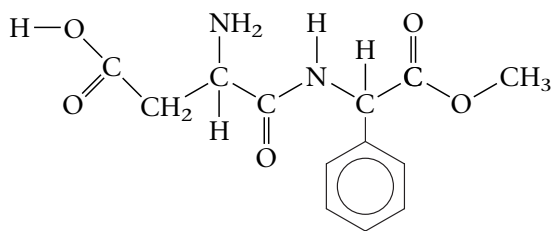
The C—O—C bond formed between monosaccharide rings in a disaccharide

More complex carbohydrates are formed by the combining of monosaccharides. For example, two monosaccharides can combine to form a **disaccharide**. **Sucrose**, common table sugar, is a disaccharide formed from glucose and fructose by elimination of water to form a C—O—C bond between the rings that is called a **glycoside linkage** (see **Figure 21.10c**). When sucrose is consumed in food, this reaction is reversed (the glycoside bond is broken). An enzyme in saliva catalyzes the breakdown of sucrose into its two monosaccharides.

Active Reading Question

What elements are involved in a glycoside linkage?

CELEBRITY CHEMICAL

Aspartame ($C_{14}H_{18}N_2O_5$)

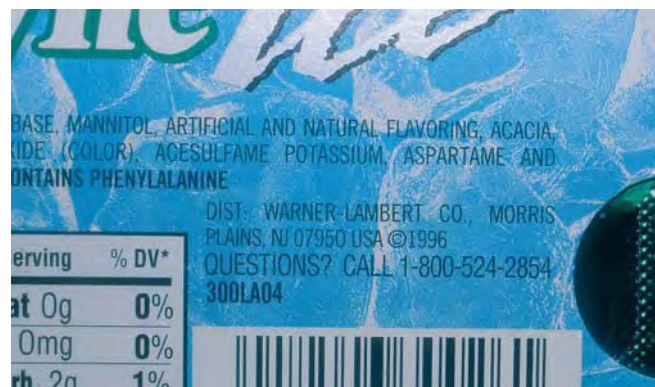
Most humans love things that taste sweet—chocolate candy, cake, ice cream, and so on. The problem with sweet things is they have lots of calories and tend to make us gain weight. Wouldn't it be great to have sweets without the calories? Of course, this dream has now become reality through the magic of chemistry—artificial sweeteners, or molecules that taste sweet but do not add to the calorie count. Many molecules have a sweet taste—even sweeter than sugar. The earliest of these compounds were discovered in the late 1800s when tasting a new compound was accepted procedure. Most of the current artificial sweeteners were discovered by accident, however, when scientists were not observing proper lab hygiene.

Today a widely used artificial sweetener is a compound called aspartame, which is formed by coupling together two amino acids naturally found in humans: aspartic acid and phenylalanine

Aspartame is 200 times sweeter than sugar and leaves no unpleasant aftertaste. It is now found in most diet drinks and many other low-calorie products. One disadvantage of aspartame is that it

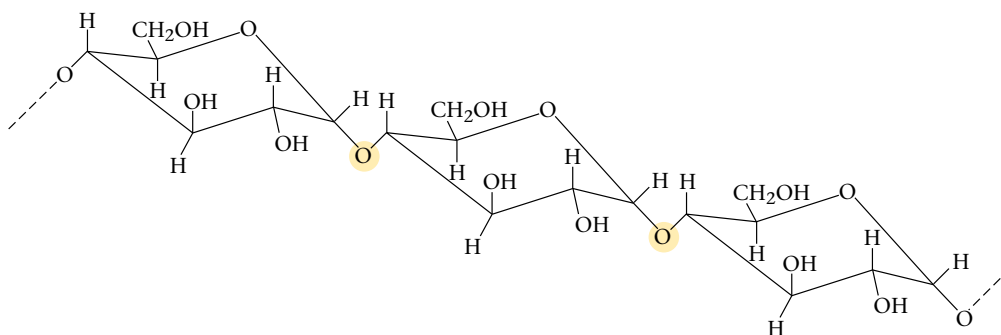
decomposes at baking temperatures and thus cannot be used in products that are heated.

Like all artificial sweeteners, aspartame has had some critics. When this substance breaks down in the body it forms aspartic acid, phenylalanine, and methanol. The first two compounds are naturally present in the body and pose no hazard to most people. However, about one person in 15,000 suffers from a disease called phenylketonuria, which makes an individual sensitive to excess phenylalanine. People with this condition must avoid aspartame. Also, the fact that methanol is a product of the breakdown of aspartame makes some people concerned, because methanol is a toxic substance. Nevertheless, studies have shown that the small amount of methanol produced from aspartame poses no risk. In fact, a glass of natural fruit juice contains as much methanol as is produced from typical use of aspartame.



The label on a product containing aspartame

Large polymers containing many monosaccharide units are called **polysaccharides** and can form when each ring forms two glycoside linkages, as shown in **Figure 21.11**. Three of the most important of these polymers are starch, cellulose, and glycogen. All of these substances are polymers of glucose; they differ in the way the glucose rings are linked together.

**Polysaccharide**

Polymers containing many monosaccharide units

Figure 21.11

Starch is made by hooking together many glucose rings. Only a small part of the chain is shown here.

Starch

The main carbohydrate reservoir in plants; a polymer of glucose

Cellulose

The major structural component of woody plants and natural fibers, such as cotton; a polymer of glucose

Glycogen

The main carbohydrate reservoir in animals; a polymer of glucose

Deoxyribonucleic acid (DNA)

A huge nucleotide polymer having a double-helical structure with complementary bases on the two strands; functions include protein synthesis and the storage and transport of genetic material

Ribonucleic acid (RNA)

Large nucleotide polymer that, along with DNA, functions to transport genetic material

Nucleotide

A monomer of DNA and RNA consisting of a nitrogen-containing base, a five-carbon sugar, and a phosphate group

Starch (Figure 21.11) is the carbohydrate reservoir in plants. It is the form in which glucose is stored by the plant for later use as cellular fuel—both by the plants themselves and by organisms that eat plants.

Cellulose, the major structural component of woody plants and natural fibers such as cotton, is also a polymer of glucose. However, in cellulose the way in which the glucose rings are linked is different from the way they are linked in starch. This difference in linkage has very important consequences. The human digestive system contains enzymes that can catalyze breakage of the glycoside bonds between the glucose molecules in starch. These enzymes are *not* effective on the glycoside bonds of cellulose, however, presumably because the different structure causes a poor fit between the enzyme's active site and the carbohydrate. Interestingly, the enzymes necessary to cleave the glycoside linkages in cellulose are found in bacteria that exist in the digestive tracts of termites, cows, deer, and many other animals. Therefore, unlike humans, these animals can derive nutrition from the cellulose in wood, hay, and other similar substances.

Glycogen is the main carbohydrate reservoir in animals. For example, it is found in muscles, where it can be broken down into glucose units when energy is required for physical activity.

B. Nucleic Acids

Life is possible only because each cell, when it divides, can transmit the vital information about what it is to the next generation of cells. It has been known for a long time that this process involves the chromosomes in the nucleus of the cell. Only since 1953, however, have scientists understood the molecular basis of this intriguing cellular “talent.”

The substance that stores and transmits the genetic information is a polymer called **deoxyribonucleic acid (DNA)**, a huge molecule with a molar mass as high as several billion grams. Together with other similar nucleic acids called the **ribonucleic acids (RNA)**, DNA carries the information needed for the synthesis of the various proteins the cell requires to carry out its life functions. The RNA molecules, which are found in the cytoplasm outside the cell nucleus, are much smaller than DNA polymers, with molar mass values of only 20,000 to 40,000 grams.

Active Reading Question

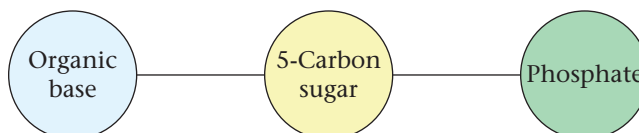
Which are larger, DNA polymers or RNA polymers?

Both DNA and RNA are polymers—they are constructed by the hooking together of many smaller units. The fundamental unit in these polymers is called a **nucleotide**.

A nucleotide has three parts:

1. A nitrogen-containing organic base
2. A five-carbon sugar
3. A phosphate group

A nucleotide can be represented as follows:



information

Note that the prefix *de-* refers to the absence of the substance that follows. Thus deoxyribose means ribose with a missing oxygen (see Figure 21.12).

In the DNA polymer, the five-carbon sugar is deoxyribose (see **Figure 21.12a**); in the RNA polymer, it is ribose (see **Figure 21.12b**). This difference in the sugar molecules present in the polymers is responsible for the names DNA (deoxyribonucleic acid) and RNA (ribonucleic acid).

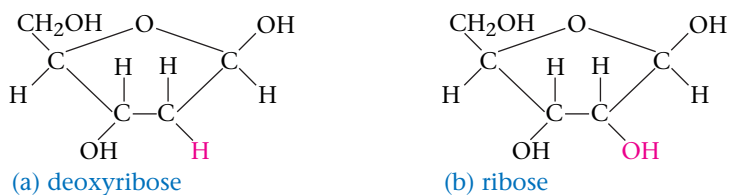


Figure 21.12

The structure of the pentoses (a) deoxyribose and (b) ribose. Deoxyribose is the sugar molecule present in DNA; ribose is found in RNA. The difference in these sugars is the substitution, in ribose, of an OH for one H in deoxyribose.

The organic base molecules in the nucleotides of DNA and RNA are shown in **Figure 21.13**. Notice that some of these bases are found only in DNA, some are found only in RNA, and some are found in both. To form the DNA and RNA polymers, the nucleotides are hooked together. DNA polymers can contain up to a *billion* nucleotide units. **Figure 21.14** shows a small portion of a DNA chain.

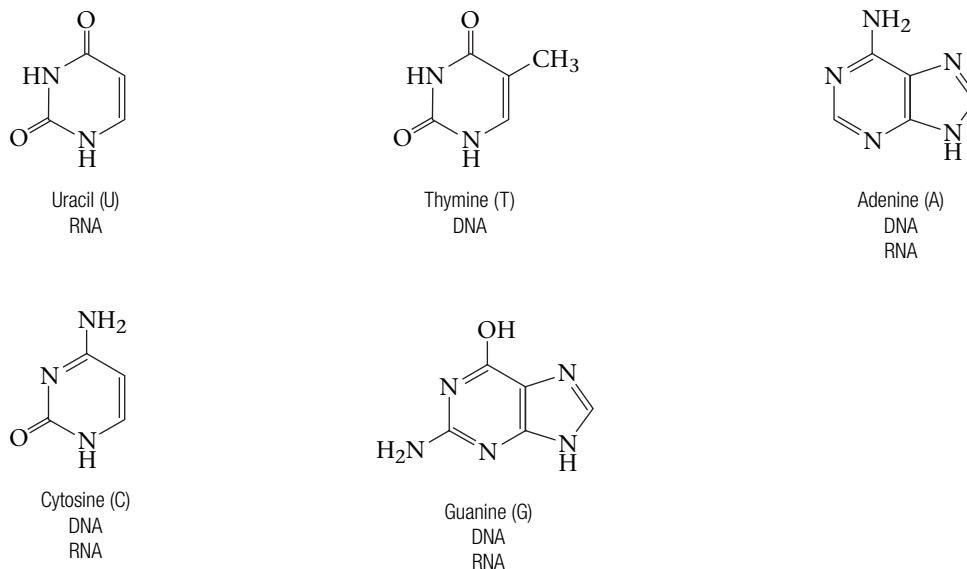


Figure 21.13

The organic bases found in DNA and RNA. Note that uracil is found only in RNA and that thymine is found only in DNA.



A molecular model of part of the DNA structure

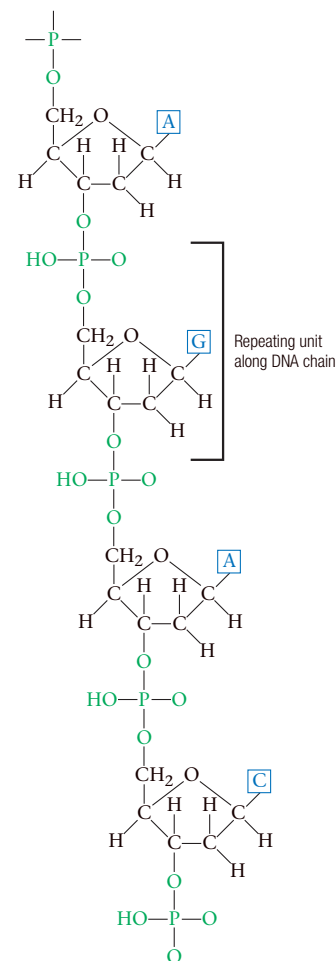


Figure 21.14

A portion of a typical nucleic acid chain, half of the DNA double helix shown in Figure 21.15.

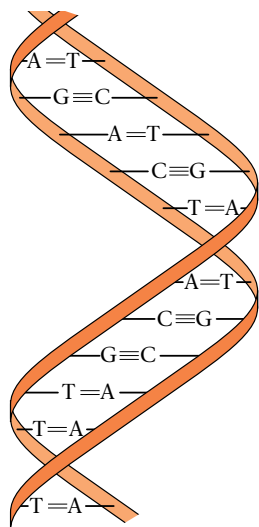


Figure 21.15
The DNA double helix contains two sugar-phosphate backbones with the bases from the two strands hydrogen-bonded to each other.

The key to DNA's functioning is its *double-helical structure with complementary bases on the two strands*. **Figure 21.15**. Complementary bases form hydrogen bonds to each other, as shown in **Figure 21.16**. Note that the structures of cytosine and guanine make them perfect partners (complementary) for hydrogen bonding and that they are *always* found as pairs on the two strands of DNA. Thymine and adenine form similar hydrogen-bonding pairs.

Active Reading Question

Why are thymine and adenine always found as pairs on the two strands of DNA?

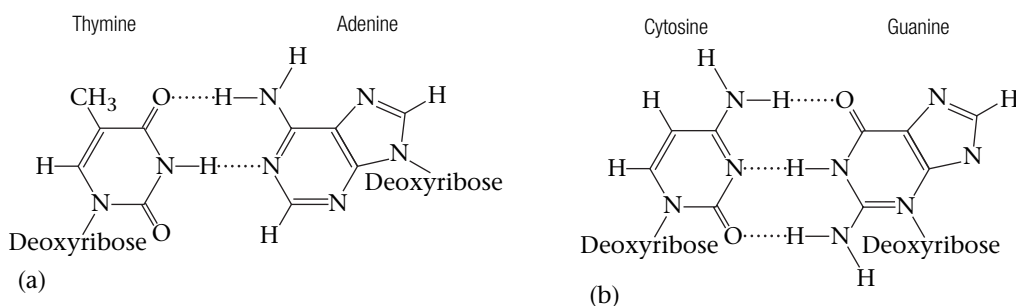


Figure 21.16
The (a) thymine–adenine and (b) cytosine–guanine pairs show complementarity. The hydrogen-bonding interactions are shown by dotted lines.

There is much evidence to suggest that the two strands of DNA unwind during cell division and that new complementary strands are constructed on the unraveled strands (see **Figure 21.17**). Because the bases on the strands always pair in the same way—cytosine with guanine and thymine with adenine—each unraveled strand serves as a template for attaching each complementary base (along with the rest of its nucleotide). This process results in two new double-helix DNA structures that are identical to the original one. Each new double strand contains one strand from the original DNA double helix and one newly synthesized strand. This replication of DNA makes possible the transmission of genetic information when cells divide.

DNA and Protein Synthesis

Besides replication, the other major function of DNA is **protein synthesis**. The proteins consumed by an organism in its food are typically not the specific proteins that organism needs to maintain its existence. The nutrient proteins are broken down into their constituent amino acids, which are then used to construct those proteins that the organism needs. The information for constructing each protein needed by a particular organism is stored in that organism's DNA.

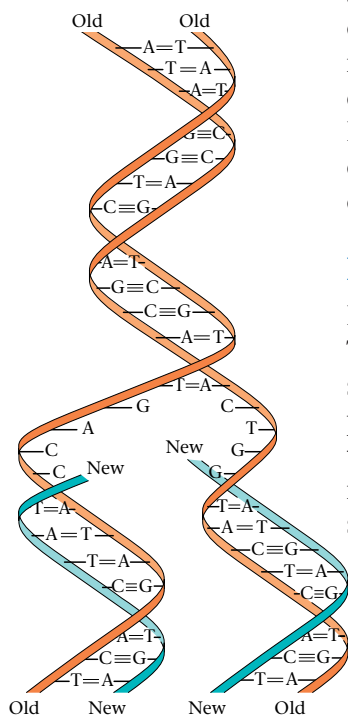


Figure 21.17
During cell division, the original DNA double helix unwinds and new complementary strands are constructed on each original strand. In this way, the two cells resulting from the division have exact copies of the DNA of the original cell.

Table 21.5

Some Common Fatty Acids and Their Major Sources

Name	Formula	Major Source
Saturated		
arachidic acid	$\text{CH}_3(\text{CH}_2)_{18}-\text{COOH}$	peanut oil
butyric acid	$\text{CH}_3(\text{CH}_2)_2-\text{COOH}$	butter
caproic acid	$\text{CH}_3(\text{CH}_2)_4-\text{COOH}$	butter
lauric acid	$\text{CH}_3(\text{CH}_2)_{10}-\text{COOH}$	coconut oil
stearic acid	$\text{CH}_3(\text{CH}_2)_{16}-\text{COOH}$	animal and vegetable fats
Unsaturated		
oleic acid	$\text{CH}_3(\text{CH}_2)_7\text{CH}=\text{CH}(\text{CH}_2)_7-\text{COOH}$	corn oil
linoleic acid	$\text{CH}_3(\text{CH}_2)_4\text{CH}=\text{CH}-\text{CH}_2-\text{CH}=\text{CH}(\text{CH}_2)_7-\text{COOH}$	linseed oil
linolenic acid	$\text{CH}_3\text{CH}_2\text{CH}=\text{CH}-\text{CH}_2\text{CH}=\text{CH}-\text{CH}_2-\text{CH}=\text{CH}-(\text{CH}_2)_7\text{COOH}$	linseed oil

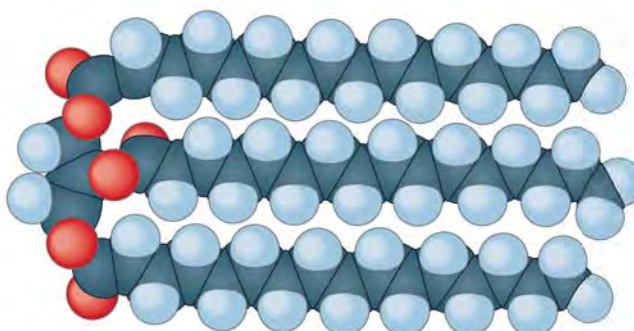
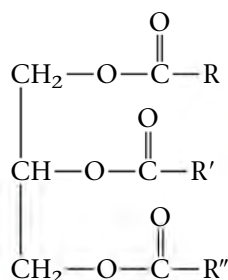
Triglyceride

A fat that is an ester of glycerol

Saponification

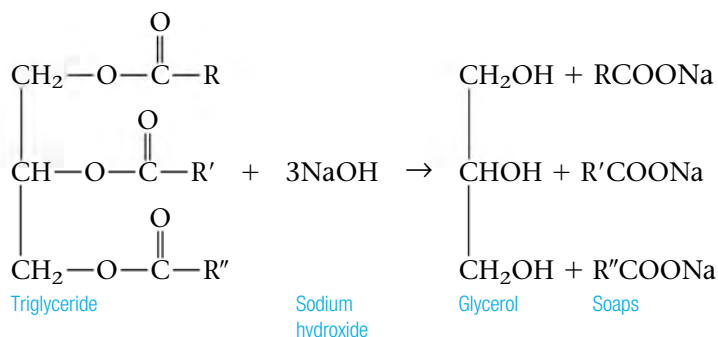
The process of breaking down a triglyceride by treatment with aqueous sodium hydroxide to produce glycerol and the fatty acid salts; the fatty acid salts produced are soaps.

Fats that are esters of glycerol are called **triglycerides** and have the general structure



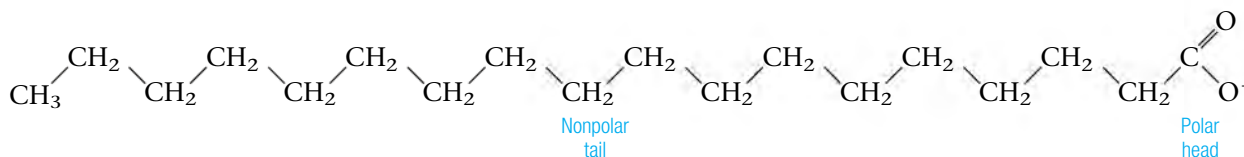
where the three R groups may be the same or different and may be saturated or unsaturated. Vegetable fats tend to be unsaturated and usually occur as oily liquids; most animal fats are saturated (contain only C—C single bonds) and occur as solids at room temperature.

Triglycerides can be broken down by treatment with aqueous sodium hydroxide. The products are glycerol and the fatty acid salts; the latter are known as soaps. This process is called **saponification**.

**Information**

Unsaturated fats contain one or more C=C bonds.

Much of what we call greasy dirt is nonpolar. Grease, for example, consists mostly of long-chain hydrocarbons. However, water, the solvent most commonly available to us, is very polar and does not dissolve “greasy dirt.” We need to add something to the water that is somehow compatible with both the polar water and the nonpolar grease. Fatty-acid anions are perfect for this role, because they have a long nonpolar tail and a polar head. For example, the stearate anion can be represented as



Such ions can be dispersed in water because they form **micelles** (see **Figure 21.19**). These aggregates of fatty-acid anions have the water-incompatible tails in the interior; the anionic parts (the polar heads) point outward and interact with the polar water molecules. A soap solution does not contain *individual* fatty-acid anions dispersed in the water but rather groups of ions (micelles).

Soap dissolves grease by taking the grease molecules into the nonpolar interior of the micelle (see **Figure 21.20**), so they can be carried away by the water. Soap thus acts to suspend the normally incompatible grease in the water. Because of this ability to assist water in suspending nonpolar materials, soap is also called a *wetting agent*, or **surfactant**.

Information

Like dissolves like.

Surfactant

A wetting agent that assists water in suspending nonpolar materials; soap is a surfactant.

Active Reading Question

Why is soap sometimes called a *wetting agent*?

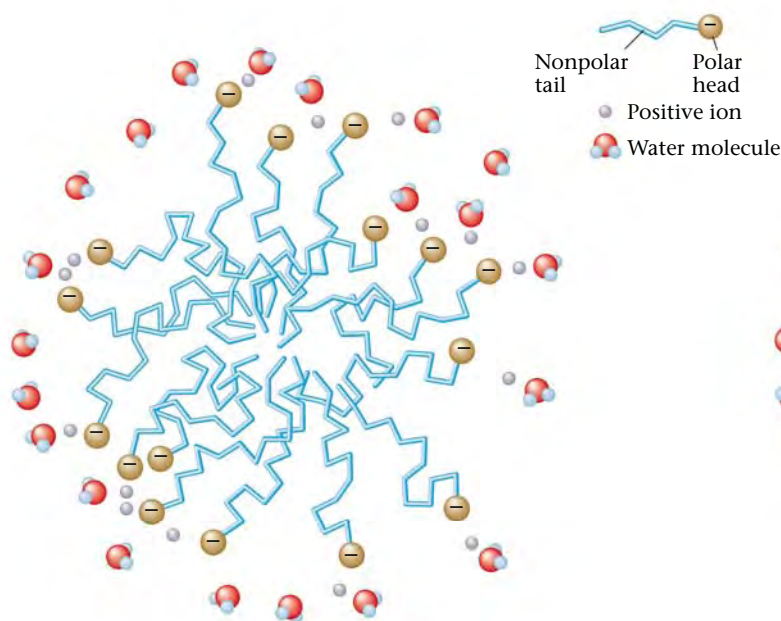


Figure 21.19

A two-dimensional “slice” of the structure of a micelle of fatty-acid anions. Cations surround the negatively charged micelle.

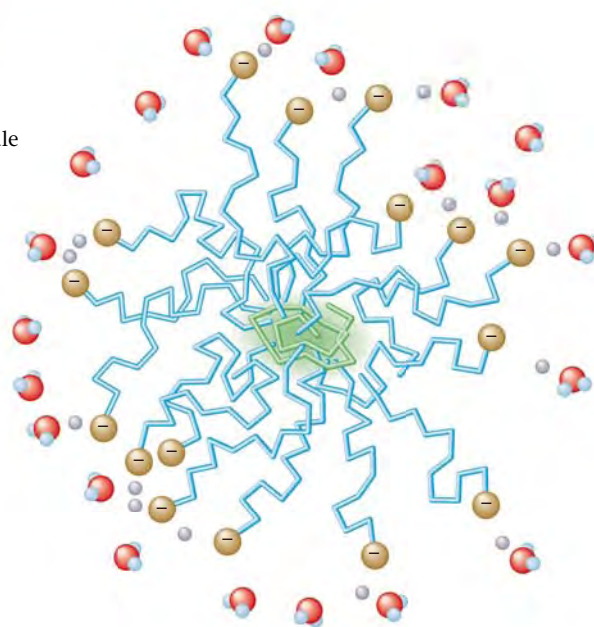


Figure 21.20

Soap micelles absorb grease molecules into their interiors so that the molecules are suspended in the water and can be washed away.

Colorful Milk**Materials**

- shallow colorless container
- whole milk
- food coloring
- cotton swab
- detergent

Procedure

1. Half-fill a shallow colorless container with whole milk.
2. Add 1–2 drops of food coloring to the surface of the milk. Do not stir.
3. Add 1–2 drops of a different food coloring to the surface in a different area. Do not stir.
4. Without moving or shaking the container, observe the food coloring and record what you see.

5. Dip one end of a cotton swab in a liquid dishwashing detergent. Touch the center of the surface of the milk with this end of the swab. Be careful not to stir the milk.
6. Record your observations. If the motion stops, add another drop of the detergent (see step 5).

Results/Analysis

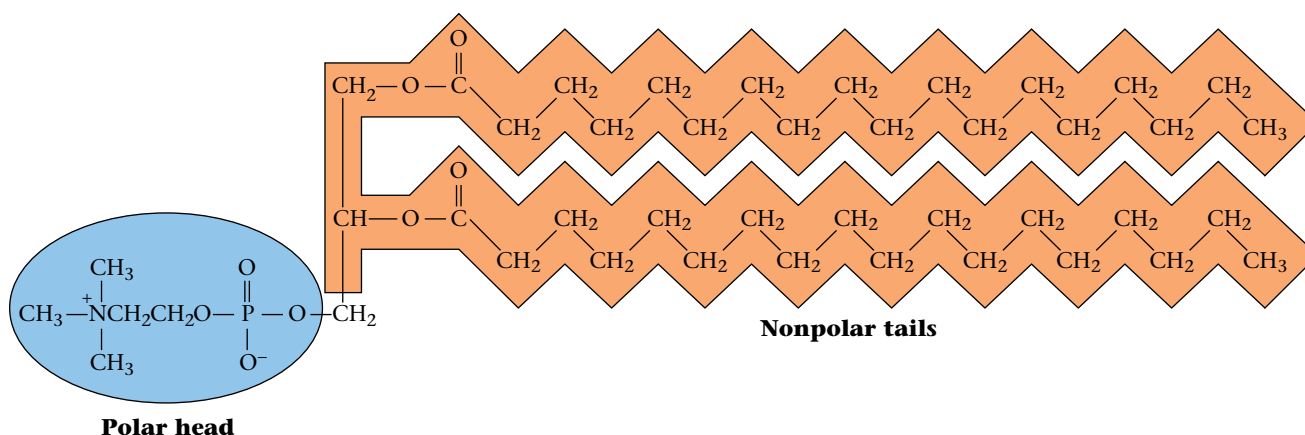
1. Explain your observations from step 4.
2. Compare what the food coloring did in the milk to what happens when you add food coloring to water (see “Chemistry in Action: Mysterious Mixing” on page 39).
3. How did the detergent affect the food coloring in the milk? Explain.

A major disadvantage is that soap anions form precipitates in hard water (water that contains large concentrations of ions such as Ca^{2+} and Mg^{2+}). These precipitates occur because the Ca^{2+} and Mg^{2+} ions form insoluble solids with soap anions. These precipitates (“soap scum”) dull clothes and drastically reduce soap’s cleaning efficiency. To help alleviate this problem, a huge industry has developed to produce artificial soaps, called detergents. Detergents are similar to natural soaps in that they have a long nonpolar tail and an ionic head. However, detergent anions have the advantage of not forming insoluble solids with Ca^{2+} and Mg^{2+} ions.

Phospholipids are similar in structure to fats in that they are esters of glycerol. Unlike fats, however, they contain only two fatty acids. The third group bound to glycerol is a phosphate group, which gives phospholipids two distinct parts: the long nonpolar “tail” and the polar substituted-phosphate “head” (see **Figure 21.21**).

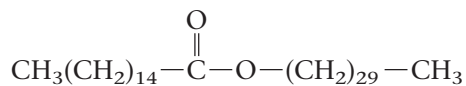
Phospholipid

An ester of glycerol containing two fatty acids; consists of a long nonpolar “tail” and a polar substituted-phosphate “head”

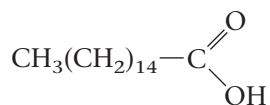
**Figure 21.21**

Lecithin, a phospholipid, with its long nonpolar tails and polar substituted-phosphate head

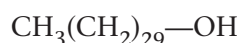
Waxes are another class of lipids. Like fats and phospholipids, waxes are esters, but unlike these other lipids, they involve monohydroxy alcohols instead of glycerol. For example, beeswax, a substance secreted by the wax glands of bees, is mainly myricyl palmitate,



formed from palmitic acid,

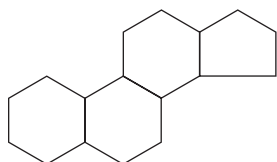


and carnaubyl alcohol



Waxes are solids that furnish waterproof coatings on leaves and fruit and on the skins and feathers of animals. Waxes are also important commercially. For example, whale oil is largely composed of the wax cetyl palmitate. It has been used in so many products, including cosmetics and candles, that the blue whale has been hunted almost to extinction.

Steroids are a class of lipids that have a characteristic carbon ring structure of the type



Steroids comprise four groups: cholesterol, adrenocorticoid hormones, sex hormones, and bile acids.

Whale oil is largely composed of the wax cetyl palmitate.



Beeswax in a honeycomb from Indonesia

Waxes

A class of lipids that involve monohydroxy alcohols instead of glycerol



CHEMISTRY EXPLORERS

Percy Julian • 1899–1975

Percy Julian was born in Birmingham, Alabama, in 1899, the son of a railway mail clerk and a schoolteacher, and the grandson of slaves. He earned a master's degree from Harvard University, where he graduated at the top of his class. After serving as associate professor of chemistry at Howard University for two years, he moved to Austria and received a Ph.D. in organic chemistry in 1931 from the University of Vienna.

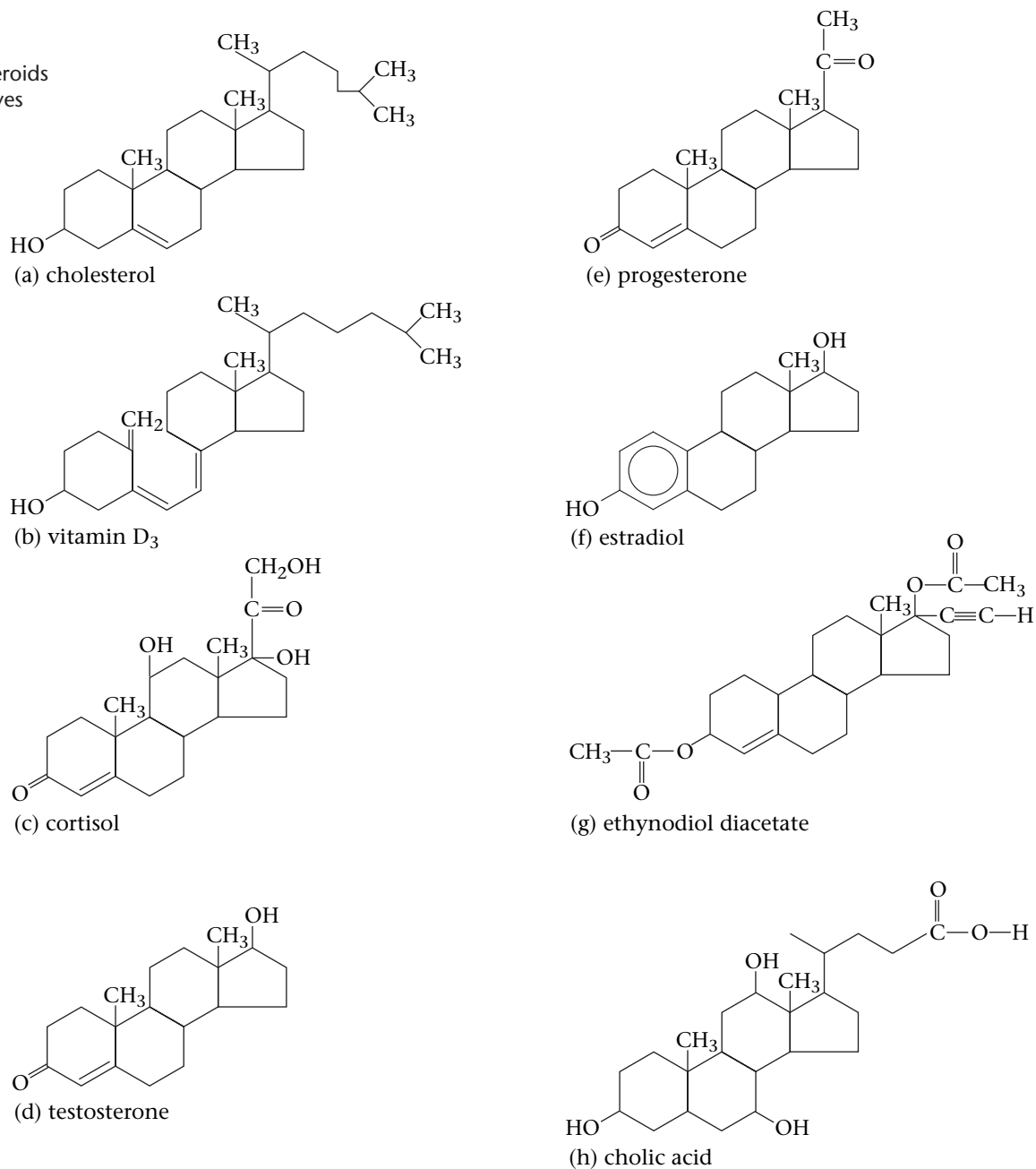
Julian is noted for his work with steroids from natural products. Specifically, he synthesized chemicals used in the treatment of glaucoma and rheumatoid arthritis. In 1953, he founded The Julian Laboratories, Inc.

Percy Julian was awarded over 100 chemistry patents, was elected to the National Academy of Sciences in 1973, and was inducted into the National Inventors Hall of Fame in 1990.



Cholesterol (see **Figure 21.22a**) is found in virtually all organisms and is the starting material for the formation of the many other steroid-based molecules, such as vitamin D (see **Figure 21.22b**). Although cholesterol is essential for human life, it has been implicated in the formation of plaque on the walls of arteries (a process called atherosclerosis, or hardening of the arteries), which can lead eventually to clogging. This effect seems especially important in the arteries that supply blood to the heart. Blockage of these arteries leads to heart damage that often results in death from a heart attack.

Figure 21.22
Several common steroids
and steroid derivatives



The **adrenocorticoid hormones**, such as cortisol (see **Figure 21.22c**), are synthesized in the adrenal glands (glands that lie next to each kidney) and are involved in various regulatory functions.

Of the **sex hormones**, the most important male hormone is *testosterone* (see **Figure 21.22d**), a hormone that controls the growth of the reproductive organs and hair and the development of the muscle structure and deep voice that are characteristic of males. There are two types of female sex hormones of particular significance: *progesterone* (see **Figure 21.22e**), and a group of estrogens, one of which is *estradiol* (see **Figure 21.22f**). Changes in the concentrations of these hormones cause the periodic changes in the ovaries and uterus that are responsible for the menstrual cycle. During pregnancy, a high level of progesterone is maintained, which prevents ovulation.

