

Chemistry of Cooking

Sorangel Rodriguez-Velazquez

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Introduction

People around the world are fascinated about the preparation of food for eating. There are countless cooking books, TV shows, celebrity chefs and kitchen gadgets that make cooking an enjoyable activity for everyone. The chemistry of cooking course seeks to understand the science behind our most popular meals by studying the behavior of atoms and molecules present in food. This book is intended to give students a basic understanding of the chemistry involved in cooking such as caramelization, Maillard reaction, acid-base reactions, catalysis, and fermentation. Students will be able to use chemistry language to describe the process of cooking, apply chemistry knowledge to solve questions related to food, and ultimately create their own recipes.



Essential Ideas

Essential Ideas Introduction



Molecules and atoms in food are essential for our existence, providing sustenance, keeping us healthy, and contributing to the taste and smell of our favorite meals.
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Your alarm goes off and, after hitting “snooze” once or twice, you pry yourself out of bed. You drink a glass of water, then you shower, get dressed, check your phone for messages and grab the salad you made the night before for lunch. On your way to school, you stop to get a cup of coffee to help you get going, and an egg sandwich for breakfast, almost making you late for the first day of chemistry class. As you find a seat in the classroom, you read the question projected on the screen: “Welcome to class! Why should we study chemistry?”

Do you have an answer? You may be studying chemistry because it fulfills an academic requirement, but if you consider your daily activities, you might find chemistry interesting for other reasons. Preparing all the food you eat during your day involves chemistry. Making coffee, cooking eggs, and toasting bread involve chemistry. The ingredients we use—like flour, sugar and butter, the cooking technology that makes cooking easier and fun, waiting for bananas to ripen, melting sugar to make caramel—all of these and more involve chemical substances and processes. Whether you are aware or not, chemistry is part of your everyday food. In this course, you will learn many of the essential principles underlying the chemistry of cooking.

Attribution

Phases and Classification of Matter

Matter is defined as anything that occupies space and has mass, and it is all around us. Solids and liquids are more obviously matter: We can see that they take up space, and their weight tells us that they have mass. Gases are also matter; if gases did not take up space, a balloon would stay collapsed rather than inflate when filled with gas.

Solids, liquids, and gases are the three states of matter commonly found on earth (Figure 1). A solid is rigid and possesses a definite shape. A liquid flows and takes the shape of a container, except that it forms a flat or slightly curved upper surface when acted upon by gravity. (In zero gravity, liquids assume a spherical shape.) Both liquid and solid samples have volumes that are very nearly independent of pressure. A gas takes both the shape and volume of its container.

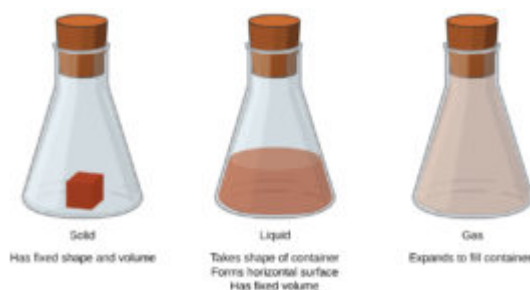


Figure 1 The three most common states or phases of matter are solid, liquid, and gas.

A fourth state of matter, plasma, occurs naturally in the interiors of stars. A plasma is a gaseous state of matter that contains appreciable numbers of electrically charged particles (Figure 2). The presence of these charged particles imparts unique properties to plasmas that justify their classification as a state of matter distinct from gases. In addition to stars, plasmas are found in some other high-temperature environments (both natural and man-made), such as lightning strikes, certain television screens, and specialized analytical instruments used to detect trace amounts of metals.



Figure 2. A plasma torch can be used to cut metal.

(credit: "Hypertherm"/Wikimedia Commons)

Some samples of matter appear to have properties of solids, liquids, and/or gases at the same time. This can occur when the sample is composed of many small pieces. For example, we can pour sand as if it were a liquid because it is composed of many small grains of solid sand. Matter can also have properties of more than one state when it is a mixture, such as with clouds. Clouds appear to behave somewhat like gases, but they are actually mixtures of air (gas) and tiny particles of water (liquid or solid).

The *mass* of an object is a measure of the amount of matter in it. One way to measure an object's mass is to measure the force it takes to accelerate the object. It takes much more force to accelerate a car than a bicycle because the car has much more mass. A more common way to determine the mass of an object is to use a balance to compare its mass with a standard mass.

Although weight is related to mass, it is not the same thing. Weight refers to the force that gravity exerts on an object. This force is directly proportional to the mass of the object. The weight of an object changes as the force of gravity changes, but its mass does not. An astronaut's mass does not change just because she goes to the moon. But her weight on the moon is only one-sixth her earth-bound weight because the moon's gravity is only one-sixth that of the earth's. She may feel "weightless" during her trip when she experiences negligible external forces (gravitational or any other), although she is, of course, never "massless."

The law of conservation of matter summarizes many scientific observations about matter: It states that *there is no detectable change in the total quantity of matter present when matter converts from one type to another (a chemical change) or changes among solid, liquid, or gaseous states (a physical change)*. Brewing beer and the operation of batteries provide examples of the conservation of matter (Figure 3). During the brewing of beer, the ingredients (water, yeast, grains, malt, hops, and sugar) are converted into beer (water, alcohol, carbonation, and flavoring substances) with no actual loss of substance. This is most clearly seen during the bottling process, when glucose turns into ethanol and carbon dioxide, and the total mass of the substances does not change. This can also be seen in a lead-acid car battery: The original substances (lead, lead oxide, and sulfuric acid), which are capable of producing electricity, are changed into other substances (lead sulfate and water) that do not produce electricity, with no change in the actual amount of matter.

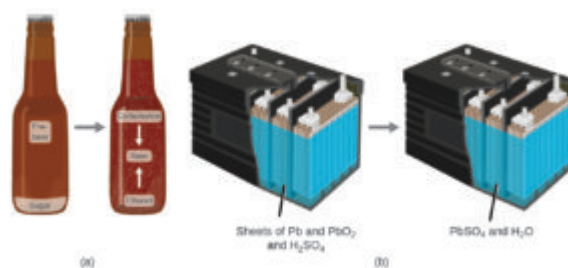


Figure 3. (a) The mass of beer precursor materials is the same as the mass of beer produced: Sugar has become alcohol and carbonation. (b) The mass of the lead, lead oxide plates, and sulfuric acid that goes into the production of electricity is exactly equal to the mass of lead sulfate and water that is formed.

Although this conservation law holds true for all conversions of matter, convincing examples are few and far between because, outside of the controlled conditions in a laboratory, we seldom collect all of the material that is produced during a particular conversion. For example, when you eat, digest, and assimilate food, all of the matter in the original food is preserved. But because some of the matter is incorporated into your body, and much is excreted as various types of waste, it is challenging to verify by measurement.

Atoms and Molecules

An atom is the smallest particle of an element that has the properties of that element and can enter into a chemical combination. Consider the element gold, for example. Imagine cutting a gold nugget in half, then cutting one of the halves in half, and repeating this process until a piece of gold remained that was so small

that it could not be cut in half (regardless of how tiny your knife may be). This minimally sized piece of gold is an atom (from the Greek *atomos*, meaning “indivisible”) (Figure 4). This atom would no longer be gold if it were divided any further.

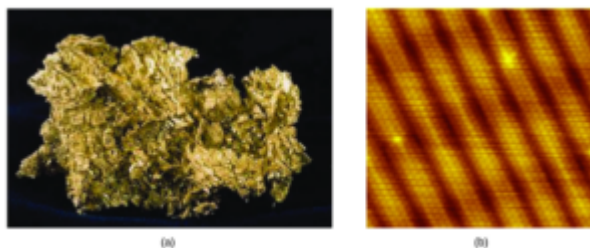


Figure 4. (a) This photograph shows a gold nugget. (b) A scanning-tunneling microscope (STM) can generate views of the surfaces of solids, such as this image of a gold crystal. Each sphere represents one gold atom. (credit a: modification of work by United States Geological Survey; credit b: modification of work by “Erwinrossen”/Wikimedia Commons)

The first suggestion that matter is composed of atoms is attributed to the Greek philosophers Leucippus and Democritus, who developed their ideas in the 5th century BCE. However, it was not until the early nineteenth century that John Dalton (1766–1844), a British schoolteacher with a keen interest in science, supported this hypothesis with quantitative measurements. Since that time, repeated experiments have confirmed many aspects of this hypothesis, and it has become one of the central theories of chemistry. Other aspects of Dalton’s atomic theory are still used but with minor revisions (details of Dalton’s theory are provided in the chapter on atoms and molecules).

An atom is so small that its size is difficult to imagine. One of the smallest things we can see with our unaided eye is a single thread of a spider web: These strands are about 1/10,000 of a centimeter (0.00001 cm) in diameter. Although the cross-section of one strand is almost impossible to see without a microscope, it is huge on an atomic scale. A single carbon atom in the web has a diameter of about 0.000000015 centimeter, and it would take about 7000 carbon atoms to span the diameter of the strand. To put this in perspective, if a carbon atom were the size of a dime, the cross-section of one strand would be larger than a football field, which would require about 150 million carbon atom “dimes” to cover it. (Figure 5) shows increasingly close microscopic and atomic-level views of ordinary cotton.

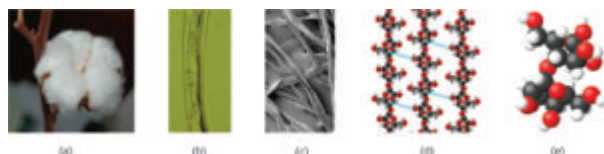


Figure 5. These images provide an increasingly closer view: (a) a cotton boll, (b) a single cotton fiber viewed under an optical microscope (magnified 40 times), (c) an image of a cotton fiber obtained with an electron microscope (much higher magnification than with the optical microscope); and (d and e) atomic-level models of the fiber (spheres of different colors represent atoms of different elements). (credit c: modification of work by “Featheredtar”/Wikimedia Commons)

An atom is so light that its mass is also difficult to imagine. A billion lead atoms (1,000,000,000 atoms) weigh about 3×10^{-13} grams, a mass that is far too light to be weighed on even the world’s most sensitive balances. It would require over 300,000,000,000,000 lead atoms (300 trillion, or 3×10^{14}) to be weighed, and they would weigh only 0.0000001 gram.

It is rare to find collections of individual atoms. Only a few elements, such as the gases helium, neon, and argon, consist of a collection of individual atoms that move about independently of one another. Other elements, such as the gases hydrogen, nitrogen, oxygen, and chlorine, are composed of units that consist

of pairs of atoms (Figure 6). One form of the element phosphorus consists of units composed of four phosphorus atoms. The element sulfur exists in various forms, one of which consists of units composed of eight sulfur atoms. These units are called molecules. A molecule consists of two or more atoms joined by strong forces called chemical bonds. The atoms in a molecule move around as a unit, much like the cans of soda in a six-pack or a bunch of keys joined together on a single key ring. A molecule may consist of two or more identical atoms, as in the molecules found in the elements hydrogen, oxygen, and sulfur, or it may consist of two or more different atoms, as in the molecules found in water. Each water molecule is a unit that contains two hydrogen atoms and one oxygen atom. Each glucose molecule is a unit that contains 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms. Like atoms, molecules are incredibly small and light. If an ordinary glass of water were enlarged to the size of the earth, the water molecules inside it would be about the size of golf balls.

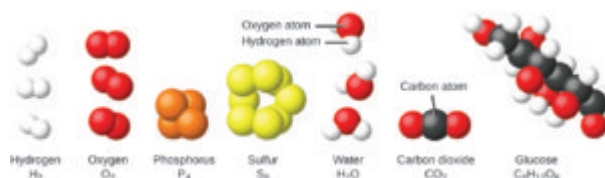


Figure 6. The elements hydrogen, oxygen, phosphorus, and sulfur form molecules consisting of two or more atoms of the same element. The compounds water, carbon dioxide, and glucose consist of combinations of atoms of different elements.

Classifying Matter

We can classify matter into several categories. Two broad categories are mixtures and pure substances. A pure substance has a constant composition. All specimens of a pure substance have exactly the same makeup and properties. Any sample of sucrose (table sugar) consists of 42.1% carbon, 6.5% hydrogen, and 51.4% oxygen by mass. Any sample of sucrose also has the same physical properties, such as melting point, color, and sweetness, regardless of the source from which it is isolated.