The **bile acids** are produced from cholesterol in the liver and stored in the gallbladder. The primary human bile acid is *cholic acid* (see **Figure 21.22h**), a substance that aids in the digestion of fats by emulsifying them in the intestine. Bile acids can also dissolve cholesterol ingested in food and are therefore important in controlling cholesterol in the body.



SECTION 21.2

REVIEW QUESTIONS

- 1 How is a protein different from a carbohydrate?
- 2 Where does the name *carbohydrate* come from?
- 3 How are starch and cellulose similar? Why can we not digest cellulose?
- 4 What is a nucleotide? What are the fundamental parts of a nucleotide?
- 5 In the double helix that forms DNA, what are the substances which form the steps in the ladder? How are they held together?
- 6 What are the major functions of DNA?
- 7 What are lipids? List the types of lipids and give a function for each.
- 8 Explain how water cannot dissolve grease, and why soap in water dissolves grease.



RESEARCH LINKS
CLASSZONE.COM

Key Terms

21.1 Biochemistry
 Biotechnology
 Essential elements
 Trace elements
 Cell
 Proteins
 Fibrous proteins
 Globular proteins
 Alpha-amino acids
 Side chains
 Dipeptide
 Peptide linkage
 Polypeptide
 Primary structure
 Secondary structure
 Alpha-helix
 Pleated sheet
 Tertiary structure
 Disulfide linkage
 Denaturation
 Enzymes
 Lock-and-key model
 Substrate
 Active site

Key Ideas

21.1 Introduction to Proteins

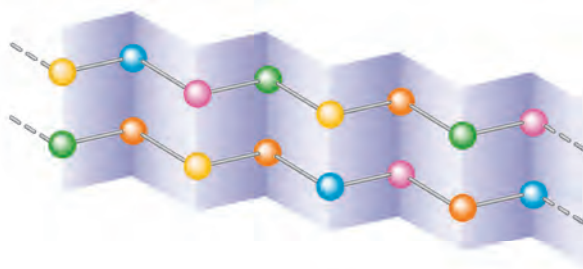
- Proteins are natural polymers made from various α -amino acids.
 - Fibrous proteins are used in the body for structural purposes.
 - Globular proteins are “worker” molecules that act as catalysts, as an O_2 storage system, as regulators of biological function, and so on.
 - Protein structure has several levels
 - ◆ Primary: the order of α -amino acids



- ◆ Secondary: arrangement of the protein in space
 - ★ α -helix



- ★ β -pleated sheet



- ◆ Tertiary: overall shape of the protein



21.2 Carbohydrates

Monosaccharides
(simple sugars)

Disaccharide

Sucrose

Glycoside linkage

Polysaccharides

Starch

Cellulose

Glycogen

DNA (deoxyribonucleic acid)

RNA (ribonucleic acid)

Nucleotide

Protein synthesis

Gene

mRNA (messenger RNA)

tRNA (transfer RNA)

Lipids

Fats

Fatty acids

Saponification

Micelles

Surfactant

Phospholipids

Waxes

Steroids

Cholesterol

Adrenocorticoid hormones

Sex hormones

Bile acids

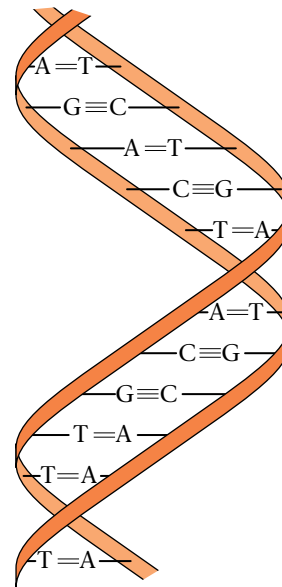
21.2 Carbohydrate, Nucleic Acids, and Lipids

■ Carbohydrates

- Serve as food sources for most organisms and structural materials for plants
- Monosaccharides
 - ◆ Commonly 5 and 6 carbon polyhydroxy ketones and aldehydes
 - ◆ Combine to form more complex carbohydrates such as sucrose (table sugar), starch, and cellulose

■ Nucleic acids

- Deoxyribonucleic acid (DNA)
 - ◆ Has a double helical structure



- ◆ During cell division the double helix unravels to enable a copy of the DNA to be made.
- ◆ Contains segments called genes that store the primary structure of various proteins

■ Lipids

- Water soluble substances found in cells divided into units
- Four classes
 - ◆ Fats
 - ◆ Phospholipids
 - ◆ Waxes
 - ◆ Steroids

All exercises with **blue numbers** have answers in the back of this book.

21.1 Introduction to Proteins

A. Proteins

1. What element is present in the human body in the largest percentage by mass? Name ten other elements essential to life and list their uses in the body.
2. What are proteins? What portion of the human body (by mass) consists of proteins?
3. Describe the range of molar masses shown by proteins in the body. Are such molar masses consistent with proteins being polymers?
4. What uses do *fibrous* and *globular* proteins have in the body? Describe the general shape of these types of proteins.

B. Primary Structure of Proteins

5. Sketch the general formula for an α -amino acid, indicating the α -amino carbon atom. Circle the portion of the structure that all α -amino acids have in common.
6. Write the formulas of two amino acids with polar side chains and of two amino acids with nonpolar side chains. Explain *why* the side chains you have chosen have their respective polarities.
7. Why is it important to distinguish between hydrophilic and hydrophobic side chains? What solvent forms the natural environment for proteins in the body?
8. Write the amino acid sequences possible for a tripeptide containing the amino acids *cys*, *ala*, and *phe*, with each amino acid being used only once in each tripeptide.
9. Given the formulas of the common amino acids in Figure 21.2, sketch the structures of the following simple peptides. Circle the peptide bonds in your structures. Label the terminal amino and carboxyl groups clearly.
 - a. ile-ala-gly
 - b. gln-ser
 - c. ser-gln
 - d. cys-asn-gly

10. Given the formulas of the common amino acids in Figure 21.2, sketch the structures of the six different tripeptides possible containing phenylalanine, alanine, and glycine (each taken only once). Circle the peptide bonds in your structures. Label the terminal amino and carboxyl groups clearly.

C. Secondary Structure of Proteins

11. In general terms, what does the secondary structure of a protein represent?
12. How is the secondary structure of a protein related to its function in the body? Give examples.
13. Describe the secondary protein structure known as the pleated sheet. Give two examples of materials containing proteins with this structure.
14. Describe the secondary protein structure known as the α -helix. What types of proteins have this secondary structure?

D. Tertiary Structure of Proteins

15. In general terms, what does the tertiary structure of a protein describe? Clearly distinguish between the secondary and tertiary structures.
16. Describe the function of the amino acid cysteine in influencing the tertiary structure of a protein.

E. Functions of Proteins

17. What is meant by *denaturation* of a protein? Give three examples of situations in which proteins are denatured.
18. Describe the structure of the protein *collagen*. What function does collagen have in the body?
19. What name is given to proteins that catalyze biochemical reactions in the cell?
20. What are antibodies? Give an example of an important antibody.
21. Give several examples of proteins that serve a protective function in the body.
22. How does a “permanent wave” change the structure of the protein of hair? What amino acid undergoes reaction during a permanent wave? Which level of structure for the hair protein is affected?

F. Enzymes

23. How does the efficiency of an enzyme compare to that of inorganic catalysts? Are enzymes more or less efficient?

24. What general name is given to the molecule acted on by an enzyme? What does it mean to say that an enzyme is *specific* for this molecule?
25. What name is given to the specific portion of the enzyme molecule where catalysis actually occurs?
26. Describe the lock-and-key model for enzymes. Why are the shapes of the enzyme and its substrate important in this model?

21.2 Carbohydrates, Nucleic Acids, and Lipids

A. Carbohydrates

27. What functional groups are present in the simple sugars (monosaccharides)?
28. Sugars can be referred to as polyhydroxy carbonyl compounds. Explain this terminology.
29. In aqueous solution, the monosaccharides generally adopt a ring structure, rather than having straight chains. Sketch representations for the ring forms of glucose and fructose.
30. Sketch the straight-chain structures of the following monosaccharides.
 - a. glucose
 - b. ribose
 - c. ribulose
 - d. galactose

B. Nucleic Acids

31. What molecule stores and transmits genetic information in the cell? How large is this molecule? Where in the cell is this molecule found?
32. Name the five nitrogen bases found in DNA and RNA. Which base is found commonly in RNA but not in DNA? Which base is found commonly in DNA but not in RNA?

33. Describe the double-helical structure of DNA. What type of bonding occurs *within* the chain of each strand of the double helix? What type of bonding exists *between* strands to link them together?
34. Describe the *complementary base pairing* between the two individual strands of DNA that forms the overall double-helical structure. How is complementary base pairing involved in the replication of the DNA molecule during cell division?
35. What name is applied to the specific section of the DNA molecule that contains the code for construction of a particular protein?
36. Explain the distinction between, and the functions of, messenger RNA and transfer RNA.

C. Lipids

37. Under what circumstances is a biomolecule classified as a *lipid*? Is the classification based on a particular characteristic structure or on some physical property?
38. Referring to Table 21.5, list several saturated and several unsaturated fatty acids. What is the general source of most unsaturated fatty acids?
39. What is *saponification*? Write a general equation for the saponification of a triglyceride. What are the ionic products of a saponification reaction more commonly called?
40. What is a *micelle*? How do the micelles formed by soap molecules suspend greasy dirt in a solution?
41. What steroid serves as the starting material in the body for the synthesis of other steroids? What dangers are involved in having too large a concentration of this substance in the body?

Critical Thinking

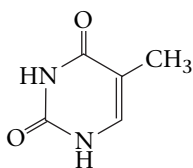
For exercises 42–63 choose one of the following terms to match the description given.

- a. aldohexose
- b. saliva
- c. antibody
- d. cellulose
- e. CH_2O
- f. cysteine
- g. denaturation
- h. disaccharides
- i. disulfide
- j. DNA
- k. enzymes
- l. fibrous
- m. globular
- n. glycogen
- o. glycoside linkage
- p. hormone
- q. hydrophobic
- r. inhibition
- s. ketohexoses
- t. oxytocin
- u. pleated sheet
- v. polypeptide
- w. polysaccharides
- x. primary structure
- y. substrate
- z. sucrose

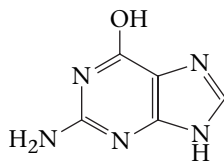
- 42. polymer consisting of many amino acids
- 43. linkage that forms between two cysteine species
- 44. peptide hormone that triggers milk secretion
- 45. proteins with roughly spherical shape
- 46. sequence of amino acids in a protein
- 47. silk protein secondary structure
- 48. water-repelling amino acid side chain
- 49. amino acid responsible for permanent wave in hair
- 50. biological catalysts
- 51. breakdown of a protein's tertiary and/or secondary structure
- 52. molecule acted on by an enzyme
- 53. occurs when an enzyme's active site is blocked by a foreign molecule
- 54. special protein synthesized in response to foreign substance
- 55. substance that has a specific effect on a particular target organ
- 56. animal polymer of glucose
- 57. $-\text{C}-\text{O}-\text{C}-$ bond between rings in disaccharide sugars
- 58. empirical formula leading to the name *carbohydrate*
- 59. where enzymes catalyzing the breakdown of glycoside link are found
- 60. six-carbon ketone sugars
- 61. structural component of plants, polymer of glucose
- 62. sugars consisting of two monomer units
- 63. six-carbon aldehyde sugars

- 1 The order or sequences of amino acids in the protein chain is called the
- primary structure.
 - secondary structure.
 - tertiary structure.
 - polypeptide.
- 2 A pleated sheet is a type of
- the primary structure of a protein.
 - the secondary structure of a protein.
 - the tertiary structure of a protein.
 - disulfide linkage.
- 3 The lock-and-key model applies to
- the structure of carbohydrates.
 - the shape of proteins.
 - the functioning of enzymes.
 - glycoside linkages.
- 4 Table sugar (sucrose) is a disaccharide formed from
- ribose and glucose.
 - glucose and fructose.
 - fructose and ribose.
 - glucose and galactose.
- 5 Which of the following is not a carbohydrate?
- starch
 - cellulose
 - glycogen
 - dipeptide
- 6 The fundamental units in both DNA and RNA are called
- nucleotides.
 - amino acids.
 - lipids.
 - substrates.
- 7 Which if the following is not a steroid or steroid derivative?
- cholesterol
 - cortisol
 - adenosine
 - cholic acid

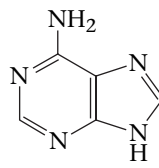
- 8 Write the complementary pairs of bases.



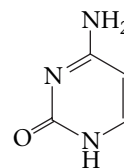
Thymine



Guanine



Adenine



Cytosine

- 9 The phrase *like dissolves like* means that polar solvents dissolve polar solutes, and nonpolar solvents dissolve nonpolar solutes. Soap is dissolved in water (a polar molecule) to dissolve grease (nonpolar).
- Explain how soap is able to dissolve in a polar solvent and dissolve nonpolar molecules.
 - Sketch a diagram showing how a soap solution is able to dissolve grease.
 - Explain what “hard” water is and why a soap solution is less effective in it.

Appendices

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Appendix A



Laboratory Safety Handbook

One of the first things a chemist learns is that working in a laboratory can be an exciting experience. But the laboratory can also be a dangerous place if safety rules are not followed. You must be responsible for your safety and the safety of others.

General Safety Rules

1. Read the directions for the experiment several times. Follow the directions exactly as they are written. If you have questions about any part of the procedure, ask your teacher for assistance.
2. Never perform any activities that are not authorized by your teacher. Do not work in the laboratory without teacher supervision.
3. Use the safety equipment provided for you. Safety goggles and a lab apron should be worn whenever you are working with chemicals or heating any substance.
4. Never eat or drink in the lab. Never inhale chemicals. Do not taste any chemicals.
5. Take care not to spill any material in the lab. If a spill occurs, immediately ask your teacher about the proper cleanup procedure. Do not pour chemicals or other substances into the sink or trash container.
6. Tie back long hair to keep it away from chemicals, burners, and other lab equipment.
7. Remove or tie back articles of loose clothing and jewelry that can hang down and come in contact with chemicals or flames.
8. Know the locations of the safety showers and eyewashes and know how to use them.

First Aid in the Lab

Report any accident or injury, no matter how minor, to your teacher.

Heating and Fire Safety

1. Maintain a clean work area and keep materials away from flames.
2. Never reach across an open flame.
3. Make sure you know how to use a Bunsen burner. Never leave a heat source unattended.
4. When heating a test tube, always point it away from you and others in case the substance bumps out of the tube.
5. Never heat a liquid in a sealed container. The increase in pressure could cause the container to break, injuring you or others.

Injury	Response
Burns	Apply cold water. Call your teacher immediately.
Cuts and Abrasions	Stop any bleeding by applying direct pressure. Call your teacher immediately.
Foreign Matter in Eye	Use eyewash to flush with plenty of water. Call your teacher immediately.
Fainting	Leave the person lying down. Loosen any tight clothing. Call your teacher immediately. Keep people away.
Poisoning	Note the suspected substance. Call your teacher immediately.
Spills on Skin	Flush with large amounts of water or use the safety shower. Call your teacher immediately.

6. Before picking up a container that has been heated, hold the back of your hand close to the container. If you can feel heat on your hand the container is too hot to handle. Use a clamp, tongs, or a hot pad. Remember that hot glass looks just like cold glass.
7. If a fire should break out in the classroom or if your clothing catches fire, smother it with a fire blanket or roll on the ground. NEVER RUN.

Working in the Laboratory Safely

1. Always use heat resistant (Pyrex) glassware for heating.
2. Use materials only from properly labeled containers. Read labels carefully before using chemicals.
3. Do not return chemicals to stock bottles. Discard extra chemicals as your teacher directs.
4. When diluting acids, *always add acid to water*.
5. Do not use the sink to discard matches or any other solid material.
6. Use glycerine or a lubricant when inserting glass tubing or thermometers through rubber stoppers. Protect your hand with a paper towel or a folded cloth.
7. Store backpacks, coats, and other personal items away from the immediate lab bench area.

Cleaning Up the Laboratory

1. After working in the lab, clean up your work area and return all equipment to its proper place.
2. Turn off all water and burners. Check that you have turned off the gas jet as well as the burner.
3. Dispose of chemicals and other materials as directed by your teacher.
4. Wash your hands thoroughly before leaving the lab.

Appendix B

Math Review

In solving chemistry problems you will use, over and over again, relatively few mathematical procedures. In this section we review the few algebraic manipulations that you will need.

Solving an Equation

In the course of solving a chemistry problem, we often construct an algebraic equation that includes the unknown quantity (the thing we want to calculate). An example is

$$(1.5)V = (0.23)(0.08206)(298)$$

We need to “solve this equation for V .” That is, we need to isolate V on one side of the equal sign and all the numbers on the other side. How can we do this? The key idea in solving an algebraic equation is that *doing the same thing on both sides of the equal sign* does not change the equality. That is, it is always “legal” to do the same thing to both sides of the equation. Here we want to solve for V , so we must get the number 1.5 on the other side of the equal sign. We can do this by dividing *both sides* by 1.5.

$$\frac{(1.5)V}{1.5} = \frac{(0.23)(0.08206)(298)}{1.5}$$

Now the 1.5 in the denominator on the left cancels the 1.5 in the numerator:

$$\frac{\cancel{1.5}V}{\cancel{1.5}} = \frac{(0.23)(0.08206)(298)}{1.5}$$

to give

$$V = \frac{(0.23)(0.08206)(298)}{1.5}$$

Using the procedures in “Using Your Calculator” for chain calculations, we can now obtain the value for V with a calculator.

$$V = 3.7$$

Solutions

a. $\frac{PV}{V} = \frac{k}{V}$
 $P = \frac{k}{V}$

c. $\frac{PV}{RT} = \frac{nRT}{RT}$
 $\frac{PV}{RT} = n$

e. $\frac{^{\circ}\text{F} - 32}{^{\circ}\text{C}} \times ^{\circ}\text{C} = \frac{9}{5}^{\circ}\text{C}$
 $\frac{5}{9} (^{\circ}\text{F} - 32) = \frac{5}{9} \times \frac{9}{5}^{\circ}\text{C}$
 $\frac{5}{9} (^{\circ}\text{F} - 32) = ^{\circ}\text{C}$

b. $1.5x + 6 - 6 = 3 - 6$
 $1.5x = -3$
 $\frac{1.5x}{1.5} = \frac{-3}{1.5}$
 $x = -\frac{3}{1.5} = -2$

d. $\frac{P_1V_1}{T_1} \times T_2 = \frac{P_2V_2}{T_2} \times T_2$
 $\frac{P_1V_1T_2}{T_1P_2} = \frac{P_2V_2}{P_2}$
 $\frac{P_1V_1T_2}{T_1P_2} = V_2$

f. $\frac{^{\circ}\text{F} - 32}{^{\circ}\text{C}} \times ^{\circ}\text{C} = \frac{9}{5}^{\circ}\text{C}$
 $^{\circ}\text{F} - 32 + 32 = \frac{9}{5}^{\circ}\text{C} + 32$
 $^{\circ}\text{F} = \frac{9}{5}^{\circ}\text{C} + 32$

Sometimes it is necessary to solve an equation that consists of symbols. For example, consider the equation

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

Let's assume we want to solve for T_2 . That is, we want to isolate T_2 on one side of the equation. There are several possible ways to proceed, keeping in mind that we always do the same thing on both sides of the equal sign. First we multiply both sides by T_2 .

$$T_2 \times \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \times T_2$$

This cancels T_2 on the right. Next we multiply both sides by T_1 .

$$T_2 \times \frac{P_1V_1}{T_1} \times T_1 = P_2V_2T_1$$

This cancels T_1 on the left. Now we divide both sides by P_1V_1 .

$$T_2 \times \frac{P_1V_1}{P_1V_1} = \frac{P_2V_2T_1}{P_1V_1}$$

This yields the desired equation,

$$T_2 = \frac{P_2V_2T_1}{P_1V_1}$$

For practice, solve each of the following equations for the variable indicated.

- $PV = k$; solve for P
- $1.5x + 6 = 3$; solve for x
- $PV = nRT$; solve for n
- $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$; solve for V_2
- $\frac{^{\circ}\text{F} - 32}{^{\circ}\text{C}} = \frac{9}{5}$; solve for $^{\circ}\text{C}$
- $\frac{^{\circ}\text{F} - 32}{^{\circ}\text{C}} = \frac{9}{5}$; solve for $^{\circ}\text{F}$

Appendix C

Scientific Notation

The numbers we must work with in scientific measurements are often very large or very small; thus it is convenient to express them using powers of 10. For example, the number 1,300,000 can be expressed as 1.3×10^6 , which means multiply 1.3 by 10 six times, or

$$1.3 \times 10^6 = 1.3 \times \underbrace{10 \times 10 \times 10 \times 10 \times 10 \times 10}_{10^6 = 1 \text{ million}}$$

A number written in scientific notation always has the form:

A number (between 1 and 10) times
the appropriate power of 10

To represent a large number such as 20,500 in scientific notation, we must move the decimal point in such a way as to achieve a number between 1 and 10 and then multiply the result by a power of 10 to compensate for moving the decimal point. In this case, we must move the decimal point four places to the left.

$$\begin{array}{cccc} 2 & 0 & 5 & 0 & 0 \\ \hline & 4 & 3 & 2 & 1 \end{array}$$

to give a number between 1 and 10:

2.05

where we retain only the significant figures (the number 20,500 has three significant figures). To compensate for moving the decimal point four places to the left, we must multiply by 10^4 . Thus

$$20,500 = 2.05 \times 10^4$$

As another example, the number 1985 can be expressed as 1.985×10^3 . To end up with the number 1.985, which is between 1 and 10, we had to move the decimal point three places to the left. To compensate for that, we must multiply by 10^3 . Some other examples are given below.

Number	Exponential Notation
5.6	5.6×10^0 or 5.6×1
39	3.9×10^1
943	9.43×10^2
1126	1.126×10^3

So far, we have considered numbers greater than 1. How do we represent a number such as 0.0034 in exponential notation? First, to achieve a number between 1 and 10, we start with 0.0034 and move the decimal point three places to the right.

$$\begin{array}{cccc} 0 & . & 0 & 0 & 3 & 4 \\ \hline & & 1 & 2 & 3 & \end{array}$$

This yields 3.4. Then, to compensate for moving the decimal point to the right, we must multiply by a power of 10 with a negative exponent, in this case, 10^{-3} . Thus

$$0.0034 = 3.4 \times 10^{-3}$$

In a similar way, the number 0.00000014 can be written as 1.4×10^{-7} , because going from 0.00000014 to 1.4 requires that we move the decimal point seven places to the right.

Mathematical Operations with Exponents

We next consider how various mathematical operations are performed using exponents. First we cover the various rules for these operations; then we consider how to perform them on your calculator.

Multiplication and Division

When two numbers expressed in exponential notation are multiplied, the initial numbers are multiplied and the exponents of 10 are *added*.

$$(M \times 10^m)(N \times 10^n) = (MN) \times 10^{m+n}$$

For example (to two significant figures, as required),

$$(3.2 \times 10^4)(2.8 \times 10^3) = 9.0 \times 10^7$$

When the numbers are multiplied, if a result greater than 10 is obtained for the initial number, the decimal point is moved one place to the left and the exponent of 10 is increased by 1.

$$\begin{aligned} (5.8 \times 10^2)(4.3 \times 10^8) &= 24.9 \times 10^{10} \\ &= 2.49 \times 10^{11} \\ &= 2.5 \times 10^{11} \text{ (two significant figures)} \end{aligned}$$

Division of two numbers expressed in exponential notation involves normal division of the initial numbers and subtraction of the exponent of the divisor from that of the dividend. For example,

$$\frac{4.8 \times 10^8}{2.1 \times 10^3} = \frac{4.8}{2.1} \times 10^{(8-3)} = 2.3 \times 10^5$$

Divisor

If the initial number resulting from the division is less than 1, the decimal point is moved one place to the right and the exponent of 10 is decreased by 1. For example,

$$\begin{aligned} \frac{6.4 \times 10^3}{8.3 \times 10^5} &= \frac{6.4}{8.3} \times 10^{(3-5)} = 0.77 \times 10^{-2} \\ &= 7.7 \times 10^{-3} \end{aligned}$$

Addition and Subtraction

In order for us to add or subtract numbers expressed in exponential notation, *the exponents of the numbers must be the same*. For example, to add 1.31×10^5 and 4.2×10^4 , we must rewrite one number so that the exponents of both are the same. The number 1.31×10^5 can be written 13.1×10^4 because decreasing the exponent by 1 compensates for moving the decimal point one place to the right. Now we can add the numbers.

$$\begin{array}{r} 13.1 \times 10^4 \\ + 4.2 \times 10^4 \\ \hline 17.3 \times 10^4 \end{array}$$

In correct exponential notation, the result is expressed as

$$1.73 \times 10^5.$$

To perform addition or subtraction with numbers expressed in exponential notation, we add or subtract only the initial numbers. The exponent of the result is the same as the exponents of the numbers being added or

subtracted. To subtract 1.8×10^2 from 8.99×10^3 , we first convert 1.8×10^2 to 0.18×10^3 so that both numbers have the same exponent. Then we subtract.

$$\begin{array}{r} 8.99 \times 10^3 \\ - 0.18 \times 10^3 \\ \hline 8.81 \times 10^3 \end{array}$$

Powers and Roots

When a number expressed in exponential notation is taken to some power, the initial number is taken to the appropriate power and the exponent of 10 is *multiplied* by that power.

$$(N \times 10^n)^m = N^m \times 10^{m \times n}$$

For example,

$$\begin{aligned} (7.5 \times 10^2)^2 &= (7.5)^2 \times 10^{2 \times 2} \\ &= 56. \times 10^4 \\ &= 5.6 \times 10^5 \end{aligned}$$

When a root is taken of a number expressed in exponential notation, the root of the initial number is taken and the exponent of 10 is divided by the number representing the root. For example, we take the square root of a number as follows:

$$\sqrt{N \times 10^n} = (N \times 10^n)^{1/2} = \sqrt{N} \times 10^{n/2}$$

For example,

$$\begin{aligned} (2.9 \times 10^6)^{1/2} &= \sqrt{2.9} \times 10^{6/2} \\ &= 1.7 \times 10^3 \end{aligned}$$

Using a Calculator to Perform Mathematical Operations on Exponents

In dealing with exponents, you must first learn to enter them into your calculator. First the number is keyed in and then the exponent. There is a special key that must be pressed just before the exponent is entered. This key is often labeled $\boxed{\text{EE}}$ or $\boxed{\text{exp}}$. For example, the number 1.56×10^6 is entered as follows:

Press	Display
1.56	1.56
EE or exp	1.56 00
6	1.56 06

To enter a number with a negative exponent, use the change-of-sign key $\boxed{+/-}$ after entering the exponent number. For example, the number 7.54×10^{-3} is entered as follows:

Press	Display
7.54	7.54
EE or exp	7.54 00
3	7.54 03
+/-	7.54 -03

Once a number with an exponent is entered into your calculator, the mathematical operations are performed exactly the same as with a "regular" number. For example, the numbers 1.0×10^3 and 1.0×10^2 are multiplied as follows:

Press	Display
1.0	1.0
EE or exp	1.0 00
3	1.0 03
\times	1 03
1.0	1.0
EE or exp	1.0 00
2	1.0 02
=	1 05

The answer is correctly represented as 1.0×10^5 .

The numbers 1.50×10^5 and 1.1×10^4 are added as follows:

Press	Display
1.5	1.50
EE or exp	1.50 00
5	1.50 05
+	1.5 05
1.1	1.1
EE or exp	1.1 00
4	1.1 04
=	1.61 05

The answer is correctly represented as 1.61×10^5 . Note that when exponential numbers are added, the calculator automatically takes into account any difference in exponents.

To take the power, root, or reciprocal of an exponential number, enter the number first, then press the appropriate key or keys. For example, the square root of 5.6×10^3 is obtained as follows:

Press	Display
5.6	5.6
EE or exp	5.6 00
3	5.6 03
$\sqrt{\text{X}}$	7.4833148 01

The answer is correctly represented as 7.5×10^1 .

Practice by performing the following operations that involve exponential numbers. The answers follow the exercises.

- $7.9 \times 10^2 \times 4.3 \times 10^4$
- $\frac{5.4 \times 10^3}{4.6 \times 10^5}$
- $1.7 \times 10^2 + 1.63 \times 10^3$
- $4.3 \times 10^{-3} + 1 \times 10^{-4}$
- $(8.6 \times 10^{-6})^2$
- $\frac{1}{8.3 \times 10^2}$
- $\log(1.0 \times 10^{-7})$
- $-\log(1.3 \times 10^{-5})$
- $\sqrt{6.7 \times 10^9}$

Solutions

- 3.4×10^7
- 1.2×10^{-2}
- 1.80×10^3
- 4.4×10^{-3}
- 7.4×10^{-11}
- 1.2×10^{-3}
- 7.00
- 4.89
- 8.2×10^4

Appendix D

Graphing

In interpreting the results of a scientific experiment, it is often useful to make a graph. If possible, the function to be graphed should be in a form that gives a straight line. The equation for a straight line (a *linear equation*) can be represented in the general form

$$y = mx + b$$

where y is the *dependent variable*, x is the *independent variable*, m is the *slope*, and b is the *intercept* with the y -axis.

To illustrate the characteristics of a linear equation, $y = 3x + 4$ is plotted in Figure A.1. For this equation $m = 3$ and $b = 4$. Note that the y -intercept occurs when $x = 0$. In this case the y intercept is 4, as can be seen from the equation ($b = 4$).

The slope of a straight line is defined as the ratio of the rate of change in y to that in x :

$$m = \text{slope} = \frac{\Delta y}{\Delta x}$$

For the equation $y = 3x + 4$, y changes three times as fast as x (because x has a coefficient of 3). Thus the slope in this case is 3. This can be verified from the graph. For the triangle shown in Figure A.1,

$$\Delta y = 50 - 14 = 36 \text{ and } \Delta x = 15 - 3 = 12$$

Thus

$$\text{Slope} = \frac{\Delta y}{\Delta x} = \frac{36}{12} = 3$$

This example illustrates a general method for obtaining the slope of a line from the graph of that line. Simply draw a triangle with one side parallel to the y axis and the other side parallel to the x axis, as shown in Figure A.1. Then determine the lengths of the sides to get Δy and Δx , respectively, and compute the ratio $\Delta y/\Delta x$.

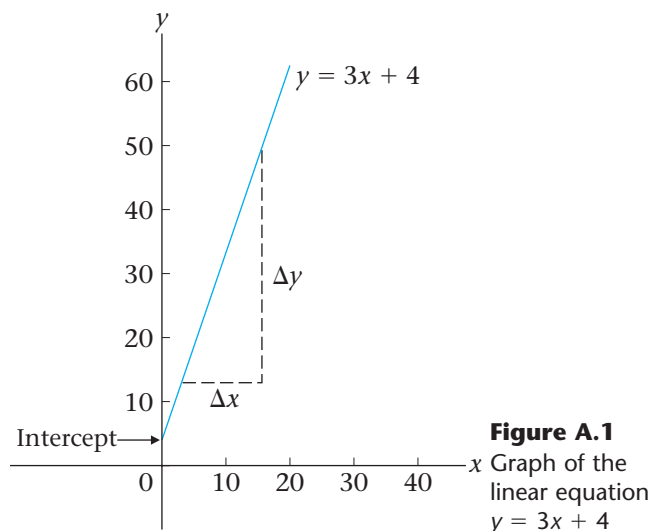


Figure A.1

Graph of the linear equation $y = 3x + 4$

Appendix E

Naming Common Ions

Common Simple Cations and Anions

Cation	Name	Anion	Name*
H ⁺	hydrogen	H ⁻	hydride
Li ⁺	lithium	F ⁻	fluoride
Na ⁺	sodium	Cl ⁻	chloride
K ⁺	potassium	Br ⁻	bromide
Cs ⁺	cesium	I ⁻	iodide
Be ²⁺	beryllium	O ²⁻	oxide
Mg ²⁺	magnesium	S ²⁻	sulfide
Ca ²⁺	calcium		
Ba ²⁺	barium		
Al ³⁺	aluminum		
Ag ⁺	silver		

*The root is given in color.

Common Type II Cations

Ion	Systematic Name	Older Name
Fe ³⁺	iron(III)	ferric
Fe ²⁺	iron(II)	ferrous
Cu ²⁺	copper(II)	cupric
Cu ⁺	copper(I)	cuprous
Co ³⁺	cobalt(III)	cobaltic
Co ²⁺	cobalt(II)	cobaltous
Sn ⁴⁺	tin(IV)	stannic
Sn ²⁺	tin(II)	stannous
Pb ⁴⁺	lead(IV)	plumbic
Pb ²⁺	lead(II)	plumbous
Hg ²⁺	mercury(II)	mercuric
Hg ₂ ²⁺	mercury(I)	mercurous

*Mercury(I) ions always occur bound together in pairs to form Hg₂²⁺.

Names of Common Polyatomic Ions

Ion	Name	Ion	Name
NH_4^+	ammonium	CO_3^{2-}	carbonate
NO_2^-	nitrite	HCO_3^-	hydrogen carbonate (bicarbonate is a widely used common name)
NO_3^-	nitrate	ClO^-	hypochlorite
SO_3^{2-}	sulfite	ClO_2^-	chlorite
SO_4^{2-}	sulfate	ClO_3^-	chlorate
HSO_4^-	hydrogen sulfate (bisulfate is a widely used common name)	ClO_4^-	perchlorate
OH^-	hydroxide	$\text{C}_2\text{H}_3\text{O}_2^-$	acetate
CN^-	cyanide	MnO_4^-	permanganate
PO_4^{3-}	phosphate	$\text{Cr}_2\text{O}_7^{2-}$	dichromate
HPO_4^{2-}	hydrogen phosphate	CrO_4^{2-}	chromate
H_2PO_4^-	dihydrogen phosphate	O_2^{2-}	peroxide

Appendix F

SI Units and Conversion Factors

These conversion factors are given with more significant figures than those typically used in the body of the text.

Length

SI Unit: Meter (m)

1 meter	=	1.0936 yards
1 centimeter	=	0.39370 inch
1 inch	=	2.54 centimeters (exactly)
1 kilometer	=	0.62137 mile
1 mile	=	5280. feet
	=	1.6093 kilometers

Volume

SI Unit: Cubic Meter (m³)

1 liter	=	10 ⁻³ m ³
	=	1 dm ³
	=	1.0567 quarts
1 gallon	=	4 quarts
	=	8 pints
	=	3.7854 liters
1 quart	=	32 fluid ounces
	=	0.94635 liter

Mass

SI Unit: Kilogram (kg)

1 kilogram	=	1000 grams
	=	2.2046 pounds
1 pound	=	453.59 grams
	=	0.45359 kilogram
	=	16 ounces
1 atomic mass unit	=	1.66057 × 10 ⁻²⁷ kilograms

Pressure

SI Unit: Pascal (Pa)

1 atmosphere	=	101.325 kilopascals
	=	760. torr (mm Hg)
	=	14.70 pounds per square inch

Energy

SI Unit: Joule (J)

1 joule	=	0.23901 calorie
1 calorie	=	4.184 joules

Appendix G

Solubility Rules

General Rules for Solubility of Ionic Compounds (Salts) in Water at 25 °C

1. Most nitrate (NO_3^-) salts are soluble.
2. Most salts of Na^+ , K^+ , and NH_4^+ are soluble.
3. Most chloride salts are soluble. Notable exceptions are AgCl , PbCl_2 , and Hg_2Cl_2 .
4. Most sulfate salts are soluble. Notable exceptions are BaSO_4 , PbSO_4 , and CaSO_4 .
5. Most hydroxide compounds are only slightly soluble.* The important exceptions are NaOH and KOH . $\text{Ba}(\text{OH})_2$ and $\text{Ca}(\text{OH})_2$ are only moderately soluble.
6. Most sulfide (S^{2-}), carbonate (CO_3^{2-}), and phosphate (PO_4^{3-}) salts are only slightly soluble.*

*The terms *insoluble* and *slightly soluble* really mean the same thing: such a tiny amount dissolves that it is not possible to detect it with the naked eye.

Practice Problem Exercises

Chapter 2

Practice Problem Exercise 2.1

Items (a) and (c) are physical properties. When the solid gallium melts, it forms liquid gallium. There is no change in composition. Items (b) and (d) reflect the ability to change composition and are thus chemical properties. Statement (b) means that platinum does not react with oxygen to form some new substance.

Practice Problem Exercise 2.2

- Milk turns sour because new substances are formed. This is a chemical change.
- Melting the wax is a physical change (a change of state). When the wax burns, new substances are formed. This is a chemical change.

Practice Problem Exercise 2.3

- Maple syrup is a homogeneous mixture of sugar and other dissolved substances dispersed uniformly in water.
- Helium and oxygen form a homogeneous mixture.
- Oil and vinegar salad dressing is a heterogeneous mixture. (Note the two distinct layers the next time you look at a bottle of dressing.)
- Common salt is a pure substance (sodium chloride), so it always has the same composition. (Note that other substances such as iodine are often added to commercial preparations of table salt, which is mostly sodium chloride. Thus commercial table salt is a homogeneous mixture.)

Chapter 3

Practice Problem Exercise 3.1

- P_4O_{10}
- UF_6
- $AlCl_3$

Practice Problem Exercise 3.2

- In the symbol ${}^{90}_{38}\text{Sr}$, the number 38 is the atomic number, which represents the number of protons in the nucleus of a strontium atom. Because the atom is neutral overall, it must also have 38 electrons. The number 90 (the mass number) represents the number of protons plus the number of neutrons. Thus the number of neutrons is $A - Z = 90 - 38 = 52$.
- The atom ${}^{201}_{80}\text{Hg}$ has 80 protons, 80 electrons, and $201 - 80 = 121$ neutrons.

Practice Problem Exercise 3.3

The atomic number for phosphorus is 15 and the mass number is $15 + 17 = 32$. Thus the symbol for the atom is ${}^{32}_{15}\text{P}$.

Practice Problem Exercise 3.5

Element	Symbol	Atomic Number	Metal or Nonmetal	Family Name
a. argon	Ar	18	nonmetal	noble gas
b. chlorine	Cl	17	nonmetal	halogen
c. barium	Ba	56	metal	alkaline earth metals
d. cesium	Cs	55	metal	alkali metals

Practice Problem Exercise 3.6

- KI
 $(1+) + (1-) = 0$
- Mg_3N_2
 $3(2+) + 2(3-) = (6+) + (6-) = 0$
- Al_2O_3
 $2(3+) + 3(2-) = 0$

Chapter 4

Practice Problem Exercise 4.1

- rubidium oxide
- potassium sulfide
- strontium iodide

Practice Problem Exercise 4.3

- The compound PbBr_2 must contain Pb^{2+} , named lead(II), to balance the charges of the two Br^- ions. Thus the name is lead(II) bromide. The compound PbBr_4 must contain Pb^{4+} , named lead(IV), to balance the charges of the four Br^- ions. The name is therefore lead(IV) bromide.
- The compound FeS contains the S^{2-} ion (sulfide) and thus the iron cation present must be Fe^{2+} , iron(II). The name is iron(II) sulfide. The compound Fe_2S_3 contains three S^{2-} ions and two iron cations of unknown charge. We can determine the iron charge from the following:

$$2(?) + 3(2-) = 0$$

$$\begin{array}{ccc} \uparrow & & \uparrow \\ \text{Iron} & & \text{S}^{2-} \\ \text{charge} & & \text{charge} \end{array}$$

In this case, ? must represent 3 because

$$2(3+) + 3(2-) = 0$$

Thus Fe_2S_3 contains Fe^{3+} and S^{2-} and its name is iron(III) sulfide.

- The compound AlBr_3 contains Al^{3+} and Br^- . Because aluminum forms only one ion (Al^{3+}), no Roman numeral is required. The name is aluminum bromide.

- d. The compound Na_2S contains Na^+ and S^{2-} ions. The name is sodium sulfide. (Because sodium forms only Na^+ , no Roman numeral is needed.)
- e. The compound CoCl_3 contains three Cl^- ions. Thus the cobalt cation must be Co^{3+} , which is named cobalt(III) because cobalt is a transition metal and can form more than one type of cation. Thus the name of CoCl_3 is cobalt(III) chloride.

Practice Problem Exercise 4.4

Compound	Individual Names	Prefixes	Name
a. CCl_4	carbon chloride	none	carbon tetrachloride
b. NO_2	nitrogen oxide	none	nitrogen dioxide
c. IF_5	iodine fluoride	none	iodine pentafluoride

Practice Problem Exercise 4.5

- a. silicon dioxide
- b. dioxygen difluoride
- c. xenon hexafluoride

Practice Problem Exercise 4.6

- a. chlorine trifluoride
- b. vanadium(V) fluoride
- c. copper(I) chloride
- d. manganese(IV) oxide
- e. magnesium oxide
- f. water

Practice Problem Exercise 4.7

- a. calcium hydroxide
- b. sodium phosphate
- c. potassium permanganate
- d. ammonium dichromate
- e. cobalt(II) perchlorate (Perchlorate has a $1-$ charge, so the cation must be Co^{2+} to balance the two ClO_4^- ions.)
- f. potassium chlorate
- g. copper(II) nitrite [This compound contains two NO_2^- (nitrite) ions and thus must contain a Cu^{2+} cation.]

Practice Problem Exercise 4.8

Compound	Name
a. NaHCO_3	sodium hydrogen carbonate Contains Na^+ and HCO_3^- ; often called sodium (common name) bicarbonate
b. BaSO_4	barium sulfate Contains Ba^{2+} and SO_4^{2-}
c. CsClO_4	cesium perchlorate Contains Cs^+ and ClO_4^-
d. BrF_5	bromine pentafluoride Both nonmetals (Type III binary)
e. NaBr	sodium bromide Contains Na^+ and Br^- (Type I binary)
f. KOCI	potassium hypochlorite Contains K^+ and OCI^-
g. $\text{Zn}_3(\text{PO}_4)_2$	zinc(II) phosphate

Contains Zn^{2+} and PO_4^{3-} ; Zn is a transition metal and may require a Roman numeral. However, because Zn forms only the Zn^{2+} cation, the II is left out. Thus the name of the compound is given as zinc phosphate.

Practice Problem Exercise 4.9

- | Name | Chemical Formula |
|-------------------------|---|
| a. ammonium sulfate | $(\text{NH}_4)_2\text{SO}_4$
Two ammonium ions (NH_4^+) are required for each sulfate ion (SO_4^{2-}) to achieve charge balance. |
| b. vanadium(V) fluoride | VF_5
The compound contains V^{5+} ions and requires five F^- ions for charge balance. |
| c. disulfur dichloride | S_2Cl_2
The prefix <i>di-</i> indicates two of each atom. |
| d. rubidium peroxide | Rb_2O_2
Because rubidium is in Group 1, it forms only 1^+ ions. Thus two Rb^+ ions are needed to balance the $2-$ charge on the peroxide ion (O_2^{2-}). |
| e. aluminum oxide | Al_2O_3
Aluminum forms only 3^+ ions. Two Al^{3+} ions are required to balance the charge on three O^{2-} ions. |

Chapter 5

Practice Problem Exercise 5.2

$$357 = 3.57 \times 10^2$$

$$0.0055 = 5.5 \times 10^{-3}$$

Practice Problem Exercise 5.3

- a. Three significant figures. The leading zeros (to the left of the 1) do not count, but the trailing zeros do.
- b. Five significant figures. The one captive zero and the two trailing zeros all count.
- c. This is an exact number obtained by counting the cars. It has an unlimited number of significant figures.

Practice Problem Exercise 5.5

- a. $12.6 \times 0.53 = 6.678 = 6.7$
Limiting
- b. $12.6 \times 0.53 = 6.7$; $\frac{6.7}{-4.59} = 2.11 = 2.1$
Limiting
- c. 25.36
 $\frac{-4.15}{21.21} = \frac{21.21}{2.317} = 9.15408 = 9.154$

Practice Problem Exercise 5.6

$$0.750 \cancel{\text{L}} \times \frac{1.06 \text{ qt}}{1 \cancel{\text{L}}} = 0.795 \text{ qt}$$

Practice Problem Exercise 5.7

$$225 \frac{\text{mi}}{\text{h}} \times \frac{1760 \text{ yd}}{\text{mi}} \times \frac{1 \cancel{\text{m}}}{1.094 \text{ yd}} \times \frac{1 \text{ km}}{1000 \cancel{\text{m}}} = 362 \frac{\text{km}}{\text{h}}$$

Practice Problem Exercise 5.9

The best way to solve this problem is to convert 172 K to Celsius degrees. To do this we will use the formula $T_{\text{C}} = T_{\text{K}} - 273$.

In this case

$$T_{\text{C}} = T_{\text{K}} - 273 = 172 - 273 = -101$$

So 172 K = -101 °C, which is a lower temperature than -75 °C. Thus 172 K is colder than -75 °C.

Practice Problem Exercise 5.11

The problem is 41 °C = ? °F.

Using the formula

$$T_{\text{F}} = 1.80(T_{\text{C}}) + 32$$

we have

$$T_{\text{F}} = ? \text{ °F} = 1.80(41) + 32 = 74 + 32 = 106$$

That is, 41 °C = 106 °F.

Practice Problem Exercise 5.12

This problem can be stated as 239 °F = ? °C.

Using the formula

$$T_{\text{C}} = \frac{T_{\text{F}} - 32}{1.80}$$

we have

$$T_{\text{C}} = ? \text{ °C} = \frac{239 - 32}{1.80} = \frac{207}{1.80} = 115$$

That is, 239 °F = 115 °C.

Practice Problem Exercise 5.14

We obtain the density of the cleaner by dividing its mass by its volume.

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{2.81 \text{ g}}{35.8 \text{ mL}} = 0.785 \text{ g/mL}$$

This density identifies the liquid as isopropyl alcohol.

Chapter 6

Practice Problem Exercise 6.1

The average mass of nitrogen is 14.01 amu. The appropriate equivalence statement is 1 N atom = 14.01 amu, which yields the conversion factor we need:

$$23 \text{ N atoms} \times \frac{14.0 \text{ amu}}{1 \text{ N atom}} = 322.2 \text{ amu.}$$

(exact)

Practice Problem Exercise 6.2

The average mass of oxygen is 16.00 amu, which gives the equivalence statement 1 O atom = 16.00 amu. The number of oxygen atoms present is

$$288 \text{ amu} \times \frac{1 \text{ O atom}}{16.00 \text{ amu}} = 18.0 \text{ O atoms}$$

Practice Problem Exercise 6.4

Note that the sample of 5.00×10^{20} atoms of chromium is less than 1 mol (6.022×10^{23} atoms) of chromium. The fraction of a mole that it represents can be determined as follows:

$$5.00 \times 10^{20} \text{ atoms Cr} \times \frac{1 \text{ mol Cr}}{6.022 \times 10^{23} \text{ atoms Cr}} = 8.30 \times 10^{-4} \text{ mol Cr}$$

Because the mass of 1 mol of chromium atoms is 52.00 g, the mass of 5.00×10^{20} atoms can be determined as follows:

$$8.30 \times 10^{-4} \text{ mol Cr} \times \frac{52.00 \text{ g Cr}}{1 \text{ mol Cr}} = 4.32 \times 10^{-2} \text{ g Cr}$$

Practice Problem Exercise 6.5

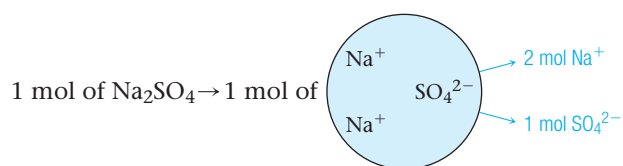
Each molecule of $\text{C}_2\text{H}_3\text{Cl}$ contains two carbon atoms, three hydrogen atoms, and one chlorine atom, so 1 mol of $\text{C}_2\text{H}_3\text{Cl}$ molecules contains 2 mol of C atoms, 3 mol of H atoms, and 1 mol of Cl atoms.

$$\begin{array}{r} \text{Mass of 2 mol of C atoms: } 2 \times 12.01 = 24.02 \text{ g} \\ \text{Mass of 3 mol of H atoms: } 3 \times 1.008 = 3.024 \text{ g} \\ \text{Mass of 1 mol of Cl atoms: } 1 \times 35.45 = 35.45 \text{ g} \\ \hline 62.494 \text{ g} \end{array}$$

The molar mass of $\text{C}_2\text{H}_3\text{Cl}$ is 62.49 g (rounding to the correct number of significant figures).

Practice Problem Exercise 6.6

The formula for sodium sulfate is Na_2SO_4 . One mole of Na_2SO_4 contains 2 mol of sodium ions and 1 mol of sulfate ions.



$$\begin{array}{r} \text{Mass of 2 mol of Na}^+ = 2 \times 22.99 = 45.98 \text{ g} \\ \text{Mass of 1 mol of SO}_4^{2-} = 32.07 + 4(16.00) = 96.07 \text{ g} \\ \text{Mass of 1 mol of Na}_2\text{SO}_4 = 142.05 \text{ g} \end{array}$$

The molar mass for sodium sulfate is 142.05 g.

A sample of sodium sulfate with a mass of 300.0 g represents more than 1 mol. (Compare 300.0 g to the molar mass of Na_2SO_4 .) We calculate the number of moles of Na_2SO_4 present in 300.0 g as follows:

$$300.0 \text{ g Na}_2\text{SO}_4 \times \frac{1 \text{ mol Na}_2\text{SO}_4}{142.5 \text{ g Na}_2\text{SO}_4} = 2.112 \text{ mol Na}_2\text{SO}_4$$

Practice Problem Exercise 6.8

First we must compute the mass of 1 mol of C_2F_4 molecules (the molar mass). Because 1 mol of C_2F_4 contains 2 mol of C atoms and 4 mol of F atoms, we have:

$$2 \text{ mol C} \times \frac{12.01 \text{ g}}{\text{mol}} = 24.02 \text{ g C}$$

$$4 \text{ mol F} \times \frac{19.00 \text{ g}}{\text{mol}} = 76.00 \text{ g F}$$

$$\text{Mass of 1 mol of C}_2\text{F}_4: 100.02 \text{ g} = \text{molar mass}$$

Using the equivalence statement $100.02 \text{ g C}_2\text{F}_4 = 1 \text{ mol C}_2\text{F}_4$, we calculate the moles of C_2F_4 units in 135 g of Teflon.

$$135 \text{ g C}_2\text{F}_4 \text{ units} \times \frac{1 \text{ mol C}_2\text{F}_4}{100.02 \text{ g C}_2\text{F}_4} = 1.35 \text{ mol C}_2\text{F}_4 \text{ units}$$

Next, using the equivalence statement $1 \text{ mol} = 6.022 \times 10^{23}$ units, we calculate the number in C_2F_4 units in 135 g of Teflon.

$$135 \text{ mol } \text{C}_2\text{F}_4 \times \frac{6.022 \times 10^{23} \text{ units}}{1 \text{ mol}} = 8.13 \times 10^{23} \text{ C}_2\text{F}_4 \text{ units}$$

Practice Problem Exercise 6.9

The molar mass of penicillin F is computed as follows:

$$\text{C: } 14 \text{ mol} \times 12.01 \frac{\text{g}}{\text{mol}} = 168.1 \text{ g}$$

$$\text{H: } 20 \text{ mol} \times 1.008 \frac{\text{g}}{\text{mol}} = 20.16 \text{ g}$$

$$\text{N: } 2 \text{ mol} \times 14.01 \frac{\text{g}}{\text{mol}} = 28.02 \text{ g}$$

$$\text{S: } 1 \text{ mol} \times 32.07 \frac{\text{g}}{\text{mol}} = 32.07 \text{ g}$$

$$\text{O: } 4 \text{ mol} \times 16.00 \frac{\text{g}}{\text{mol}} = 64.00 \text{ g}$$

Mass of 1 mol of $\text{C}_{14}\text{H}_{20}\text{N}_2\text{SO}_4 = 312.39 \text{ g} = 312.4 \text{ g}$

$$\text{Mass percent of C} = \frac{168.1 \text{ g C}}{312.4 \text{ g C}_{14}\text{H}_{20}\text{N}_2\text{SO}_4} \times 100\% = 53.81\%$$

$$\text{Mass percent of H} = \frac{20.16 \text{ g H}}{312.4 \text{ g C}_{14}\text{H}_{20}\text{N}_2\text{SO}_4} \times 100\% = 6.453\%$$

$$\text{Mass percent of N} = \frac{28.02 \text{ g N}}{312.4 \text{ g C}_{14}\text{H}_{20}\text{N}_2\text{SO}_4} \times 100\% = 8.969\%$$

$$\text{Mass percent of S} = \frac{32.07 \text{ g S}}{312.4 \text{ g C}_{14}\text{H}_{20}\text{N}_2\text{SO}_4} \times 100\% = 10.27\%$$

$$\text{Mass percent of O} = \frac{64.00 \text{ g O}}{312.4 \text{ g C}_{14}\text{H}_{20}\text{N}_2\text{SO}_4} \times 100\% = 20.49\%$$

Check: The percentages add up to 99.99%.

Practice Problem Exercise 6.12

Step 1 0.6884 g lead and 0.2356 g chlorine

$$\text{Step 2 } 0.6884 \text{ g Pb} \times \frac{1 \text{ mol Pb}}{207.2 \text{ g Pb}} = 0.003322 \text{ mol Pb}$$

$$0.2356 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 0.006646 \text{ mol Cl}$$

$$\text{Step 3 } \frac{0.003322 \text{ mol Pb}}{0.003322} = 1.000 \text{ mol Pb}$$

$$\frac{0.006646 \text{ mol Cl}}{0.003322} = 2.001 \text{ mol Cl}$$

These numbers are very close to integers, so step 4 is unnecessary. The empirical formula is PbCl_2 .

Practice Problem Exercise 6.13

Step 1 0.8007 g C, 0.9333 g N, 0.2016 g H, and 2.133 g O

$$\text{Step 2 } 0.8007 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.06667 \text{ mol C}$$

$$0.9333 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.06662 \text{ mol N}$$

$$0.2016 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 0.2000 \text{ mol H}$$

$$2.133 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.1333 \text{ mol O}$$

$$\text{Step 3 } \frac{0.06667 \text{ mol C}}{0.06667} = 1.001 \text{ mol C}$$

$$\frac{0.06662 \text{ mol N}}{0.06667} = 1.000 \text{ mol N}$$

$$\frac{0.2000 \text{ mol H}}{0.06662} = 3.002 \text{ mol H}$$

$$\frac{0.1333 \text{ mol O}}{0.06662} = 2.001 \text{ mol O}$$

The empirical formula is CNH_3O_2 .

Practice Problem Exercise 6.14

Step 1 In 100.00 g of Nylon-6 the masses of elements present are 63.68 g C, 12.38 g N, 9.80 g H, and 14.14 g O.

$$\text{Step 2 } 63.68 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 5.302 \text{ mol C}$$

$$12.38 \text{ g N} \times \frac{1 \text{ mol N}}{14.01 \text{ g N}} = 0.8837 \text{ mol N}$$

$$9.80 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 9.72 \text{ mol H}$$

$$14.14 \text{ g O} \times \frac{1 \text{ mol O}}{16.00 \text{ g O}} = 0.8838 \text{ mol O}$$

$$\text{Step 3 } \frac{5.302 \text{ mol C}}{0.8836} = 6.000 \text{ mol C}$$

$$\frac{0.8837 \text{ mol N}}{0.8837} = 1.000 \text{ mol N}$$

$$\frac{9.72 \text{ mol H}}{0.8837} = 11.0 \text{ mol H}$$

$$\frac{0.8838 \text{ mol O}}{0.8837} = 1.000 \text{ mol O}$$

The empirical formula for Nylon-6 is $\text{C}_6\text{NH}_{11}\text{O}$.

Practice Problem Exercise 6.15

Step 1 First we convert the mass percents to mass in grams. In 100.0 g of the compound, there are 71.65 g of chlorine, 24.27 g of carbon, and 47.0 g of hydrogen.

Step 2 We use these masses to compute the moles of atoms present.

$$71.65 \text{ g Cl} \times \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} = 2.021 \text{ mol Cl}$$

$$24.27 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 2.021 \text{ mol C}$$

$$4.07 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} = 4.04 \text{ mol H}$$

Step 3 Dividing each mole value by 2.021 (the smallest number of moles present), we obtain the empirical formula ClCH_2 .

To determine the molecular formula, we must compare the empirical formula mass to the molar mass. The empirical formula mass is 49.48.

$$\begin{array}{l} \text{Cl:} \quad 35.45 \\ \text{C:} \quad 12.01 \\ 2 \text{ H: } 2 \times (1.008) \\ \text{ClCH}_2: \quad 49.48 = \text{empirical formula mass} \end{array}$$

The molar mass is known to be 98.96. We know that

$$\text{Molar mass} = n \times (\text{empirical formula mass})$$

So we can obtain the value of n as follows:

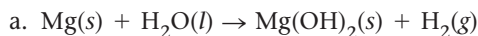
$$\frac{\text{Molar mass}}{\text{Empirical formula mass}} = \frac{98.96}{49.48} = 2$$

$$\text{Molecular formula} = (\text{ClCH}_2)_2 = \text{Cl}_2\text{C}_2\text{H}_4$$

This substance is composed of molecules with the formula $\text{Cl}_2\text{C}_2\text{H}_4$.

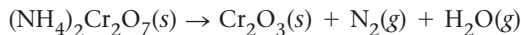
Chapter 7

Practice Problem Exercise 7.1

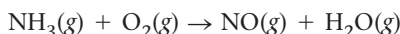


Note that magnesium (which is in Group 2) always forms the Mg^{2+} cation and thus requires two OH^- anions for a zero net charge.

b. Ammonium dichromate contains the polyatomic ions NH_4^+ and $\text{Cr}_2\text{O}_7^{2-}$ (you should have these memorized). Because NH_4^+ has a 1+ charge, two NH_4^+ cations are required for each $\text{Cr}_2\text{O}_7^{2-}$, with its 2- charge, to give the formula $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$. Chromium(III) oxide contains Cr^{3+} ions—signified by chromium(III)—and O^{2-} (the oxide ion). To achieve a net charge of zero, the solid must contain two Cr^{3+} ions for every three O^{2-} ions, so the formula is Cr_2O_3 . Nitrogen gas contains diatomic molecules and is written $\text{N}_2(g)$, and gaseous water is written $\text{H}_2\text{O}(g)$. Thus the unbalanced equation for the decomposition of ammonium dichromate is



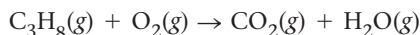
c. Gaseous ammonia, $\text{NH}_3(g)$, and gaseous oxygen, $\text{O}_2(g)$, react to form nitrogen monoxide gas, $\text{NO}(g)$, plus gaseous water, $\text{H}_2\text{O}(g)$. The unbalanced equation is



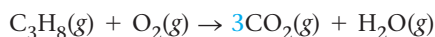
Practice Problem Exercise 7.3

Step 1 The reactants are propane, $\text{C}_3\text{H}_8(g)$, and oxygen, $\text{O}_2(g)$; the products are carbon dioxide, $\text{CO}_2(g)$, and water, $\text{H}_2\text{O}(g)$. All are in the gaseous state.

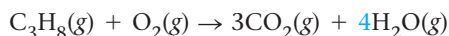
Step 2 The unbalanced equation for the reaction is



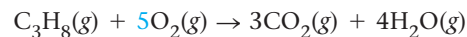
Step 3 We start with C_3H_8 because it is the most complicated molecule. C_3H_8 contains three carbon atoms per molecule, so a coefficient of 3 is needed for CO_2 .



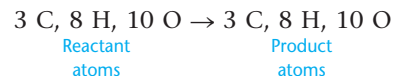
Also, each C_3H_8 molecule contains eight hydrogen atoms, so a coefficient of 4 is required for H_2O .



The final element to be balanced is oxygen. Note that the left side of the equation now has two oxygen atoms, and the right side has ten. We can balance the oxygen by using a coefficient of 5 for O_2 .

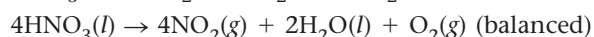
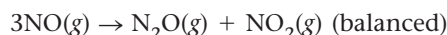
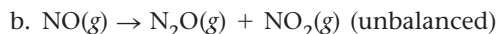
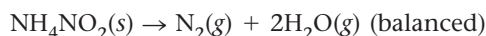
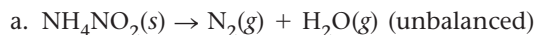


Step 4 Check:



We cannot divide all coefficients by a given integer to give smaller integer coefficients.

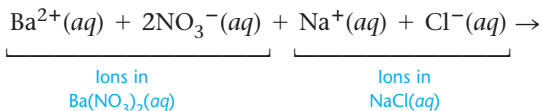
Practice Problem Exercise 7.4



Chapter 8

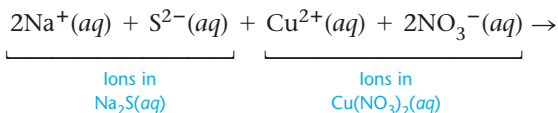
Practice Problem Exercise 8.2

a. The ions present are

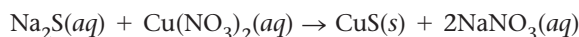


Exchanging the anions gives the possible solid products BaCl_2 and NaNO_3 . Using Table 8.1, we see that both substances are very soluble (rules 1, 2, and 3). Thus no solid forms.

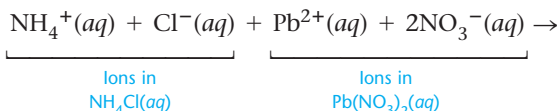
b. The ions present in the mixed solution before any reaction occurs are



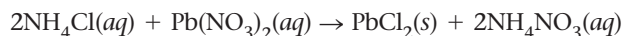
Exchanging the anions gives the possible solid products CuS and NaNO_3 . According to rules 1 and 2 in Table 8.1, NaNO_3 is soluble, and by rule 6, CuS should be insoluble. Thus CuS will precipitate. The balanced equation is



c. The ions present are

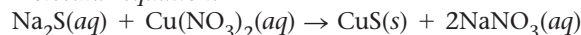


Exchanging the anions gives the possible solid products NH_4NO_3 and PbCl_2 . NH_4NO_3 is soluble (rules 1 and 2) and PbCl_2 is insoluble (rule 3). Thus PbCl_2 will precipitate. The balanced equation is

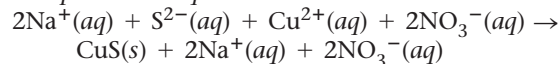


Practice Problem Exercise 8.3

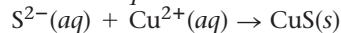
a. Molecular equation:



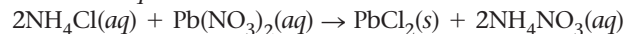
Complete ionic equation:



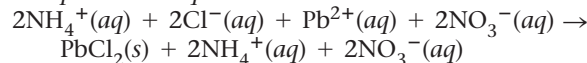
Net ionic equation:



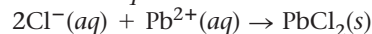
b. Molecular equation:



Complete ionic equation:

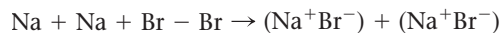


Net ionic equation:

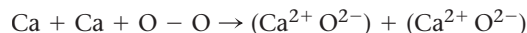


Practice Problem Exercise 8.5

a. The compound NaBr contains the ions Na^+ and Br^- . Thus each sodium atom loses one electron ($\text{Na} \rightarrow \text{Na}^+ + e^-$), and each bromine atom gains one electron ($\text{Br} + e^- \rightarrow \text{Br}^-$).



b. The compound CaO contains the Ca^{2+} and O^{2-} ions. Thus each calcium atom loses two electrons ($\text{Ca} \rightarrow \text{Ca}^{2+} + 2e^-$), and each oxygen atom gains two electrons ($\text{O} + 2e^- \rightarrow \text{O}^{2-}$).



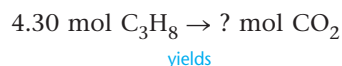
Practice Problem Exercise 8.6

- oxidation–reduction reaction; combustion reaction
- synthesis reaction; oxidation–reduction reaction; combustion reaction
- synthesis reaction; oxidation–reduction reaction
- decomposition reaction; oxidation–reduction reaction
- precipitation reaction (and double–displacement)
- synthesis reaction; oxidation–reduction reaction
- acid–base reaction (and double–displacement)
- combustion reaction; oxidation–reduction reaction

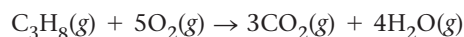
Chapter 9

Practice Problem Exercise 9.3

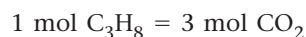
The problem can be stated as follows:



From the balanced equation



we derive the equivalence statement



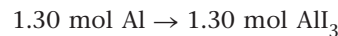
The appropriate conversion factor (moles of C_3H_8 must cancel) is $3 \text{ mol CO}_2/1 \text{ mol C}_3\text{H}_8$, and the calculation is

$$4.30 \text{ mol C}_3\text{H}_8 \times \frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} = 12.9 \text{ mol CO}_2$$

Thus we can say

$$4.30 \text{ mol C}_3\text{H}_8 \text{ yields } 12.9 \text{ mol CO}_2$$

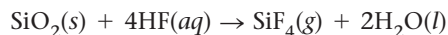
Practice Problem Exercise 9.5



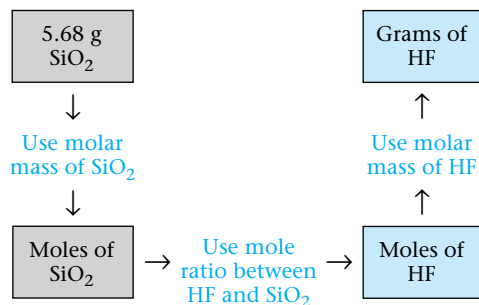
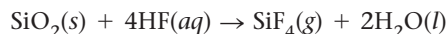
$$1.30 \text{ mol AlI}_3 \times \frac{407.7 \text{ g AlI}_3}{1 \text{ mol AlI}_3} = 530. \text{ g AlI}_3$$

Practice Problem Exercise 9.6

a. We first write the balanced equation.



The map of the steps required is



We convert 5.68 g of SiO_2 to moles as follows:

$$5.68 \text{ g SiO}_2 \times \frac{1 \text{ mol SiO}_2}{60.09 \text{ g SiO}_2} = 9.45 \times 10^{-2} \text{ mol SiO}_2$$

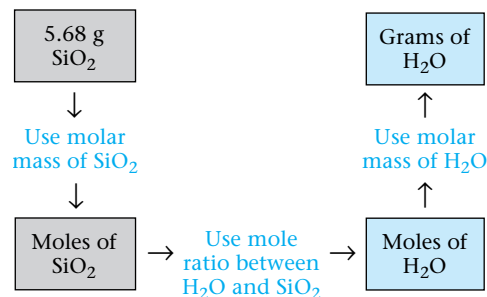
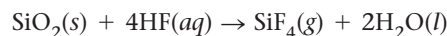
Using the balanced equation, we obtain the appropriate mole ratio and convert to moles of HF.

$$9.45 \times 10^{-2} \text{ mol SiO}_2 \times \frac{4 \text{ mol HF}}{1 \text{ mol SiO}_2} = 3.78 \times 10^{-1} \text{ mol HF}$$

Finally, we calculate the mass of HF by using its molar mass.

$$3.78 \times 10^{-1} \text{ mol HF} \times \frac{20.01 \text{ g HF}}{1 \text{ mol HF}} = 7.56 \text{ g HF}$$

b. The map for this problem is



We have already accomplished the first conversion in part a. Using the balanced equation, we obtain moles of H₂O as follows:

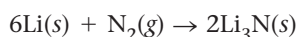
$$9.45 \times 10^{-2} \cancel{\text{mol SiO}_2} \times \frac{2 \text{ mol H}_2\text{O}}{1 \cancel{\text{mol SiO}_2}} = 1.89 \times 10^{-1} \text{ mol H}_2\text{O}$$

The mass of water formed is

$$1.89 \times 10^{-1} \cancel{\text{mol H}_2\text{O}} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \cancel{\text{mol H}_2\text{O}}} = 3.41 \text{ g H}_2\text{O}$$

Practice Problem Exercise 9.8

Step 1 The balanced equation for the reaction is



Step 2 To determine the limiting reactant, we must convert the masses of lithium (atomic mass = 6.941 g) and nitrogen (molar mass = 28.02 g) to moles.

$$56.0 \text{ g Li} \times \frac{1 \text{ mol Li}}{6.941 \text{ g Li}} = 8.07 \text{ mol Li}$$

$$56.0 \text{ g N}_2 \times \frac{1 \text{ mol N}_2}{28.02 \text{ g N}_2} = 2.00 \text{ mol N}_2$$

Step 3 Using the mole ratio from the balanced equation, we can calculate the moles of lithium required to react with 2.00 mol of nitrogen.

$$2.00 \cancel{\text{mol N}_2} \times \frac{6 \text{ mol Li}}{1 \cancel{\text{mol N}_2}} = 12.0 \text{ mol Li}$$

Therefore, 12.0 mol of Li is required to react with 2.00 mol of N₂. However, we have only 8.07 mol of Li, so lithium is limiting. It will be consumed before the nitrogen runs out.

Step 4 Because lithium is the limiting reactant, we must use the 8.07 mol of Li to determine how many moles of Li₃N can be formed.

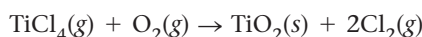
$$8.07 \cancel{\text{mol Li}} \times \frac{2 \text{ mol Li}_3\text{N}}{6 \cancel{\text{mol Li}}} = 2.69 \text{ mol Li}_3\text{N}$$

Step 5 We can now use the molar mass of Li₃N (34.83 g) to calculate the mass of Li₃N formed.

$$2.69 \cancel{\text{mol Li}_3\text{N}} \times \frac{34.83 \text{ g Li}_3\text{N}}{1 \cancel{\text{mol Li}_3\text{N}}} = 93.7 \text{ g Li}_3\text{N}$$

Practice Problem Exercise 9.9

a. **Step 1** The balanced equation is



Step 2 The numbers of moles of reactants are

$$6.71 \times 10^3 \cancel{\text{g TiCl}_4} \times \frac{1 \text{ mol TiCl}_4}{189.68 \cancel{\text{g TiCl}_4}} = 3.54 \times 10^1 \text{ mol TiCl}_4$$

$$2.45 \times 10^3 \cancel{\text{g O}_2} \times \frac{1 \text{ mol O}_2}{32.00 \cancel{\text{g O}_2}} =$$

$$7.66 \times 10^1 \text{ mol O}_2$$

Step 3 In the balanced equation both TiCl₄ and O₂ have coefficients of 1, so

$$1 \text{ mol TiCl}_4 = 1 \text{ mol O}_2$$

and

$$3.54 \times 10^1 \cancel{\text{mol TiCl}_4} \times \frac{1 \text{ mol O}_2}{1 \cancel{\text{mol TiCl}_4}} =$$

$$3.54 \times 10^1 \text{ mol O}_2 \text{ required}$$

We have 7.66×10^1 mol of O₂, so the O₂ is an excess and the TiCl₄ is limiting. This makes sense. TiCl₄ and O₂ react in a 1:1 mole ratio, so the TiCl₄ is limiting because fewer moles of TiCl₄ are present than moles of O₂.

Step 4 We will now use the moles of TiCl₄ (the limiting reactant) to determine the moles of TiO₂ that would form if the reaction produced 100% of the expected yield (the theoretical yield).

$$3.54 \times 10^1 \cancel{\text{mol TiCl}_4} \times \frac{1 \text{ mol TiO}_2}{1 \cancel{\text{mol TiCl}_4}} =$$

$$3.54 \times 10^1 \text{ mol TiO}_2$$

The mass of TiO₂ expected for 100% yield is

$$3.54 \times 10^1 \cancel{\text{mol TiO}_2} \times \frac{79.88 \text{ g TiO}_2}{1 \cancel{\text{mol TiO}_2}} =$$

$$2.83 \times 10^3 \text{ g TiO}_2$$

This amount represents the theoretical yield.

b. Because the reaction is said to give only a 75.0% yield of TiO₂, we use the definition of percent yield.

$$\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\% = \% \text{ yield}$$

to write the equation

$$\frac{\text{Actual yield}}{2.83 \times 10^3 \text{ g TiO}_2} \times 100\% = 75.0\% \text{ yield}$$

We now want to solve for the actual yield. First we divide both sides by 100%.

$$\frac{\text{Actual yield}}{2.83 \times 10^3 \text{ g TiO}_2} \times \frac{100\%}{100\%} = \frac{75.0}{100} = 0.750$$

Then we multiply both sides by 2.83×10^3 g TiO₂.

$$2.83 \times 10^3 \text{ g TiO}_2 \times \frac{\text{Actual yield}}{2.83 \times 10^3 \text{ g TiO}_2} = 0.750 \times 2.83 \times 10^3 \text{ g TiO}_2$$

$$\text{Actual yield} = 0.750 \times 2.83 \times 10^3 \text{ g TiO}_2$$

$$= 2.12 \times 10^3 \text{ g TiO}_2$$

Thus 2.12×10^3 g of TiO₂(s) is actually obtained in this reaction.