We can divide pure substances into two classes: elements and compounds. Pure substances that cannot be broken down into simpler substances by chemical changes are called elements. Iron, silver, gold, aluminum, sulfur, oxygen, and copper are familiar examples of the more than 100 known elements, of which about 90 occur naturally on the earth, and two dozen or so have been created in laboratories.

Pure substances that can be broken down by chemical changes are called compounds. This breakdown may produce either elements or other compounds, or both. Mercury(II) oxide, an orange, crystalline solid, can be broken down by heat into the elements mercury and oxygen (Figure 7). When heated in the absence of air, the compound sucrose is broken down into the element carbon and the compound water. (The initial stage of this process, when the sugar is turning brown, is known as caramelization—this is what imparts the characteristic sweet and nutty flavor to caramel apples, caramelized onions, and caramel). Silver(I) chloride is a white solid that can be broken down into its elements, silver and chlorine, by absorption of light. This property is the basis for the use of this compound in photographic films and photochromic eyeglasses (those with lenses that darken when exposed to light).

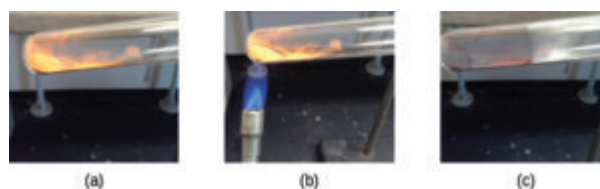


Figure 7. (a) The compound mercury(II) oxide, (b) when heated, (c) decomposes into silvery droplets of liquid mercury and invisible oxygen gas. (credit: modification of work by Paul Flowers)

Many compounds break down when heated. This [site](#) shows the breakdown of mercury oxide, HgO. You can also view an example of the [photochemical decomposition of silver chloride](#) (AgCl), the basis of early photography.

The properties of combined elements are different from those in the free, or uncombined, state. For example, white crystalline sugar (sucrose) is a compound resulting from the chemical combination of the element carbon, which is a black solid in one of its uncombined forms, and the two elements hydrogen and oxygen, which are colorless gases when uncombined. Free sodium, an element that is a soft, shiny, metallic solid, and free chlorine, an element that is a yellow-green gas, combine to form sodium chloride (table salt), a compound that is a white, crystalline solid.

A mixture is composed of two or more types of matter that can be present in varying amounts and can be separated by physical changes, such as evaporation (you will learn more about this later). A mixture with a composition that varies from point to point is called a heterogeneous mixture. Italian dressing is an example of a heterogeneous mixture (Figure 8a). Its composition can vary because we can make it from varying amounts of oil, vinegar, and herbs. It is not the same from point to point throughout the mixture—one drop may be mostly vinegar, whereas a different drop may be mostly oil or herbs because the oil and vinegar separate and the herbs settle. Other examples of heterogeneous mixtures are chocolate chip cookies (we can see the separate bits of chocolate, nuts, and cookie dough) and granite (we can see the quartz, mica, feldspar, and more).

A homogeneous mixture, also called a solution, exhibits a uniform composition and appears visually the same throughout. An example of a solution is a sports drink, consisting of water, sugar, coloring, flavoring, and electrolytes mixed together uniformly (Figure 8b). Each drop of a sports drink tastes the same because each drop contains the same amounts of water, sugar, and other components. Note that the composition of a sports drink can vary—it could be made with somewhat more or less sugar, flavoring, or other components, and still be a sports drink. Other examples of homogeneous mixtures include air, maple syrup, gasoline, and a solution of salt in water.

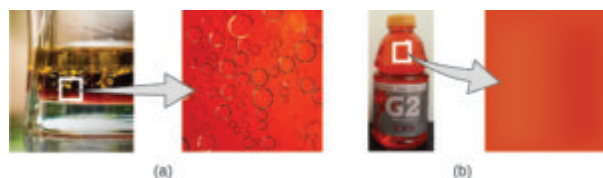


Figure 8. (a) Oil and vinegar salad dressing is a heterogeneous mixture because its composition is not uniform throughout. (b) A commercial sports drink is a homogeneous mixture because its composition is uniform throughout. (credit a “left”: modification of work by John Mayer; credit a “right”: modification of work by Umberto Salvagnin; credit b “left”: modification of work by Jeff Bedford)

Although there are just over 100 elements, tens of millions of chemical compounds result from different combinations of these elements. Each compound has a specific composition and possesses definite chemical and physical properties by which we can distinguish it from all other compounds. And, of course, there are innumerable ways to combine elements and compounds to form different mixtures. A summary of how to distinguish between the various major classifications of matter is shown in (Figure 9).



Figure 9. Depending on its properties, a given substance can be classified as a homogeneous mixture, a heterogeneous mixture, a compound, or an element.

Eleven elements make up about 99% of the earth’s crust and atmosphere (Table 1). Oxygen constitutes nearly one-half and silicon about one-quarter of the total quantity of these elements. A majority of elements on earth are found in chemical combinations with other elements; about one-quarter of the elements are also found in the free state.

Table 1 Elemental Composition of Earth

Element	Symbol	Percent Mass	Element	Symbol	Percent Mass
oxygen	O	49.20	chlorine	Cl	0.19
silicon	Si	25.67	phosphorus	P	0.11
aluminum	Al	7.50	manganese	Mn	0.09
iron	Fe	4.71	carbon	C	0.08
calcium	Ca	3.39	sulfur	S	0.06
sodium	Na	2.63	barium	Ba	0.04
potassium	K	2.40	nitrogen	N	0.03
magnesium	Mg	1.93	fluorine	F	0.03
hydrogen	H	0.87	strontium	Sr	0.02
titanium	Ti	0.58	all others	-	0.47

NOTE: DECOMPOSITION OF WATER / PRODUCTION OF HYDROGEN

Water consists of the elements hydrogen and oxygen combined in a 2 to 1 ratio. Water can be broken down into hydrogen and oxygen gases by the addition of energy. One way to do this is with a battery or power supply, as shown in (Figure 10).

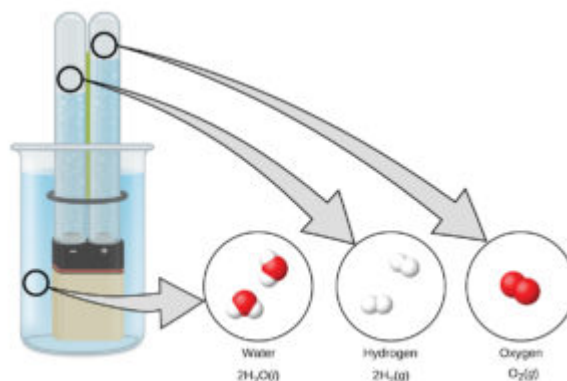


Figure 10. The decomposition of water is shown at the macroscopic, microscopic, and symbolic levels. The battery provides an electric current (microscopic) that decomposes water. At the macroscopic level, the liquid separates into the gases hydrogen (on the left) and oxygen (on the right). Symbolically, this change is presented by showing how liquid H_2O separates into H_2 and O_2 gases.

The breakdown of water involves a rearrangement of the atoms in water molecules into different molecules, each composed of two hydrogen atoms and two oxygen atoms, respectively. Two water molecules form one oxygen molecule and two hydrogen molecules. The representation for what occurs, $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$, will be explored in more depth in later chapters.

The two gases produced have distinctly different properties. Oxygen is not flammable but is required for

combustion of a fuel, and hydrogen is highly flammable and a potent energy source. How might this knowledge be applied in our world? One application involves research into more fuel-efficient transportation. Fuel-cell vehicles (FCV) run on hydrogen instead of gasoline (Figure 11). They are more efficient than vehicles with internal combustion engines, are nonpolluting, and reduce greenhouse gas emissions, making us less dependent on fossil fuels. FCVs are not yet economically viable, however, and current hydrogen production depends on natural gas. If we can develop a process to economically decompose water, or produce hydrogen in another environmentally sound way, FCVs may be the way of the future.

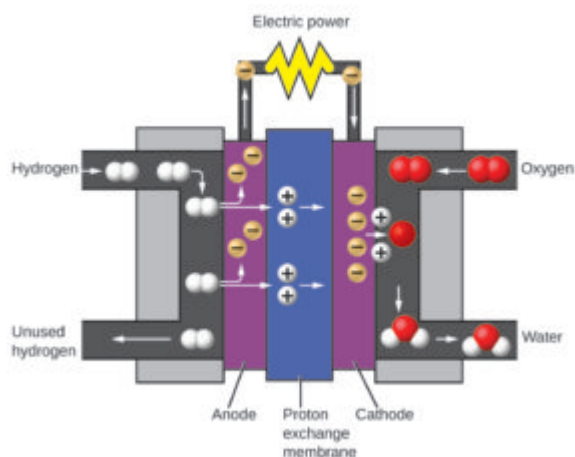


Figure 11 A fuel cell generates electrical energy from hydrogen and oxygen via an electrochemical process and produces only water as the waste product.

CHEMISTRY OF CELL PHONES

Imagine how different your life would be without cell phones (Figure 12) and other smart devices. Cell phones are made from numerous chemical substances, which are extracted, refined, purified, and assembled using an extensive and in-depth understanding of chemical principles. About 30% of the elements that are found in nature are found within a typical smart phone. The case/body/frame consists of a combination of sturdy, durable polymers comprised primarily of carbon, hydrogen, oxygen, and nitrogen [acrylonitrile butadiene styrene (ABS) and polycarbonate thermoplastics], and light, strong, structural metals, such as aluminum, magnesium, and iron. The display screen is made from a specially toughened glass (silica glass strengthened by the addition of aluminum, sodium, and potassium) and coated with a material to make it conductive (such as indium tin oxide). The circuit board uses a semiconductor material (usually silicon); commonly used metals like copper, tin, silver, and gold; and more unfamiliar elements such as yttrium, praseodymium, and gadolinium. The battery relies upon lithium ions and a variety of other materials, including iron, cobalt, copper, polyethylene oxide, and polyacrylonitrile.

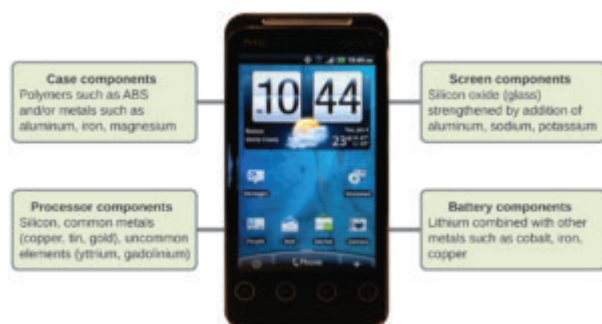


Figure 12 Almost one-third of naturally occurring elements are used to make a cell phone. (credit: modification of work by John Taylor)

Key Concepts and Summary

Matter is anything that occupies space and has mass. The basic building block of matter is the atom, the smallest unit of an element that can enter into combinations with atoms of the same or other elements. In many substances, atoms are combined into molecules. On earth, matter commonly exists in three states: solids, of fixed shape and volume; liquids, of variable shape but fixed volume; and gases, of variable shape and volume. Under high-temperature conditions, matter also can exist as a plasma. Most matter is a mixture: It is composed of two or more types of matter that can be present in varying amounts and can be separated by physical means. Heterogeneous mixtures vary in composition from point to point; homogeneous mixtures have the same composition from point to point. Pure substances consist of only one type of matter. A pure substance can be an element, which consists of only one type of atom and cannot be broken down by a chemical change, or a compound, which consists of two or more types of atoms.

Phases and Classification of Matter Exercises

What properties distinguish solids from liquids? Liquids from gases? Solids from gases? How does a heterogeneous mixture differ from a homogeneous mixture? How are they similar?

How does a homogeneous mixture differ from a pure substance? How are they similar?

How does an element differ from a compound? How are they similar?

How do molecules of elements and molecules of compounds differ? In what ways are they similar?

How does an atom differ from a molecule? In what ways are they similar?

Many of the items you purchase are mixtures of pure compounds. Select three of these commercial products and prepare a list of the ingredients that are pure compounds.