Chapter 10

Practice Problem Exercise 10.1

The conversion factor needed is $\frac{1 \text{ cal}}{4.184 \text{ J}}$, and the conversion is

$$28.4 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 6.79 \text{ cal}$$

Practice Problem Exercise 10.2

We know that it takes 4.184 J of energy to change the temperature of each gram of water by 1 °C, so we must multiply 4.184 by the mass of water (454 g) and the temperature change (98.6 °C - 5.4 °C = 93.2 °C).

$$4.184 \frac{\text{J}}{\text{g} \cdot \text{C}} \times 454 \text{ g} \times 93.2 \text{ C} = 1.77 \times 10^5 \text{ J}$$

Practice Problem Exercise 10.3

From Table 10.1, the specific heat capacity for solid gold is 0.13 J/g °C. Because it takes 0.13 J to change the temperature of *one* gram of gold by *one* Celsius degree, we must multiply 0.13 by the sample size (5.63 g) and the change in temperature (32 °C - 21 °C = 11 °C).

$$0.13 \frac{\text{J}}{\text{g} \cdot \text{C}} \times 5.63 \text{ g} \times 11 \text{ C} = 8.1 \text{ J}$$

We can change this energy to units of calories as follows:

$$8.1 \text{ J} \times \frac{1 \text{ cal}}{4.184 \text{ J}} = 1.9 \text{ cal}$$

Practice Problem Exercise 10.4

Table 10.1 lists the specific heat capacities of several metals. We want to calculate the specific heat capacity (*s*) for this metal and then use Table 10.1 to identify the metal. Using the equation

$$Q = s \times m \times \Delta T$$

we can solve for *s* by dividing both sides by *m* (the mass of the sample) and by ΔT :

$$\frac{Q}{m \times \Delta T} = s$$

In this case,

$$Q = \text{energy (heat) required} = 10.1 \text{ J}$$

$$m = 2.8 \text{ g}$$

$$\Delta T = \text{temperature change} = 36 \text{ C} - 21 \text{ C} = 15 \text{ C}$$

so

$$s = \frac{Q}{m \times \Delta T} = \frac{10.1 \text{ J}}{(2.8 \text{ g})(15 \text{ C})} = 0.24 \text{ J/g C}$$

Table 10.1 shows that silver has a specific heat capacity of 0.24 J/g °C. The metal is silver.

Practice Problem Exercise 10.5

We are told that 1652 kJ of energy is *released* when 4 mol of Fe reacts. We first need to determine what number of moles 1.00 g Fe represents.

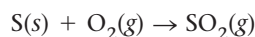
$$1.00 \text{ g Fe} \times \frac{1 \text{ mol}}{55.85 \text{ g}} = 1.79 \times 10^{-2} \text{ mol Fe}$$

$$1.79 \times 10^{-2} \text{ mol Fe} \times \frac{1652 \text{ kJ}}{4 \text{ mol Fe}} = 7.39 \text{ kJ}$$

Thus 7.39 kJ of energy (as heat) is released when 1.00 g of iron reacts.

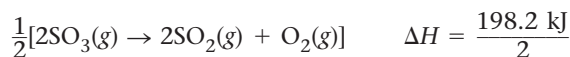
Practice Problem Exercise 10.6

Noting the reactants and products in the desired reaction

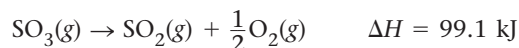


We need to reverse the second equation and multiply it

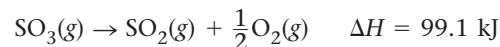
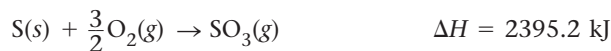
by $\frac{1}{2}$. This reverses the sign and cuts the amount of energy by a factor of 2.



or



Now we add this reaction to the first reaction



Chapter 11

Practice Problem Exercise 11.1

- Circular pathways for electrons in the Bohr model
- Three-dimensional probability maps that represent the likelihood that the electron will occupy a given point in space. The details of electron motion are not described by an orbital.
- The surface that contains 90% of the total electron probability. That is, the electron is found somewhere inside this surface 90% of the time.
- A sublevel is a set of orbitals of a given type of orbital within a principal energy level. For example, there are three sublevels in principal level 3; they consist of the 3s orbital, the three 3p orbitals, and the five 3d orbitals.

Practice Problem Exercise 11.2

Element	Electron Configuration	Orbital Diagram
		1s 2s 2p 3s 3p
Al	$1s^2 2s^2 2p^6 3s^2 3p^1$	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow$ $\uparrow\downarrow$ \uparrow \square \square
	$[\text{Ne}]3s^2 3p^1$	
Si	$[\text{Ne}]3s^2 3p^2$	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow$ $\uparrow\downarrow$ \uparrow \uparrow \square
P	$[\text{Ne}]3s^2 3p^3$	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow$ $\uparrow\downarrow$ \uparrow \uparrow \uparrow
S	$[\text{Ne}]3s^2 3p^4$	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ \uparrow \uparrow
Cl	$[\text{Ne}]3s^2 3p^5$	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow\uparrow\downarrow$ \uparrow
Ar	$[\text{Ne}]3s^2 3p^6$	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow\uparrow\downarrow\uparrow\downarrow$

Practice Problem Exercise 11.3

Fluorine (F): In Group 7 and Period 2, it is the fifth "2p element." The configuration is $1s^2 2s^2 2p^5$, or $[\text{He}]2s^2 2p^5$.

Silicon (Si): In Group 4 and Period 3, it is the second of the "3p elements." The configuration is $1s^2 2s^2 2p^6 3s^2 3p^2$, or $[\text{Ne}]3s^2 3p^2$.

Cesium (Cs): In Group 1 and Period 6, it is the first of the "6s elements." The configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^1$, or $[\text{Xe}]6s^1$.

Lead (Pb): In Group 4 and Period 6, it is the second of the "6p elements." The configuration is $[\text{Xe}]6s^2 4f^{14} 5d^{10} 6p^2$.

Iodine (I): In Group 7 and Period 5, it is the fifth of the "5p elements." The configuration is $[\text{Kr}]5s^2 4d^{10} 5p^5$.

Chapter 12

Practice Problem Exercise 12.1

Using the electronegativity values given in Figure 12.4, we choose the bond in which the atoms exhibit the largest difference in electronegativity. (Electronegativity values are shown in parentheses.)

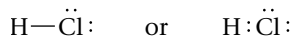
- a. $\text{H}-\text{C} > \text{H}-\text{P}$
(2.1)(2.5) (2.1)(2.1)
- b. $\text{O}-\text{I} > \text{O}-\text{F}$
(3.5)(2.5) (3.5)(4.0)
- c. $\text{S}-\text{O} > \text{N}-\text{O}$
(2.5)(3.5) (3.0)(3.5)
- d. $\text{N}-\text{H} > \text{Si}-\text{H}$
(3.0)(2.1) (1.8)(2.1)

Practice Problem Exercise 12.2

H has one electron, and Cl has seven valence electrons. This gives a total of eight valence electrons. We first draw in the bonding pair:



We have six electrons yet to place. The H already has two electrons, so we place three lone pairs around the chlorine to satisfy the octet rule.



Practice Problem Exercise 12.3

Step 1 O_3 : $3(6) = 18$ valence electrons

Step 2 $\text{O}-\text{O}-\text{O}$

Step 3 $:\ddot{\text{O}}=\ddot{\text{O}}-\ddot{\text{O}}:$ and $:\ddot{\text{O}}-\ddot{\text{O}}=\ddot{\text{O}}:$

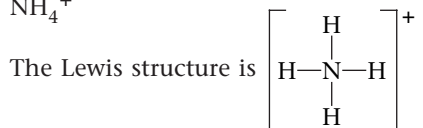
This molecule shows resonance (it has two valid Lewis structures).

Practice Problem Exercise 12.4

See table at the top of page A22.

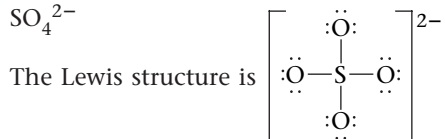
Practice Problem Exercise 12.6

a. NH_4^+



(See Practice Problem Exercise 12.4.) There are four pairs of electrons around the nitrogen. This requires a tetrahedral arrangement of electron pairs. The NH_4^+ ion has a tetrahedral molecular structure (case 3 in Table 12.4), because all electron pairs are shared.

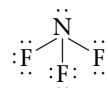
b. SO_4^{2-}



(See Practice Problem Exercise 12.4) The four electron pairs around the sulfur require a tetrahedral arrangement. The SO_4^{2-} has a tetrahedral molecular structure (case 3 in Table 12.4).

c. NF_3

The Lewis structure is



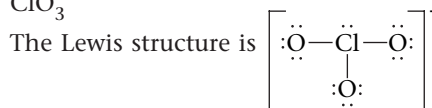
(See Practice Problem Exercise 12.4) The four pairs of electrons on the nitrogen require a tetrahedral arrangement. In this case only three of the pairs are shared with the fluorine atoms, leaving one lone pair. Thus the molecular structure is a trigonal pyramid (case 4 in Table 12.4).

d. H_2S

The Lewis structure is $\text{H}-\ddot{\text{S}}-\text{H}$

(See Practice Problem Exercise 12.4) The four pairs of electrons around the sulfur require a tetrahedral arrangement. In this case two pairs are shared with hydrogen atoms, leaving two lone pairs. Thus the molecular structure is bent or V-shaped (case 5 in Table 12.4).

e. ClO_3^-



(See Practice Problem Exercise 12.4.) The four pairs of electrons require a tetrahedral arrangement. In this case, three pairs are shared with oxygen atoms, leaving one lone pair. Thus the molecular structure is a trigonal pyramid (case 4 in Table 12.4).

f. BeF_2

The Lewis structure is $:\ddot{\text{F}}-\text{Be}-\ddot{\text{F}}:$

The two electron pairs on beryllium require a linear arrangement. Because both pairs are shared by fluorine atoms, the molecular structure is also linear (case 1 in Table 12.4).

Chapter 13

Practice Problem Exercise 13.1

We know that $1.000 \text{ atm} = 760.0 \text{ mm Hg}$. So

$$525 \text{ mm Hg} \times \frac{1.000 \text{ atm}}{760.0 \text{ mm Hg}} = 0.691 \text{ atm}$$

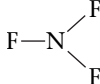
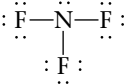


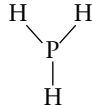
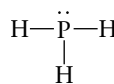
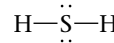
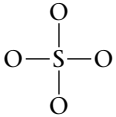
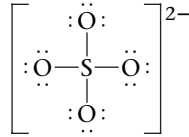
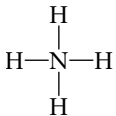
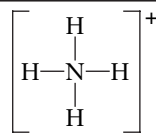
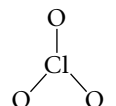
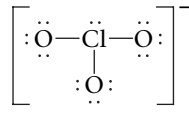
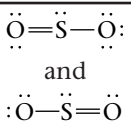
Practice Problem Exercise 13.2

Initial Conditions	Final Conditions
$P_1 = 635 \text{ torr}$	$P_2 = 785 \text{ torr}$
$V_1 = 1.51 \text{ L}$	$V_2 = ?$

Solving Boyle's law ($P_1V_1 = P_2V_2$) for V_2 gives

$$\begin{aligned} V_2 &= V_1 \times \frac{P_1}{P_2} \\ &= 1.51 \text{ L} \times \frac{635 \text{ torr}}{785 \text{ torr}} = 1.22 \text{ L} \end{aligned}$$

Note that the volume decreased, as the increase in pressure led us to expect.

Molecule or Ion	Total Valence Electrons	Draw Single Bonds	Calculate Number of Electrons Remaining	Use Remaining Electrons to Achieve Noble Gas Configurations	Check	
					Atom	Electrons
a. NF_3	$5 + 3(7) = 26$		$26 - 6 = 20$		N F	8 8
b. O_2	$2(6) = 12$	$\text{O}-\text{O}$	$12 - 2 = 10$		O	8
c. CO	$4 + 6 = 10$	$\text{C}-\text{O}$	$10 - 2 = 8$		C O	8 8
d. PH_3	$5 + 3(1) = 8$		$8 - 6 = 2$		P H	8 8
e. H_2S	$2(1) + 6 = 8$	$\text{H}-\text{S}-\text{H}$	$8 - 4 = 4$		S H	8 2
f. SO_4^{2-}	$6 + 4(6) + 2 = 32$		$32 - 8 = 24$		S O	8 8
g. NH_4^+	$5 + 4(1) - 1 = 8$		$8 - 8 = 0$		N H	8 2
h. ClO_3^-	$7 + 3(6) + 1 = 26$		$26 - 6 = 20$		Cl O	8 2
i. SO_2	$6 + 2(6) = 18$	$\text{O}-\text{S}-\text{O}$	$18 - 4 = 14$		S O	8 8

Answer to Practice Problem Exercise 12.4

Practice Problem Exercise 13.5

Because the temperature of the gas inside the bubble decreases (at constant pressure), the bubble gets smaller. The conditions are

Initial Conditions

$$T_1 = 28^\circ\text{C} = 28 + 273 = 301\text{ K}$$

$$V_1 = 23\text{ cm}^3$$

Final Conditions

$$T_2 = 18^\circ\text{C} = 18 + 273 = 291\text{ K}$$

$$V_2 = ?$$

Solving Charles's law,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

for V_2 gives

$$V_2 = V_1 \times \frac{T_2}{T_1} = 23\text{ cm}^3 \times \frac{291\text{ K}}{301\text{ K}} = 22\text{ cm}^3$$

Practice Problem Exercise 13.7

Because the temperature and pressure of the two samples are the same, we can use Avogadro's law in the form

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

The following information is given:

Sample 1

$$V_1 = 36.7\text{ L}$$

$$n_1 = 1.5\text{ mol}$$

Sample 2

$$V_2 = 16.5\text{ L}$$

$$n_2 = ?$$

We can now solve Avogadro's law for the value of n_2 (the moles of n_2 in Sample 2):

$$n_2 = n_1 \times \frac{V_2}{V_1} = 1.5 \text{ mol} \times \frac{16.5 \cancel{\text{L}}}{36.7 \cancel{\text{L}}} = 0.67 \text{ mol}$$

Here n_2 is smaller than n_1 , which makes sense in view of the fact that V_2 is smaller than V_1 .

Note: We isolate n_2 from Avogadro's law as given above by multiplying both sides of the equation by n_2 and then by n_1/V_1 ,

$$\left(n_2 \times \frac{n_1}{V_1}\right) \frac{V_1}{n_1} = \left(n_2 \times \frac{n_1}{V_1}\right) \frac{V_2}{n_2}$$

to give $n_2 = n_1 \times V_2/V_1$.

Practice Problem Exercise 13.8

We are given the following information:

$$\begin{aligned} P &= 1.00 \text{ atm} \\ V &= 2.70 \times 10^6 \text{ L} \\ n &= 1.10 \times 10^5 \text{ mol} \end{aligned}$$

We solve for T by dividing both sides of the ideal gas law by nR :

$$\frac{PV}{nR} = \frac{nRT}{nR}$$

to give

$$T = \frac{PV}{nR} = \frac{(1.00 \text{ atm})(2.70 \times 10^6 \cancel{\text{L}})}{(1.10 \times 10^5 \cancel{\text{mol}})(0.08206 \frac{\cancel{\text{L atm}}}{\text{K mol}})} = 299 \text{ K}$$

The temperature of the helium is 299 K, or $299 - 273 = 26^\circ\text{C}$.

Practice Problem Exercise 13.9

We are given the following information about the radon sample:

$$\begin{aligned} n &= 1.5 \text{ mol} \\ V &= 21.0 \text{ L} \\ T &= 33^\circ\text{C} = 33 + 273 = 306 \text{ K} \\ P &= ? \end{aligned}$$

We solve the ideal gas law ($PV = nRT$) for P by dividing both sides of the equation by V :

$$P = \frac{nRT}{V} = \frac{(1.5 \cancel{\text{mol}})(0.08206 \frac{\cancel{\text{L atm}}}{\text{K mol}})(306 \cancel{\text{K}})}{21.0 \cancel{\text{L}}} = 1.8 \text{ atm}$$

Practice Problem Exercise 13.10

To solve this problem we take the ideal gas law and separate those quantities that change from those that remain constant (on opposite sides of the equation). In this case volume and temperature change, and number of moles and pressure (and, of course, R) remain constant. So $PV = nRT$ becomes $V/T = nR/P$, which leads to

$$\frac{V_1}{T_1} = \frac{nR}{P} \text{ and } \frac{V_2}{T_2} = \frac{nR}{P}$$

Combining these gives

$$\frac{V_1}{T_1} = \frac{nR}{P} = \frac{V_2}{T_2} \text{ or } \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

We are given

Initial Conditions

$$\begin{aligned} T_1 &= 5^\circ\text{C} = 5 + 273 = 278 \text{ K} \\ V_1 &= 3.8 \text{ L} \end{aligned}$$

Final Conditions

$$\begin{aligned} T_2 &= 86^\circ\text{C} = 86 + 273 = 359 \text{ K} \\ V_2 &= ? \end{aligned}$$

Thus

$$V_2 = \frac{T_2 V_1}{T_1} = \frac{(359 \cancel{\text{K}})(3.8 \text{ L})}{278 \cancel{\text{K}}} = 4.9 \text{ L}$$

Check: Is the answer sensible? In this case the temperature was increased (at constant pressure), so the volume should increase. The answer makes sense.

Note that this problem could be described as a "Charles's law problem." The real advantage of using the ideal gas law is that you need to remember only one equation to do virtually any problem involving gases.

Practice Problem Exercise 13.11

We are given the following information:

Initial Conditions

$$\begin{aligned} P_1 &= 0.747 \text{ atm} \\ T_1 &= 13^\circ\text{C} = 13 + 273 = 286 \text{ K} \\ V_1 &= 11.0 \text{ L} \end{aligned}$$

Final Conditions

$$\begin{aligned} P_2 &= 1.18 \text{ atm} \\ T_2 &= 56^\circ\text{C} = 56 + 273 = 329 \text{ K} \\ V_2 &= ? \end{aligned}$$

In this case the number of moles remains constant. Thus we can say

$$\frac{P_1 V_1}{T_1} = nR \text{ and } \frac{P_2 V_2}{T_2} = nR$$

or

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Solving for V_2 gives

$$\begin{aligned} V_2 &= V_1 \times \frac{T_2}{T_1} \times \frac{P_1}{P_2} = (11.0 \text{ L}) \left(\frac{329 \cancel{\text{K}}}{286 \cancel{\text{K}}} \right) \left(\frac{0.747 \cancel{\text{atm}}}{1.18 \cancel{\text{atm}}} \right) \\ &= 8.01 \text{ L} \end{aligned}$$

Practice Problem Exercise 13.12

As usual when dealing with gases, we can use the ideal gas equation $PV = nRT$. First consider the information given:

$$\begin{aligned} P &= 0.91 \text{ atm} = P_{\text{total}} \\ V &= 2.0 \text{ L} \\ T &= 25^\circ\text{C} = 25 + 273 = 298 \text{ K} \end{aligned}$$

Given this information, we can calculate the number of moles of gas in the mixture:

$n_{\text{total}} = n_{\text{N}_2} + n_{\text{O}_2}$: Solving for n in the ideal gas equation gives

$$n_{\text{total}} = \frac{P_{\text{total}} V}{RT} = \frac{(0.91 \text{ atm})(2.0 \cancel{\text{L}})}{(0.08206 \frac{\cancel{\text{L atm}}}{\text{K mol}})(298 \cancel{\text{K}})} = 0.074 \text{ mol}$$

We also know that 0.050 mol N₂ is present. Because

$$n_{\text{total}} = n_{\text{N}_2} + n_{\text{O}_2} = 0.074 \text{ mol}$$

\uparrow
 (0.050 mol)

We can calculate the moles of O₂ present.

$$\begin{aligned} 0.050 \text{ mol} + n_{\text{O}_2} &= 0.074 \text{ mol} \\ n_{\text{O}_2} &= 0.074 \text{ mol} - 0.050 \text{ mol} \\ &= 0.024 \text{ mol} \end{aligned}$$

Now that we know the moles of oxygen present, we can calculate the partial pressure of oxygen from the ideal gas equation.

$$\begin{aligned} P_{\text{O}_2} &= \frac{n_{\text{O}_2}RT}{V} = \frac{(0.024 \text{ mol})\left(0.08206 \frac{\text{L atm}}{\text{K mol}}\right)(298 \text{ K})}{2.0 \text{ L}} \\ &= 0.29 \text{ atm} \end{aligned}$$

Although it is not requested, note that the partial pressure of the N₂ must be 0.62 atm, because

$$\underbrace{0.62 \text{ atm}}_{P_{\text{N}_2}} + \underbrace{0.29 \text{ atm}}_{P_{\text{O}_2}} = \underbrace{0.91 \text{ atm}}_{P_{\text{total}}}$$

Practice Problem Exercise 13.13

The volume is 0.500 L, the temperature is 25 °C (25 + 273 = 298 K), and the total pressure is given as 0.950 atm. Of this total pressure, 24 torr is due to the water vapor. We can calculate the partial pressure of the H₂ because we know that

$$P_{\text{total}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}} = 0.950 \text{ atm}$$

\uparrow
 24 torr

Before we carry out the calculation, however, we must convert the pressures to the same units. Converting P_{H₂O} to atmospheres gives

$$24 \text{ torr} \times \frac{1.000 \text{ atm}}{760.0 \text{ torr}} = 0.032 \text{ atm}$$

Thus

$$P_{\text{total}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}} = 0.950 \text{ atm} = P_{\text{H}_2} + 0.032 \text{ atm}$$

and

$$P_{\text{H}_2} = 0.950 \text{ atm} - 0.032 \text{ atm} = 0.918 \text{ atm}$$

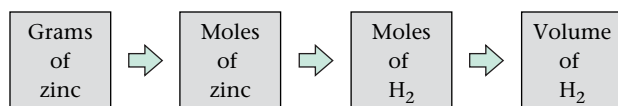
Now that we know the partial pressure of the hydrogen gas, we can use the ideal gas equation to calculate the moles of H₂.

$$\begin{aligned} n_{\text{H}_2} &= \frac{P_{\text{H}_2}V}{RT} = \frac{(0.918 \text{ atm})(0.500 \text{ L})}{\left(0.08206 \frac{\text{L atm}}{\text{K mol}}\right)(298 \text{ K})} \\ &= 0.0188 \text{ mol} = 1.88 \times 10^{-2} \text{ mol} \end{aligned}$$

The sample of gas contains 1.88 × 10⁻² mol of H₂, which exerts a partial pressure of 0.918 atm.

Practice Problem Exercise 13.14

We will solve this problem by taking the following steps:



Step 1 Using the atomic mass of zinc (65.38 g), we calculate the moles of zinc in 26.5 g.

$$26.5 \text{ g Zn} \times \frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} = 0.405 \text{ mol Zn}$$

Step 2 Using the balanced equation, we next calculate the moles of H₂ produced.

$$0.405 \text{ mol Zn} \times \frac{1 \text{ mol H}_2}{1 \text{ mol Zn}} = 0.405 \text{ mol H}_2$$

Step 3 Now that we know the moles of H₂, we can compute the volume of H₂ by using the ideal gas law, where

$$\begin{aligned} P &= 1.50 \text{ atm} \\ V &= ? \\ n &= 0.405 \text{ mol} \\ R &= 0.08206 \text{ L atm/K mol} \\ T &= 19 \text{ °C} = 19 + 273 = 292 \text{ K} \\ V &= \frac{nRT}{P} = \frac{(0.405 \text{ mol})\left(0.08206 \frac{\text{L atm}}{\text{K mol}}\right)(292 \text{ K})}{1.50 \text{ atm}} \\ &= 6.47 \text{ L of H}_2 \end{aligned}$$

Practice Problem Exercise 13.15

Although there are several possible ways to do this problem, the most convenient method involves using the molar volume at STP. First we use the ideal gas equation to calculate the moles of NH₃ present:

$$n = \frac{PV}{RT}$$

where P = 15.0 atm, V = 5.00 L, and T = 25 °C + 273 = 298 K.

$$n = \frac{(15.0 \text{ atm})(5.00 \text{ L})}{\left(0.08206 \frac{\text{L atm}}{\text{K mol}}\right)(298 \text{ K})} = 3.07 \text{ mol}$$

We know that at STP each mole of gas occupies 22.4 L. Therefore, 3.07 mol has the volume

$$3.07 \text{ mol} \times \frac{22.4 \text{ L}}{1 \text{ mol}} = 68.8 \text{ L}$$

The volume of the ammonia at STP is 68.8 L.

Chapter 14

Practice Problem Exercise 14.2

Energy to melt the ice:

$$15 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18 \text{ g H}_2\text{O}} = 0.83 \text{ mol H}_2\text{O}$$

$$0.83 \text{ mol H}_2\text{O} \times 6.02 \frac{\text{kJ}}{\text{mol H}_2\text{O}} = 5.0 \text{ kJ}$$

Energy to heat the water from 0 °C to 100 °C:

$$4.18 \frac{\text{J}}{\text{g °C}} \times 15 \text{ g} \times 100 \text{ °C} = 6300 \text{ J}$$

$$6300 \text{ J} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 6.3 \text{ kJ}$$

Energy to vaporize the water at 100 °C:

$$0.83 \text{ mol H}_2\text{O} \times 40.6 \frac{\text{kJ}}{\text{mol H}_2\text{O}} = 34 \text{ kJ}$$

Total energy required:

$$5.0 \text{ kJ} + 6.3 \text{ kJ} + 34 \text{ kJ} = 45 \text{ kJ}$$

Practice Problem Exercise 14.4

- Contains SO_3 molecules—a molecular solid.
- Contains Ba^{2+} and O^{2-} ions—an ionic solid.
- Contains Au atoms—an atomic solid.

Chapter 15

Practice Problem Exercise 15.1

$$\text{Mass percent} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

For this sample, the mass of solution is 135 g and the mass of the solute is 4.73 g, so

$$\begin{aligned} \text{Mass percent} &= \frac{4.73 \text{ g solute}}{135 \text{ g solution}} \times 100\% \\ &= 3.50\% \end{aligned}$$

Practice Problem Exercise 15.2

Using the definition of mass percent, we have

$$\begin{aligned} \frac{\text{Mass of solute}}{\text{Mass of solution}} &= \\ \frac{\text{grams of solute}}{\text{grams of solute} + \text{grams of solvent}} \times 100\% &= 40.0\% \end{aligned}$$

There are 425 g of solute (formaldehyde). Substituting, we have

$$\frac{425 \text{ g}}{425 \text{ g} + \text{grams of solvent}} \times 100\% = 40.0\%$$

We must now solve for grams of solvent (water). This will take some patience, but we can do it if we proceed step by step. First we divide both sides by 100%.

$$\frac{425 \text{ g}}{425 \text{ g} + \text{grams of solvent}} \times \frac{100\%}{100\%} = \frac{40.0\%}{100\%} = 0.400$$

Now we have

$$\frac{425 \text{ g}}{425 \text{ g} + \text{grams of solvent}} = 0.400$$

Next we multiply both sides by (425 g + grams of solvent).

$$\begin{aligned} (425 \text{ g} + \text{grams of solvent}) \times \frac{425 \text{ g}}{425 \text{ g} + \text{grams of solvent}} \\ = 0.400 \times (425 \text{ g} + \text{grams of solvent}) \end{aligned}$$

This gives

$$425 \text{ g} = 0.400 \times (425 \text{ g} + \text{grams of solvent})$$

Carrying out the multiplication gives

$$425 \text{ g} = 170. \text{ g} + 0.400 (\text{grams of solvent})$$

Now we subtract 170. g from both sides,

$$425 \text{ g} - 170. \text{ g} = \cancel{170. \text{ g}} - \cancel{170. \text{ g}} + 0.400 (\text{grams of solvent})$$

$$255 \text{ g} = 0.400 (\text{grams of solvent})$$

and divide both sides by 0.400.

$$\frac{255 \text{ g}}{0.400} = \frac{0.400}{0.400} (\text{grams of solvent})$$

We finally have the answer:

$$\begin{aligned} \frac{255 \text{ g}}{0.400} &= 638 \text{ g} = \text{grams of solvent} \\ &= \text{mass of water needed} \end{aligned}$$

Practice Problem Exercise 15.4

The moles of ethanol can be obtained from its molar mass (46.1).

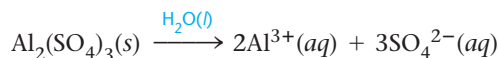
$$\begin{aligned} 1.00 \text{ g C}_2\text{H}_5\text{OH} \times \frac{1 \text{ mol C}_2\text{H}_5\text{OH}}{46.1 \text{ g C}_2\text{H}_5\text{OH}} \\ = 2.17 \times 10^{-2} \text{ mol C}_2\text{H}_5\text{OH} \end{aligned}$$

$$\text{Volume in liters} = 101 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 0.101 \text{ L}$$

$$\begin{aligned} \text{Molarity of C}_2\text{H}_5\text{OH} &= \frac{\text{moles of C}_2\text{H}_5\text{OH}}{\text{liters of solution}} \\ &= 2.17 \times \frac{10^{-2} \text{ mol}}{0.101 \text{ L}} = 0.215 \text{ M} \end{aligned}$$

Practice Problem Exercise 15.5

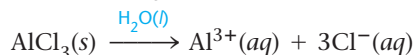
For a. and b.: When Na_2CO_3 and $\text{Al}_2(\text{SO}_4)_3$ dissolve in water, they produce ions as follows:



Therefore, in a 0.10 M Na_2CO_3 solution, the concentration of Na^+ ions is $2 \times 0.10 \text{ M} = 0.20 \text{ M}$ and the concentration of CO_3^{2-} ions is 0.10 M. In a 0.010 M $\text{Al}_2(\text{SO}_4)_3$ solution, the concentration of Al^{3+} ions is $2 \times 0.010 \text{ M} = 0.020 \text{ M}$ and the concentration of SO_4^{2-} ions is $3 \times 0.010 \text{ M} = 0.030 \text{ M}$.

Practice Problem Exercise 15.6

When solid AlCl_3 dissolves, it produces ions as follows:



Thus a $1.0 \times 10^{-3} \text{ M}$ AlCl_3 solution contains $1.0 \times 10^{-3} \text{ M}$ Al^{3+} ions and $3.0 \times 10^{-3} \text{ M}$ Cl^- ions.

To calculate the moles of Cl^- ions in 1.75 L of the $1.0 \times 10^{-3} \text{ M}$ AlCl_3 solution, we must multiply the volume by the molarity.

$$\begin{aligned} 1.75 \text{ L solution} \times 3.0 \times 10^{-3} \text{ M Cl}^- \\ = 1.75 \text{ L solution} \times \frac{3.0 \times 10^{-3} \text{ mol Cl}^-}{\text{L solution}} \\ = 5.25 \times 10^{-3} \text{ mol Cl}^- = 5.3 \times 10^{-3} \text{ mol Cl}^- \end{aligned}$$

Practice Problem Exercise 15.7

We must first determine the number of moles of formaldehyde in 2.5 L of 12.3 M formalin. Remember that volume of solution (in liters) times molarity gives moles of solute.

In this case, the volume of solution is 2.5 L and the molarity is 12.3 mol of HCHO per liter of solution.

$$2.5 \text{ L solution} \times \frac{12.3 \text{ mol HCHO}}{1 \text{ L solution}} = 31 \text{ mol HCHO}$$

Next, using the molar mass of HCHO (30.0 g), we convert 31 mol of HCHO to grams.

$$31 \text{ mol HCHO} \times \frac{30.0 \text{ g HCHO}}{1 \text{ mol HCHO}} = 9.3 \times 10^2 \text{ g HCHO}$$

Therefore, 2.5 L of 12.3 M formalin contains 9.3×10^2 g of formaldehyde. We must weigh out 930 g of formaldehyde and dissolve it in enough water to make 2.5 L of solution.

Practice Problem Exercise 15.8

We are given the following information:

$$M_1 = 12 \frac{\text{mol}}{\text{L}} \qquad M_2 = 0.25 \frac{\text{mol}}{\text{L}}$$

$$V_1 = ? \text{ (what we need to find)} \qquad V_2 = 0.75 \text{ L}$$

Using the fact that the moles of solute do not change upon dilution, we know that

$$M_1 \times V_1 = M_2 \times V_2$$

Solving for V_1 by dividing both sides by M_1 gives

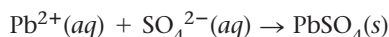
$$V_1 = \frac{M_2 \times V_2}{M_1} = \frac{0.25 \frac{\text{mol}}{\text{L}} \times 0.75 \text{ L}}{12 \frac{\text{mol}}{\text{L}}}$$

and

$$V_1 = 0.016 \text{ L} = 16 \text{ mL}$$

Practice Problem Exercise 15.10

Step 1 When the aqueous solutions of Na_2SO_4 (containing Na^+ and SO_4^{2-} ions) and $\text{Pb}(\text{NO}_3)_2$ (containing Pb^{2+} and NO_3^- ions) are mixed, solid PbSO_4 is formed.



Step 2 We must first determine whether Pb^{2+} or SO_4^{2-} is the limiting reactant by calculating the moles of Pb^{2+} and SO_4^{2-} ions present. Because 0.0500 M $\text{Pb}(\text{NO}_3)_2$ contains 0.0500 M Pb^{2+} ions, we can calculate the moles of Pb^{2+} ions in 1.25 L of this solution as follows:

$$1.25 \cancel{\text{L}} \times \frac{0.0500 \text{ mol Pb}^{2+}}{\cancel{\text{L}}} = 0.0625 \text{ mol Pb}^{2+}$$

The 0.0250 M Na_2SO_4 solution contains 0.0250 M SO_4^{2-} ions, and the number of moles of SO_4^{2-} ions in 2.00 L of this solution is

$$2.00 \cancel{\text{L}} \times \frac{0.0250 \text{ mol SO}_4^{2-}}{\cancel{\text{L}}} = 0.0500 \text{ mol SO}_4^{2-}$$

Step 3 Pb^{2+} and SO_4^{2-} react in a 1:1 ratio, so the amount of SO_4^{2-} ions is limiting because a smaller number of SO_4^{2-} moles is present.

Step 4 The Pb^{2+} ions are present in excess, and only 0.0500 mol of solid PbSO_4 will be formed.

Step 5 We calculate the mass of PbSO_4 by using the molar mass of PbSO_4 (303.3 g).

$$0.0500 \text{ mol PbSO}_4 \times \frac{303.3 \text{ g PbSO}_4}{1 \text{ mol PbSO}_4} = 15.2 \text{ g PbSO}_4$$

Practice Problem Exercise 15.11

Step 1 Because nitric acid is a strong acid, the nitric acid solution contains H^+ and NO_3^- ions. The KOH solution contains K^+ and OH^- ions. When these solutions are mixed, the H^+ and OH^- react to form water.



Step 2 The number of moles of OH^- present in 125 mL of 0.050 M KOH is

$$125 \cancel{\text{ mL}} \times \frac{1 \cancel{\text{ L}}}{1000 \cancel{\text{ mL}}} \times \frac{0.050 \text{ mol OH}^-}{\cancel{\text{ L}}} = 6.3 \times 10^{-3} \text{ mol OH}^-$$

Step 3 H^+ and OH^- react in a 1:1 ratio, so we need 6.3×10^{-3} mole of H^+ from the 0.100 M HNO_3 .

Step 4 6.3×10^{-3} mole of OH^- requires 6.3×10^{-3} mole of H^+ to form 6.3×10^{-3} mole of H_2O .

Therefore,

$$V \times \frac{0.100 \text{ mol H}^+}{\text{L}} = 6.3 \times 10^{-3} \text{ mol H}^+$$

where V represents the volume in liters of 0.100 M HNO_3 required. Solving for V , we have

$$V = \frac{6.3 \times 10^{-3} \text{ mol H}^+}{0.100 \text{ mol H}^+/\text{L}} = 6.3 \times 10^{-2} \text{ L} = 6.3 \times 10^{-2} \cancel{\text{ L}} \times \frac{1000 \text{ mL}}{\cancel{\text{ L}}} = 63 \text{ mL}$$

Practice Problem Exercise 15.13

From the definition of normality, $N = \text{equiv/L}$, we need to calculate (1) the equivalents of KOH and (2) the volume of the solution in liters. To find the number of equivalents, we use the equivalent weight of KOH, which is 56.1 g (see Table 15.2).

$$23.6 \text{ g KOH} \times \frac{1 \text{ equiv KOH}}{56.1 \text{ g KOH}} = 0.421 \text{ equiv KOH}$$

Next we convert the volume to liters.

$$755 \cancel{\text{ mL}} \times \frac{1 \text{ L}}{1000 \cancel{\text{ mL}}} = 0.755 \text{ L}$$

Finally, we substitute these values into the equation that defines normality.

$$\text{Normality} = \frac{\text{equiv}}{\text{L}} = \frac{0.421 \text{ equiv}}{0.755 \text{ L}} = 0.558 \text{ N}$$

Practice Problem Exercise 15.14

To solve this problem, we use the relationship

$$N_{\text{acid}} \times V_{\text{acid}} = N_{\text{base}} \times V_{\text{base}}$$

where

$$N_{\text{acid}} = 0.50 \frac{\text{equiv}}{\text{L}}$$

$$V_{\text{acid}} = ?$$

$$N_{\text{base}} = 0.80 \frac{\text{equiv}}{\text{L}}$$

$$V_{\text{base}} = 0.250 \text{ L}$$

We solve the equation

$$N_{\text{acid}} \times V_{\text{acid}} = N_{\text{base}} \times V_{\text{base}}$$

for V_{acid} by dividing both sides by N_{acid} .

$$\frac{N_{\text{acid}} \times V_{\text{acid}}}{N_{\text{acid}}} = \frac{N_{\text{base}} \times V_{\text{base}}}{N_{\text{acid}}}$$

$$V_{\text{acid}} = \frac{N_{\text{base}} \times V_{\text{base}}}{N_{\text{acid}}} = \frac{(0.80 \frac{\text{equiv}}{\text{L}}) \times (0.250 \text{ L})}{0.50 \frac{\text{equiv}}{\text{L}}}$$

$$V_{\text{acid}} = 0.40 \text{ L}$$

Therefore, 0.40 L of 0.50 N H_2SO_4 is required to neutralize 0.250 L of 0.80 N KOH.

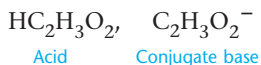
Chapter 16

Practice Problem Exercise 16.2

The conjugate acid–base pairs are



and



The members of both pairs differ by one H^+ .

Practice Problem Exercise 16.3

Because $[\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$, we can solve for $[\text{H}^+]$.

$$[\text{H}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-2}} = 5.0 \times 10^{-13} \text{ M}$$

This solution is basic: $[\text{OH}^-] = 2.0 \times 10^{-2} \text{ M}$ is greater than $[\text{H}^+] = 5.0 \times 10^{-13} \text{ M}$.

Practice Problem Exercise 16.5

a. Because $[\text{H}^+] = 1.0 \times 10^{-3} \text{ M}$, we get $\text{pH} = 3.00$ by using the regular steps.

b. Because $[\text{OH}^-] = 5.0 \times 10^{-5} \text{ M}$, we can find $[\text{H}^+]$ from the K_w expression.

$$[\text{H}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{5.0 \times 10^{-5}} = 2.0 \times 10^{-10} \text{ M}$$

Then we follow the regular steps to get pH from $[\text{H}^+]$.

1. Enter (-).

2. Push \log .

3. Enter 2.0×10^{-10} .

$$\text{pH} = 9.70$$

Practice Problem Exercise 16.7

$$\text{pOH} + \text{pH} = 14.00$$

$$\text{pOH} = 14.00 - \text{pH} = 14.00 - 3.5$$

$$\text{pOH} = 10.5$$

Practice Problem Exercise 16.8

$$\text{pH} = 3.50$$

$$-\text{pH} = -3.50$$

$$\text{2nd } \log -3.50 = 3.2 \times 10^{-4}$$

$$[\text{H}^+] = 3.2 \times 10^{-4} \text{ M}$$

Practice Problem Exercise 16.9

$$\text{pOH} = 10.50$$

$$-\text{pOH} = -10.50$$

$$\text{2nd } \log -10.50 = 3.2 \times 10^{-11}$$

$$[\text{OH}^-] = 3.2 \times 10^{-11} \text{ M}$$

Practice Problem Exercise 16.10

Because HCl is a strong acid, it is completely dissociated:

$$5.0 \times 10^{-3} \text{ M HCl} \rightarrow$$

$$5.0 \times 10^{-3} \text{ M H}^+ \text{ and } 5.0 \times 10^{-3} \text{ M Cl}^-$$

so $[\text{H}^+] = 5.0 \times 10^{-3} \text{ M}$.

$$\text{pH} = -\log(5.0 \times 10^{-3}) = 2.30$$

Practice Problem Exercise 16.11

The reaction is



75.00 mL of 0.1500 M KOH contains

$$\text{moles OH}^- = 75.00 \text{ mL} \times \frac{1.000 \text{ L}}{1.000 \times 10^3 \text{ mL}} \times$$

$$0.1500 \frac{\text{mol OH}^-}{\text{L}} = 1.125 \times 10^{-2} \text{ mol OH}^-$$

To neutralize the KOH solution, we need an equal number of moles of H^+ from the 0.3000 M HCl solution.

$$1.125 \times 10^{-2} \text{ mol H}^+ = V \times 0.3000 \frac{\text{mol H}^+}{\text{L}}$$

$$V = \frac{1.125 \times 10^{-2} \text{ mol H}^+}{0.3000 \frac{\text{mol H}^+}{\text{L}}} = 3.750 \times 10^{-2} \text{ L}$$

$$3.750 \times 10^{-2} \text{ L} \times \frac{1.000 \times 10^3 \text{ mL}}{1.000 \text{ L}} = 37.50 \text{ mL}$$

The volume of 0.3000 M HCl required is 37.50 mL.

Chapter 17

Practice Problem Exercise 17.1

Applying the law of chemical equilibrium gives

$$K = \frac{[\text{NO}_2]^4 [\text{H}_2\text{O}]^6}{[\text{NH}_3]^4 [\text{O}_2]^7}$$

Coefficient of NO_2
Coefficient of H_2O
Coefficient of NH_3
Coefficient of O_2

Practice Problem Exercise 17.3

a. $K = [\text{O}_2]^3$ The solids are not included.

b. $K = [\text{N}_2\text{O}][\text{H}_2\text{O}]^2$ The solid is not included. Water is gaseous in this reaction, so it is included.

c. $K = \frac{1}{[\text{CO}_2]}$ The solids are not included.

d. $K = \frac{1}{[\text{SO}_3]}$ Water and H_2SO_4 are pure liquids and so are not included.

Practice Problem Exercise 17.4

When rain is imminent, the concentration of water vapor in the air increases. This shifts the equilibrium to the right, forming $\text{CoCl}_2 \cdot 6\text{H}_2\text{O}(s)$, which is pink.

Practice Problem Exercise 17.5

- No change. Both sides of the equation contain the same number of gaseous components. The system cannot change its pressure by shifting its equilibrium position.
- Shifts to the left. The system can increase the number of gaseous components present, and so increase the pressure, by shifting to the left.
- Shifts to the right to increase the number of gaseous components and thus its pressure.

Practice Problem Exercise 17.6

- Shifts to the right away from added SO_2 .
- Shifts to the right to replace removed SO_3 .
- Shifts to the right to decrease its pressure.
- Shifts to the right. Energy is a product in this case, so a decrease in temperature favors the forward reaction (which produces energy).

Practice Problem Exercise 17.8

- $\text{BaSO}_4(s) \rightleftharpoons \text{Ba}^{2+}(aq) + \text{SO}_4^{2-}(aq); K_{\text{sp}} = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$
- $\text{Fe}(\text{OH})_3(s) \rightleftharpoons \text{Fe}^{3+}(aq) + 3\text{OH}^-(aq); K_{\text{sp}} = [\text{Fe}^{3+}][\text{OH}^-]^3$
- $\text{Ag}_3\text{PO}_4(s) \rightleftharpoons 3\text{Ag}^+(aq) + \text{PO}_4^{3-}(aq); K_{\text{sp}} = [\text{Ag}^+]^3[\text{PO}_4^{3-}]$

Practice Problem Exercise 17.9

$$[\text{Ba}^{2+}] = [\text{SO}_4^{2-}] = 3.9 \times 10^{-5} \text{ M}$$

$$K_{\text{sp}} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = (3.9 \times 10^{-5})^2 = 1.5 \times 10^{-9}$$

Practice Problem Exercise 17.10

$$\text{PbCrO}_4(s) \rightleftharpoons \text{Pb}^{2+}(aq) + \text{CrO}_4^{2-}(aq)$$

$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{CrO}_4^{2-}] = 2.0 \times 10^{-16}$$

$$[\text{Pb}^{2+}] = x$$

$$[\text{CrO}_4^{2-}] = x$$

$$2.0 \times 10^{-16} = x^2 = K_{\text{sp}}$$

$$x = 1.4 \times 10^{-8} \text{ M} = \text{Solubility of PbCrO}_4(s)$$

Chapter 18

Practice Problem Exercise 18.1

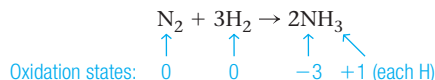
- CuO contains Cu^{2+} and O^{2-} ions, so copper is oxidized ($\text{Cu} \rightarrow \text{Cu}^{2+} + 2e^-$) and oxygen is reduced ($\text{O} + 2e^- \rightarrow \text{O}^{2-}$).
- CsF contains Cs^+ and F^- ions. Thus cesium is oxidized ($\text{Cs} \rightarrow \text{Cs}^+ + e^-$) and fluorine is reduced ($\text{F} + e^- \rightarrow \text{F}^-$).

Practice Problem Exercise 18.2

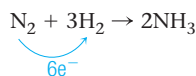
- SO_3
We assign oxygen first. Each O is assigned an oxidation state of -2 , giving a total of -6 (3×-2) for the three oxygen atoms. Because the molecule has zero charge overall, the sulfur must have an oxidation state of $+6$.
Check: $+6 + 3(-2) = 0$
- SO_4^{2-}
As in part a, each oxygen is assigned an oxidation state of -2 , giving a total of -8 (4×-2) on the four oxygen atoms. The anion has a net charge of -2 , so the sulfur must have an oxidation state of $+6$.
Check: $+6 + 4(-2) = -2$
 SO_4^{2-} has a charge of -2 , so this is correct.
- N_2O_5
We assign oxygen before nitrogen because oxygen is more electronegative. Thus each O is assigned an oxidation state of -2 , giving a total of -10 (5×-2) on the five oxygen atoms. Therefore, the oxidation states of the *two* nitrogen atoms must total $+10$, because N_2O_5 has no overall charge. Each N is assigned an oxidation state of $+5$.
Check: $2(+5) + 5(-2) = 0$
- PF_3
First we assign the fluorine an oxidation state of -1 , giving a total of -3 (3×-1) on the three fluorine atoms. Thus P must have an oxidation state of $+3$.
Check: $+3 + 3(-1) = 0$
- C_2H_6
In this case it is best to recognize that hydrogen is always $+1$ in compounds with nonmetals. Thus each H is assigned an oxidation state of $+1$, which means that the six H atoms account for a total of $+6$ ($6 \times +1$). Therefore, the *two* carbon atoms must account for -6 , and each carbon is assigned an oxidation state of -3 .
Check: $2(-3) + 6(+1) = 0$

Practice Problem Exercise 18.4

We can tell whether this is an oxidation–reduction reaction by comparing the oxidation states of the elements in the reactants and products:

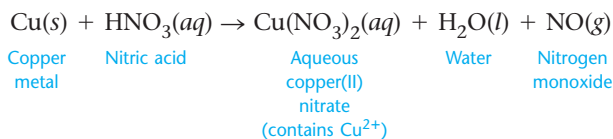


Nitrogen goes from 0 to -3 . Thus it gains three electrons and is reduced. Each hydrogen atom goes from 0 to $+1$ and is thus oxidized, so this is an oxidation–reduction reaction. The oxidizing agent is N_2 (it takes electrons from H_2). The reducing agent is H_2 (it gives electrons to N_2).

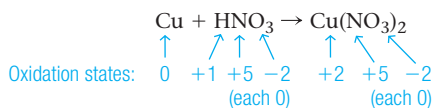


Practice Problem Exercise 18.6

The unbalanced equation for this reaction is

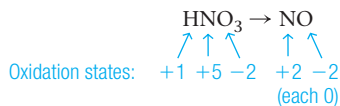


Step 1 The oxidation half-reaction is



The copper goes from 0 to +2 and thus is oxidized.

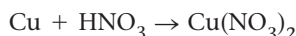
The reduction reaction is



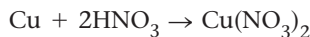
In this case nitrogen goes from +5 in HNO_3 to +2 in NO and so is reduced. Notice two things about these reactions:

1. The HNO_3 must be included in the oxidation half-reaction to supply NO_3^- in the product $\text{Cu}(\text{NO}_3)_2$.
2. Although water is a product in the overall reaction, it does not need to be included in either half-reaction at the beginning. It will appear later as we balance the equation.

Step 2 Balance the oxidation half-reaction.

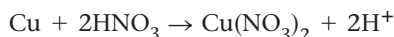


a. Balance nitrogen first.

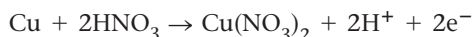


b. Balancing nitrogen also caused oxygen to balance.

c. Balance hydrogen using H^+ .

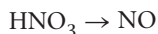


d. Balance the charge using e^- .



This is the balanced oxidation half-reaction.

Balance the reduction half-reaction.

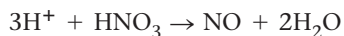


a. All elements are balanced except hydrogen and oxygen.

b. Balance oxygen using H_2O .



c. Balance hydrogen using H^+ .



d. Balance the charge using e^- .

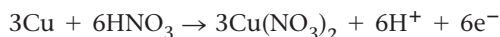


This is the balanced reduction half-reaction.

Step 3 We equalize electrons by multiplying the oxidation half-reaction by 3:



which gives



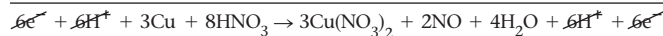
Multiplying the reduction half-reaction by 2



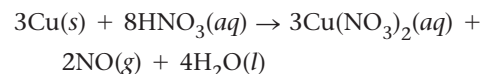
gives



Step 4 We can now add the balanced half-reactions, which both involve a six-electron change.



Canceling species common to both sides gives the balanced overall equation:



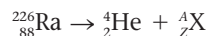
Step 5 Check the elements and charges.



Chapter 19

Practice Problem Exercise 19.1

a. An alpha particle is a helium nucleus, ${}^4_2\text{He}$. We can initially represent the production of an alpha particle by ${}^{226}_{88}\text{Ra}$ as follows:



Because we know that both A and Z are conserved, we can write

$$A + 4 = 226 \text{ and } Z + 2 = 88$$

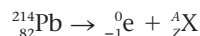
Solving for A gives 222 and for Z gives 86, so ${}^A_Z\text{X}$ is ${}^{222}_{86}\text{X}$. Because Rn has $Z = 86$, ${}^A_Z\text{X}$ is ${}^{222}_{86}\text{Rn}$. The overall balanced equation is



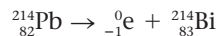
Check: $Z = 88 \quad Z = 86 + 2 = 88$

$$\rightarrow \quad A = 226 \quad A = 222 + 4 = 226$$

b. Using a similar strategy, we have



Because $Z - 1 = 82$, $Z = 83$, and because $A + 0 = 214$, $A = 214$. Therefore, ${}^A_Z\text{X} = {}^{214}_{83}\text{Bi}$. The balanced equation is



Check: $Z = 82 \quad Z = 83 - 1 = 82$

$$\rightarrow \quad A = 214 \quad A = 214 + 0 = 214$$

Practice Problem Exercise 19.2

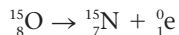
a. The missing particle must be ${}^4_2\text{He}$ (an α particle), because



is a balanced equation.

$$\begin{array}{l} \text{Check: } Z = 86 \quad Z = 84 + 2 = 86 \\ \quad \quad \quad \quad \quad \quad \rightarrow \\ A = 222 \quad A = 218 + 4 = 222 \end{array}$$

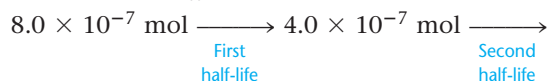
b. The missing species must be ^{15}X or $^{15}_7\text{N}$, because the balanced equation is



$$\begin{array}{l} \text{Check: } Z = 8 \quad Z = 7 + 1 = 8 \\ \quad \quad \quad \quad \quad \quad \rightarrow \\ A = 15 \quad A = 15 + 0 = 15 \end{array}$$

Practice Problem Exercise 19.3

Let's do this problem by thinking about the number of half-lives required to go from 8.0×10^{-7} mol to 1.0×10^{-7} mol of $^{228}_{88}\text{Ra}$.



It takes three half-lives, then, for the sample to go from 8.0×10^{-7} mol of $^{228}_{88}\text{Ra}$ to 1.0×10^{-7} mol of $^{228}_{88}\text{Ra}$. From Table 19.3, we know that the half-life of $^{228}_{88}\text{Ra}$ is 6.7 years. Therefore, the elapsed time is $3(6.7 \text{ years}) = 20.1$ years, or 2.0×10^1 years when we use the correct number of significant figures.

Chapter 20

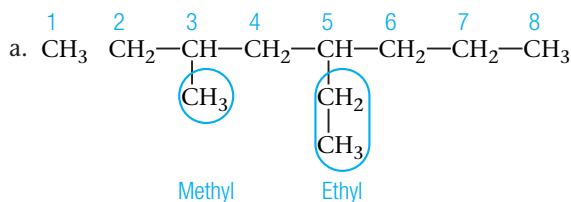
Practice Problem Exercise 20.1

The alkane with ten carbon atoms can be represented as $\text{CH}_3-(\text{CH}_2)_8-\text{CH}_3$ and its formula is $\text{C}_{10}\text{H}_{22}$. The alkane with fifteen carbons,

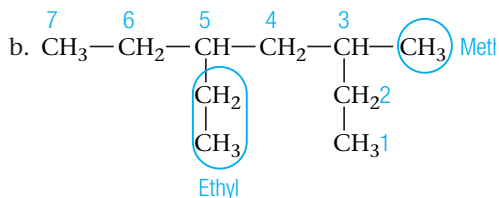


has the formula $\text{C}_{15}\text{H}_{32}$.

Practice Problem Exercise 20.4



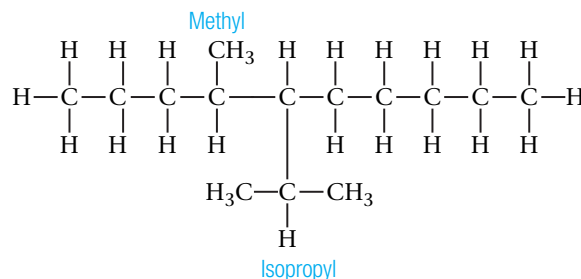
This molecule is 5-ethyl-3-methyloctane.



This molecule is 5-ethyl-3-methylheptane. Note that this chain could be numbered from the opposite direction to give the name 3-ethyl-5-methylheptane. These two names are equally correct.

Practice Problem Exercise 20.5

The root name decane indicates a ten-carbon chain. There is a methyl group at the number 4 position and an isopropyl group at the number 5 position. The structural formula is



Practice Problem Exercise 20.6

- The longest chain has eight carbon atoms with a double bond, so the root name is octene. The double bond exists between carbons 3 and 4, so the name is 3-octene. There is a methyl group on the number-2 carbon. The name is 2-methyl-3-octene.
- The carbon chain has five carbons with a triple bond between carbons 1 and 2. The name is 1-pentyne.

Practice Problem Exercise 20.7

- 2-chloronitrobenzene or o-chloronitrobenzene
- 4-phenyl-2-hexene

Practice Problem Exercise 20.8

- 1-pentanol; primary alcohol
- 2-methyl-2-propanol (but this alcohol is usually called tertiary butyl alcohol); tertiary alcohol
- 5-bromo-2-hexanol; secondary alcohol

Practice Problem Exercise 20.9

- 4-ethyl-3-hexanone
Because the compound is named as a hexanone, the carbonyl group is assigned the lowest possible number.
- 7-isopropyldecanal

Selected End-of-Chapter Problems

Chapter 1

- (a) cellular processes, drug mechanisms; (b) understanding forensic reports, examination of the crime scene; (c) drug interactions, dosages; (d) properties of paints and thinners, properties of sculpture media; (e) film-developing process, properties of papers and coatings; (f) fertilizers, plant hormones; (g) drug interactions, dosages
- Many examples exist. Chemical and biological weapons are produced in some countries. Although the development of plastics has been a benefit in many endeavors, their use depletes fossil fuel reserves and increases our solid waste problems. Although biotechnology has produced many exciting new drugs and treatments, the testing procedures that have been developed for determining whether a person has a genetic likelihood of developing a particular disease might be used to make it difficult or impossible for that person to obtain health or life insurance.
- Answers will depend on the student's choices.
- Answers will depend on student's experience.
- David and Susan first recognized the problem (unexplained medical problems). A possible explanation was then proposed (the glaze on their china might be causing the lead poisoning). The explanation was tested by experiment (it was determined that the china did contain lead). A full scenario is given in the text.
- (a) quantitative—a number (measurement) is indicated explicitly; (c) quantitative—a numerical measurement is indicated; (e) quantitative—a number (measurement) is implied
- A natural law is a summary of observed, measurable behavior that occurs repeatedly and consistently. A theory is an attempt to explain such behavior.
- Most applications of chemistry are oriented toward the interpretation of observations that solve problems. Although memorization of facts may *aid* in these endeavors, it is the ability to combine, relate, and synthesize information that is most important in the study of chemistry.
- In real-life situations, the problems and applications likely to be encountered are not simple textbook examples. You must be able to observe an event, hypothesize a cause, and then test this hypothesis. You must be able to carry forward what has been learned in class to new, different situations.

Chapter 2

- Atoms are very tiny.
- In general, the properties of a compound are very different from the properties of its constituent elements.

For example, the properties of water are very different from the properties of the elements (hydrogen gas and oxygen gas) that make up water.

- Because gases are mostly open space, they can be compressed easily to smaller volumes. In solids and liquids, most of the sample's bulk volume is filled with the molecules, leaving little empty space.
- chemical
- the orange color
- (a) chemical; (c) chemical; (e) chemical; (g) chemical; (i) physical
- (a) mixture; (c) mixture
- (a) homogeneous; (c) heterogeneous
- Consider a mixture of table salt (sodium chloride) and sand. Salt is soluble in water, sand is not. The mixture is added to water and stirred to dissolve the salt, and then filtered. The salt solution passes through the filter, but the sand remains on the filter. The water can then be evaporated from the salt.
- The solution is heated to vaporize (boil) the water. The water vapor is then cooled so that it condenses back to the liquid state, and the liquid is collected. After all the water is vaporized from the original sample, pure sodium chloride remains. The process consists of physical changes.
- compound
- physical
- chemical

Chapter 3

- Alchemists discovered several previously unknown elements (mercury, sulfur, antimony) and were the first to prepare several common acids.
- over 115 elements are known; 88 occur naturally; the others are man-made. Table 3.1 lists the most common elements on earth.
- The four most abundant elements in living creatures are oxygen, carbon, hydrogen, and nitrogen. In the nonliving world, the most abundant elements are oxygen, silicon, aluminum, and iron.
- Sb (antimony), Cu (copper), Au (gold), Pb (lead), Hg (mercury), K (potassium), Ag (silver), Na (sodium), Sn (tin), W (tungsten), Fe (iron)
- (a) Ne; (c) K; (e) Ba
- (a) copper; (c) calcium; (e) chromium; (g) chlorine
- According to Dalton, a given compound always consists of the same number and type of atoms; thus the composition of the compound on a mass percentage basis is always the same.
- According to Dalton, all atoms of the same element are identical; in particular, every atom of a given element has the same mass as every other atom of that

atoms bonded together; as a whole, it carries a negative electrical charge. An oxyanion is a negative ion containing a particular element and one or more oxygen atoms.

- 32.** (a) gold(III) bromide; (c) magnesium hydrogen phosphate; (e) ammonia
- 33.** (a) ammonium carbonate; (c) calcium phosphate; (e) manganese(IV) oxide
- 34.** (a) K_2O ; (c) FeO ; (e) ZnO
- 37.** $Ca(NO_3)_2$ $CaSO_4$ $Ca(HSO_4)_2$ $Ca(H_2PO_4)_2$ CaO $CaCl_2$
 $Sr(NO_3)_2$ $SrSO_4$ $Sr(HSO_4)_2$ $Sr(H_2PO_4)_2$ SrO $SrCl_2$
 NH_4NO_3 $(NH_4)_2SO_4$ NH_4HSO_4 $NH_4H_2PO_4$ $(NH_4)_2O$ NH_4Cl
 $Al(NO_3)_3$ $Al_2(SO_4)_3$ $Al(HSO_4)_3$ $Al(H_2PO_4)_3$ Al_2O_3 $AlCl_3$
 $Fe(NO_3)_3$ $Fe_2(SO_4)_3$ $Fe(HSO_4)_3$ $Fe(H_2PO_4)_3$ Fe_2O_3 $FeCl_3$
 $Ni(NO_3)_2$ $NiSO_4$ $Ni(HSO_4)_2$ $Ni(H_2PO_4)_2$ NiO $NiCl_2$
 $AgNO_3$ Ag_2SO_4 $AgHSO_4$ AgH_2PO_4 Ag_2O $AgCl$
 $Au(NO_3)_3$ $Au_2(SO_4)_3$ $Au(HSO_4)_3$ $Au(H_2PO_4)_3$ Au_2O_3 $AuCl_3$
 KNO_3 K_2SO_4 $KHSO_4$ KH_2PO_4 K_2O KCl
 $Hg(NO_3)_2$ $HgSO_4$ $Hg(HSO_4)_2$ $Hg(H_2PO_4)_2$ HgO $HgCl_2$
 $Ba(NO_3)_2$ $BaSO_4$ $Ba(HSO_4)_2$ $Ba(H_2PO_4)_2$ BaO $BaCl_2$
- 39.**(a) $Al(13e^-) \rightarrow Al^{3+}(10e^-) + 3e^-$; (c) $Cu(29e^-) \rightarrow Cu^+(28e^-) + 1e^-$; (e) $Zn(30e^-) \rightarrow Zn^{2+}(28e^-) + 2e^-$
- 40.** (a) none likely; (c) Te^{2-} ; (e) Br^-
- 41.** (a) Na_2S ; (c) BaO ; (e) $CuBr_2$; (g) Al_2O_3
- 42.**(a) beryllium oxide; (c) sodium sulfide; (e) hydrogen chloride (gaseous); hydrochloric acid (aqueous); (g) silver(I) sulfide; usually called silver sulfide
- 43.** (a) incorrect; (c) incorrect
- 44.** (a) iron(II) bromide; (c) cobalt(III) sulfide; (e) mercury(I) chloride
- 45.** (a) xenon hexafluoride; (c) arsenic triiodide; (e) dichlorine monoxide
- 46.** (a) iron(III) acetate; (c) potassium peroxide; (e) copper(II) permanganate
- 47.** (a) lithium dihydrogen phosphate; (c) lead(II) nitrate; (e) sodium chlorite
- 48.** (a) $CaCl_2$; (c) Al_2S_3 ; (e) H_2S ; (g) MgI_2
- 49.** (a) SO_2 ; (c) XeF_4 ; (e) PCl_5
- 50.** (a) NaH_2PO_4 ; (c) $Cu(HCO_3)_2$; (e) BaO_2
- 51.** (a) $AgClO_4$; (c) $NaClO$ or $NaOCl$; (e) NH_4NO_2 ; (g) NH_4HCO_3

Chapter 5

- 2.** 4512
- 4.** (a) -5; (b) 6; (c) -4; (d) 4
- 5.** (a) 9.367421×10^6 ; (c) 5.519×10^{-4} ; (f) 6.319×10^1 ; (h) 7.21×10^{-1}
- 6.** (a) 483; (c) 6.1; (e) 4,221,000; (g) 9999; (i) 101,600; (k) 97,100

- 7.** (a) 1.423×10^5 ; (c) 2.27×10^4 ; (e) 2.51×10^2 ; (g) 9.7752 (9.7752×10^0)
- 8.** (a) 3.1×10^3 ; (c) 1 or 1×10^0 ; (e) 1×10^7 ; (g) 1.00×10^{-7}
- 9.** (a) 10^3 ; (c) 10^{-3} ; (e) 10^{-9}
- 10.** (a) mega; (c) nano; (e) centi
- 11.** 100 miles **12.** quart
- 14.** the woman **16.** d
- 18.** d **20.** 40 quarters
- 23.** The scale of the ruler is marked to the nearest tenth of a centimeter. Writing 2.850 would imply that the scale was marked to the nearest hundredth of a centimeter (and that the zero in the thousandths place had been estimated).
- 24.** (a) 4; (c) 4; (e) 3; (g) 6
- 25.** (a) probably 2; (c) infinite (definition)
- 26.** (a) 1,570,000 or 1.57×10^6 ; (c) 84,600 or 8.46×10^4
- 27.** (a) 3.42×10^{-4} ; (c) 1.7992×10^1
- 29.** 3 **31.** none
- 32.** (a) 641.0; (c) 77.34
- 33.** (a) 124; (c) 1.14×10^{-2}
- 34.** (a) 2.045; (c) 5.19×10^{-5}
- 38.** 1 lb/\$0.79
- 39.** (a) 2.44 yd; (c) 115 in; (e) 648.1 mi; (g) 0.0362 km
- 40.** (a) 0.2543 kg; (c) 6.06 lb; (e) 1.177 lb; (g) 2.5×10^2 g
- 42.** 3.1×10^2 km; 3.1×10^5 m; 1.0×10^6 ft
- 47.** (a) 118 K; (c) 221 K; (e) 226 K
- 48.** (a) 2 °C; (c) -273 °C; (e) 9727 °C
- 49.** (a) -40 °F; (c) 241 °C
- 50.** (a) 144 K; (c) 664 °F
- 54.** Density is a characteristic property of a pure substance.
- 56.** (a) 20.1 g/cm³; (c) 0.907 g/cm³
- 57.** 0.843 g/mL **59.** float
- 60.** 11.7 mL
- 62.** (a) 966 g; (b) 394 g; (c) 567 g; (d) 135 g
- 63.** (a) cm; (c) km
- 65.** \$1 **68.** 959 g
- 69.** (a) 4; positive; (c) 0; zero
- 71.** (a) 1; (c) 4; (e) infinite (definition)
- 72.** 38.2 g/mL

Chapter 6

- 2.** 307 corks; 116 stoppers; 2640 g
- 4.** The average atomic mass is a weighted average including a contribution based on the mass of the individual isotopes of an element and their abundance in nature.
- 5.** (a) 640. amu; (c) 1642 amu

6. (a) one; (c) 100
 9. Avogadro's number (6.022×10^{23})
 11. 177 g 13. 0.50 mol O
 14. (a) 0.133 mol Au; (c) 2.44×10^{-3} mol Ba;
 (e) 5.20×10^{-13} mol Ni
 15. (a) 112 g Fe; (c) 0.240 g Pt; (e) 0.0248 g Mg
 16. (a) 1.05×10^{19} atoms; (c) 0.0467 mol cobalt; (e) 249 g
 18. adding (summing)
 19. (a) 82.98 g; (c) 97.94 g; (e) 82.09 g
 20. (a) 336.2 g; (c) 278.1 g; (e) 260.5 g
 21. (a) 6.14×10^{-4} mol SO_3 ; (c) 0.495 mol CHCl_3 ;
 (e) 0.167 mol LiOH
 22. (a) 3.55×10^{-5} mol; (c) 2704 mol
 23. (a) 612 g AlI_3 ; (c) 721 g $\text{C}_6\text{H}_{12}\text{O}_6$
 24. (a) 0.0559 g CO_2 ; (c) 0.361 g NH_4NO_3
 25. (a) 3.84×10^{24} molecules; (c) 8.76×10^{16} molecules
 26. (a) 0.0141 mol S; (c) 0.0258 mol S
 27. (a) 32.37% Na, 22.58% S, 45.05% O; (c) 58.91% Na,
 41.09% S; (e) 55.26% K, 14.59% P, 30.15% O;
 (g) 28.73% K, 1.481% H, 22.76% P, 47.03% O
 28. (a) 28.45% Cu; (c) 44.06% Fe; (e) 18.84% Co;
 (g) 88.12% Sn
 29. (a) 49.32% C; (c) 49.47% C; (e) 71.92% C; (g) 85.03% C
 30. (a) 33.73% NH_4^+ ; (c) 64.94% Au^{3+}
 32. The empirical formula represents the smallest whole-
 number ratio of the elements present in a compound. The
 molecular formula indicates the actual number of atoms
 of each element found in a molecule of the substance.
 33. (a) NaO; (c) $\text{C}_{12}\text{H}_{12}\text{N}_2\text{O}_3$
 34. (a) these have the same empirical formulas; (c) these
 have the same empirical formulas
 36. CaO 37. BaSO_4
 39. Co_2S_3 41. CuO
 43. Na_3N ; NaN_3 45. molar mass
 48. $\text{C}_6\text{H}_{24}\text{O}_6$
 50. empirical formula, $\text{C}_3\text{H}_3\text{O}$; molecular formula, $\text{C}_6\text{H}_6\text{O}_2$
 52. 5.00 g Al, 0.185 mol, 1.12×10^{23} atoms; 0.140 g Fe,
 0.00250 mol, 1.51×10^{21} atoms; 270 g Cu, 4.3 mol,
 2.6×10^{24} atoms; 0.00250 g Mg, 1.03×10^{-4} mol,
 6.19×10^{19} atoms; 0.062 g Na, 2.7×10^{-3} mol,
 1.6×10^{21} atoms; 3.95×10^{-18} g U, 1.66×10^{-20} mol,
 1.00×10^4 atoms
 55. Cu_2O , CuO 57. 2.12 g Fe
 60. $\text{C}_3\text{H}_7\text{N}_2\text{O}$ 61. HgO
 62. BaCl_2

Chapter 7

2. decrease in mass; change in texture
 4. change in odor as acetic acid is produced