Classify each of the following as an element, a compound, or a mixture: (a) copper (b) water (c) nitrogen (d) sulfur (e) air (f) sucrose (g) a substance composed of molecules each of which contains two iodine atoms (h) gasoline

Classify each of the following as an element, a compound, or a mixture: (a) iron (b) oxygen (c) mercury oxide (d) pancake syrup (e) carbon dioxide (f) a substance composed of molecules each of which contains one hydrogen atom and one chlorine atom (g) baking soda (h) baking powder

A sulfur atom and a sulfur molecule are not identical. What is the difference?

How are the molecules in oxygen gas, the molecules in hydrogen gas, and water molecules similar? How do they differ?

We refer to astronauts in space as weightless, but not without mass. Why?

As we drive an automobile, we don't think about the chemicals consumed and produced.

Prepare a list of the principal chemicals consumed and produced during the operation of an automobile.

Matter is everywhere around us. Make a list by name of fifteen different kinds of matter that you encounter every day. Your list should include (and label at least one example of each) the following: a solid, a liquid, a gas, an element, a compound, a homogeneous mixture, a heterogeneous mixture, and a pure substance.

When elemental iron corrodes it combines with oxygen in the air to ultimately form red brown iron(III) oxide which we call rust. (a) If a shiny iron nail with an initial mass of 23.2 g is weighed after being coated in a layer of rust, would you expect the mass to have increased, decreased, or remained the same? Explain. (b) If the mass of the iron nail increases to 24.1 g, what mass of oxygen combined with the iron?

As stated in the text, convincing examples that demonstrate the law of conservation of matter outside of the laboratory are few and far between. Indicate whether the mass would increase,

decrease, or stay the same for the following scenarios where chemical reactions take place:(a) Exactly one pound of bread dough is placed in a baking tin. The dough is cooked in an oven at 350 °F releasing a wonderful aroma of freshly baked bread during the cooking process. Is the mass of the baked loaf less than, greater than, or the same as the one pound of original dough? Explain.(b) When magnesium burns in air a white flaky ash of magnesium oxide is produced. Is the mass of magnesium oxide less than, greater than, or the same as the original piece of magnesium? Explain.(c) Antoine Lavoisier, the French scientist credited with first stating the law of conservation of matter, heated a mixture of tin and air in a sealed flask to produce tin oxide. Did the mass of the sealed flask and contents decrease, increase, or remain the same after the heating?

Yeast converts glucose to ethanol and carbon dioxide during anaerobic fermentation as depicted in the simple chemical equation here:glucose→ethanol+carbon dioxide

(a) If 200.0 g of glucose is fully converted, what will be the total mass of ethanol and carbon dioxide produced?(b) If the fermentation is carried out in an open container, would you expect the mass of the container and contents after fermentation to be less than, greater than, or the same as the mass of the container and contents before fermentation? Explain.(c) If 97.7 g of carbon dioxide is produced, what mass of ethanol is produced?

Glossary

atom

smallest particle of an element that can enter into a chemical combination

compound

pure substance that can be decomposed into two or more elements

element

substance that is composed of a single type of atom; a substance that cannot be decomposed by a chemical change

gas

state in which matter has neither definite volume nor shape

heterogeneous mixture

combination of substances with a composition that varies from point to point

homogeneous mixture

(also, solution) combination of substances with a composition that is uniform throughout

liquid

state of matter that has a definite volume but indefinite shape

law of conservation of matter

when matter converts from one type to another or changes form, there is no detectable change in the total amount of matter present

mass

fundamental property indicating amount of matter

matter

anything that occupies space and has mass

mixture

matter that can be separated into its components by physical means

molecule

bonded collection of two or more atoms of the same or different elements

plasma

gaseous state of matter containing a large number of electrically charged atoms and/or molecules

pure substance

homogeneous substance that has a constant composition

solid

state of matter that is rigid, has a definite shape, and has a fairly constant volume

weight

force that gravity exerts on an object

Attribution

Physical and Chemical Properties

The characteristics that enable us to distinguish one substance from another are called properties. A physical property is a characteristic of matter that is not associated with a change in its chemical composition. Familiar examples of physical properties include density, color, hardness, melting and boiling points, and electrical conductivity. We can observe some physical properties, such as density and color, without changing the physical state of the matter observed. Other physical properties, such as the melting temperature of iron or the freezing temperature of water, can only be observed as matter undergoes a physical change. A physical change is a change in the state or properties of matter without any accompanying change in its chemical composition (the identities of the substances contained in the matter). We observe a physical change when wax melts, when sugar dissolves in coffee, and when steam condenses into liquid water (Figure 1). Other examples of physical changes include magnetizing and demagnetizing metals (as is done with common antitheft security tags) and grinding solids into powders (which can sometimes yield noticeable changes in color). In each of these examples, there is a change in the physical state, form, or properties of the substance, but no change in its chemical composition.



Figure 1 (a) Wax undergoes a physical change when solid wax is heated and forms liquid wax. (b) Steam condensing inside a cooking pot is a physical change, as water vapor is changed into liquid water. (credit a: modification of work by "95jb14"/Wikimedia Commons; credit b: modification of work by "mjneuby"/Flickr)

The change of one type of matter into another type (or the inability to change) is a chemical property. Examples of chemical properties include flammability, toxicity, acidity, reactivity (many types), and heat of combustion. Iron, for example, combines with oxygen in the presence of water to form rust; chromium does not oxidize (Figure 2). Nitroglycerin is very dangerous because it explodes easily; neon poses almost no hazard because it is very unreactive.



Figure 2 (a) One of the chemical properties of iron is that it rusts; (b) one of the chemical properties of chromium is that it does not. (credit a: modification of work by Tony Hisgett; credit b: modification of work by

To identify a chemical property, we look for a chemical change. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The explosion of nitroglycerin is a chemical change because the gases produced are very different kinds of matter from the original substance. Other examples of chemical changes include reactions that are performed in a lab (such as copper reacting with nitric acid), all forms of combustion (burning), and food being cooked, digested, or rotting (Figure 3).

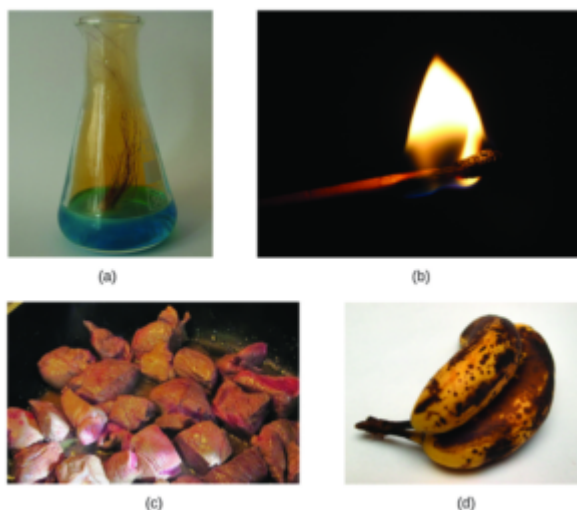


Figure 3 (a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change. (d) A banana turning brown is a chemical change as new, darker (and less tasty) substances form. (credit b: modification of work by Jeff Turner; credit c: modification of work by Gloria Cabada-Leman; credit d: modification of work by Roberto Verzo)

Properties of matter fall into one of two categories. If the property depends on the amount of matter present, it is an extensive property. The mass and volume of a substance are examples of extensive properties; for instance, a gallon of milk has a larger mass and volume than a cup of milk. The value of an extensive property is directly proportional to the amount of matter in question. If the property of a sample of matter does not depend on the amount of matter present, it is an intensive property. Temperature is an example of an intensive property. If the gallon and cup of milk are each at 20 °C (room temperature), when they are combined, the temperature remains at 20 °C. As another example, consider the distinct but related properties of heat and temperature. A drop of hot cooking oil spattered on your arm causes brief, minor discomfort, whereas a pot of hot oil yields severe burns. Both the drop and the pot of oil are at the same temperature (an intensive property), but the pot clearly contains much more heat (extensive property).

NOTE: HAZARD DIAMOND

You may have seen the symbol shown in Figure 4 on containers of chemicals in a laboratory or workplace. Sometimes called a “fire diamond” or “hazard diamond,” this chemical hazard diamond provides valuable information that briefly summarizes the various dangers of which to be aware when working with a particular substance.

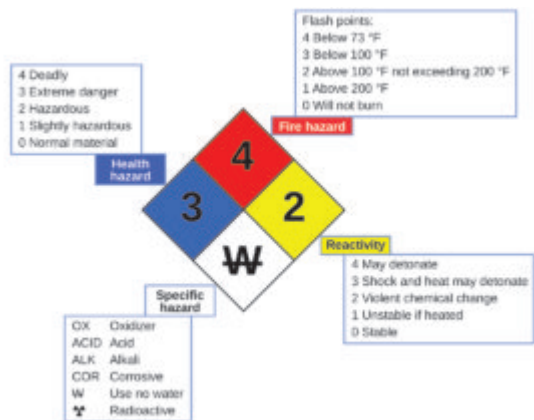


Figure 4 The National Fire Protection Agency (NFPA) hazard diamond summarizes the major hazards of a chemical substance.

The National Fire Protection Agency (NFPA) 704 Hazard Identification System was developed by NFPA to provide safety information about certain substances. The system details flammability, reactivity, health, and other hazards. Within the overall diamond symbol, the top (red) diamond specifies the level of fire hazard (temperature range for flash point). The blue (left) diamond indicates the level of health hazard. The yellow (right) diamond describes reactivity hazards, such as how readily the substance will undergo detonation or a violent chemical change. The white (bottom) diamond points out special hazards, such as if it is an oxidizer (which allows the substance to burn in the absence of air/oxygen), undergoes an unusual or dangerous reaction with water, is corrosive, acidic, alkaline, a biological hazard, radioactive, and so on. Each hazard is rated on a scale from 0 to 4, with 0 being no hazard and 4 being extremely hazardous.

While many elements differ dramatically in their chemical and physical properties, some elements have similar properties. We can identify sets of elements that exhibit common behaviors. For example, many elements conduct heat and electricity well, whereas others are poor conductors. These properties can be used to sort the elements into three classes: metals (elements that conduct well), nonmetals (elements that conduct poorly), and metalloids (elements that have properties of both metals and nonmetals).

The periodic table is a table of elements that places elements with similar properties close together (Figure 5). You will learn more about the periodic table as you continue your study of chemistry.

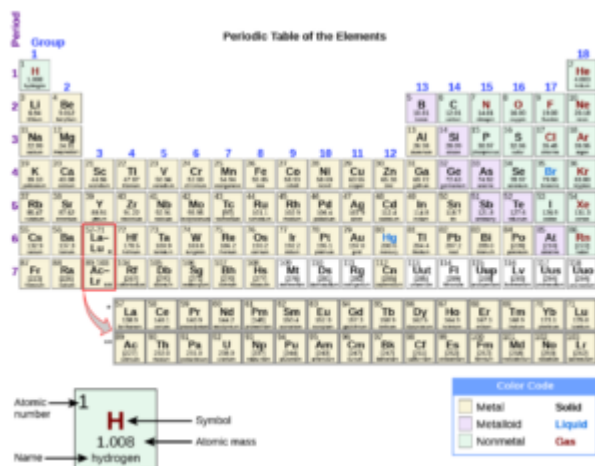


Figure 5 The periodic table shows how elements may be grouped according to certain similar properties. Note the background color denotes whether an element is a metal, metalloid, or nonmetal, whereas the element symbol color indicates whether it is a solid, liquid, or gas.

Key Concepts and Summary

All substances have distinct physical and chemical properties, and may undergo physical or chemical changes. Physical properties, such as hardness and boiling point, and physical changes, such as melting or freezing, do not involve a change in the composition of matter. Chemical properties, such as flammability and acidity, and chemical changes, such as rusting, involve production of matter that differs from that present beforehand.

Measurable properties fall into one of two categories. Extensive properties depend on the amount of matter present, for example, the mass of gold. Intensive properties do not depend on the amount of matter present, for example, the density of gold. Heat is an example of an extensive property, and temperature is an example of an intensive property.