6. atoms
 8. water
 10. $\text{C}_3\text{H}_8(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$
 12. $(\text{NH}_4)_2\text{CO}_3(\text{s}) \rightarrow \text{NH}_3(\text{g}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$
 14. $\text{CO}(\text{g}) + \text{H}_2(\text{g}) \rightarrow \text{CH}_3\text{OH}(\text{l})$
 16. $\text{Ca}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{Ca}(\text{OH})_2(\text{s}) + \text{H}_2(\text{g})$
 18. $\text{Mg}(\text{OH})_2(\text{s}) + \text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{l})$
 20. $\text{H}_2\text{S}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$
 22. $\text{Fe}_2\text{O}_3(\text{s}) + \text{CO}(\text{g}) \rightarrow \text{Fe}(\text{l}) + \text{CO}_2(\text{g})$
 24. $\text{O}_2(\text{g}) \rightarrow \text{O}_3(\text{g})$
 26. $\text{NH}_3(\text{g}) + \text{HNO}_3(\text{g}) \rightarrow \text{NH}_4\text{NO}_3(\text{aq})$
 28. $\text{Xe}(\text{g}) + \text{F}_2(\text{g}) \rightarrow \text{XeF}_4(\text{g})$
 30. $\text{Ag}(\text{s}) + \text{HNO}_3(\text{aq}) \rightarrow \text{AgNO}_3(\text{aq}) + \text{H}_2(\text{g})$
 32. whole numbers
 33. (a) $2\text{H}_2\text{O}_2(\text{aq}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$;
 (c) $2\text{FeO}(\text{s}) + \text{C}(\text{s}) \rightarrow 2\text{Fe}(\text{l}) + \text{CO}_2(\text{g})$
 34. (a) $\text{CaF}_2(\text{s}) + \text{H}_2\text{SO}_4(\text{l}) \rightarrow \text{CaSO}_4(\text{s}) + 2\text{HF}(\text{g})$;
 (c) $\text{TiCl}_4(\text{l}) + 4\text{Na}(\text{s}) \rightarrow 4\text{NaCl}(\text{s}) + \text{Ti}(\text{s})$
 35. (a) $\text{SiI}_4(\text{s}) + 2\text{Mg}(\text{s}) \rightarrow \text{Si}(\text{s}) + 2\text{MgI}_2(\text{s})$;
 (c) $8\text{Ba}(\text{s}) + \text{S}_8(\text{s}) \rightarrow 8\text{BaS}(\text{s})$
 36. (a) $\text{Ba}(\text{NO}_3)_2(\text{aq}) + \text{Na}_2\text{CrO}_4(\text{aq}) \rightarrow$
 $\text{BaCrO}_4(\text{s}) + 2\text{NaNO}_3(\text{aq})$;
 (b) $\text{C}_2\text{H}_5\text{OH}(\text{l}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 3\text{H}_2\text{O}(\text{l})$
 37. $\text{Al}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{Al}_2\text{O}_3(\text{s})$
 39. $\text{C}_{12}\text{H}_{22}\text{O}_{11}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow 4\text{C}_2\text{H}_5\text{OH}(\text{l}) + 4\text{CO}_2(\text{g})$
 41. $2\text{Al}_2\text{O}_3(\text{s}) + 3\text{C}(\text{s}) \rightarrow 4\text{Al}(\text{s}) + 3\text{CO}_2(\text{g})$
 43. $2\text{Li}(\text{s}) + \text{S}(\text{s}) \rightarrow \text{Li}_2\text{S}(\text{s})$; $2\text{Na}(\text{s}) + \text{S}(\text{s}) \rightarrow \text{Na}_2\text{S}(\text{s})$;
 $2\text{K}(\text{s}) + \text{S}(\text{s}) \rightarrow \text{K}_2\text{S}(\text{s})$; $2\text{Rb}(\text{s}) + \text{S}(\text{s}) \rightarrow \text{Rb}_2\text{S}(\text{s})$;
 $2\text{Cs}(\text{s}) + \text{S}(\text{s}) \rightarrow \text{Cs}_2\text{S}(\text{s})$; $2\text{Fr}(\text{s}) + \text{S}(\text{s}) \rightarrow \text{Fr}_2\text{S}(\text{s})$
 44. $2\text{KClO}_3(\text{s}) \rightarrow 2\text{KCl}(\text{s}) + 3\text{O}_2(\text{g})$
 48. $\text{K}_2\text{CrO}_4(\text{aq}) + \text{BaCl}_2(\text{aq}) \rightarrow \text{BaCrO}_4(\text{s}) + 2\text{KCl}(\text{aq})$

Chapter 8

2. Driving forces are types of changes in a system that
 pull a reaction in the direction of product formation;
 they include formation of a solid, formation of water,
 formation of a gas, and transfer of electrons.
 3. The net charge of a precipitate must be zero. The total
 number of positive charges equals the total number
 of negative charges.
 7. "Insoluble" and "slightly soluble" have roughly the
 same meanings. However, if a substance is highly toxic
 and found in a water supply, for example, the differ-
 ence could be crucial.
 8. (a) soluble (most nitrate salts are soluble); (c) soluble
 (most sodium salts are soluble); (e) insoluble; (g) insol-
 ible (most sulfide salts are only slightly soluble)
 9. (a) rule 5: most hydroxide compounds are only slightly
 soluble; (c) rule 6: most phosphate salts are only
 slightly soluble

10. (a) calcium sulfate; (c) lead(II) phosphate; (e) no precipitate: all potassium and sodium salts are soluble
11. (a) no precipitate; (c) $\text{H}_2\text{SO}_4(aq) + \text{Pb}(\text{NO}_3)_2(aq) \rightarrow \text{PbSO}_4(s) + 2\text{HNO}_3(aq)$; (e) no precipitate
12. (a) $2\text{AgNO}_3(aq) + \text{H}_2\text{SO}_4(aq) \rightarrow \text{Ag}_2\text{SO}_4(s) + 2\text{HNO}_3(aq)$
13. (a) $(\text{NH}_4)_2\text{S}(aq) + \text{CoCl}_2(aq) \rightarrow \text{CoS}(s) + 2\text{NH}_4\text{Cl}(aq)$
14. (a) $\text{Ca}^{2+}(aq) + \text{SO}_4^{2-}(aq) \rightarrow \text{CaSO}_4(s)$; (c) $\text{Ag}^+(aq) + \text{I}^-(aq) \rightarrow \text{AgI}(s)$; (e) $\text{Hg}_2^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{Hg}_2\text{Cl}_2(s)$
15. $\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s)$; $\text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{PbCl}_2(s)$; $\text{Hg}_2^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{Hg}_2\text{Cl}_2(s)$
18. The strong bases are those hydroxide compounds that dissociate fully when dissolved in water. The strong bases that are highly soluble in water (NaOH, KOH) are also strong electrolytes.
19. 1000; 1000
21. $\text{RbOH}(s) \rightarrow \text{Rb}^+(aq) + \text{OH}^-(aq)$; $\text{CsOH}(s) \rightarrow \text{Cs}^+(aq) + \text{OH}^-(aq)$
22. (a) $\text{HCl}(aq) + \text{RbOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{RbCl}(aq)$;
(c) $\text{HBr}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaBr}(aq)$
23. (a) $\text{H}_2\text{SO}_4(aq) + 2\text{NaOH}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{Na}_2\text{SO}_4(aq)$;
(c) $\text{HClO}_4(aq) + \text{KOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{KClO}_4(aq)$
25. the metal loses electrons; the nonmetal gains electrons
26. Each potassium atom would lose one electron to become a K^+ ion. Each sulfur atom would gain two electrons to become a S^{2-} ion. Two potassium ions would have to react for every one sulfur atom.
27. AlBr_3 is made up of Al^{3+} ions and Br^- ions. Aluminum atoms each lose 3 electrons and bromine atoms each gain 1 electron (Br_2 gains 2 electrons).
29. (a) $2\text{Fe}(s) + 3\text{S}(s) \rightarrow \text{Fe}_2\text{S}_3(s)$; (c) $2\text{Sn}(s) + \text{O}_2(g) \rightarrow 2\text{SnO}(s)$; (e) $2\text{Cs}(s) + 2\text{H}_2\text{O}(l) \rightarrow 2\text{CsOH}(s) + \text{H}_2(g)$
31. Examples of the formation of water: $\text{H}_2\text{SO}_4(aq) + 2\text{NaOH}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{Na}_2\text{SO}_4(aq)$; $\text{HClO}_4(aq) + \text{KOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{KClO}_4(aq)$. Examples of formation of a gaseous product: $\text{Mg}(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + \text{H}_2(g)$; $2\text{KClO}_3(s) \rightarrow 2\text{KCl}(s) + 3\text{O}_2(g)$
33. (a) oxidation–reduction; (b) oxidation–reduction; (c) acid–base; (d) acid–base; (e) precipitation; (f) precipitation; (g) oxidation–reduction; (h) oxidation–reduction; (i) acid–base
35. A decomposition reaction is one in which a given compound is broken down into simpler compounds or constituent elements. These may be classified in other ways. For example, the reaction $2\text{HgO}(s) \rightarrow 2\text{Hg}(l) + \text{O}_2(g)$ is both a decomposition reaction, and an oxidation–reduction reaction.
36. (a) $\text{C}_2\text{H}_5\text{OH}(l) + 3\text{O}_2(g) \rightarrow 2\text{CO}_2(g) + 3\text{H}_2\text{O}(g)$
37. (a) $2\text{C}_2\text{H}_6(g) + 7\text{O}_2(g) \rightarrow 4\text{CO}_2(g) + 6\text{H}_2\text{O}(g)$
38. (a) $2\text{Co}(s) + 3\text{S}(s) \rightarrow \text{Co}_2\text{S}_3(s)$; (c) $\text{FeO}(s) + \text{CO}_2(g) \rightarrow \text{FeCO}_3(s)$
39. (a) $2\text{NI}_3(s) \rightarrow \text{N}_2(g) + 3\text{I}_2(s)$; (c) $\text{C}_6\text{H}_{12}\text{O}_6(s) \rightarrow 6\text{C}(s) + 6\text{H}_2\text{O}(g)$
41. (a) silver ion: $\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s)$; lead(II) ion: $\text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{PbCl}_2(s)$; mercury(I) ion: $\text{Hg}_2^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{Hg}_2\text{Cl}_2(s)$; (c) hydroxide ion: $\text{Fe}^{3+}(aq) + 3\text{OH}^-(aq) \rightarrow \text{Fe}(\text{OH})_3(s)$; sulfide ion: $2\text{Fe}^{3+}(aq) + 3\text{S}^{2-}(aq) \rightarrow \text{Fe}_2\text{S}_3(s)$; phosphate ion: $\text{Fe}^{3+}(aq) + \text{PO}_4^{3-}(aq) \rightarrow \text{FePO}_4(s)$; (e) sulfide ion: $\text{Hg}_2^{2+}(aq) + \text{S}^{2-}(aq) \rightarrow \text{Hg}_2\text{S}(s)$; carbonate ion: $\text{Hg}_2^{2+}(aq) + 2\text{CO}_3^{2-}(aq) \rightarrow \text{Hg}_2\text{CO}_3(s)$; chloride ion: $\text{Hg}_2^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{Hg}_2\text{Cl}_2(s)$
43. (a) $\text{HNO}_3(aq) + \text{KOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{KNO}_3(aq)$;
(c) $\text{HClO}_4(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaClO}_4(aq)$
45. (a) $\text{AgNO}_3(aq) + \text{HCl}(aq) \rightarrow \text{AgCl}(s) + \text{HNO}_3(aq)$;
(c) $\text{FeSO}_4(aq) + \text{K}_2\text{CO}_3(aq) \rightarrow \text{FeCO}_3(s) + \text{K}_2\text{SO}_4(aq)$;
(e) $\text{Pb}(\text{NO}_3)_2(aq) + \text{Li}_2\text{CO}_3(aq) \rightarrow \text{PbCO}_3(s) + 2\text{LiNO}_3(aq)$
46. (a) $\text{Ag}^+(aq) + \text{Cl}^-(aq) \rightarrow \text{AgCl}(s)$;
(c) $\text{Pb}^{2+}(aq) + 2\text{Cl}^-(aq) \rightarrow \text{PbCl}_2(s)$
48. (a) 1; (c) 2; (e) 3
49. (a) 2; (c) 3
50. (a) $2\text{C}_3\text{H}_8\text{OH}(l) + 9\text{O}_2(g) \rightarrow 6\text{CO}_2(g) + 8\text{H}_2\text{O}(g)$;
oxidation–reduction, combustion;
(c) $3\text{HCl}(aq) + \text{Al}(\text{OH})_3(s) \rightarrow \text{AlCl}_3(aq) + 3\text{H}_2\text{O}(l)$;
acid–base, double-displacement

Chapter 9

2. Balanced chemical equations tell us in what proportions on a mole basis substances combine; since the molar masses of $\text{C}(s)$ and $\text{O}_2(g)$ are different, 1 g of each reactant cannot represent the same number of moles of each reactant.
4. (a) $3\text{MnO}_2(s) + 4\text{Al}(s) \rightarrow 3\text{Mn}(s) + 2\text{Al}_2\text{O}_3(s)$. Three formula units of manganese(IV) oxide react with four aluminum atoms, producing three manganese atoms and two formula units of aluminum oxide. Three moles of solid manganese(IV) oxide react with four moles of solid aluminum, producing three moles of solid manganese and two moles of solid aluminum oxide. (c) $3\text{NO}_2(g) + \text{H}_2\text{O}(l) \rightarrow 2\text{HNO}_3(aq) + \text{NO}(g)$. Three molecules of nitrogen dioxide react with one molecule of water to produce two molecules of nitric acid and one molecule of nitrogen monoxide. Three moles of gaseous nitrogen dioxide react with one mole of liquid water to produce two moles of aqueous nitric acid and one mole of nitrogen monoxide gas.
7. $2\text{Ag}(s) + \text{H}_2\text{S}(g) \rightarrow \text{Ag}_2\text{S}(s) + \text{H}_2(g)$. For Ag_2S : 1 mol $\text{Ag}_2\text{S}/2$ mol Ag; for H_2 : 1 mol $\text{H}_2/2$ mol Ag
8. (a) 0.125 mol Fe, 0.0625 mol CO_2 ; (c) 0.500 mol H_3BO_3 , 0.125 mol Na_2SO_4
9. (a) 0.50 mol NH_4Cl (27 g); (c) 0.50 mol H_3PO_3 (41 g), 1.5 mol HCl (55 g)
10. (a) 0.469 mol O_2 ; (c) 0.625 mol CH_3CHO
12. Stoichiometry is the process of using a chemical equation to calculate the relative masses of reactants and products involved in a reaction.

59. (a) 41 kJ; (c) 47 kJ
 60. (a) exothermic; (c) endothermic

Chapter 11

- This experiment is explained in Chapter 3.
- The different forms of electromagnetic radiation all exhibit the same wavelike behavior and move through space at the same speed (the speed of light). The types of electromagnetic radiation differ in their frequency and wavelength and in the resulting amount of energy carried per photon.
- The speed of electromagnetic radiation represents how fast energy is transferred through space; the frequency of the radiation tells us how many waves pass a given point in a certain amount of time.
- The wave-particle nature of light refers to the fact that a beam of electromagnetic energy can be considered not only as a continuous wave, but also as a stream of discrete packets of energy moving through space.
- The colors are due to the releasing of energy.
- It is emitted as a photon.
- When excited hydrogen atoms emit their excess energy, the photons of radiation that are emitted always have the exact same wavelength and energy. Thus the hydrogen atom possesses only certain allowed energy states.
- The ground state of an atom is its lowest possible energy state.
- According to Bohr, electrons move in fixed circular orbits around the nucleus. An electron can move to a larger orbit if a photon of applied energy equals the difference in energy between the two orbits.
- Bohr's theory explained the line spectrum of hydrogen exactly. It did not explain measurements made for atoms other than hydrogen.
- An orbit refers to an exact circular pathway for the electron around the nucleus. An orbital represents a region in space in which there is a high probability of finding an electron.
- Pictures representing orbitals are probability maps. They do not imply that the electron moves on the surface of, or within, the region of the picture. The mathematical probability of finding an electron never becomes zero moving out from the nucleus (the electron could possibly be "anywhere").
- The 2s orbital is similar in shape to the 1s orbital, but the 2s orbital is larger.
- increases
- The Pauli exclusion principle states that an orbital can hold a maximum of two electrons, and those two electrons must have opposite spin.
- increases
- Valence electrons are those in the outermost energy level. They can interact with the (valence) electrons of another atom in a chemical reaction.

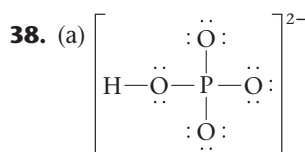
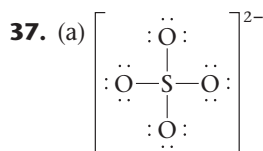
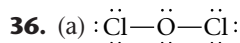
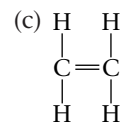
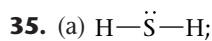
- (a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$; (c) $1s^2$
- (a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$; (c) $1s^2 2s^2 2p^5$
- (a) $(\uparrow\downarrow) (\uparrow\downarrow) (\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) (\uparrow\downarrow) (\uparrow) (\uparrow) (\uparrow) (\uparrow)$;
 $1s \quad 2s \quad 2p \quad 3s \quad 3p$
 (c) $(\uparrow\downarrow) (\uparrow\downarrow) (\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) (\uparrow\downarrow) (\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)$;
 $1s \quad 2s \quad 2p \quad 3s \quad 3p \quad 4s$
 $(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) (\uparrow\downarrow)(\uparrow\downarrow)(\uparrow)$
 $3d \quad 4p$
- (a) 1; (c) 7
- The properties of Rb and Sr suggest that they are members of Groups 1 and 2, respectively. This means that the outer electrons are filling the 5s orbital. The 5s orbital, then, must be lower in energy and fills before the 4d orbitals.
- (a) $[\text{Ar}]4s^2$; (c) $[\text{Kr}]5s^2 4d^1$
- (a) $[\text{Ne}]3s^2 3p^3$; (c) $[\text{Ne}]3s^2$
- (a) 1; (c) 0
- (a) 5f; (c) 4f
- (a) $[\text{Rn}]7s^2 6d^1 5f^3$; (c) $[\text{Xe}]6s^2 4f^{14} 5d^{10}$
- The metallic elements lose electrons and form positive ions (cations); the nonmetallic elements gain electrons and form negative ions (anions).
- The elements of a given period (horizontal row) have valence electrons in the same subshells. Nuclear charge increases across a period from left to right. Atoms at the left side have smaller nuclear charges and bind their valence electrons less tightly.
- The nuclear charge increases from left to right within a period. The greater the nuclear charge, the more tightly the valence electrons are "pulled."
- (a) Na; (c) Br
- (a) Li; (c) Cl
- (a) Xe < Sn < Sr < Rb
- speed of light
- orbital
- (a) $(\uparrow\downarrow) (\uparrow\downarrow) (\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) (\uparrow\downarrow) (\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) (\uparrow)$;
 $1s \quad 2s \quad 2p \quad 3s \quad 3p \quad 4s$
 $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$; $[\text{Ar}]4s^1$;
 (c) $(\uparrow\downarrow) (\uparrow\downarrow) (\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow)(\uparrow\downarrow) (\uparrow)(\uparrow)(\uparrow)$;
 $1s \quad 2s \quad 2p \quad 3s \quad 3p$
 $1s^2 2s^2 2p^6 3s^2 3p^2$; $[\text{Ne}]3s^2 3p^2$
- (a) ns^2 ; (c) $ns^2 np^4$
- (a) 2 (d electrons are not counted as valence); (c) 2
- Light is emitted from the hydrogen atom only at certain wavelengths. If the energy levels of hydrogen were continuous, a hydrogen atom would emit energy at all possible wavelengths.
- (a) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$;
 (c) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$
- (a) Ca

Chapter 12

- ionic; covalent
- By bonding with each other, the H and F atoms share

a pair of valence electrons. Because the atoms have different attractions for the electrons, this sharing is not equal. The bonding in HF is polar covalent.

7. A bond is polar if the shared electron pair is attracted more strongly by one of the bonded atoms than the other. The polarity is due to differences in electronegativity values.
8. (a) $K < Ca < Sc$
9. (a) covalent; (c) polar covalent
10. (a) no polar covalent bonds; (c) no polar covalent bonds
11. (a) $H-O$; (c) $H-F$
12. (a) $Na-O$; (c) $K-Cl$
14. The presence of strong dipoles and a large overall dipole moment makes water a very polar substance. Properties of water that are dependent on its dipole moment include its freezing point, melting point, vapor pressure, and ability to dissolve many substances.
15. (a) H
16. (a) $\delta^+P \rightarrow F^{\delta-}$; (c) $\delta^+P \rightarrow C^{\delta-}$
17. (a) $\delta^+P \rightarrow S^{\delta-}$; (c) $\delta^+S \rightarrow N^{\delta-}$
18. gaining
20. (a) Li: $[He]2s^1$; Li^+ : $[He]$; (c) Cs: $[Xe]6s^1$; Cs^+ : $[Xe]$
21. (a) Ca^{2+} ; (c) Br^-
22. (a) Na_2S ; (c) $MgBr_2$; (e) KH
23. (a) Al^{3+} , $[Ne]$; S^{2-} , $[Ar]$; (c) Rb^+ , $[Kr]$; O^{2-} , $[Ne]$
25. An ionic solid such as NaCl consists of an array of alternating positively and negatively charged ions. In most ionic solids, the ions are packed as tightly as possible.
27. In forming an anion, an atom gains additional electrons in its outermost (valence) shell. Having additional electrons in the valence shell increases the repulsive forces between electrons, and the outermost shell becomes larger to accommodate this repulsion.
28. (a) F^- is larger than Li^+ . The F^- ion has a filled $n = 2$ shell. A lithium atom has lost the electron from the $n = 2$ shell, leaving $n = 1$ as its outermost shell. (c) Ca is larger than Ca^{2+} . The calcium ion has $n = 3$ as its outer shell, and the calcium atom has $n = 4$ as its outer shell.
29. (a) I^- ; (c) Cl^-
31. When atoms form covalent bonds, they try to attain a valence-electron configuration similar to that of the nearest noble gas element. Thus hydrogen in a bond would have two electrons (like helium, the duet rule) and others would have eight electrons (like neon and argon, the octet rule).
33. (a) 32; (c) 30
34. (a) $H-\ddot{N}-H$; (c) $\begin{array}{c} \ddot{Cl}-\ddot{N}-\ddot{Cl} \\ | \\ \ddot{Cl} \end{array}$



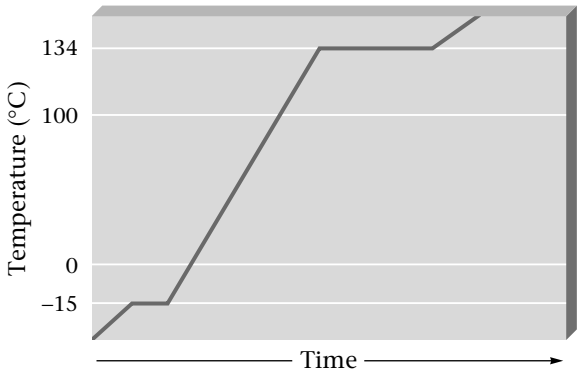
40. Trigonal pyramid structure with four electron pairs (three bonding pairs and one lone pair). The bond angle is somewhat less than 109.5° (due to the lone pair).
42. Tetrahedral structure with four bonding electron pairs. The bond angle is 109.5° .
44. The general molecular structure is determined by the number of electron pairs around the central atom (bonding and lone pairs).
46. In NF_3 the nitrogen is surrounded by four electron pairs (tetrahedral geometry); in BF_3 the boron is surrounded by three electron pairs (trigonal planar).
47. (a) four electron pairs (tetrahedral geometry)
48. (a) tetrahedral
49. (a) $<109.5^\circ$ (about 105°); (c) 109.5°
50. (a) tetrahedral
52. The acetylene molecule would be linear, with bond angles of 180° , because of the presence of the triple bond between carbon atoms.
54. (a) $S-F$
55. The bond energy is the energy required to break the bond.
56. (a) Be
57. (a) O
58. (a) Na^+ ; (c) K^+ ; (e) S^{2-} ; (g) Al^{3+}
59. (a) nonlinear (V-shaped or bent); (c) trigonal planar

Chapter 13

1. A solid is rigid and incompressible and has a definite shape and volume. A liquid is less rigid than a solid; although it has a definite volume, a liquid takes the shape of its container. A gas has no fixed volume or shape; it takes the volume and shape of its container and is affected more by changes in its pressure and temperature than is a liquid or a solid.
4. (a) 1.038 atm; (c) 0.989 atm

5. (a) 758.1 mm Hg; (c) 748 mm
 6. (a) 1.03×10^5 Pa; (c) 1.125×10^5 Pa
 7. (a) 59.7 mL 8. (a) 146 mL
 10. 20.0 atm
 12. Measure the volume of a gas at several different temperatures (keep pressure and moles of gas constant). Plot these data and extrapolate the line (continue past the points you have measured). The temperature at which the gas has zero volume is absolute zero.
 13. 49.2 mL 14. (a) 273 °C
 15. (a) -238 °C
 17. The volume of the neon gas would become half the initial value at 149 K. The volume would double at 596 K.
 20. 10.8 L
 21. Real gases behave most ideally at relatively high temperatures and low pressures.
 23. (a) 5.02 L 24. (a) 0.201 L
 26. 4,430 g He; 2,230 g H₂
 27. 655 K 30. 464 mL
 32. The pressure of the water vapor must be subtracted from the total pressure.
 34. 5.0 atm N₂; 3.3 atm O₂; 1.7 atm CO₂
 36. $P_{\text{O}_2} = 732$ torr, 1.98×10^{-2} mol O₂
 38. 14.7 L O₂
 40. 5.03 L (dry volume)
 43. 15.6 L, 0.225 atm O₂, 0.256 atm N₂, 0.163 atm CO₂, 0.356 atm Ne
 45. 0.709 L Cl₂
 46. A theory is successful if it explains known experimental observations; no
 48. We assume that the volume of the gas molecules in a sample is negligible compared to the bulk volume. This helps explain why gases are so compressible.
 49. Pressure is due to atoms or molecules of a gas colliding with the walls of the container.
 51. The temperature is a measure of the average kinetic energy of a sample of atoms or molecules. At high temperatures, the gas particles have, on average, greater kinetic energy and are moving faster than the same gas particles at a lower temperature.
 54. When the volume of a gas is decreased, the gas particles take up a greater percentage of the volume of the container. The assumption in the KMT that gas particles take up a negligible volume is less reliable.
 55. 125 balloons
 58. The volume would be 3.0 L, so the balloon will burst.
 61. 5.8 L O₂; 3.9 L SO₂
 62. 32 L; $P_{\text{He}} = 0.86$ atm; $P_{\text{Ar}} = 0.017$ atm; $P_{\text{Ne}} = 0.12$ atm

Chapter 14

2. Dipole–dipole forces are stronger at shorter distances; they are relatively short-range forces.
 4. Hydrogen bonding can exist when H is bonded to O, N, or F. This additional intermolecular force requires more energy to separate the molecules for boiling.
 6. All atoms and molecules exhibit London dispersion forces between each other.
 7. (a) London dispersion forces; (c) dipole–dipole forces and London dispersion forces
 9. Strong hydrogen bonding occurs in both ethanol and water. Thus the molecules have great attraction for one another and are able to approach each other very closely; in fact, they must approach each other more closely in the mixture than either can approach a like molecule in the separate liquids.
 11. Water, as perspiration, helps cool the human body. Some nuclear power plants use water to cool the reactor core.
 13. Water pipes can be broken during cold weather. Also, ice can float on water, which is crucial for aquatic life (and ultimately for humans).
 15. Sloped portions represent changes in temperature. The flat portions represent phase changes.
 17. These processes are explained in detail in the text.
 19. Intermolecular
 21. 
 22. 8.35 kJ; 84.6 kJ; 23.1 kJ
 25. Vapor pressure is the pressure of vapor present at equilibrium above a liquid in a sealed container at a given temperature. Molecules in the liquid evaporate, but as the number of molecules in the vapor state increases, some of these rejoin (condense). Eventually dynamic equilibrium is reached between evaporation and condensation.
 27. See Figure 14.5 and the corresponding explanation.
 28. (a) H₂S. H₂O exhibits hydrogen bonding and H₂S does not.
 30. Hydrogen bonding can exist in ethyl alcohol but not dimethyl ether. Therefore ethyl alcohol would be less volatile (higher boiling point).

- 31.** At higher altitudes the boiling temperatures of liquids are lower because there is less atmospheric pressure above the liquid. The temperature at which food cooks is determined by the temperature to which water in the food can be heated before it escapes as steam. By boiling at a lower temperature, the food will have to cook longer.
- 34.** Ionic solids (NaCl); molecular solids (ice); atomic solids (copper)
- 36.** Ionic solids tend to be harder and have higher melting and boiling points. Compare table salt (NaCl) to table sugar (sucrose), for example.
- 38.** Ionic solids have strong electrostatic forces (due to the attraction of positively and negatively charged ions). These forces require a great deal of energy to be overcome.
- 42.** The presence of a second metal's atoms in a metal's lattice changes the properties (usually making the alloy stronger because irregularities arise, keeping the crystal from being easily deformed). Steels with relatively high carbon content are exceptionally strong. Steels produced by alloying iron with nickel, chromium, and cobalt are more resistant to corrosion (rusting) than iron itself.
- 53.** Diethyl ether should have a higher vapor pressure because it does not exhibit hydrogen bonding while 1-butanol does.
- 54.** (a) KBr (exhibits ionic bonding)
- 56.** Steel is a general term applied to alloys consisting mostly of iron. Steel is much stronger and harder than iron, which is relatively soft, malleable, and ductile.
- 57.** Dipole–dipole interactions are typically 1% as strong as covalent bonds.
- 58.** London dispersion forces arise from instantaneous dipoles; they are weaker than dipole–dipole forces or covalent bonds.
- 59.** (a) London dispersion forces; (c) London dispersion forces
- 61.** In NH_3 , strong hydrogen bonding can exist. Because CH_4 molecules are nonpolar, only the relatively weak London dispersion forces exist.
- 13.** (a) 0.719 g NaCl; (c) 0.494 g NaCl
- 15.** 19.6% CaCl_2
- 17.** 4.8 g heptane; 2.7 g pentane; 86 g hexane
- 19.** 0.221 mol Ca^{2+} ; 0.442 mol Cl^-
- 21.** (a) 2.0 M; (c) 0.67 M
- 22.** (a) 0.253 M; (c) 0.434 M
- 24.** 0.0902 M
- 25.** 0.623 M Fe^{3+} ; 1.87 M Cl^-
- 27.** (a) 0.0133 mol; 0.838 g; (c) 0.00505 mol; 0.490 g
- 28.** 6.04 g NH_4Cl
- 29.** (a) 4.60×10^{-3} mol Al^{3+} ; 1.38×10^{-2} mol Cl^- ; (c) 2.19×10^{-3} mol Cu^{2+} ; 4.38×10^{-3} mol Cl^-
- 31.** (a) 0.0837 M; (c) 0.0964 M
- 33.** Transfer 6.21 mL of the 3.02 M solution to a 125-mL volumetric flask and add water to the calibration mark.
- 36.** 7.5 mL
- 38.** 0.976 g PbCrO_4
- 40.** 0.107 M NaOH
- 41.** (a) 3.85 mL NaOH; (c) 4.29 mL NaOH
- 43.** 1.53 equivalents OH^- ion. By definition, one equivalent of OH^- ion exactly neutralizes one equivalent of H^+ ion.
- 44.** (a) 0.277 N
- 45.** (a) 0.134 N
- 46.** 0.359 M; 1.08 N
- 47.** 47.4 mL
- 49.** Colligative properties are properties of a solution that depend only on the number, not the identity, of the solute particles.
- 52.** 0.167 M Co^{3+} ; 0.117 M Ni^{2+} ; 0.567 M Cl^-
- 54.** 1.93 g
- 56.** 2.56 M
- 57.** (a) 0.0909 M; (c) 0.192 M
- 58.** (a) 0.50 N $\text{HC}_2\text{H}_3\text{O}_2$

Chapter 16

- 2.** In the Arrhenius definition, an acid is a substance that produces hydrogen ions (H^+) when dissolved in water. An Arrhenius base produces hydroxide ions (OH^-) when dissolved in water. These definitions are not complete because they limit the bases to compounds with a hydroxide ion, and the only allowable solvent is water.
- 4.** A conjugate acid–base pair differs by one hydrogen ion, H^+ . For example, $\text{HC}_2\text{H}_3\text{O}_2$ (acetic acid) differs from its conjugate base $\text{C}_2\text{H}_3\text{O}_2^-$ (acetate ion) by a single hydrogen ion.
- 6.** When an acid is dissolved in water, the hydronium ion (H_3O^+) is formed. The hydronium ion is the conjugate acid of water.
- 7.** (a) The conjugate base of H_2SO_4 is the HSO_4^- ion. The conjugate acid of SO_4^{2-} is also the HSO_4^- ion. (c) The conjugate base of HClO_4 is the ClO_4^- ion. The conjugate acid of Cl^- is HCl.
- 8.** (a) NH_3 (base), NH_4^+ (acid); H_2O (acid), OH^- (base)

Chapter 15

- 1.** A homogeneous mixture is a combination of two (or more) pure substances that is uniform in composition and appearance throughout. Examples of homogeneous mixtures include rubbing alcohol (70% isopropyl alcohol and 30% water) and gasoline (a mixture of hydrocarbons).
- 4.** The intermolecular forces must be similar in the two substances for them to mix. Polar and ionic solids may dissolve in water.
- 6.** unsaturated
- 8.** large
- 9.** Increasing the surface area of a solid increases the amount of solid that comes in contact with the solvent.
- 12.** (a) 5.00% CaCl_2 ; (c) 5.00% CaCl_2

9. (a) H_2SO_4 ; (c) HClO_4
10. (a) CO_3^{2-} ; (c) Cl^-
11. (a) $\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^-$;
(c) $\text{O}^{2-} + \text{H}_2\text{O} \rightarrow \text{OH}^- + \text{OH}^-$
12. (a) $\text{HClO}_4 + \text{H}_2\text{O} \rightarrow \text{ClO}_4^- + \text{H}_3\text{O}^+$;
(c) $\text{HSO}_3^- + \text{H}_2\text{O} \rightarrow \text{SO}_3^{2-} + \text{H}_3\text{O}^+$
14. If an acid is weak in aqueous solution, it does not easily transfer protons to water (and does not fully ionize). Thus the anion of the acid must attract protons rather strongly.
15. A strong acid loses its protons easily and fully ionizes in water. Thus the acid's conjugate base is poor at attracting protons (that is, a very weak base). A weak acid does not ionize fully, so its conjugate base must have a rather strong attraction for a proton (that is, it must be a relatively strong base).
17. (a) $\text{CH}_3\text{COO}^- (\text{C}_2\text{H}_3\text{O}_2^-)$ is a relatively strong base;
(c) HS^- is relatively strong base
18. (a) HSO_4^- is a moderately strong acid; (c) HCN is a weak acid
20. $\text{H}_2\text{O} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$; $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$
22. (a) $[\text{H}^+] = 2.5 \times 10^{-10} \text{ M}$; basic;
(c) $[\text{H}^+] = 1.4 \times 10^{-13} \text{ M}$; basic
23. (a) $[\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$; neutral;
(c) $[\text{OH}^-] = 1.4 \times 10^{-14} \text{ M}$; acidic
24. (a) $[\text{H}^+] = 1.2 \times 10^{-3} \text{ M}$ is more acidic
25. (a) $[\text{H}^+] = 1.04 \times 10^{-8} \text{ M}$
27. household ammonia (pH 12); blood (pH 7–8); milk (pH 6–7); vinegar (pH 3); lemon juice (pH 2–3); stomach acid (pH 2)
29. The pH of a solution is defined as the negative logarithm of the hydrogen ion concentration; that is, $\text{pH} = -\log[\text{H}^+]$. Mathematically, as the $[\text{H}^+]$ increases, pH must decrease.
30. (a) $\text{pH} = 3.000$ (acidic); (c) $\text{pH} = 10.037$ (basic)
31. (a) 7.000 (neutral); (c) 10.017 (basic)
32. (a) $\text{pH} = 9.68$; solution is basic;
(c) $\text{pH} = 12.19$; solution is basic
33. (a) $[\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$; $\text{pH} = 7.00$; $\text{pOH} = 7.00$;
(c) $[\text{OH}^-] = 2.3 \times 10^{-4} \text{ M}$; $\text{pH} = 10.37$; $\text{pOH} = 3.63$
34. (a) 0.091 M ; (c) $1.0 \times 10^{-6} \text{ M}$
35. (a) $[\text{H}^+] = 8.1 \times 10^{-11} \text{ M}$; (c) $[\text{H}^+] = 1.4 \times 10^{-13} \text{ M}$
36. (a) $[\text{H}^+] = 7.6 \times 10^{-6} \text{ M}$; $[\text{OH}^-] = 1.3 \times 10^{-9} \text{ M}$;
(c) $[\text{H}^+] = [\text{OH}^-] = 1.0 \times 10^{-7}$
37. We can measure the pH of a solution with a pH meter, pH paper, and indicators. The pH meter is the most accurate.
40. The solution contains water molecules, H_3O^+ ions (protons), and NO_3^- ions. No HNO_3 molecules are present because HNO_3 is a strong acid.
42. (a) $[\text{H}^+] = 1.0 \times 10^{-4} \text{ M}$; $\text{pH} = 4.00$;
(c) $[\text{H}^+] = 4.21 \times 10^{-5} \text{ M}$; $\text{pH} = 4.376$

43. 500.0 mL
46. A buffered solution consists of a mixture of a weak acid and its conjugate base; one example of a buffered solution is a mixture of acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) and sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$).
48. The weak acid component of the buffered solution reacts with the strong base. For example, hydroxide reacts with the acetic acid as follows:
 $\text{OH}^- + \text{HC}_2\text{H}_3\text{O}_2 \rightarrow \text{H}_2\text{O} + \text{C}_2\text{H}_3\text{O}_2^-$.
49. (a) no; (c) yes
53. (a) basic solution; (c) not a basic solution (acidic)
54. (a) strong acid; (c) strong acid
55. (a) this is a conjugate acid–base pair;
(c) this is a conjugate acid–base pair
56. (a) CH_3NH_2 (base), CH_3NH_3^+ (acid); H_2O (acid), OH^- (base)
58. (a) $[\text{OH}^-] = 3.2 \times 10^{-6} \text{ M}$

Chapter 17

2. A nitrogen–nitrogen triple bond and three hydrogen–hydrogen bonds must be broken. Six nitrogen–hydrogen bonds must form.
4. The activation energy is the minimum energy required by two molecules for their collision to result in a reaction. If this energy requirement is not met, a collision will not result in a reaction.
6. The catalysts found in living cells are called enzymes. They are necessary for life because most of the cell's chemical processes would proceed too slowly at room temperature without a catalyst.
8. For a chemical reaction to occur, the reactants must collide with one another. As the temperature is increased, the average kinetic energy is increased. The particles are moving faster and the collisions will occur more frequently and have higher energies.
11. A state of equilibrium is reached when two opposing processes are exactly balanced. The development of a vapor pressure above a liquid in a closed container is an example of a physical equilibrium. Any chemical reaction that appears to “stop” before completion is an example of chemical equilibrium.
13. A system has reached equilibrium when no more product forms, even though significant amounts of all reactants are present. This indicates that the reverse process is now occurring at the same rate as the forward process. In other words, every time a product molecule is formed, another product molecule reacts to give back the reactants. Reactions that come to equilibrium are indicated by a double arrow.
15. We recognize a state of chemical equilibrium by the fact that the concentrations of reactants and products no longer change with time. However, this does not mean the reaction has “stopped.” The concentrations remain the same because the forward and reverse reactions are occurring at the same rate.

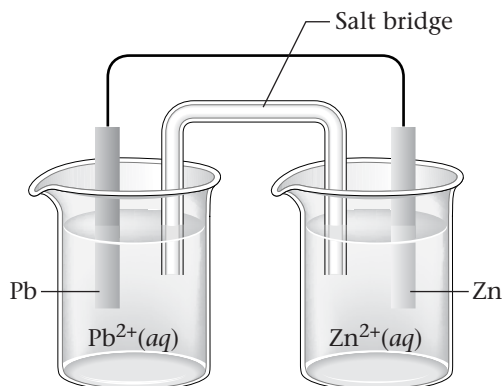
17. Depending on the amounts of reactants that were present originally, differing amounts of products and reactants will be present at equilibrium. However, the ratio of the products to the reactants (which is represented by the equilibrium constant) will always be the same. For example, 4/2 and 6/3 involve different numbers, but each of these ratios is equal to 2.
18. (a) $K = \frac{[\text{NCl}_3(\text{g})]^2}{[\text{N}_2(\text{g})][\text{Cl}_2(\text{g})]^3}$
19. (a) $K = \frac{[\text{CH}_3\text{OH}(\text{g})]}{[\text{CO}(\text{g})][\text{H}_2(\text{g})]^2}$
21. $K = 2.5 \times 10^{10}$
24. Equilibrium constants represent ratios of the concentrations of the products and reactants present at equilibrium. The concentration of a pure liquid or solid is constant and is determined by the density of the solid or liquid.
25. (a) $K = [\text{H}_2\text{O}(\text{g})][\text{CO}_2(\text{g})]$
26. (a) $K = [\text{N}_2(\text{g})][\text{Br}_2(\text{g})]^3$
28. When more of one reactant is added to an equilibrium system, the system shifts to the right and adjusts to use up some of the added reactant. A net increase in the amount of product results, compared with the equilibrium position before the additional reactant was added. The numerical value of the equilibrium constant does not change when a reactant is added. The concentrations of all reactants and products adjust until they satisfy the value of K .
30. If heat is applied (the temperature is raised) to an endothermic reaction, the equilibrium shifts to the right. More product will be present (and less reactant will be present) at equilibrium than if the temperature had not been increased. The value of K increases.
32. (a) no change; (b) shifts left; (c) shifts right; (d) shifts right
34. For an endothermic reaction, an increase in temperature will shift the position of equilibrium to the right (toward the products).
37. A small equilibrium constant means that at equilibrium the concentration of the products is small compared with the concentration of reactants. The position of equilibrium lies far to the left. A reaction with a small equilibrium constant is therefore not useful as a means of producing the products. To become useful, the equilibrium position would have to be shifted to the right.
38. $K = 3.2 \times 10^{11}$ 40. $K = 2.1 \times 10^{-3}$
42. $[\text{N}_2\text{O}_4] = 5.4 \times 10^{-4} \text{ M}$
45. Stirring or grinding the solute increases the speed with which the solute dissolves. However, the amount of solute that dissolves is set by the equilibrium constant for that process.
47. (a) $\text{PbBr}_2(\text{s}) \rightleftharpoons \text{Pb}^{2+}(\text{aq}) + 2\text{Br}^{-}(\text{aq});$
 $K_{\text{sp}} = [\text{Pb}^{2+}(\text{aq})][\text{Br}^{-}(\text{aq})]^2;$
 (c) $\text{PbCO}_3(\text{s}) \rightleftharpoons \text{Pb}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq});$
 $K_{\text{sp}} = [\text{Pb}^{2+}(\text{aq})][\text{CO}_3^{2-}(\text{aq})]$
49. $K_{\text{sp}} = 1.6 \times 10^{-11}$ 51. $1.7 \times 10^{-4} \text{ g/L}$
52. $K_{\text{sp}} = 1.9 \times 10^{-4}; 10. \text{ g/L}$
55. An increase in temperature increases the fraction of molecules with energy greater than the activation energy.
56. To say a reaction is reversible means that the reaction may occur in either direction.
58. In an exothermic process heat can be considered a product of the reaction. Adding heat (by increasing the temperature) opposes the forward process.
59. $9.0 \times 10^{-3} \text{ M}$
61. At higher temperatures, the average kinetic energy of the reactant molecules is larger. Therefore, the probability that the molecules will have enough energy for a reaction to take place is greater. At the molecular level a higher temperature means a given molecule will be moving faster.
63. (a) $K = \frac{[\text{HBr}(\text{g})]^2}{[\text{H}_2(\text{g})][\text{Br}_2(\text{g})]}$
 (b) $K = \frac{[\text{H}_2\text{S}(\text{g})]^2}{[\text{H}_2(\text{g})]^2[\text{S}_2(\text{g})]}$
 (c) $K = \frac{[\text{HCN}(\text{g})]^2}{[\text{H}_2(\text{g})][\text{C}_2\text{N}_2(\text{g})]}$
64. (a) $K = \frac{1}{[\text{O}_2(\text{g})]^3}$ 66. $K_{\text{sp}} = 1.4 \times 10^{-8}$

Chapter 18

2. Reduction can be defined as the gaining of electrons by an atom, molecule, or ion. Reduction may also be defined as a decrease in oxidation state for an element. This decrease occurs because of the gaining of one or more electrons (thus the definitions are the same). The following equation shows the reduction of sulfur: $\text{S}(\text{s}) + 2\text{e}^{-} \rightarrow \text{S}^{2-}(\text{aq})$.
3. (a) sodium is oxidized, nitrogen is reduced;
 (c) aluminum is oxidized, bromine is reduced
4. (a) Mg is oxidized, Br_2 is reduced;
 (c) Cl_2 is reduced, Br is oxidized
5. The assignment of oxidation states is a bookkeeping method by which charges are assigned to the various atoms in a compound. This method allows us to keep track of electrons transferred between species in oxidation–reduction reactions.
7. A neutral molecule overall has a charge of zero.
8. The sum of all oxidation states of the atoms in a polyatomic ion must equal the overall charge on the ion.
9. (a) N, +3; Cl, -1; (c) P, +5; Cl, -1
10. (a) H, +1; Br, -1; (c) Br, 0
11. (a) H, +1; O, -2; N, +5; (c) H, +1; O, -2; S, +6
12. (a) Cu, +2; Cl, -1; (c) H, +1; O, -2; Cr, +6
13. (a) H, +1; C, -4; (c) K, +1; H, +1; C, +4; O, -2
15. When an atom gains an electron (which is negative), it gains a negative charge, which will lower its oxidation state.

tion state. For example, in the reaction $\text{Cl} + \text{e}^- \rightarrow \text{Cl}^-$, the oxidation state goes from 0 to -1 .

- 16.** An oxidizing agent causes another species to be oxidized (to lose electrons). To do so, the oxidizing agent must take in electrons (be reduced). Thus an oxidizing agent is reduced (and a reducing agent is oxidized).
- 18.** (a) Zn is oxidized (0 to $+2$); H is reduced ($+1$ to 0); (c) N is reduced (0 to -3); Br is oxidized (0 to $+1$)
- 19.** (a) copper is oxidized (0 to $+2$); silver is reduced ($+1$ to 0); (c) sulfur is oxidized (0 to $+4$); iron is reduced ($+3$ to 0)
- 20.** Silver is reduced [$+1$ in $\text{AgBr}(s)$, 0 in $\text{Ag}(s)$]; bromine is oxidized [-1 in AgBr , 0 in $\text{Br}_2(g)$]
- 23.** Under ordinary conditions, it is impossible to have “free” electrons that are not part of some atom, ion, or molecule. Thus the total number of electrons lost by the species being oxidized must equal the total number of electrons gained by the species being reduced.
- 24.** (a) $\text{N}_2 + 6\text{e}^- \rightarrow 2\text{N}^{3-}$; (c) $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$
- 25.** (a) $4\text{e}^- + 4\text{H}^+ + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$; (c) $\text{e}^- + 2\text{H}^+ + \text{VO}^{2+} \rightarrow \text{V}^{3+} + 2\text{H}_2\text{O}$
- 26.** (a) $2\text{MnO}_4^-(aq) + 16\text{H}^+(aq) + 5\text{Zn}(s) \rightarrow 2\text{Mn}^{2+}(aq) + 8\text{H}_2\text{O}(l) + 5\text{Zn}^{2+}(aq)$;
(c) $2\text{NO}_3^-(aq) + 4\text{H}^+(aq) + \text{Zn}(s) \rightarrow \text{NO}_2(g) + 2\text{H}_2\text{O}(l) + \text{Zn}^{2+}(aq)$
- 27.** (a) $\text{IO}_3^-(aq) + 6\text{H}^+(aq) + 5\text{I}^-(aq) \rightarrow 3\text{I}_2(aq) + 3\text{H}_2\text{O}(l)$
- 29.** A salt bridge completes the electrical circuit in a cell. It allows ion flow from one half of the cell to another without allowing bulk mixing of the solutions. It is typically a U-shaped tube filled with an electrolyte that is not involved in the oxidation–reduction reaction. Any method that allows ion flow without large-scale mixing of solutions (such as a porous cup or frit) can be used in place of a salt bridge.
- 32.** $\text{Pb}^{2+}(aq)$ ion is reduced; $\text{Zn}(s)$ is oxidized. The reaction at the anode is $\text{Zn}(s) \rightarrow \text{Zn}^{2+}(aq) + 2\text{e}^-$. The reaction at the cathode is $\text{Pb}^{2+}(aq) + 2\text{e}^- \rightarrow \text{Pb}(s)$.



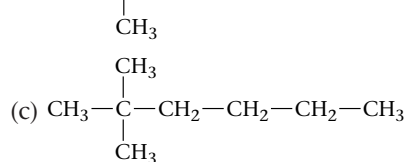
- 34.** Both normal and alkaline cells contain zinc as an electrode; zinc corrodes more slowly under alkaline (basic) conditions than in a normal dry cell (which is acidic).
Anode: $\text{Zn}(s) + 2\text{OH}^-(aq) \rightarrow \text{ZnO}(s) + \text{H}_2\text{O}(l) + 2\text{e}^-$.
Cathode: $2\text{MnO}_2(s) + \text{H}_2\text{O}(l) + 2\text{e}^- \rightarrow \text{Mn}_2\text{O}_3(s) + 2\text{OH}^-(aq)$.
- 36.** Aluminum reacts very quickly with the oxygen in air to form a thin coating of Al_2O_3 on its surface. This coating is transparent and clings tightly to the aluminum. In addition, the aluminum oxide is much less reactive than aluminum and protects the surface from further reaction.
- 38.** In cathodic protection of steel tanks and pipes, a more reactive metal than iron is connected to the item to be protected. The active metal is then oxidized instead of the iron of the tank or pipe.
- 40.** The main recharging reaction for the lead storage battery is $2\text{PbSO}_4(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Pb}(s) + \text{PbO}_2(s) + 2\text{H}_2\text{SO}_4(aq)$. A major side reaction is the electrolysis of water, $2\text{H}_2\text{O}(l) \rightarrow 2\text{H}_2(g) + \text{O}_2(g)$.
- 42.** The item to be plated should be made the cathode in a cell containing a solution of ions of the desired plating metal.
- 44.** lose
- 45.** hydrogen, oxygen
- 46.** (a) $4\text{Fe}(s) + 3\text{O}_2(g) \rightarrow 2\text{Fe}_2\text{O}_3(s)$; iron is oxidized, oxygen is reduced
- 47.** (a) $2\text{MnO}_4^-(aq) + 6\text{H}^+(aq) + 5\text{H}_2\text{O}_2(aq) \rightarrow 2\text{Mn}^{2+}(aq) + 8\text{H}_2\text{O}(l) + 5\text{O}_2(g)$
- 48.** (a) zinc is oxidized, nitrogen is reduced;
(c) potassium is oxidized, oxygen is reduced
- 49.** (a) Mn, $+4$; O, -2 ; (c) H, $+1$; S, $+4$; O, -2
- 50.** (a) oxygen is oxidized, chlorine is reduced;
(c) carbon is oxidized, chlorine is reduced
- 2.** The radius of a typical atomic nucleus is on the order of 10^{-13} cm. This is about 100,000 times smaller than the radius of an atom overall.
- 3.** The atomic number (Z) represents the number of protons in the nucleus. The mass number (A) represents the total number of protons and neutrons in the nucleus. For example, for $^{13}_6\text{C}$, with six protons and seven neutrons, we have $Z = 6$ and $A = 13$.
- 5.** The atomic number (Z) is written as a left subscript, while the mass number (A) is written as a left superscript. The general symbol for a nuclide is ^A_ZX . As an example, the isotope of oxygen with 8 neutrons would be $^{16}_8\text{O}$.
- 7.** When a nucleus produces a beta particle, the atomic number of the parent nucleus is increased by one unit.
- 8.** Gamma rays are high-energy photons of electromagnetic radiation; they are normally not considered to be particles. When a nucleus produces only gamma radiation, the atomic number and mass number remain the same.
- 10.** Electron capture occurs when one of the inner orbital electrons is pulled into, and becomes part of, the nucleus.

12. (a) ${}^{192}_{83}\text{Bi}$
13. (a) ${}^0_{-1}\text{e}$
14. (a) ${}^{136}_{53}\text{I} \rightarrow {}^0_{-1}\text{e} + {}^{136}_{54}\text{Xe}$
15. (a) ${}^{226}_{88}\text{Ra} \rightarrow {}^{222}_{86}\text{Rn} + {}^4_2\text{He}$; (c) ${}^{239}_{94}\text{Pu} \rightarrow {}^{235}_{92}\text{U} + {}^4_2\text{He}$
17. The target nucleus and the bombarding particles often repel each other greatly. This is especially true if the bombarding particles are positively charged. Accelerators are needed to speed up the particles to overcome this repulsion.
19. ${}^{27}_{13}\text{Al} + {}^4_2\text{He} \rightarrow {}^{30}_{15}\text{P} + {}^1_0\text{n}$
21. The half-life for a nucleus is the time required for one-half of the original sample of nuclei to decay. A given isotope of an element always has the same half-life. Different isotopes of the same element may have very different half-lives. Nuclei of different elements have different half-lives.
24. highest to lowest activity: ${}^{87}\text{Sr} > {}^{99}\text{Tc} > {}^{24}\text{Na} > {}^{99}\text{Mo} > {}^{133}\text{Xe} > {}^{131}\text{I} > {}^{32}\text{P} > {}^{51}\text{Cr} > {}^{59}\text{Fe}$
25. For ${}^{223}\text{Ra}$, approximately 125 mg (after 36 days or 3 half-lives); for ${}^{224}\text{Ra}$, approximately 4 mg (after 29 days or 8 half-lives); for ${}^{225}\text{Ra}$, approximately 250 mg (after 30 days, or 2 half-lives).
28. Carbon-14 is produced in the upper atmosphere by the bombardment of nitrogen with neutrons from space: ${}^{14}_7\text{N} + {}^1_0\text{n} \rightarrow {}^{14}_6\text{C} + {}^1_1\text{H}$.
30. When a plant dies, it no longer replenishes itself with carbon-14 from the atmosphere. As the carbon-14 it has undergoes decay, the amount decreases over time. This assumes that the concentration of carbon-14 in the atmosphere is constant over time.
32. These isotopes are listed in Table 19.4.
34. Nuclear fusion is the combining of two light nuclei to form a heavier, more stable nucleus. Splitting a heavy nucleus into nuclei with smaller mass numbers is nuclear fission.
36. ${}^1_0\text{n} + {}^{235}_{92}\text{U} \rightarrow {}^{142}_{56}\text{Ba} + {}^{91}_{36}\text{Kr} + 3 {}^1_0\text{n}$
38. A critical mass of fissionable material is the amount needed to sustain a chain reaction. Enough neutrons have to be produced to cause a constant fission of more material. A sample with less than the critical mass is still radioactive, but it cannot sustain a chain reaction.
40. The type of nuclear explosion produced by a nuclear weapon cannot occur in a nuclear reactor. The concentration of fissionable materials in a reactor is not sufficient to form a supercritical mass.
42. ${}^1_0\text{n} + {}^{238}_{92}\text{U} \rightarrow {}^{239}_{92}\text{U}$; ${}^{239}_{92}\text{U} \rightarrow {}^{239}_{93}\text{Np} + {}^0_{-1}\text{e}$; ${}^{239}_{93}\text{Np} \rightarrow {}^{239}_{94}\text{Pu} + {}^0_{-1}\text{e}$
44. The hydrogen nuclei are positively charged and therefore repel each other. Extremely high temperatures are needed to overcome this repulsion as they are shot into each other.
47. Somatic damage is damage directly to the organism itself, which causes nearly immediate sickness or death. Genetic damage is damage to the genetic machinery of the organism, which causes defects in future generations of offspring.

49. Alpha particles are much heavier than gamma rays so they can ionize biological materials very effectively. Gamma rays can penetrate long distances but seldom cause ionization.
51. The exposure limits given in Table 19.5 as causing no detectable clinical effect are 0–25 rem. The total yearly exposures from natural and human-induced radioactive sources are estimated in Table 19.6 as less than 200 millirem (0.2 rem)—well within the acceptable limits.
52. alpha, beta, beta, alpha, alpha, alpha, alpha, alpha, beta, beta, alpha, beta, beta, alpha
53. (a) cobalt is a component of vitamin B₁₂; (c) red blood cells contain hemoglobin, an iron-protein compound
55. Al-27: 13 protons, 14 neutrons; Al-28: 13 protons, 15 neutrons; Al-29: 13 protons, 16 neutrons
57. Breeder reactors convert nonfissionable U-238 into fissionable Pu-239. A combination of U-238 and U-235 is used. Excess neutrons from the U-235 fission are absorbed by U-238, converting it into Pu-239. The chemical and physical properties of Pu-239 make it very difficult and expensive to handle.

Chapter 20

2. Four. Carbon atoms have four valence electrons. By making four bonds, carbon atoms complete their valence octet.
4. four; tetrahedral
6. The geometry around each carbon is a tetrahedral, making the angles 109.5°. To be “straight,” the angles would have to be 180°.
8. (a) $\begin{array}{cccc} \text{H} & \text{H} & \text{H} & \text{H} \\ | & | & | & | \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ | & | & | & | \\ \text{H} & \text{H} & \text{H} & \text{H} \end{array}$
- (c) $\begin{array}{cccc} \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \\ | & | & | & | & | \\ \text{H}-\text{C}-\text{C}-\text{C}-\text{C}-\text{C}-\text{H} \\ | & | & | & | & | \\ \text{H} & \text{H} & \text{H} & \text{H} & \text{H} \end{array}$
10. a branch or substituent
13. The root name is derived from the number of carbon atoms in the longest continuous chain of carbon atoms.
14. The position of a substituent is indicated by a number that corresponds to the carbon atom in the longest chain to which the substituent is attached.
16. (a) 3-ethylpentane; (c) 2,2-dimethylpropane
17. (a) $\text{CH}_3-\text{CH}-\text{CH}_2-\text{CH}_2-\text{CH}_2-\text{CH}_3$;



19. Number of Carbon Atoms	Use
5–12	gasoline
10–18	kerosene, jet fuel
15–25	diesel fuel, heating oil
>25	asphalt

23. Combustion represents the vigorous reaction of a hydrocarbon (or other substances) with oxygen. It has been used as a source of heat and light.

25. Dehydrogenation reactions involve the removal of hydrogen atoms from adjacent carbon atoms in an alkane (or other substance). When two hydrogen atoms are removed from an alkane, a double bond is formed.

26. (a) $2\text{C}_6\text{H}_{14}(l) + 19\text{O}_2(g) \rightarrow 12\text{CO}_2(g) + 14\text{H}_2\text{O}(g)$

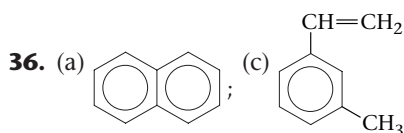
28. An alkyne is a hydrocarbon containing a carbon-carbon triple bond. The general formula is $\text{C}_n\text{H}_{2n-2}$.

29. (a) $\text{CH}_3\text{CH}_2\text{CH}_3(g)$

30. (a) 2-butene; (c) 1-butyne

33. A set of equivalent Lewis structures can be drawn for benzene. Each structure differs only in the location of the three double bonds in the ring. Experimentally benzene does not have the chemical properties expected for molecules having any double bonds.

35. *ortho*:- adjacent substituents (1,2-); *meta*:- two substituents with one unsubstituted carbon between them (1,3-); *para*:- two substituents with two unsubstituted carbon atoms between them (1,4-)



37. (a) 1,2-dimethylbenzene; (c) anthracene

39. (a) ether; (c) alcohol

41. Primary alcohols have *one* hydrocarbon fragment (alkyl group) bonded to the carbon atom where the —OH group is attached. Secondary alcohols have *two* such alkyl groups attached, and tertiary alcohols contain three such alkyl groups. Examples are

Ethanol (primary) $\text{CH}_3\text{—CH}_2\text{—OH}$

2-Propanol (secondary) $\text{CH}_3\text{—CH(OH)—CH}_3$

2-Methyl-2-propanol (tertiary) $\begin{array}{c} \text{OH} \\ | \\ \text{CH}_3 \\ | \\ \text{CH}_3\text{—C—CH}_3 \\ | \\ \text{OH} \end{array}$

42. (a) 1-pentanol (primary); (c) 3-pentanol (secondary)

43. (a) $\text{CH}_3\text{—CH}_2\text{—CH}_2\text{—CH}_2\text{—CH}_2\text{—OH}$; (primary);

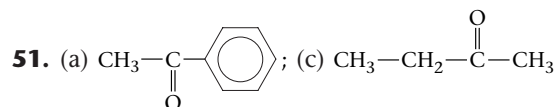
(c) $\text{H}_3\text{C—CH}_2\text{—CH(OH)—CH}_2\text{—CH}_3$ (secondary)

45. $\text{C}_6\text{H}_{12}\text{O}_6 \rightarrow 2\text{CH}_3\text{CH}_2\text{OH} + 2\text{CO}_2$ (yeast is a catalyst). The yeast is killed if the concentration of alcohol is greater than 13%. To make more concentrated ethanol solutions, distillation is needed.

47. methanol (CH_3OH): starting material for synthesis of acetic acid and many plastics; isopropyl alcohol (2-propanol, $\text{CH}_3\text{—CH(OH)—CH}_3$): rubbing alcohol

49. The location of the carbonyl group (C=O). Aldehydes contain the carbonyl group at the end of a hydrocarbon chain. Ketones contain the carbonyl group in the interior of a hydrocarbon chain.

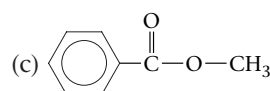
50. (a) 3-hexanone (ethyl propyl ketone);
(c) 3,4-dimethylpentanal



52. Carboxylic acids are typically weak acids;
 $\text{CH}_3\text{CH}_2\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{CH}_2\text{COO}_2^- + \text{H}_3\text{O}^+$

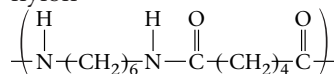
54. (a) 3-methylbutanoic acid; (c) 2-hydroxypropanoic acid

55. (a) $\text{CH}_3\text{—CH}_2\text{—CH(CH}_3\text{)—CH}_2\text{—C(=O)—OH}$

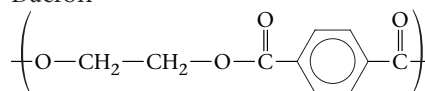


57. A polyester is a polymer in which the polymerization forms an ester group. The repeating unit is an ester; thus there are many esters, or “poly-esters.” Dacron is an example.

59. nylon



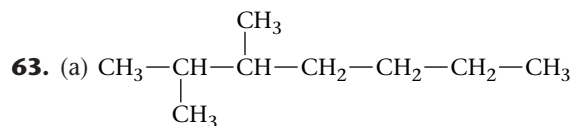
Dacron



60. The general formula is $\text{C}_n\text{H}_{2n+2}$. Each successive alkane differs from the previous or following alkane by CH_2 (a methylene unit).

62. (a) 2-chlorobutane;

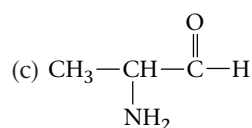
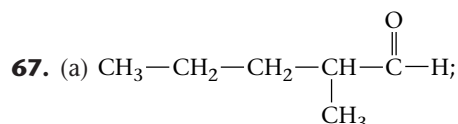
(c) triiodomethane (common name: iodoform)



(c) $\text{CH}_2\text{=C(Cl)—CH}_2\text{—CH}_2\text{—CH}_2\text{—CH}_3$

64. (b) $\text{CH}_3\text{—CH}_2\text{—CH}_2\text{—CH}_3$

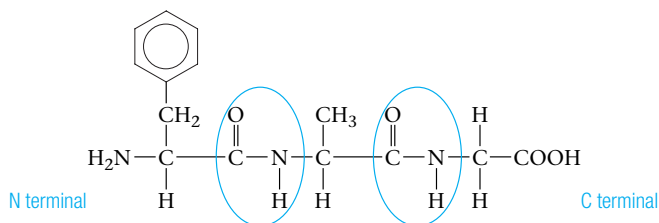
65. 1,2,3-trihydroxypropane (1,2,3-propanetriol)



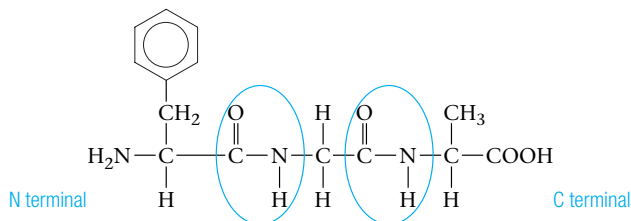
68. (a) carboxylic acid; (c) ester

Chapter 21

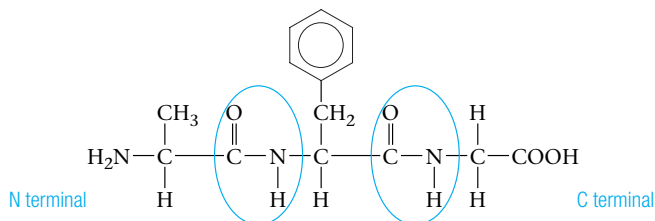
- Proteins represent biopolymers of α -amino acids. Proteins make up about 15% of the body by mass.
- Fibrous proteins provide the structural material of many tissues in the body. They are the chief constituents of hair, cartilage, and muscles. Fibrous proteins consist of lengthwise bundles of polypeptide chains (a fiber). Globular proteins consist of polypeptide chains folded into a spherical shape; they are found in the bloodstream where they transport and store various substances, act as antibodies (fight infections), act as enzymes (catalysts), and participate in the body's various regulatory systems.
- The structures of the amino acids are given in Figure 21.2. A side chain is nonpolar if it is mostly hydrocarbon in nature (like alanine). Polar side chains may contain the hydroxyl group ($-\text{OH}$), the sulfhydryl group ($-\text{SH}$), or a second amino ($-\text{NH}_2$) or carboxyl ($-\text{COOH}$) group.
- cys-ala-phe; cys-phe-ala; phe-ala-cys; phe-cys-ala; ala-cys-phe; ala-phe-cys
- phe-ala-gly



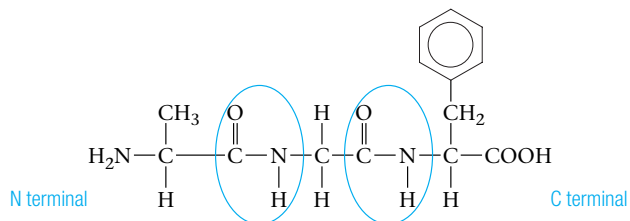
phe-gly-ala



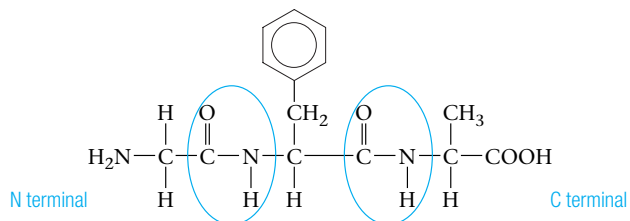
ala-phe-gly



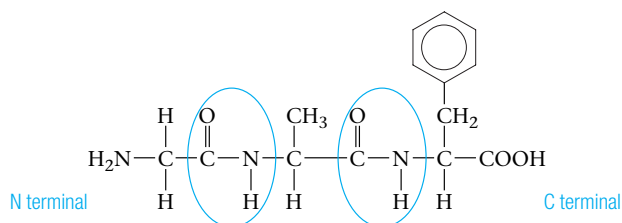
ala-gly-phe



gly-phe-ala



gly-ala-phe



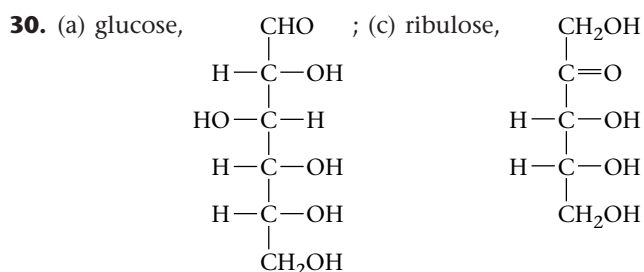
- Long, thin, resilient proteins (such as hair) typically contain elongated, elastic α -helical protein molecules. Other proteins (such as silk) that form sheets or plates typically contain protein molecules having the beta pleated-sheet structure. Proteins without a structural function in the body (such as hemoglobin) typically have a globular structure.
- The polypeptide chain forms a coil or spiral. Such proteins are found in wool, hair, and tendons.
- Cysteine contains the sulfhydryl ($-\text{SH}$) group in its side chain. It can therefore form disulfide linkages ($-\text{S}-\text{S}-$) with other cysteine molecules in the same polypeptide chain. This produces a kink or a knot in the chain, which leads to the tertiary structure (three-dimensional shape). For example, cysteine is responsible for the curling of hair.
- Collagen has an α -helical secondary structure. Collagen functions as the raw material from which tendons are constructed.
- Antibodies are proteins made in the body in response to foreign substances such as bacteria or viruses. Interferon is an important antibody because it offers general protection against viruses (many antibodies target specific invaders).

22. The cross-linkages between adjacent polypeptide chains of one protein are broken chemically and then re-formed chemically in a new location. The primary cross-linkage involved is a disulfide linkage between cysteine units in the polypeptide chain. The tertiary structure is mainly affected, although the secondary structure can also be affected if the waving lotion is left on too long (making the hair very “frizzy”).

24. The molecule acted on by an enzyme is referred to as the enzyme’s substrate. If an enzyme is said to be specific for a particular substrate, it will catalyze the reactions of only that molecule.

26. The lock-and-key model postulates that the structures of the enzyme and its substrate are complementary. In this way, the active site of the enzyme and the portion of the substrate to be acted on can fit closely together. The structures of these portions of the molecules are unique to the particular enzyme–substrate pair. They fit together like a particular key is necessary to open a specific lock.

28. Sugars contain an aldehyde or ketone functional group (carbonyl group, C=O) as well as several hydroxyl groups (–OH).



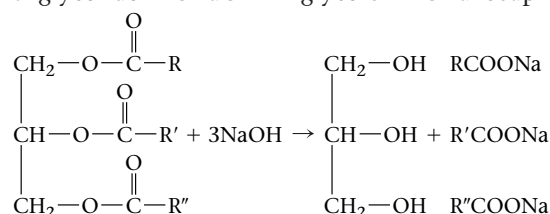
32. uracil (RNA only); cytosine (DNA, RNA); thymine (DNA only); adenine (DNA, RNA); guanine (DNA, RNA)

34. A comparison of the two strands of a DNA molecule shows that a given base in one strand is always paired with a particular base in the other strand. Because of the shapes and side atoms along the rings of the nitrogen bases, only certain parts can hydrogen-bond with one another in the double helix. Adenine is always paired with thymine; cytosine is always paired with guanine. When a DNA helix unwinds for replication during cell division, only the appropriate complementary bases are able to approach and bond to the nitrogen bases of each strand. For example, when the two strands of a guanine–cytosine pair in the original DNA separate, only a new cytosine molecule can bond to the original guanine, and only a new guanine molecule can bond to the original cytosine.

36. Messenger RNA molecules are synthesized to be complementary to a portion (gene) of the DNA molecule in the cell. They serve as the template or pattern on which a protein will be constructed (a particular group of nitrogen bases on mRNA is able to accommodate and specify a particular amino acid in a particular location in the protein). Transfer RNA molecules are much smaller than mRNA, and their structure accommodates only a single specific amino acid molecule. They “find” their specific amino acid in the cellular fluids and bring it to mRNA, where it is added to the protein molecule being synthesized.

37. Rather than having some common structural feature, substances are classified as lipids based on their solubility characteristics. Lipids are water-insoluble substances that can be extracted from cells by nonpolar organic solvents such as benzene or carbon tetrachloride.

39. Saponification is the production of a *soap* by treatment of a triglyceride with a strong base such as NaOH.
 $\text{triglyceride} + 3\text{NaOH} \rightarrow \text{glycerol} + 3\text{Na}^+\text{soap}^-$



41. Cholesterol is a naturally occurring steroid from which the body synthesizes other needed steroids. Because cholesterol is insoluble in water, having too large a concentration of this substance in the bloodstream may lead to buildup on the walls of blood vessels, causing their eventual blockage.

- 43.** i
- 45.** m
- 47.** u
- 49.** f
- 51.** g
- 53.** r
- 55.** p
- 57.** o
- 59.** b
- 61.** d
- 63.** a

Glossary

Absolute zero $-273\text{ }^{\circ}\text{C}$

Acid a substance that produces hydrogen ions in aqueous solution; a proton donor.

Acid-base indicator a chemical that changes color according to the pH of a solution.

Acidic oxide a covalent oxide that dissolves in water to give an acidic solution.

Actinide series a group of fourteen elements following actinium on the periodic table, in which the $5f$ orbitals are being filled.

Activation energy the minimum energy required in order to cause a chemical reaction.

Active site the part of the enzyme to which the specific substrate is bound as a reaction is catalyzed.