Physical and Chemical Properties Exercises

Classify each of the following changes as physical or chemical:

- (a) condensation of steam
- (b) burning of gasoline
- (c) souring of milk
- (d) dissolving of sugar in water
- (e) melting of gold

Classify each of the following changes as physical or chemical:

- (a) coal burning
- (b) ice melting
- (c) mixing chocolate syrup with milk
- (d) explosion of a firecracker
- (e) magnetizing of a screwdriver

The volume of a sample of oxygen gas changed from 10 mL to 11 mL as the temperature changed. Is this a chemical or physical change?

A 2.0-liter volume of hydrogen gas combined with 1.0 liter of oxygen gas to produce 2.0 liters of water vapor. Does oxygen undergo a chemical or physical change?

Explain the difference between extensive properties and intensive properties.

Identify the following properties as either extensive or intensive.

- (a) volume
- (b) temperature
- (c) humidity
- (d) heat
- (e) boiling point

The density (d) of a substance is an intensive property that is defined as the ratio of its mass (m) to its volume (V).

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad d = \frac{m}{V}$$

Considering that mass and volume are both extensive properties, explain why their ratio, density, is intensive.

Glossary

chemical change

change producing a different kind of matter from the original kind of matter

chemical property

behavior that is related to the change of one kind of matter into another kind of matter

extensive property

property of a substance that depends on the amount of the substance

intensive property

property of a substance that is independent of the amount of the substance

physical change

change in the state or properties of matter that does not involve a change in its chemical composition

physical property

characteristic of matter that is not associated with any change in its chemical composition

Attribution

Measurements

Measurements provide the macroscopic information that is the basis of most of the hypotheses, theories, and laws that describe the behavior of matter and energy in both the macroscopic and microscopic domains of chemistry. Every measurement provides three kinds of information: the size or magnitude of the measurement (a number); a standard of comparison for the measurement (a unit); and an indication of the uncertainty of the measurement. While the number and unit are explicitly represented when a quantity is written, the uncertainty is an aspect of the measurement result that is more implicitly represented and will be discussed later.

The number in the measurement can be represented in different ways, including decimal form and scientific notation. (Scientific notation is also known as exponential notation; a review of this topic can be found in Appendix B.) For example, the maximum takeoff weight of a Boeing 777-200ER airliner is 298,000 kilograms, which can also be written as 2.98×10^5 kg. The mass of the average mosquito is about 0.0000025 kilograms, which can be written as 2.5×10^{-6} kg.

Units, such as liters, pounds, and centimeters, are standards of comparison for measurements. When we buy a 2-liter bottle of a soft drink, we expect that the volume of the drink was measured, so it is two times larger than the volume that everyone agrees to be 1 liter. The meat used to prepare a 0.25-pound hamburger is measured so it weighs one-fourth as much as 1 pound. Without units, a number can be meaningless, confusing, or possibly life threatening. Suppose a doctor prescribes phenobarbital to control a patient's seizures and states a dosage of "100" without specifying units. Not only will this be confusing to the medical professional giving the dose, but the consequences can be dire: 100 mg given three times per day can be effective as an anticonvulsant, but a single dose of 100 g is more than 10 times the lethal amount.

We usually report the results of scientific measurements in SI units, an updated version of the metric system, using the units listed in Table 1. Other units can be derived from these base units. The standards for these units are fixed by international agreement, and they are called the International System of Units or SI Units (from the French, *Le Système International d'Unités*). SI units have been used by the United States National Institute of Standards and Technology (NIST) since 1964.

Table 1		
Base Units of the SI System		
Property Measured	Name of Unit	Symbol of Unit
length	meter	m
mass	kilogram	kg
time	second	s
temperature	kelvin	K

Table 1

Base Units of the SI System		
Property Measured	Name of Unit	Symbol of Unit
electric current	ampere	A
amount of substance	mole	mol
luminous intensity	candela	cd

Sometimes we use units that are fractions or multiples of a base unit. Ice cream is sold in quarts (a familiar, non-SI base unit), pints (0.5 quart), or gallons (4 quarts). We also use fractions or multiples of units in the SI system, but these fractions or multiples are always powers of 10. Fractional or multiple SI units are named using a prefix and the name of the base unit. For example, a length of 1000 meters is also called a kilometer because the prefix *kilo* means “one thousand,” which in scientific notation is 10^3 (1 kilometer = 1000 m = 10^3 m). The prefixes used and the powers to which 10 are raised are listed in Table 2.

Table 2

Common Unit Prefixes			
Prefix	Symbol	Factor	Example
femto	f	10^{-15}	1 femtosecond (fs) = 1×10^{-15} s (0.000000000000001 s)
pico	p	10^{-12}	1 picometer (pm) = 1×10^{-12} m (0.000000000001 m)
nano	n	10^{-9}	4 nanograms (ng) = 4×10^{-9} g (0.000000004 g)
micro	μ	10^{-6}	1 microliter (μ L) = 1×10^{-6} L (0.000001 L)
milli	m	10^{-3}	2 millimoles (mmol) = 2×10^{-3} mol (0.002 mol)
centi	c	10^{-2}	7 centimeters (cm) = 7×10^{-2} m (0.07 m)
deci	d	10^{-1}	1 deciliter (dL) = 1×10^{-1} L (0.1 L)
kilo	k	10^3	1 kilometer (km) = 1×10^3 m (1000 m)
mega	M	10^6	3 megahertz (MHz) = 3×10^6 Hz (3,000,000 Hz)
giga	G	10^9	8 gigayears (Gyr) = 8×10^9 yr (8,000,000,000 Gyr)
tera	T	10^{12}	5 terawatts (TW) = 5×10^{12} W (5,000,000,000,000 W)

Need a refresher or more practice with scientific notation? Visit this [site](#) to go over the basics of scientific notation.

SI Base Units

The initial units of the metric system, which eventually evolved into the SI system, were established in

France during the French Revolution. The original standards for the meter and the kilogram were adopted there in 1799 and eventually by other countries. This section introduces four of the SI base units commonly used in chemistry. Other SI units will be introduced in subsequent chapters.

Length

The standard unit of length in both the SI and original metric systems is the meter (m). A meter was originally specified as 1/10,000,000 of the distance from the North Pole to the equator. It is now defined as the distance light in a vacuum travels in 1/299,792,458 of a second. A meter is about 3 inches longer than a yard (Figure 1); one meter is about 39.37 inches or 1.094 yards. Longer distances are often reported in kilometers (1 km = 1000 m = 10^3 m), whereas shorter distances can be reported in centimeters (1 cm = 0.01 m = 10^{-2} m) or millimeters (1 mm = 0.001 m = 10^{-3} m).

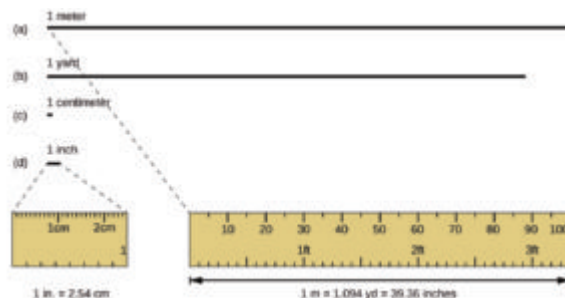


Figure 1 The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.

Mass

The standard unit of mass in the SI system is the kilogram (kg). A kilogram was originally defined as the mass of a liter of water (a cube of water with an edge length of exactly 0.1 meter). It is now defined by a certain cylinder of platinum-iridium alloy, which is kept in France (Figure 2). Any object with the same mass as this cylinder is said to have a mass of 1 kilogram. One kilogram is about 2.2 pounds. The gram (g) is exactly equal to 1/1000 of the mass of the kilogram (10^{-3} kg).



Figure 2 This replica prototype kilogram is housed at the National Institute of Standards and Technology (NIST) in Maryland. (credit: National Institutes of Standards and Technology)

Temperature

Temperature is an intensive property. The SI unit of temperature is the kelvin (K). The IUPAC convention is to use kelvin (all lowercase) for the word, K (uppercase) for the unit symbol, and neither the word “degree” nor the degree symbol (°). The degreeCelsius (°C) is also allowed in the SI system, with both the word “degree” and the degree symbol used for Celsius measurements. Celsius degrees are the same magnitude as those of kelvin, but the two scales place their zeros in different places. Water freezes at 273.15 K (0 °C) and boils at 373.15 K (100 °C) by definition, and normal human body temperature is approximately 310 K (37 °C). The conversion between these two units and the Fahrenheit scale will be discussed later in this chapter.

Time

The SI base unit of time is the second (s). Small and large time intervals can be expressed with the appropriate prefixes; for example, 3 microseconds = 0.000003 s = 3×10^6 s. Alternatively, hours, days, and years can be used.

Derived SI Units

We can derive many units from the seven SI base units. For example, we can use the base unit of length to define a unit of volume, and the base units of mass and length to define a unit of density.

Volume

Volume is the measure of the amount of space occupied by an object. The standard SI unit of volume is defined by the base unit of length (Figure 3). The standard volume is a cubic meter (m^3), a cube with an edge length of exactly one meter. To dispense a cubic meter of water, we could build a cubic box with edge lengths of exactly one meter. This box would hold a cubic meter of water or any other substance.

A more commonly used unit of volume is derived from the decimeter (0.1 m, or 10 cm). A cube with edge lengths of exactly one decimeter contains a volume of one cubic decimeter (dm^3). A **liter (L)** is the more common name for the cubic decimeter. One liter is about 1.06 quarts.

A cubic centimeter (cm^3) is the volume of a cube with an edge length of exactly one centimeter. The abbreviation **cc** (for **cubic centimeter**) is often used by health professionals. A cubic centimeter is also called a **milliliter (mL) and is 1/1000 of a liter**.

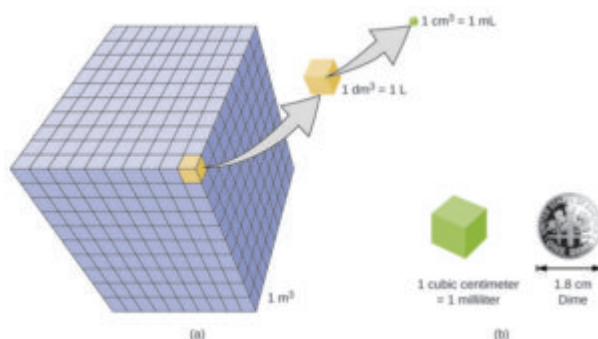


Figure 3 (a) The relative volumes are shown for cubes of $1 m^3$, $1 dm^3$ (1 L), and $1 cm^3$ (1 mL) (not to scale). (b) The diameter of a dime is compared relative to the edge length of a $1\text{-}cm^3$ (1-mL) cube. (image credit: OpenStax College, CC BY)

Density

We use the mass and volume of a substance to determine its density. Thus, the units of density are defined by the base units of mass and length.

The **density** of a substance is the ratio of the mass of a sample of the substance to its volume. The SI unit for density is the kilogram per cubic meter (kg/m^3). For many situations, however, this is an inconvenient unit, and we often use grams per cubic centimeter (g/cm^3) for the densities of solids and liquids, and grams per liter (g/L) for gases. Although there are exceptions, most liquids and solids have densities that range from about 0.7 g/cm^3 (the density of gasoline) to 19 g/cm^3 (the density of gold). The density of air is about 1.2 g/L . Table 3 shows the densities of some common substances.

Densities of Common Substances		
Solids	Liquids	Gases (at 25 °C and 1 atm)
ice (at 0 °C) 0.92 g/cm^3	water 1.0 g/cm^3	dry air 1.20 g/L
oak (wood) $0.60\text{--}0.90 \text{ g/cm}^3$	ethanol 0.79 g/cm^3	oxygen 1.31 g/L
iron 7.9 g/cm^3	acetone 0.79 g/cm^3	nitrogen 1.14 g/L
copper 9.0 g/cm^3	glycerin 1.26 g/cm^3	carbon dioxide 1.80 g/L
lead 11.3 g/cm^3	olive oil 0.92 g/cm^3	helium 0.16 g/L
silver 10.5 g/cm^3	gasoline $0.70\text{--}0.77 \text{ g/cm}^3$	neon 0.83 g/L
gold 19.3 g/cm^3	mercury 13.6 g/cm^3	radon 9.1 g/L

While there are many ways to determine the density of an object, perhaps the most straightforward method involves separately finding the mass and volume of the object, and then dividing the mass of the sample by its volume. In the following example, the mass is found directly by weighing, but the volume is found indirectly through length measurements.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

Example 1: Calculation of Density

Gold—in bricks, bars, and coins—has been a form of currency for centuries. In order to swindle people into paying for a brick of gold without actually investing in a brick of gold, people have considered filling the centers of hollow gold bricks with lead to fool buyers into thinking that the entire brick is gold. It does not work: Lead is a dense substance, but its density is not as great as that of gold, 19.3 g/cm^3 . What is the density of lead if a cube of lead has an edge length of 2.00 cm and a mass of 90.7 g ?