Addition polymerization the process in which monomers simply “add together” to form polymers.

Addition reaction a reaction in which new atoms form single bonds to the carbon atoms in unsaturated hydrocarbons that were involved in double or triple bonds.

Air pollution contamination of the atmosphere, mainly by the gaseous products of transportation and the production of electricity.

Alcohol a hydrocarbon derivative in which the hydroxyl group ($-\text{OH}$) is the functional group.

Aldehyde an organic compound containing the carbonyl group bonded to at least one hydrogen atom.

Alkali metal a Group 1 metal.

Alkaline earth metal a Group 2 metal.

Alkane a saturated hydrocarbon with the general formula $\text{C}_n\text{H}_{2n+2}$.

Alkene an unsaturated hydrocarbon containing a carbon-carbon double bond. The general formula is C_nH_{2n} .

Alkyne an unsaturated hydrocarbon containing a carbon-carbon triple bond. The general formula is $\text{C}_n\text{H}_{2n-2}$.

Alloy a mixture of elements that has metallic properties.

Alloy steel a form of steel containing carbon plus metals such as chromium, cobalt, manganese, and molybdenum.

Alpha (α) particle a helium nucleus produced in radioactive decay.

Alpha-particle production a common mode of decay for heavy radioactive nuclides resulting in a loss of 4 in mass number and a loss of 2 in atomic number.

Amine a hydrocarbon derived from ammonia in which one or more of the hydrogen atoms are replaced by organic groups.

α -Amino acid an organic acid in which an amino group, a hydrogen atom, and an R group are attached to the carbon atom next to the carboxyl group.

Ampere the unit of measurement for electric current; 1 ampere is equal to 1 coulomb of charge per second.

Amphoteric substance a substance that can behave either as an acid or as a base.

Anion a negatively charged ion.

Anode the electrode in a galvanic cell at which oxidation occurs.

Aqueous solution a solution with water as a solvent.

Aromatic hydrocarbon one of a special class of cyclic unsaturated hydrocarbons, the simplest of which is benzene.

Arrhenius concept of acids and bases a concept postulating that acids produce hydrogen ions in solution, whereas bases produce hydroxide ions in solution.

Atmosphere the mixture of gases that surrounds the earth's surface.

Atom the fundamental unit of which elements are composed.

Atomic mass unit (amu) a small unit of mass equal to 1.66×10^{-24} grams.

Atomic number the number of protons in the nucleus of a given atom.

Atomic radius half of the distance between the atomic nuclei in a molecule consisting of identical atoms.

Atomic solid a solid that contains atoms at the lattice points.

Auto-ionization the transfer of a proton from one molecule to another of the same substance.

Average atomic mass (weight) the weighted average of the masses of all of the isotopes of an element.

Avogadro's law equal volumes of gases at the same temperature and pressure contain the same number of particles (atoms or molecules).

Avogadro's number the number of atoms in exactly 12 grams of pure ^{12}C , equal to 6.022×10^{23} .

Ball-and-stick model a molecular model that distorts the sizes of atoms but shows bond relationships clearly.

Barometer a device that measures atmospheric pressure.

Base a substance that produces hydroxide ions in solution; a proton acceptor.

Basic oxide an ionic oxide that dissolves in water to produce a basic solution.

Battery a group of galvanic cells connected in series.

Beta (β) particle an electron produced in radioactive decay.

Beta-particle production a decay process for radioactive nuclides in which the mass number remains constant and the atomic number increases by one. The net effect is to change a neutron to a proton.

Binary compound a two-element compound.

Binary ionic compound a two-element compound consisting of a cation and an anion.

Binding energy (nuclear) the energy required to decompose a nucleus into its component nucleons.

Biochemistry the study of the chemistry of living systems.

Biomolecule a molecule that functions in maintaining and/or reproducing life.

Bond (chemical bond) the force that holds two or more atoms together and makes them function as a unit.

Bond energy the energy required to break a given chemical bond.

Bond length the distance between the nuclei of the two atoms that are connected by a bond.

- Bonding pair** a pair of electrons that are shared between two atoms forming a covalent or polar-covalent bond.
- Boyle's law** the pressure of a given sample of a gas is inversely related to the volume of the gas at constant temperature. $PV = k$
- Breeder reactor** a nuclear reactor in which fissionable fuel is produced while the reactor runs.
- Brønsted-Lowry model** a model proposing that an acid is a proton donor and that a base is a proton acceptor.
- Buffer capacity** the ability of a buffered solution to absorb protons or hydroxide ions without a significant change in pH.
- Buffered solution** a solution that resists a change in pH when either an acid or base is added.
- Buret** a device for the accurate measurement of the delivery of a given volume of liquid or solution.
- Calorie** the amount of energy needed to raise the temperature of one gram of water by one Celsius degree.
- Calorimeter** a device used to determine the heat associated with a chemical or physical change.
- Calorimetry** the science of measuring heat flow.
- Carbohydrate** a polyhydroxyl ketone or polyhydroxyl aldehyde or a polymer composed of these.
- Carbon steel** an alloy of iron containing up to about 1.5% carbon.
- Carboxyl group** the $-\text{COOH}$ group in an organic acid.
- Carboxylic acid** a hydrocarbon derivative ($-\text{COOH}$) containing the carboxyl group.
- Catalyst** a substance that speeds up a reaction without being consumed.
- Cathode** the electrode in a galvanic cell at which reduction occurs.
- Cathode rays** the "rays" emanating from the negative electrode (cathode) in a partially evacuated tube; a stream of electrons.
- Cathodic protection** the connection of an active metal, such as magnesium, to steel to protect the steel from corrosion.
- Cation** a positively charged ion.
- Cell** the smallest unit in living things that exhibits the properties normally associated with life, such as reproduction, metabolism, mutation, and sensitivity to external stimuli.
- Cell potential (electromotive force)** the driving force in a galvanic cell that pushes electrons from the reducing agent in one compartment to the oxidizing agent in the other.
- Cellulose** the major structural component of woody plants and natural fibers, such as cotton; a polymer of glucose.
- Chain reaction (nuclear)** a self-sustaining fission process caused by the production of neutrons that proceed to split other nuclei.
- Charles's law** the volume of a given sample of a gas is directly related to the temperature of the gas at constant pressure. $V = bT$
- Chemical change** the change of substances into other substances through a reorganization of the atoms; a chemical reaction.
- Chemical equation** a representation of a chemical reaction using the formulas of the starting substances that react and the new substances that are formed.
- Chemical equilibrium** a dynamic state in which the concentrations of the reactants and the products remain constant over time, as long as the conditions are not changed.
- Chemical formula** a representation of a molecule in which the symbols for the elements are used to indicate the types of atoms present and subscripts are used to show the relative numbers of atoms.
- Chemical kinetics** the area of chemistry that concerns reaction rates.
- Chemical property** characteristic that describes the ability of a substance to change to a different substance.
- Chemical reaction** a process in which one or more substances are changed into one or more new substances by the reorganization of component atoms.
- Chemical stoichiometry** the quantities of materials consumed and produced in a chemical reaction.
- Chemistry** the science of materials and the changes that these materials undergo.
- Coal** a solid fossil fuel mostly consisting of carbon.
- Coefficient** the number written in front of the chemical formulas in a balanced chemical equation; coefficients indicate the relative numbers of reactants and products in the reaction.
- Colligative property** a property that depends only on the number of solute particles present in solution.
- Collision model** molecules must collide in order to react; used to account for the fact that reaction rate depends on temperature and concentrations of reactants.
- Combustion reaction** a chemical reaction involving oxygen as one of the reactants that produces enough heat so that a flame results.
- Complete ionic equation** a chemical equation for a reaction in solution representing all strong electrolytes as ions.
- Compound** a substance made of two or more different elements joined together in a specific way.
- Concentrated** describes a solution in which a relatively large amount of solute is dissolved in a solution.
- Condensation** the process in which a vapor is converted to a liquid.
- Condensation polymerization** the process in which a small molecule such as water is produced for each extension of the polymer chain.
- Condensed states of matter** liquids and solids.
- Conjugate acid** the substance formed when a proton is added to a base.
- Conjugate acid-base pair** two substances related to each other by the donating and accepting of a single proton.
- Conjugate base** the remaining substance when a proton is lost from an acid.
- Continuous spectrum** a spectrum that exhibits all the wavelengths of visible light.
- Control rods** in a nuclear reactor, rods composed of substances that absorb neutrons. These rods regulate the power level of the reactor.
- Conversion factor** a ratio used to convert from one unit to another.

- Copolymer** a polymer consisting of two different types of monomers.
- Core electron** an inner electron in an atom; one that is not in the outermost principal energy level of an atom.
- Corrosion** the process by which metals are oxidized in the atmosphere.
- Covalent bonding** a type of bonding in which atoms share electrons.
- Critical mass** the mass of fissionable material required to produce a chain reaction.
- Critical reaction (nuclear)** a reaction in which exactly one neutron from each fission event causes another fission event, thus sustaining the chain reaction.
- Crystalline solid** a solid characterized by the regular arrangement of its components.
- Dalton's atomic theory** a theory established by John Dalton in the early 1800s, used to explain the nature of materials.
- Dalton's law of partial pressures** for a mixture of gases in a container, the total pressure exerted is the sum of the partial pressures of each of the gases.
- Decomposition reaction** a chemical reaction in which a compound is broken down into simpler compounds, or into the component elements.
- Denaturation** the breaking down of the three-dimensional structure of a protein, resulting in the loss of its function.
- Density** the mass of a substance per a given volume of that substance.
- Deoxyribonucleic acid (DNA)** a huge nucleotide polymer having a double-helical structure with complementary bases on the two strands. Its major functions are protein synthesis and the storage and transport of genetic information.
- Diatomic molecule** a molecule composed of two atoms.
- Dilution** the process of adding solvent to a solution to lower the concentration of solute.
- Dimensional analysis** the process of using conversion factors to change from one unit to another.
- Dimer** a molecule consisting of two monomers joined together.
- Dipole-dipole attraction** the attractive force between the positively charged end of one polar molecule and the negatively charged end of another polar molecule.
- Dipole moment** a property of a molecule whereby the charge distribution can be represented by a center of positive charge and a center of negative charge.
- Disaccharide** a sugar formed from two monosaccharides joined by a glycoside linkage.
- Distillation** a method for separating the components of a liquid mixture that depends on the different boiling points of the substances.
- Double bond** a covalent or polar covalent bond in which two pairs of electrons are shared by two atoms.
- Dry cell battery** a common battery used in calculators, watches, radios, and tape players.
- Electrical conductivity** the ability to conduct an electric current.
- Electrochemistry** the study of the interchange of chemical and electrical energy.
- Electrolysis** a process that involves forcing a current through a cell to cause a chemical reaction that would not otherwise occur.
- Electrolyte** a material that dissolves in water to give a solution that conducts an electric current.
- Electrolytic cell** a cell that uses electrical energy to produce a chemical change that would not otherwise occur.
- Electromagnetic radiation** radiant energy that exhibits wavelike behavior and travels through space at the speed of light in a vacuum.
- Electron** a negatively charged subatomic particle.
- Electronegativity** the tendency of an atom in a molecule to attract shared electrons to itself.
- Element** a substance that cannot be decomposed into simpler substances by chemical or physical means. It consists of atoms all having the same atomic number.
- End point** the point in a titration at which the indicator changes color.
- Endothermic** refers to a process in which energy (as heat) flows from the surroundings into the system.
- Energy** the ability to do work or to produce heat.
- Enthalpy** at constant pressure, a change in enthalpy equals the energy flow as heat.
- Entropy** a function used to keep track of the natural tendency for the components of the universe to become disordered; a measure of disorder or randomness.
- Enzyme** a large molecule, usually a protein, that catalyzes biological reactions.
- Equilibrium** the exact balance of two processes, one of which is opposite of the other.
- Equilibrium constant** the value obtained when equilibrium concentrations of the chemical species are substituted into the equilibrium expression.
- Equilibrium expression** the expression (from the law of mass action) equal to the product of the product concentrations divided by the product of the reactant concentrations, each concentration having first been raised to a power represented by the coefficient in the balanced equation.
- Equilibrium position** a particular set of equilibrium concentrations of all reactants and products in a chemical system.
- Equivalence point (stoichiometric point)** the point in a titration at which enough titrant has been added to react exactly with the substance in solution that is being titrated.
- Equivalent of an acid** the amount of acid that can furnish one mole of hydrogen ions (H^+).
- Equivalent of a base** the amount of base that can furnish one mole of hydroxide ions (OH^-).
- Equivalent weight** the mass (in grams) of one equivalent of an acid or a base.
- Essential elements** the 30 elements currently known to be essential to human life.
- Ester** a hydrocarbon derivative produced by the reaction between a carboxylic acid and an alcohol.
- Exothermic** refers to a process in which energy (as heat) flows out of the system into the surroundings.
- Fat (glyceride)** an ester composed of glycerol and fatty acids.

- Fatty acid** a long-chain carboxylic acid.
- Filtration** a method for separating the components of a mixture containing a solid and a liquid.
- First law of thermodynamics** a law stating that the energy of the universe is constant.
- Fission** the process of splitting a heavy nucleus into two more stable nuclei with smaller mass numbers.
- Fossil fuel** fuel that consists of carbon-based molecules derived from decomposition of once-living organisms; coal, petroleum, or natural gas.
- Frequency** the number of waves (cycles) per second that pass a given point in space.
- Fuel cell** a galvanic cell for which the reactants are continuously supplied.
- Functional group** an atom or group of atoms in a hydrocarbon derivative that contains elements in addition to carbon and hydrogen.
- Fusion** the process of combining two light nuclei to form a heavier, more stable nucleus.
- Galvanic cell** a device in which chemical energy is converted spontaneously to electrical energy by means of oxidation–reduction reaction.
- Galvanizing** a process in which steel is coated with zinc to prevent corrosion.
- Gamma (γ) ray** a high-energy photon produced in radioactive decay.
- Gas** one of the three states of matter; substance with no definite shape or volume.
- Geiger–Müller counter (Geiger counter)** an instrument that measures the rate of radioactive decay by registering the ions and electrons produced as a radioactive particle passes through a gas-filled chamber.
- Gene** a given segment of the DNA molecule that contains the code for a specific protein.
- Glycogen** the main carbohydrate reservoir in animals; a polymer of glucose.
- Glycoside linkage** the C—O—C bond formed between monosaccharide rings in a disaccharide.
- Greenhouse effect** the process by which an atmosphere warms a planet.
- Ground state** the lowest possible energy state of an atom or a molecule.
- Group** a vertical column of elements on the periodic table.
- Haber process** the manufacture of ammonia from nitrogen and hydrogen, carried out at high pressure and high temperature with the aid of a catalyst.
- Half-life (of a radioactive sample)** the time required for half of the original sample of radioactive nuclides to decay.
- Half-reactions** the two parts of an oxidation–reduction reaction, one representing oxidation, the other reduction.
- Halogen** a Group 7 element.
- Halogenation** an addition reaction in which a halogen is a reactant.
- Hard water** water from natural sources that contains relatively large concentrations of calcium and magnesium ions.
- Heat** flow of energy due to a temperature difference.
- Heating/cooling curve** a plot of temperature versus time for a substance, where energy is added at a constant rate.
- Heisenberg uncertainty principle** a principle stating that there is a fundamental limitation to how precisely we can know both the position and the momentum of a particle at a given time.
- Herbicide** a pesticide applied to kill weeds.
- Hess's Law** the change in enthalpy in going from a given set of reactants to a given set of products does not depend on the number of steps in the reaction.
- Heterogeneous equilibrium** an equilibrium system in which all reactants and products are in different states.
- Heterogeneous mixture** a mixture that has different properties in different regions of the mixture.
- Heterogeneous reaction** reaction involving reactants and products in different states.
- Homogeneous equilibrium** an equilibrium system in which all reactants and products are in the same state.
- Homogeneous mixture** a mixture that is the same throughout.
- Homogeneous reaction** reaction involving reactants and products in the same state.
- Homopolymer** a polymer consisting of a single type of monomer.
- Hydration** the interaction between solute particles and water molecules.
- Hydrocarbon** a compound of carbon and hydrogen.
- Hydrogen bonding** unusually strong dipole–dipole attractions that occur among molecules in which hydrogen is bonded to a highly electronegative atom.
- Hydrogenation reaction** an addition reaction in which H_2 is a reactant.
- Hydronium ion** the H_3O^+ ion; a hydrated proton.
- Hypothesis** one or more assumptions put forth to explain observed phenomena.
- Ideal gas** a hypothetical gas that exactly obeys the ideal gas law. A real gas approaches ideal behavior at high temperature and/or low pressure.
- Ideal gas law** an equation relating the properties of an ideal gas, expressed as $PV = nRT$, where P = pressure, V = volume, n = moles of the gas, R = the universal gas constant, and T = temperature on the Kelvin scale. This equation expresses behavior closely approached by real gases at high temperature and/or low pressure.
- Indicator** a chemical that changes color and is used to mark the end point of a titration.
- Indicator paper** a strip of paper coated with a combination of acid–base indicators.
- Insoluble solid** a solid that dissolves to such a small degree that it is not detectable to the naked eye.
- Intermolecular forces** attraction forces that occur between molecules.
- Internal energy** the sum of the kinetic and potential energies of all particles in the system.
- Intramolecular forces** attractive forces that occur between atoms in a molecule; chemical bonds
- Ion** an atom or a group of atoms that has a positive or negative charge.

- Ionic bonding** the attraction between oppositely charged ions.
- Ionic compound** a compound that results when a metal reacts with a nonmetal to form cations and anions.
- Ionic solid** a solid containing cations and anions that dissolves in water to give a solution containing the separated ions, which are mobile and thus free to conduct an electric current.
- Ionization energy** the amount of energy required to remove an electron from a gaseous atom or ion.
- Ion-product constant (K_w)** the equilibrium constant for the auto-ionization of water; $K_w = [H^+][OH^-]$. At 25 °C, K_w equals 1.0×10^{-14} .
- Isomers** species that have the same chemical formula but different properties.
- Isotopes** atoms with the same number of protons but different numbers of neutrons.
- Joule** a unit of measurement for energy; 1 calorie = 4.184 joules.
- Ketone** a hydrocarbon derivative containing the carbonyl group bonded to two carbon atoms.
- Kinetic energy** ($\frac{1}{2}mv^2$) energy due to the motion of an object.
- Kinetic molecular theory (KMT)** a model that assumes that an ideal gas is composed of tiny particles (atoms or molecules) in constant motion.
- Lanthanide series** a group of fourteen elements following lanthanum on the periodic table, in which the 4f orbitals are being filled.
- Lattice** a three-dimensional system of points designating the positions of the centers of the components of a solid (atoms, ions, or molecules).
- Law of chemical equilibrium** a general description of the equilibrium condition; it defines the equilibrium expression.
- Law of conservation of energy** energy can be converted from one form to another but can be neither created nor destroyed.
- Law of conservation of mass** mass is neither created nor destroyed.
- Law of constant composition** a given compound always contains elements in exactly the same proportion by mass.
- Law of mass action** (also called the law of chemical equilibrium) a general description of the equilibrium condition; it defines the equilibrium expression.
- Law of multiple proportions** a law stating that when two elements form a series of compounds, the ratios of the masses of the second element that combine with one gram of the first element can always be reduced to small whole numbers.
- Lead storage battery** a battery used in cars in which the anode is lead, the cathode is lead coated with lead dioxide, and the electrolyte is sulfuric acid.
- Le Châtelier's principle** if a change is imposed on a system at equilibrium, the position of the equilibrium will shift in a direction that tends to reduce the effect of that change.
- Lewis structure** a representation of a molecule or polyatomic ion showing how valence electrons are arranged among the atoms in the molecule or ion.
- Limiting reactant (limiting reagent)** the reactant that is completely used up when a reaction is run to completion.
- Line spectrum** a spectrum showing only certain discrete wavelengths.
- Linear accelerator** a type of particle accelerator in which a changing electrical field is used to accelerate a beam of charged particles along a linear path.
- Lipids** water-insoluble substances that can be extracted from cells by nonpolar organic solvents.
- Liquid** one of the three states of matter; has a definite volume that takes the shape of its container.
- London dispersion forces** relatively weak intermolecular forces resulting from a temporarily uneven distribution of electrons that induces a dipole in a neighbor.
- Lone pair** electron pairs in a Lewis structure that are not involved in bonding.
- Main-group (representative) elements** elements in the groups labeled 1, 2, 3, 4, 5, 6, 7, and 8 on the periodic table. The group number gives the sum of the valence *s* and *p* electrons.
- Mass** the quantity of matter in an object.
- Mass number** the total number of protons and neutrons in the nucleus of a given atom.
- Mass percent** the percent by mass of a component of a mixture or of a given element in a compound.
- Matter** the material of the universe.
- Measurement** a quantitative observation
- Metal** an element that gives up electrons relatively easily and is typically lustrous, malleable, and a good conductor of heat and electricity.
- Metalloid (semi-metals)** an element that has both metallic and nonmetallic properties.
- Metallurgy** the process of separating a metal from its ore and preparing it for use.
- Millimeters of mercury (mm Hg)** a unit of measurement for pressure, also called a torr; 760 mm Hg = 760 torr = 101,325 Pa = 1 standard atmosphere.
- Mixture** a substance with variable composition.
- Model (theory)** a set of assumptions put forth to explain the observed behavior of matter. The models of chemistry usually involve assumptions about the behavior of individual atoms or molecules.
- Moderator** a substance used in a nuclear reactor to slow down the neutrons.
- Molar heat of fusion** the energy required to melt one mole of a solid.
- Molar heat of vaporization** the energy required to vaporize 1 mol of a liquid.
- Molar mass** the mass in grams of one mole of a compound.
- Molar volume** the volume of one mole of an ideal gas; equal to 22.42 liters at standard temperature and pressure.
- Molarity** moles of solute per volume of solution in liters.
- Mole (mol)** the number equal to the number of carbon atoms in exactly 12 grams of pure ^{12}C ; Avogadro's number. One mole represents 6.022×10^{23} units.

Mole ratio (stoichiometry) the ratio of moles of one substance to moles of another substance in a balanced chemical equation.

Molecular equation a chemical equation showing the complete (undissociated) forms of all reactants and products.

Molecular formula the actual formula of a compound, giving the types of atoms and the number of each type of atom.

Molecular solid a solid composed of molecules.

Molecular structure the three-dimensional arrangement of atoms in a molecule.

Molecular weight (molar mass) the mass in grams of one mole of a substance.

Molecule a collection of atoms bonded together that behave as a unit.

Monoprotic acid an acid with one acidic proton.

Natural gas a gaseous fossil fuel mostly consisting of methane and usually associated with petroleum deposits.

Natural law a statement that summarizes generally observed behavior.

Net ionic equation a chemical equation for a reaction in solution showing only those components that are directly involved in the reaction. Strong electrolytes are represented as ions.

Network solid an atomic solid containing strong directional covalent bonds.

Neutralization reaction an acid–base reaction.

Neutron a subatomic particle in the atomic nucleus with no charge.

Noble gas a Group 8 element.

Nonelectrolyte a substance that, when dissolved in water, gives a nonconducting solution.

Nonmetal an element that does not exhibit metallic characteristics. Chemically, a typical nonmetal accepts electrons from a metal.

Normal boiling point the temperature at which the vapor pressure of a liquid is exactly one atmosphere; the boiling temperature under one atmosphere of pressure.

Normal melting/freezing point the melting/freezing temperature of a liquid under one atmosphere of pressure.

Normality the number of equivalents of a substance dissolved in a liter of solution.

Nuclear atom a concept of the atom as having a dense center of positive charge (the nucleus) surrounded by moving electrons.

Nuclear transformation the change of one element into another.

Nucleon a particle in an atomic nucleus, either a neutron or a proton.

Nucleotide a monomer of DNA and RNA consisting of nitrogen-containing base, a five-carbon sugar, and a phosphate group.

Nucleus the relatively small, dense center of positive charge in an atom.

Nuclide the general term applied to each unique atom; represented by A_ZX , where X is the symbol for a particular element.

Octet rule the observation that atoms of nonmetals form the most stable molecules when they are surrounded by eight electrons (to fill their valence orbitals).

Orbital the three-dimensional region in which there is a high probability of finding an electron in an atom.

Organic acid an acid with a carbon-atom backbone and a carboxyl group.

Organic chemistry the study of carbon-containing compounds (typically containing chains of carbon atoms) and their properties.

Oxidation an increase in oxidation state; a loss of electrons.

Oxidation–reduction (redox) reaction a chemical reaction involving the transfer of electrons.

Oxidation states a concept that provides a way to keep track of electrons in oxidation–reduction reactions according to certain rules.

Oxidizing agent (electron acceptor) a reactant that accepts electrons from another reactant.

Oxyacid an acid in which the acidic proton is attached to an oxygen atom.

Oxyanion a polyatomic ion containing at least one oxygen atom and one or more atoms of at least one other element.

Ozone O_3 , a form of elemental oxygen much less common than O_2 in the atmosphere near the earth.

Partial pressures the independent pressures exerted by a gas in a mixture of gases.

Particle accelerator a device used to accelerate nuclear particles to very high speeds.

Pascal the SI unit of measurement for pressure; equal to one newton per square meter.

Pauli exclusion principle in a given atom, no two electrons can occupy the same atomic orbital and have the same spin.

Percent yield the actual yield of a product as a percentage of the theoretical yield.

Periodic table a chart showing all the elements arranged in columns in such a way that all the elements in a given column exhibit similar chemical properties.

Petroleum a thick, dark liquid composed mostly of hydrocarbon compounds.

pH meter a device used to measure the pH of a solution.

pH scale a log scale based on 10 and equal to $-\log[H^+]$; a convenient way to represent solution acidity.

Phenyl group the benzene molecule minus one hydrogen atom.

Phospholipid an ester of glycerol containing two fatty acids; consists of a long nonpolar “tail” and a polar substituted-phosphate “head”.

Photochemical smog air pollution produced by the action of light on oxygen, nitrogen oxides, and unburned fuel from auto exhaust to form ozone and other pollutants.

Photon a particle of electromagnetic radiation.

Physical charge a change in the form of a substance that does not affect the composition of a substance.

Physical property characteristic of a substance that can change without the substance’s becoming a different substance.

- Polar covalent bond** a covalent bond in which the electrons are not shared equally because one atom attracts the shared electron more strongly than the other atom.
- Polar molecule** a molecule that has a permanent dipole moment.
- Polyelectronic atom** an atom with more than one electron.
- Polymer** a large, usually chain-like molecule made from small molecules called monomers.
- Polymerization** a process in which small molecules called monomers are joined together to form a large molecule.
- Polyprotic acid** an acid with more than one acidic proton. It dissociates in a stepwise manner, one proton at a time.
- Polysaccharide** polymers containing many monosaccharide units.
- Positron production** a mode of nuclear decay in which a particle is formed that has the same mass as an electron but opposite charge. The net effect is to change a proton to a neutron.
- Potential** the “pressure” on electrons to flow from the anode to the cathode in a battery.
- Potential energy** energy due to position or composition.
- Precipitate** the solid that forms in a precipitation reaction.
- Precipitation** the formation of a solid in a chemical reaction.
- Precipitation reaction** a reaction in which a solid forms and separates from the solution as a solid.
- Precision** the degree of agreement among several measurements of the same quantity; the reproducibility of a measurement.
- Primary structure (of a protein)** the order or sequence of amino acids in the protein chain.
- Probability distribution (orbital)** a representation indicating the probabilities of finding an electron at various points in space.
- Product** the new substance formed by a chemical reaction. It is shown to the right of the reaction arrow.
- Protein** a natural polymer formed by condensation reactions between amino acids.
- Proton** a positively charged subatomic particle located in the atomic nucleus.
- Pure substance** a substance with constant composition; a pure element or a pure compound.
- Radioactive decay (radioactivity)** the spontaneous decomposition of a nucleus to form a different nucleus.
- Radiocarbon dating (carbon-14 dating)** a method for dating ancient wood or cloth on the basis of the radioactive decay of the carbon-14 nuclide.
- Radiotracer** a radioactive nuclide introduced into an organism and traced for diagnostic purposes.
- Random error** an error that has an equal probability of being high or low.
- Rate of decay** the change per unit time in the number of radioactive nuclides in a sample.
- Reactant** the starting substance of a chemical reaction shown to the left of the reaction arrow.
- Reactor core** the part of a nuclear reactor in which the fission reaction takes place.
- Reducing agent (electron donor)** a reactant that donates electrons to another substance, reducing the oxidation state of one of its atoms.
- Reduction** a decrease in oxidation state (a gain of electrons).
- Rem** a unit of radiation dosage that accounts for both the energy of the dose and its effectiveness in causing biological damage (from roentgen equivalent for man).
- Resonance** a condition occurring when more than one valid Lewis structure can be written for a particular molecule.
- Ribonucleic acid (RNA)** large nucleotide polymer that along with DNA, functions to transport genetic material.
- Salt** an ionic compound.
- Salt bridge** a U-tube containing an electrolyte that connects the two compartments of a galvanic cell, allowing ion flow without extensive mixing of the different solutions.
- Saponification** the process of breaking down a triglyceride by treatment with aqueous sodium hydroxide to produce glycerol and the fatty acid salts; the fatty acid salts produced are soaps.
- Saturated** describes a hydrocarbon in which all carbon-carbon bonds are single bonds.
- Saturated solution** a solution that contains as much solute as can be dissolved in that solution at that temperature.
- Scientific method** a process of studying natural phenomena that involves making observations, forming laws and theories, and testing theories by experimentation.
- Scientific notation** expresses a number in the form $N \times 10^M$; a convenient method for representing a very large or very small number and for easily indicating the number of significant figures.
- Scintillation counter** an instrument that measures the rate of radioactive decay by sensing flashes of light that the radiation produces in a detector.
- Second law of thermodynamics** the entropy of the universe is always increasing.
- Secondary structure (of a protein)** the arrangement in space of the chain of the protein chain (for example, α -helix, random coil, or pleated sheet).
- SI units** International System of units based on the metric system and on units derived from the metric system.
- Sigma (σ) bond** a covalent bond in which the electron pair is shared in an area centered on a line running between the atoms.
- Significant figures** the certain digits and the first uncertain digit of a measurement.
- Silica** the fundamental silicon–oxygen compound, which has the empirical formula SiO_2 and forms the basis of quartz and certain types of sand.
- Silicates** salts that contain metal cations and polyatomic silicon–oxygen anions that are usually polymeric.
- Single bond** a covalent or polar covalent bond in which one pair of electrons is shared by two atoms.
- Solid** one of the three states of matter; has a fixed shape and volume.

- Solubility** the amount of a substance that dissolves in a given volume of solvent or solution at a given temperature.
- Solubility product** the constant for the equilibrium expression representing the dissolving of an ionic solid in water.
- Soluble solid** a solid that readily dissolves in water.
- Solute** a substance dissolved in a solvent to form a solution.
- Solution** a homogeneous mixture.
- Solvent** the dissolving medium in a solution.
- Somatic damage** radioactive damage to an organism resulting in its sickness or death.
- Specific gravity** the ratio of density of a given liquid to the density of water at 4 °C.
- Specific heat** another name for specific heat capacity.
- Specific heat capacity** the amount of energy required to raise the temperature of one gram of a substance by one Celsius degree.
- Spectator ions** ions present in solution that do not participate directly in a reaction.
- Standard atmosphere** a unit of measurement for pressure equal to 760 mm Hg or 101, 325 Pa.
- Standard solution** a solution in which the concentration is accurately known.
- Standard temperature and pressure (STP)** the condition 0 °C and 1 atmosphere of pressure.
- Starch** the main carbohydrate reservoir in plants; a polymer of glucose.
- State function** a property that is independent of the pathway.
- States of matter** the three different forms in which matter can exist: solid, liquid, and gas.
- Stoichiometric quantities** quantities of reactants mixed in exactly the amounts that result in their all being used up at the same time.
- Stoichiometry** the process of using a balanced chemical equation to determine the relative masses of reactants and products involved in a reaction.
- Stoichiometry of a reaction** the relative quantities of reactants and products involved in the reaction.
- Strong acid** an acid that completely dissociates (ionizes) to produce H⁺ ions in solution.
- Strong base** a base that completely dissociates to produce OH⁻ ions in solution.
- Strong electrolyte** a substance that dissolves in water, dissociating completely into ions.
- Structural formula** the representation of a molecule in which the relative positions of the atoms are shown and the bonds are indicated by lines.
- Structural isomerism** describes what occurs when two molecules have the same atoms but different bonds
- Subcritical reaction (nuclear)** a reaction in which fewer than one of the neutrons from each fission event causes another fission event and the process dies out.
- Sublimation** the process by which a substance goes directly from the solid state to the gaseous state without passing through the liquid state.
- Substitution reaction (hydrocarbons)** a reaction in which an atom, usually a halogen, replaces a hydrogen atom in a hydrocarbon.
- Substrate** the molecule that interacts with an enzyme.
- Sucrose** table sugar; a disaccharide formed between glucose and fructose.
- Supercooling** the process of cooling a liquid to a temperature below its freezing point without its changing to a solid.
- Supercritical reaction (nuclear)** a reaction in which more than one of the neutrons from each fission event causes another fission event. The process rapidly escalates to a violent explosion.
- Superheating** the process of heating a liquid to a temperature above its boiling point without its boiling.
- Supersaturated** describes a solution that contains more solute than a saturated solution will hold at that temperature.
- Surfactant** a wetting agent that assists water in suspending nonpolar materials; soap is a surfactant.
- Surroundings** everything in the universe surrounding a thermodynamic system.
- System (thermodynamic)** that part of the universe on which attention is to be focused.
- Systematic error** an error that always occurs in the same direction.
- Temperature** measure of the random motions (average kinetic energy) of the components of a substance.
- Tertiary structure (of a protein)** the overall shape of a protein, long and narrow or globular, maintained by different types of intramolecular interactions.
- Theoretical yield** the maximum amount of a given product that can be formed when the limiting reactant is completely consumed.
- Theory (model)** a set of assumptions put forth to explain some aspect of the observed behavior of matter.
- Thermodynamics** the study of energy.
- Titration** a technique in which a solution of known concentration is used to determine the concentration of another solution.
- Titration curve (pH curve)** a plot of the pH of a given solution versus the volume of titrant added to the solution.
- Torr** another name for millimeters of mercury (mm Hg).
- Trace elements** first-row transition metals present in very small (trace) amounts in the human body.
- Transition metals** several series of elements in which inner orbitals (*d* or *f* orbitals) are being filled.
- Transuranium elements** the elements beyond uranium that are made artificially by particle bombardment.
- Triglyceride** a fat that is an ester of glycerol
- Triple bond** a covalent or polar covalent bond in which three pairs of electrons are shared by two atoms.
- Uncertainty (in measurement)** the characteristic reflecting the fact that any measurement involves estimates and cannot be exactly reproduced.
- Unit factor** an equivalence statement between units that is used for converting from one set of units to another.
- Units** the part of the measurement telling us the scale being used.
- Universal gas constant** the combined proportionality constant in the ideal gas law; 0.08206 L atm/K mol, or 8.314 J/K mol.

Unsaturated describes a hydrocarbon containing carbon-carbon multiple bonds.

Unsaturated solution a solution in which more solute can be dissolved than is dissolved already at that temperature.

Valence electrons the electrons in the outermost principal energy level of an atom.

Valence shell electron pair repulsion (VSEPR) model a model used to predict molecular geometry. Based on the idea that pairs of electrons surrounding an atom repel each other and that the atoms in a molecule are positioned to minimize this repulsion.

Vapor pressure the pressure exerted by a vapor in equilibrium with its liquid phase at a certain temperature.

Vaporization (evaporation) the process in which a liquid is converted to a gas.

Viscosity the resistance of a liquid to flow.

Volt the unit of measurement for electric potential; it is defined as one joule of work per coulomb of charge transferred.

Volume the amount of three-dimensional space occupied by a substance.

Wavelength the distance between two consecutive peaks or troughs in a wave.

Waxes a class of lipids that involve monohydroxy alcohols instead of glycerol.

Weak acid an acid that dissociates to a slight extent in aqueous solution.

Weak base a base that reacts with water to produce hydroxide ions to only a slight extent in aqueous solution.

Weak electrolyte a material that, when dissolved in water, gives a solution that conducts only a small electric current.

Weight the force exerted on an object by gravity.

Work force acting over a distance.

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