Solution:

The density of a substance can be calculated by dividing its mass by its volume. The volume of a cube is calculated by cubing the edge length.

$$\text{volume of lead cube} = 2.00 \text{ cm} \times 2.00 \text{ cm} \times 2.00 \text{ cm} = 8.00 \text{ cm}^3$$

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{90.7 \text{ g}}{8.00 \text{ cm}^3} = \frac{11.3 \text{ g}}{1.00 \text{ cm}^3} = 11.3 \text{ g/cm}^3$$

(We will discuss the reason for rounding to the first decimal place in the next section.)

Check Your Learning:

- To three decimal places, what is the volume of a cube (cm^3) with an edge length of 0.843 cm ?
- If the cube in part (a) is copper and has a mass of 5.34 g , what is the density of copper to two decimal places?

Answer:

(a) 0.599 cm³; (b) 8.91 g/cm³

To learn more about the relationship between mass, volume, and density, use this [interactive simulator](#) to explore the density of different materials, like wood, ice, brick, and aluminum.

Example 2: **Using Displacement of Water to Determine Density**

This [PhET simulation](#) illustrates another way to determine density, using displacement of water. Determine the density of the red and yellow blocks.

Solution:

When you open the density simulation and select Same Mass, you can choose from several 5.00-kg colored blocks that you can drop into a tank containing 100.00 L water. The yellow block floats (it is less dense than water), and the water level rises to 105.00 L. While floating, the yellow block displaces 5.00 L water, an amount equal to the weight of the block. The red block sinks (it is more dense than water, which has density = 1.00 kg/L), and the water level rises to 101.25 L.

The red block therefore displaces 1.25 L water, an amount equal to the volume of the block. The density of the red block is:

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{5.00 \text{ kg}}{1.25 \text{ L}} = 4.00 \text{ kg/L}$$

Note that since the yellow block is not completely submerged, you cannot determine its density from this information. But if you hold the yellow block on the bottom of the tank, the water level rises to 110.00 L, which means that it now displaces 10.00 L water, and its density can be found:

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{5.00 \text{ kg}}{10.00 \text{ L}} = 0.500 \text{ kg/L}$$

Check Your Learning:

Remove all of the blocks from the water and add the green block to the tank of water, placing it approximately in the middle of the tank. Determine the density of the green block.

Answer:

2.00 kg/L

Key Concepts and Summary

Measurements provide quantitative information that is critical in studying and practicing chemistry. Each measurement has an amount, a unit for comparison, and an uncertainty. Measurements can be represented in either decimal or scientific notation. Scientists primarily use the SI (International System) or metric systems. We use base SI units such as meters, seconds, and kilograms, as well as derived units, such as liters (for volume) and g/cm³ (for density). In many cases, we find it convenient to use unit prefixes that yield fractional and multiple units, such as microseconds (10⁻⁶ seconds) and megahertz (10⁶ hertz), respectively.

Key Equations

- $\text{density} = \frac{\text{mass}}{\text{volume}}$

Measurements Exercises

Is a meter about an inch, a foot, a yard, or a mile?

Indicate the SI base units or derived units that are appropriate for the following measurements:

- (a) the length of a marathon race (26 miles 385 yards)
- (b) the mass of an automobile
- (c) the volume of a swimming pool
- (d) the speed of an airplane
- (e) the density of gold
- (f) the area of a football field
- (g) the maximum temperature at the South Pole on April 1, 1913

Indicate the SI base units or derived units that are appropriate for the following measurements:

- (a) the mass of the moon
- (b) the distance from Dallas to Oklahoma City
- (c) the speed of sound
- (d) the density of air
- (e) the temperature at which alcohol boils
- (f) the area of the state of Delaware
- (g) the volume of a flu shot or a measles vaccination

Give the name and symbol of the prefixes used with SI units to indicate multiplication by the following exact quantities.

- (a) 10^3
- (b) 10^{-2}
- (c) 0.1
- (d) 10^{-3}
- (e) 1,000,000
- (f) 0.000001

Give the name of the prefix and the quantity indicated by the following symbols that are used with SI base units.

- (a) c
- (b) d
- (c) G
- (d) k
- (e) m
- (f) n
- (g) p
- (h) T

A large piece of jewelry has a mass of 132.6 g. A graduated cylinder initially contains 48.6 mL water. When the jewelry is submerged in the graduated cylinder, the total volume increases to 61.2 mL.

- (a) Determine the density of this piece of jewelry.
- (b) Assuming that the jewelry is made from only one substance, what substance is it likely to be? Explain.

Visit this [PhET density simulation](#) and select the Same Volume Blocks.

- (a) What are the mass, volume, and density of the yellow block?
- (b) What are the mass, volume and density of the red block?
- (c) List the block colors in order from smallest to largest mass.
- (d) List the block colors in order from lowest to highest density.

(e) How are mass and density related for blocks of the same volume?

Visit this [PhET density simulation](#) and select Custom Blocks and then My Block.

(a) Enter mass and volume values for the block such that the mass in kg is *less than* the volume in L. What does the block do? Why? Is this always the case when mass < volume?

(b) Enter mass and volume values for the block such that the mass in kg is *more than* the volume in L. What does the block do? Why? Is this always the case when mass > volume?

(c) How would (a) and (b) be different if the liquid in the tank were ethanol instead of water?

(d) How would (a) and (b) be different if the liquid in the tank were mercury instead of water?

Visit this [PhET density simulation](#) and select Mystery Blocks.

(a) Pick one of the Mystery Blocks and determine its mass, volume, density, and its likely identity.

(b) Pick a different Mystery Block and determine its mass, volume, density, and its likely identity.

(c) Order the Mystery Blocks from least dense to most dense. Explain.

Glossary

Celsius (°C)

unit of temperature; water freezes at 0 °C and boils at 100 °C on this scale

cubic centimeter (cm³ or cc)

volume of a cube with an edge length of exactly 1 cm

cubic meter (m³)

SI unit of volume

density

ratio of mass to volume for a substance or object

kelvin (K)

SI unit of temperature; 273.15 K = 0 °C

kilogram (kg)

standard SI unit of mass; 1 kg = approximately 2.2 pounds

length

measure of one dimension of an object

liter (L)

(also, cubic decimeter) unit of volume; 1 L = 1,000 cm³

meter (m)

standard metric and SI unit of length; 1 m = approximately 1.094 yards

milliliter (mL)

1/1,000 of a liter; equal to 1 cm³

second (s)

SI unit of time

SI units (International System of Units)

standards fixed by international agreement in the International System of Units (Le Système International d'Unités)

unit

standard of comparison for measurements

volume

amount of space occupied by an object

Attribution

Measurement Uncertainty, Accuracy, and Precision

Counting is the only type of measurement that is free from uncertainty, provided the number of objects being counted does not change while the counting process is underway. The result of such a counting measurement is an example of an exact number. If we count eggs in a carton, we know *exactly* how many eggs the carton contains. The numbers of defined quantities are also exact. By definition, 1 foot is exactly 12 inches, 1 inch is exactly 2.54 centimeters, and 1 gram is exactly 0.001 kilogram. Quantities derived from measurements other than counting, however, are uncertain to varying extents due to practical limitations of the measurement process used.

Significant Figures in Measurement

The numbers of measured quantities, unlike defined or directly counted quantities, are not exact. To measure the volume of liquid in a graduated cylinder, you should make a reading at the bottom of the meniscus, the lowest point on the curved surface of the liquid.



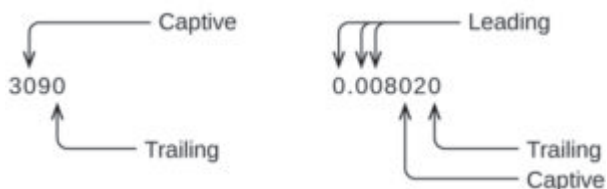
Figure 1 To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter, and then make a reading (estimate) at the bottom of the meniscus.

Refer to the illustration in Figure 1. The bottom of the meniscus in this case clearly lies between the 21 and 22 markings, meaning the liquid volume is *certainly* greater than 21 mL but less than 22 mL. The meniscus appears to be a bit closer to the 22-mL mark than to the 21-mL mark, and so a reasonable estimate of the liquid's volume would be 21.6 mL. In the number 21.6, then, the digits 2 and 1 are certain, but the 6 is an estimate. Some people might estimate the meniscus position to be equally distant from each of the markings and estimate the tenth-place digit as 5, while others may think it to be even closer to the 22-mL mark and estimate this digit to be 7. Note that it would be pointless to attempt to estimate a digit for the hundredths place, given that the tenths-place digit is uncertain. In general, numerical scales such as the one on this graduated cylinder will permit measurements to one-tenth of the smallest scale division. The scale in this case has 1-mL divisions, and so volumes may be measured to the nearest 0.1 mL.

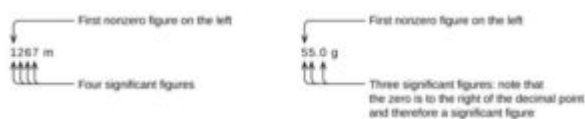
This concept holds true for all measurements, even if you do not actively make an estimate. If you place a quarter on a standard electronic balance, you may obtain a reading of 6.72 g. The digits 6 and 7 are certain, and the 2 indicates that the mass of the quarter is likely between 6.71 and 6.73 grams. The quarter

weighs *about* 6.72 grams, with a nominal uncertainty in the measurement of ± 0.01 gram. If we weigh the quarter on a more sensitive balance, we may find that its mass is 6.723 g. This means its mass lies between 6.722 and 6.724 grams, an uncertainty of 0.001 gram. Every measurement has some uncertainty, which depends on the device used (and the user's ability). All of the digits in a measurement, including the uncertain last digit, are called significant figures or significant digits. Note that zero may be a measured value; for example, if you stand on a scale that shows weight to the nearest pound and it shows "120," then the 1 (hundreds), 2 (tens) and 0 (ones) are all significant (measured) values.

Whenever you make a measurement properly, all the digits in the result are significant. But what if you were analyzing a reported value and trying to determine what is significant and what is not? Well, for starters, all nonzero digits are significant, and it is only zeros that require some thought. We will use the terms "leading," "trailing," and "captive" for the zeros and will consider how to deal with them.



Starting with the first nonzero digit on the left, count this digit and all remaining digits to the right. This is the number of significant figures in the measurement unless the last digit is a trailing zero lying to the left of the decimal point.



Captive zeros result from measurement and are therefore always significant. Leading zeros, however, are never significant—they merely tell us where the decimal point is located.



The leading zeros in this example are not significant. We could use exponential notation (as described in Appendix B) and express the number as 8.32407×10^{-3} ; then the number 8.32407 contains all of the significant figures, and 10^{-3} locates the decimal point.

The number of significant figures is uncertain in a number that ends with a zero to the left of the decimal point location. The zeros in the measurement 1,300 grams could be significant or they could simply indicate where the decimal point is located. The ambiguity can be resolved with the use of exponential notation: 1.3×10^3 (two significant figures), 1.30×10^3 (three significant figures, if the tens place was measured), or 1.300×10^3 (four significant figures, if the ones place was also measured). In cases where only the decimal-formatted number is available, it is prudent to assume that all trailing zeros are not significant.



When determining significant figures, be sure to pay attention to reported values and think about the measurement and significant figures in terms of what is reasonable or likely when evaluating whether the value makes sense. For example, the official January 2014 census reported the resident population of the

US as 317,297,725. Do you think the US population was correctly determined to the reported nine significant figures, that is, to the exact number of people? People are constantly being born, dying, or moving into or out of the country, and assumptions are made to account for the large number of people who are not actually counted. Because of these uncertainties, it might be more reasonable to expect that we know the population to within perhaps a million or so, in which case the population should be reported as 3.17×10^8 people.

Significant Figures in Calculations

A second important principle of uncertainty is that results calculated from a measurement are at least as uncertain as the measurement itself. We must take the uncertainty in our measurements into account to avoid misrepresenting the uncertainty in calculated results. One way to do this is to report the result of a calculation with the correct number of significant figures, which is determined by the following three rules for rounding numbers:

When we add or subtract numbers, we should round the result to the same number of decimal places as the number with the least number of decimal places (the least precise value in terms of addition and subtraction).

When we multiply or divide numbers, we should round the result to the same number of digits as the number with the least number of significant figures (the least precise value in terms of multiplication and division).

If the digit to be dropped (the one immediately to the right of the digit to be retained) is less than 5, we “round down” and leave the retained digit unchanged; if it is more than 5, we “round up” and increase the retained digit by 1; if the dropped digit *is* 5, we round up or down, whichever yields an even value for the retained digit. (The last part of this rule may strike you as a bit odd, but it’s based on reliable statistics and is aimed at avoiding any bias when dropping the digit “5,” since it is equally close to both possible values of the retained digit.)

The following examples illustrate the application of this rule in rounding a few different numbers to three significant figures:

- 0.028675 rounds “up” to 0.0287 (the dropped digit, 7, is greater than 5)
- 18.3384 rounds “down” to 18.3 (the dropped digit, 3, is lesser than 5)
- 6.8752 rounds “up” to 6.88 (the dropped digit is 5, and the retained digit is even)
- 92.85 rounds “down” to 92.8 (the dropped digit is 5, and the retained digit is even)

Let’s work through these rules with a few examples.

Example 1: Rounding Numbers

Round the following to the indicated number of significant figures:

- (a) 31.57 (to two significant figures)
- (b) 8.1649 (to three significant figures)
- (c) 0.051065 (to four significant figures)
- (d) 0.90275 (to four significant figures)

Solution:

- (a) 31.57 rounds “up” to 32 (the dropped digit is 5, and the retained digit is even)
- (b) 8.1649 rounds “down” to 8.16 (the dropped digit, 4, is lesser than 5)
- (c) 0.051065 rounds “down” to 0.05106 (the dropped digit is 5, and the retained digit is even)
- (d) 0.90275 rounds “up” to 0.9028 (the dropped digit is 5, and the retained digit is even)

Check Your Learning:

Round the following to the indicated number of significant figures:

- (a) 0.424 (to two significant figures)

- (b) 0.0038661 (to three significant figures)
- (c) 421.25 (to four significant figures)
- (d) 28,683.5 (to five significant figures)

Answer:

- (a) 0.42; (b) 0.00387; (c) 421.2; (d) 28,684

Example 2: Addition and Subtraction with Significant Figures

Rule: When we add or subtract numbers, we should round the result to the same number of decimal places as the number with the least number of decimal places (i.e., the least precise value in terms of addition and subtraction).

- (a) Add 1.0023 g and 4.383 g.
- (b) Subtract 421.23 g from 486 g.

Solution:

- (a) $1.0023\text{ g} + 4.383\text{ g} = 5.3853\text{ g}$
Answer is 5.385 g (round to the thousandths place; three decimal places)
- (b) $486\text{ g} - 421.23\text{ g} = 64.77\text{ g}$
Answer is 65 g (round to the ones place; no decimal places)



Check Your Learning:

- (a) Add 2.334 mL and 0.31 mL.
- (b) Subtract 55.8752 m from 56.533 m.

Answer:

- (a) 2.64 mL; (b) 0.658 m

Example 3: Multiplication and Division with Significant Figures

Rule: When we multiply or divide numbers, we should round the result to the same number of digits as the number with the least number of significant figures (the least precise value in terms of multiplication and division).

- (a) Multiply 0.6238 cm by 6.6 cm.
- (b) Divide 421.23 g by 486 mL.

Solution:

- (a) $0.6238\text{ cm} \times 6.6\text{ cm} \rightarrow$ result is 4.1 cm^2 (round to two significant figures)
four significant figures \times two significant figures \rightarrow two significant figures answer
- (b) $\frac{421.23\text{ g}}{486\text{ mL}} = 0.86728\dots\text{g/mL} \rightarrow$ result is 0.867 g/mL (round to three significant figures)
five significant figures
three significant figures \rightarrow three significant figures answer

Check Your Learning:

- (a) Multiply 2.334 cm and 0.320 cm.
- (b) Divide 55.8752 m by 56.53 s.

Answer:

- (a) 0.747 cm²
- (b) 0.9884 m/s

In the midst of all these technicalities, it is important to keep in mind the reason why we use significant figures and rounding rules—to correctly represent the certainty of the values we report and to ensure that a calculated result is not represented as being more certain than the least certain value used in the calculation.

Example 4: Calculation with Significant Figures

One common bathtub is 13.44 dm long, 5.920 dm wide, and 2.54 dm deep. Assume that the tub is rectangular and calculate its approximate volume in liters.

Solution:

$$\begin{aligned}V &= l \times w \times d \\&= 13.44 \text{ dm} \times 5.920 \text{ dm} \times 2.54 \text{ dm} \\&= 202.09459\dots\text{dm}^3 \text{ (value from calculator)} \\&= 202 \text{ dm}^3, \text{ or } 202 \text{ L (answer rounded to three significant figures)}\end{aligned}$$

Check Your Learning:

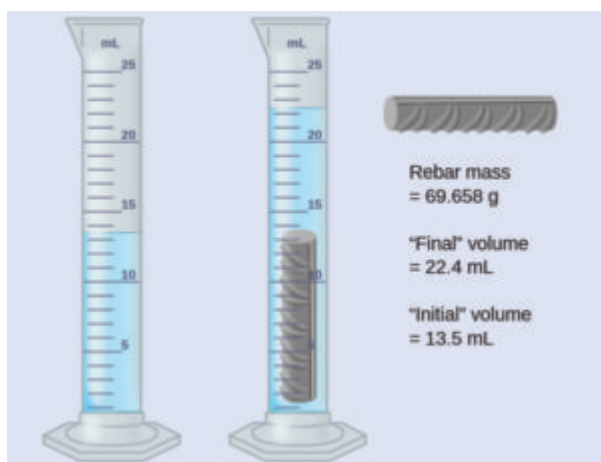
What is the density of a liquid with a mass of 31.1415 g and a volume of 30.13 cm³?

Answer:

1.034 g/mL

Example 5: Experimental Determination of Density Using Water Displacement

A piece of rebar is weighed and then submerged in a graduated cylinder partially filled with water, with results as shown.



- (a) Use these values to determine the density of this piece of rebar.
 (b) Rebar is mostly iron. Does your result in (a) support this statement? How?

Solution:

The volume of the piece of rebar is equal to the volume of the water displaced:

$$\text{volume} = 22.4 \text{ mL} - 13.5 \text{ mL} = 8.9 \text{ mL} = 8.9 \text{ cm}^3$$

(rounded to the nearest 0.1 mL, per the rule for addition and subtraction)

The density is the mass-to-volume ratio:

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{69.658 \text{ g}}{8.9 \text{ cm}^3} = 7.8 \text{ g/cm}^3$$

(rounded to two significant figures, per the rule for multiplication and division)

From Table 3 in section 1.4 the density of iron is 7.9 g/cm^3 , very close to that of rebar, which lends some support to the fact that rebar is mostly iron.

Check Your Learning:

An irregularly shaped piece of a shiny yellowish material is weighed and then submerged in a graduated cylinder, with results as shown.



- (a) Use these values to determine the density of this material.
 (b) Do you have any reasonable guesses as to the identity of this material? Explain your reasoning.

Answer:

(a) 19 g/cm^3 ; (b) It is likely gold; the right appearance for gold and very close to the density given for gold in Table 3 in section 1.4.

Accuracy and Precision

Scientists typically make repeated measurements of a quantity to ensure the quality of their findings and to know both the precision and the accuracy of their results. Measurements are said to be precise if they yield very similar results when repeated in the same manner. A measurement is considered accurate if it yields a result that is very close to the true or accepted value. Precise values agree with each other; accurate values agree with a true value. These characterizations can be extended to other contexts, such as the results of an archery competition (Figure 2).

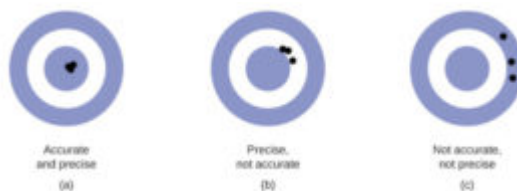


Figure 2 (a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise.

Suppose a quality control chemist at a pharmaceutical company is tasked with checking the accuracy and precision of three different machines that are meant to dispense 10 ounces (296 mL) of cough syrup into storage bottles. She proceeds to use each machine to fill five bottles and then carefully determines the actual volume dispensed, obtaining the results tabulated in Table 1.

Volume (mL) of Cough Medicine Delivered by 10-oz (296 mL) Dispensers		
Dispenser #1	Dispenser #2	Dispenser #3
283.3	298.3	296.1
284.1	294.2	295.9
283.9	296.0	296.1
284.0	297.8	296.0
284.1	293.9	296.1

Considering these results, she will report that dispenser #1 is precise (values all close to one another, within a few tenths of a milliliter) but not accurate (none of the values are close to the target value of 296 mL, each being more than 10 mL too low). Results for dispenser #2 represent improved accuracy (each volume is less than 3 mL away from 296 mL) but worse precision (volumes vary by more than 4 mL). Finally, she can report that dispenser #3 is working well, dispensing cough syrup both accurately (all volumes within 0.1 mL of the target volume) and precisely (volumes differing from each other by no more than 0.2 mL).

Key Concepts and Summary

Quantities can be exact or measured. Measured quantities have an associated uncertainty that is represented by the number of significant figures in the measurement. The uncertainty of a calculated value depends on the uncertainties in the values used in the calculation and is reflected in how the value is rounded. Measured values can be accurate (close to the true value) and/or precise (showing little variation when measured repeatedly).

Measurement Uncertainty, Accuracy, and Precision Exercises

Express each of the following numbers in exponential notation with correct significant figures:

- (a) 704
- (b) 0.03344
- (c) 547.9
- (d) 22086
- (e) 1000.00
- (f) 0.0000000651
- (g) 0.007157

Indicate whether each of the following can be determined exactly or must be measured with some degree of uncertainty:

- (a) the number of eggs in a basket
- (b) the mass of a dozen eggs
- (c) the number of gallons of gasoline necessary to fill an automobile gas tank
- (d) the number of cm in 2 m
- (e) the mass of a textbook
- (f) the time required to drive from San Francisco to Kansas City at an average speed of 53 mi/h

Indicate whether each of the following can be determined exactly or must be measured with some degree of uncertainty:

- (a) the number of seconds in an hour
- (b) the number of pages in this book
- (c) the number of grams in your weight
- (d) the number of grams in 3 kilograms
- (e) the volume of water you drink in one day
- (f) the distance from San Francisco to Kansas City

How many significant figures are contained in each of the following measurements?

- (a) 38.7 g
- (b) 2×10^{18} m
- (c) 3,486,002 kg
- (d) 9.74150×10^{-4} J
- (e) 0.0613 cm^3
- (f) 17.0 kg
- (g) 0.01400 g/mL

How many significant figures are contained in each of the following measurements?

- (a) 53 cm
- (b) 2.05×10^8 m
- (c) 86,002 J
- (d) 9.740×10^4 m/s
- (e) 10.0613 m^3
- (f) 0.17 g/mL
- (g) 0.88400 s

The following quantities were reported on the labels of commercial products. Determine the number of significant figures in each.

- (a) 0.0055 g active ingredients
- (b) 12 tablets
- (c) 3% hydrogen peroxide
- (d) 5.5 ounces
- (e) 473 mL
- (f) 1.75% bismuth
- (g) 0.001% phosphoric acid

(h) 99.80% inert ingredients

Round off each of the following numbers to two significant figures:

- (a) 0.436
- (b) 9.000
- (c) 27.2
- (d) 135
- (e) 1.497×10^{-3}
- (f) 0.445

Round off each of the following numbers to two significant figures:

- (a) 517
- (b) 86.3
- (c) 6.382×10^3
- (d) 5.0008
- (e) 22.497
- (f) 0.885

Perform the following calculations and report each answer with the correct number of significant figures.

- (a) 628×342
- (b) $(5.63 \times 10^2) \times (7.4 \times 10^3)$
- (c) $\frac{28.0}{13.483}$
- (d) 8119×0.000023
- (e) $14.98 + 27,340 + 84.7593$
- (f) $42.7 + 0.259$

Perform the following calculations and report each answer with the correct number of significant figures.

- (a) 62.8×34
- (b) $0.147 + 0.0066 + 0.012$
- (c) $38 \times 95 \times 1.792$
- (d) $15 - 0.15 - 0.6155$
- (e) $8.78 \times \left(\frac{0.0500}{0.478}\right)$
- (f) $140 + 7.68 + 0.014$
- (g) $28.7 - 0.0483$
- (h) $\frac{(88.5 - 87.57)}{45.13}$

Consider the results of the archery contest shown in this figure.

- (a) Which archer is most precise?
- (b) Which archer is most accurate?
- (c) Who is both least precise and least accurate?

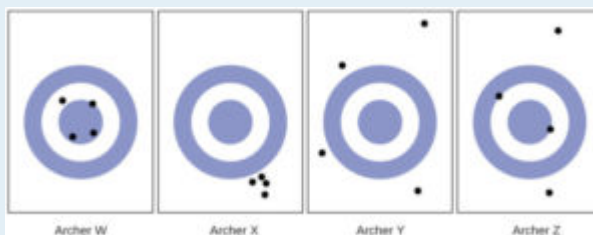


Figure 2 (a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate nor precise.

Classify the following sets of measurements as accurate, precise, both, or neither.

- (a) Checking for consistency in the weight of chocolate chip cookies: 17.27 g, 13.05 g, 19.46 g, 16.92 g
- (b) Testing the volume of a batch of 25-mL pipettes: 27.02 mL, 26.99 mL, 26.97 mL, 27.01 mL
- (c) Determining the purity of gold: 99.9999%, 99.9998%, 99.9998%, 99.9999%

Glossary

accuracy

how closely a measurement aligns with a correct value

exact number

number derived by counting or by definition

precision

how closely a measurement matches the same measurement when repeated

rounding

procedure used to ensure that calculated results properly reflect the uncertainty in the measurements used in the calculation

significant figures

(also, significant digits) all of the measured digits in a determination, including the uncertain last digit

uncertainty

estimate of amount by which measurement differs from true value

Attribution

Mathematical Treatment of Measurement Results

It is often the case that a quantity of interest may not be easy (or even possible) to measure directly but instead must be calculated from other directly measured properties and appropriate mathematical relationships. For example, consider measuring the average speed of an athlete running sprints. This is typically accomplished by measuring the *time* required for the athlete to run from the starting line to the finish line, and the *distance* between these two lines, and then computing *speed* from the equation that relates these three properties:

$$\text{speed} = \frac{\text{distance}}{\text{time}}$$

An Olympic-quality sprinter can run 100 m in approximately 10 s, corresponding to an average speed of

$$\frac{100\text{m}}{10\text{s}} = 10\text{ m/s}$$

Note that this simple arithmetic involves dividing the numbers of each measured quantity to yield the number of the computed quantity ($100/10 = 10$) and likewise dividing the units of each measured quantity to yield the unit of the computed quantity ($\text{m/s} = \text{m/s}$). Now, consider using this same relation to predict the time required for a person running at this speed to travel a distance of 25 m. The same relation between the three properties is used, but in this case, the two quantities provided are a speed (10 m/s) and a distance (25 m). To yield the sought property, time, the equation must be rearranged appropriately:

$$\text{time} = \frac{\text{distance}}{\text{speed}}$$

The time can then be computed as:

$$\frac{25\text{ m}}{10\text{ m/s}} = 2.5\text{ s}$$

Again, arithmetic on the numbers ($25/10 = 2.5$) was accompanied by the same arithmetic on the units ($\text{m/m/s} = \text{s}$) to yield the number and unit of the result, 2.5 s. Note that, just as for numbers, when a unit is divided by an identical unit (in this case, m/m), the result is “1”—or, as commonly phrased, the units “cancel.”

These calculations are examples of a versatile mathematical approach known as dimensional analysis (or the factor-label method). Dimensional analysis is based on this premise: *the units of quantities must be subjected to the same mathematical operations as their associated numbers*. This method can be applied to computations ranging from simple unit conversions to more complex, multi-step calculations involving several different quantities.

Conversion Factors and Dimensional Analysis

A ratio of two equivalent quantities expressed with different measurement units can be used as a unit conversion factor. For example, the lengths of 2.54 cm and 1 in. are equivalent (by definition), and so a unit conversion factor may be derived from the ratio,

$$\frac{2.54\text{cm}}{1\text{in.}} \quad (2.54\text{ cm} = 1\text{ in.}) \quad \text{or} \quad 2.54\text{ cm/in.}$$

Several other commonly used conversion factors are given in Table 1.

Table 1

Common Conversion Factors		
Length	Volume	Mass
1 m = 1.0936 yd	1 L = 1.0567 qt	1 kg = 2.2046 lb
1 in. = 2.54 cm (exact)	1 qt = 0.94635 L	1 lb = 453.59 g
1 km = 0.62137 mi	1 ft ³ = 28.317 L	1 (avoirdupois) oz = 28.349 g
1 mi = 1609.3 m	1 tbsp = 14.787 mL	1 (troy) oz = 31.103 g

When we multiply a quantity (such as distance given in inches) by an appropriate unit conversion factor, we convert the quantity to an equivalent value with different units (such as distance in centimeters). For example, a basketball player's vertical jump of 34 inches can be converted to centimeters by:

$$34 \text{ in.} \times \frac{2.54 \text{ cm}}{1 \text{ in.}} = 86 \text{ cm}$$

Since this simple arithmetic involves *quantities*, the premise of dimensional analysis requires that we multiply both *numbers and units*. The numbers of these two quantities are multiplied to yield the number of the product quantity, 86, whereas the units are multiplied to yield $\frac{\text{in.} \times \text{cm}}{\text{in.}}$. Just as for numbers, a ratio of identical units is also numerically equal to one, $\frac{\text{in.}}{\text{in.}} = 1$, and the unit product thus simplifies to *cm*. (When identical units divide to yield a factor of 1, they are said to “cancel.”) Using dimensional analysis, we can determine that a unit conversion factor has been set up correctly by checking to confirm that the original unit will cancel, and the result will contain the sought (converted) unit.

Example 1: Using a Unit Conversion Factor

The mass of a competition frisbee is 125 g. Convert its mass to ounces using the unit conversion factor derived from the relationship 1 oz = 28.349 g (Table 1).

Solution:

If we have the conversion factor, we can determine the mass in kilograms using an equation similar the one used for converting length from inches to centimeters.

$$x \text{ oz} = 125 \text{ g} \times \text{unit conversion factor}$$

We write the unit conversion factor in its two forms:

$$\frac{1 \text{ oz}}{28.349 \text{ g}} \text{ and } \frac{28.349 \text{ g}}{1 \text{ oz}}$$

The correct unit conversion factor is the ratio that cancels the units of grams and leaves ounces.

$$x \text{ oz} = 125 \text{ g} \times \frac{1 \text{ oz}}{28.349 \text{ g}}$$

$$= \left(\frac{125}{28.349} \right) \text{ oz}$$

$$= 4.41 \text{ oz (three significant figures)}$$

Check Your Learning:

Convert a volume of 9.345 qt to liters.

Answer:

$$8.844 \text{ L}$$

Beyond simple unit conversions, the factor-label method can be used to solve more complex problems

involving computations. Regardless of the details, the basic approach is the same—all the *factors* involved in the calculation must be appropriately oriented to insure that their *labels* (units) will appropriately cancel and/or combine to yield the desired unit in the result. This is why it is referred to as the factor-label method. As your study of chemistry continues, you will encounter many opportunities to apply this approach.

Example 2: Computing Quantities from Measurement Results and Known Mathematical Relations

What is the density of common antifreeze in units of g/mL? A 4.00-qt sample of the antifreeze weighs 9.26 lb.

Solution:

Since $\text{density} = \frac{\text{mass}}{\text{volume}}$, we need to divide the mass in grams by the volume in milliliters. In general: the number of units of B = the number of units of A unit conversion factor. The necessary conversion factors are given in Table 1 : 1 lb = 453.59 g; 1 L = 1.0567 qt; 1 L = 1,000 mL. We can convert mass from pounds to grams in one step:

$$9.26 \text{ lb} \times \frac{453.59 \text{ g}}{1 \text{ lb}} = 4.20 \times 10^3 \text{ g}$$

We need to use two steps to convert volume from quarts to milliliters.

1. Convert quarts to liters

$$4.00 \text{ qt} \times \frac{1 \text{ L}}{1.0567 \text{ qt}} = 3.78 \text{ L}$$

2. Convert liters to milliliters

$$3.78 \text{ L} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 3.78 \times 10^3 \text{ mL}$$

Then, $\text{density} = \frac{4.20 \times 10^3 \text{ g}}{3.78 \times 10^3 \text{ mL}} = 1.11 \text{ g/mL}$

Alternatively, the calculation could be set up in a way that uses three unit conversion factors sequentially as follows:

$$\frac{9.26 \text{ lb}}{4.00 \text{ qt}} \times \frac{453.59 \text{ g}}{1 \text{ lb}} \times \frac{1.0567 \text{ qt}}{1 \text{ L}} \times \frac{1 \text{ L}}{1000 \text{ mL}} = 1.11 \text{ g/mL}$$

Check Your Learning:

What is the volume in liters of 1.000 oz, given that 1 L = 1.0567 qt and 1 qt = 32 oz (exactly)?

Answer:

$$2.956 \times 10^{-2} \text{ L}$$

Example 3: Computing Quantities from Measurement Results and Known Mathematical Relations

While being driven from Philadelphia to Atlanta, a distance of about 1250 km, a 2014 Lamborghini Aventador Roadster uses 213 L gasoline.

- What (average) fuel economy, in miles per gallon, did the Roadster get during this trip?
- If gasoline costs \$3.80 per gallon, what was the fuel cost for this trip?

Solution:

(a) We first convert distance from kilometers to miles:

$$1250 \text{ km} \times \frac{0.62137 \text{ mi}}{1 \text{ km}} = 777 \text{ mi}$$

and then convert volume from liters to gallons:

$$213 \text{ L} \times \frac{1.0567 \text{ qt}}{1 \text{ L}} \times \frac{1 \text{ gal}}{4 \text{ qt}} = 56.3 \text{ gal}$$

Then,

$$\text{(average) mileage} = \frac{777 \text{ mi}}{56.3 \text{ gal}} = 13.8 \text{ miles/gallon} = 13.8 \text{ mpg}$$

Alternatively, the calculation could be set up in a way that uses all the conversion factors sequentially,

as follows:

$$\frac{1250\text{ km}}{213\text{ L}} \times \frac{0.62137\text{ mi}}{1\text{ km}} \times \frac{1\text{ L}}{1.0567\text{ qt}} \times \frac{4\text{ qt}}{1\text{ gal}} = 13.8\text{ mpg}$$

(b) Using the previously calculated volume in gallons, we find:

$$56.3\text{ gal} \times \frac{\$3.80}{1\text{ gal}} = \$214$$

Check Your Learning:

A Toyota Prius Hybrid uses 59.7 L gasoline to drive from San Francisco to Seattle, a distance of 1300 km (two significant digits).

(a) What (average) fuel economy, in miles per gallon, did the Prius get during this trip?

(b) If gasoline costs \$3.90 per gallon, what was the fuel cost for this trip?

Answer:

(a) 51 mpg; (b) \$62

Conversion of Temperature Units

We use the word temperature to refer to the hotness or coldness of a substance. One way we measure a change in temperature is to use the fact that most substances expand when their temperature increases and contract when their temperature decreases. The mercury or alcohol in a common glass thermometer changes its volume as the temperature changes. Because the volume of the liquid changes more than the volume of the glass, we can see the liquid expand when it gets warmer and contract when it gets cooler.

To mark a scale on a thermometer, we need a set of reference values: Two of the most commonly used are the freezing and boiling temperatures of water at a specified atmospheric pressure. On the Celsius scale, 0 °C is defined as the freezing temperature of water and 100 °C as the boiling temperature of water. The space between the two temperatures is divided into 100 equal intervals, which we call degrees. On the Fahrenheit scale, the freezing point of water is defined as 32 °F and the boiling temperature as 212 °F. The space between these two points on a Fahrenheit thermometer is divided into 180 equal parts (degrees).

Defining the Celsius and Fahrenheit temperature scales as described in the previous paragraph results in a slightly more complex relationship between temperature values on these two scales than for different units of measure for other properties. Most measurement units for a given property are directly proportional to one another ($y = mx$). Using familiar length units as one example:

$$\text{length in feet} = \left(\frac{1\text{ ft}}{12\text{ in}}\right) \times \text{length in inches}$$

where y = length in feet, x = length in inches, and the proportionality constant, m , is the conversion factor. The Celsius and Fahrenheit temperature scales, however, do not share a common zero point, and so the relationship between these two scales is a linear one rather than a proportional one ($y = mx + b$). Consequently, converting a temperature from one of these scales into the other requires more than simple multiplication by a conversion factor, m , it also must take into account differences in the scales' zero points (b).

The linear equation relating Celsius and Fahrenheit temperatures is easily derived from the two temperatures used to define each scale. Representing the Celsius temperature as x and the Fahrenheit temperature as y , the slope, m , is computed to be:

$$m = \frac{\Delta y}{\Delta x} = \frac{212^\circ\text{F} - 32^\circ\text{F}}{100^\circ\text{C} - 0^\circ\text{C}} = \frac{180^\circ\text{F}}{100^\circ\text{C}} = \frac{9^\circ\text{F}}{5^\circ\text{C}}$$

The y-intercept of the equation, b , is then calculated using either of the equivalent temperature pairs, (100 °C, 212 °F) or (0 °C, 32 °F), as:

$$b = y - mx = 32^\circ\text{F} - \frac{9^\circ\text{F}}{5^\circ\text{C}} \times 0^\circ\text{C} = 32^\circ\text{F}$$

The equation relating the temperature scales is then:

$$T_{\circ\text{F}} = \left(\frac{9^{\circ}\text{F}}{5^{\circ}\text{C}} \times T_{\circ\text{C}}\right) + 32_{\circ\text{C}}$$

An abbreviated form of this equation that omits the measurement units is:

$$T_{\circ\text{F}} = \frac{9}{5} \times T_{\circ\text{C}} + 32$$

Rearrangement of this equation yields the form useful for converting from Fahrenheit to Celsius:

$$T_{\circ\text{C}} = \frac{5}{9}(T_{\circ\text{F}} - 32)$$

As mentioned earlier in this chapter, the SI unit of temperature is the kelvin (K). Unlike the Celsius and Fahrenheit scales, the kelvin scale is an absolute temperature scale in which 0 (zero) K corresponds to the lowest temperature that can theoretically be achieved. The early 19th-century discovery of the relationship between a gas's volume and temperature suggested that the volume of a gas would be zero at -273.15°C . In 1848, British physicist William Thompson, who later adopted the title of Lord Kelvin, proposed an absolute temperature scale based on this concept (further treatment of this topic is provided in this text's chapter on gases).

The freezing temperature of water on this scale is 273.15 K and its boiling temperature 373.15 K. Notice the numerical difference in these two reference temperatures is 100, the same as for the Celsius scale, and so the linear relation between these two temperature scales will exhibit a slope of 1 K/ $^{\circ}\text{C}$. Following the same approach, the equations for converting between the kelvin and Celsius temperature scales are derived to be:

$$T_{\text{K}} = T_{\circ\text{C}} + 273.15$$

$$T_{\circ\text{C}} = T_{\text{K}} - 273.15$$

The 273.15 in these equations has been determined experimentally, so it is not exact. Figure 1 shows the relationship among the three temperature scales. Recall that we do not use the degree sign with temperatures on the kelvin scale.

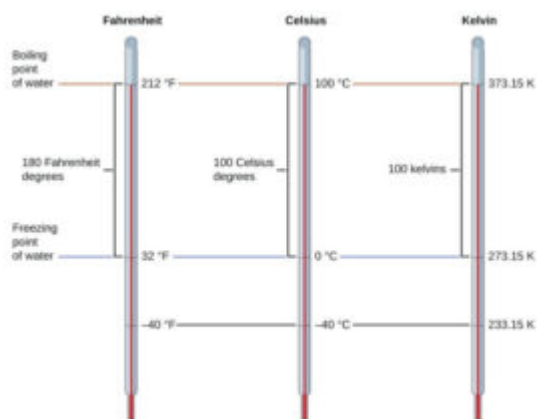
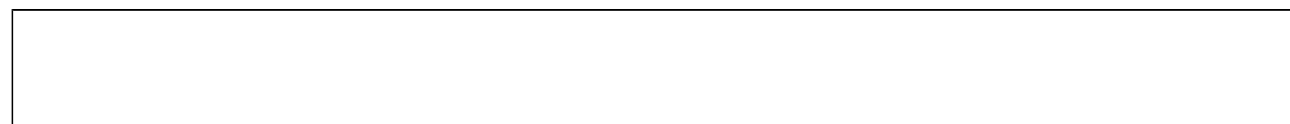


Figure 1 The Fahrenheit, Celsius, and kelvin temperature scales are compared.

Although the kelvin (absolute) temperature scale is the official SI temperature scale, Celsius is commonly used in many scientific contexts and is the scale of choice for nonscience contexts in almost all areas of the world. Very few countries (the U.S. and its territories, the Bahamas, Belize, Cayman Islands, and Palau) still use Fahrenheit for weather, medicine, and cooking.



Example 4: Conversion from Celsius

Normal body temperature has been commonly accepted as 37.0 °C (although it varies depending on time of day and method of measurement, as well as among individuals). What is this temperature on the kelvin scale and on the Fahrenheit scale?

$$K = ^\circ\text{C} + 273.15 = 37.0 + 273.15 = 310.2 \text{ K}$$

$$^\circ\text{F} = \frac{9}{5} ^\circ\text{C} + 32.0 = \left(\frac{9}{5} \times 37.0\right) + 32.0 = 66.6 + 32.0 = 98.6 ^\circ\text{F}$$

Check Your Learning:

Convert 80.92 °C to K and °F.

Answer:

354.07 K, 177.7 °F

Example 5: Conversion from Fahrenheit

Baking a ready-made pizza calls for an oven temperature of 450 °F. If you are in Europe, and your oven thermometer uses the Celsius scale, what is the setting? What is the kelvin temperature?

Solution:

$$^\circ\text{C} = \frac{5}{9} (^\circ\text{F} - 32) = \frac{5}{9} (450 - 32) = \frac{5}{9} \times 418 = 232 ^\circ\text{C} \rightarrow \text{set oven to } 230 ^\circ\text{C} \text{ (two significant figures)}$$

$$K = ^\circ\text{C} + 273.15 = 230 + 273 = 503 \text{ K} \rightarrow 5.0 \times 10^2 \text{ K (two significant figures)}$$

Check Your Learning:

Convert 50 °F to °C and K.

Answer:

10 °C, 280 K

Key Concepts and Summary

Measurements are made using a variety of units. It is often useful or necessary to convert a measured quantity from one unit into another. These conversions are accomplished using unit conversion factors, which are derived by simple applications of a mathematical approach called the factor-label method or dimensional analysis. This strategy is also employed to calculate sought quantities using measured quantities and appropriate mathematical relations.

Key Equations

- $T_{^\circ\text{C}} = \frac{5}{9} \times T_{^\circ\text{F}} - 32$
- $T_{^\circ\text{F}} = \frac{9}{5} \times T_{^\circ\text{C}} + 32$
- $T_{\text{K}} = ^\circ\text{C} + 273.15$
- $T_{^\circ\text{C}} = \text{K} - 273.15$

Mathematical Treatment of Measurement Results Exercises

Write conversion factors (as ratios) for the number of:

- (a) yards in 1 meter
- (b) liters in 1 liquid quart
- (c) pounds in 1 kilogram

Write conversion factors (as ratios) for the number of:

- (a) kilometers in 1 mile
- (b) liters in 1 cubic foot
- (c) grams in 1 ounce

The label on a soft drink bottle gives the volume in two units: 2.0 L and 67.6 fl oz. Use this information to derive a conversion factor between the English and metric units. How many significant figures can you justify in your conversion factor?

The label on a box of cereal gives the mass of cereal in two units: 978 grams and 34.5 oz. Use this information to find a conversion factor between the English and metric units. How many significant figures can you justify in your conversion factor?

Soccer is played with a round ball having a circumference between 27 and 28 in. and a weight between 14 and 16 oz. What are these specifications in units of centimeters and grams?

A woman's basketball has a circumference between 28.5 and 29.0 inches and a maximum weight of 20 ounces (two significant figures). What are these specifications in units of centimeters and grams?

How many milliliters of a soft drink are contained in a 12.0-oz can?

A barrel of oil is exactly 42 gal. How many liters of oil are in a barrel?

The diameter of a red blood cell is about 3×10^{-4} in. What is its diameter in centimeters?

The distance between the centers of the two oxygen atoms in an oxygen molecule is 1.21×10^{-8} cm. What is the distance in inches?

Is a 197-lb weight lifter light enough to compete in a class limited to those weighing 90 kg or less?

A very good 197-lb weight lifter lifted 192 kg in a move called the clean and jerk. What was the mass of the weight lifted in pounds?

Many medical laboratory tests are run using 5.0 μ L blood serum. What is this volume in milliliters?

If an aspirin tablet contains 325 mg aspirin, how many grams of aspirin does it contain?

Use scientific (exponential) notation to express the following quantities in terms of the SI base units in Table 1 in section 1.4.

- (a) 0.13 g
- (b) 232 Gg
- (c) 5.23 pm
- (d) 86.3 mg
- (e) 37.6 cm
- (f) 54 μ m
- (g) 1 Ts
- (h) 27 ps
- (i) 0.15 mK

Complete the following conversions between SI units

- (a) 612 g = _____ mg
- (b) 8.160 m = _____ cm
- (c) 3779 μ g = _____ g
- (d) 781 mL = _____ L
- >(e) 4.18 kg = _____ g
- (f) 27.8 m = _____ km
- (g) 0.13 mL = _____ L

(h) 1738 km = _____ m

(i) 1.9 Gg = _____ g

Gasoline is sold by the liter in many countries. How many liters are required to fill a 12.0-gal gas tank?

Milk is sold by the liter in many countries. What is the volume of exactly 1/2 gal of milk in liters?

A long ton is defined as exactly 2240 lb. What is this mass in kilograms?

Make the conversion indicated in each of the following:

(a) the men's world record long jump, 29 ft 4 $\frac{1}{4}$ in., to meters

(b) the greatest depth of the ocean, about 6.5 mi, to kilometers

(c) the area of the state of Oregon, 96,981 mi², to square kilometers

(d) the volume of 1 gill (exactly 4 oz) to milliliters

(e) the estimated volume of the oceans, 330,000,000 mi³, to cubic kilometers.

(f) the mass of a 3525-lb car to kilograms

(g) the mass of a 2.3-oz egg to grams

Make the conversion indicated in each of the following:

(a) the length of a soccer field, 120 m (three significant figures), to feet

(b) the height of Mt. Kilimanjaro, at 19,565 ft the highest mountain in Africa, to kilometers

(c) the area of an 8.5 t 11-inch sheet of paper in cm²

(d) the displacement volume of an automobile engine, 161 in.³, to liters

(e) the estimated mass of the atmosphere, 5.6 t 10¹⁵ tons, to kilograms

(f) the mass of a bushel of rye, 32.0 lb, to kilograms

(g) the mass of a 5.00-grain aspirin tablet to milligrams (1 grain = 0.00229 oz)

Many chemistry conferences have held a 50-Trillion Angstrom Run (two significant figures).

How long is this run in kilometers and in miles? (1 Å = 1 × 10⁻¹⁰ m)

A chemist's 50-Trillion Angstrom Run (see exercise 22) would be an archeologist's 10,900 cubit run. How long is one cubit in meters and in feet? (1 Å = 1 × 10⁻⁸ cm)

The gas tank of a certain luxury automobile holds 22.3 gallons according to the owner's manual. If the density of gasoline is 0.8206 g/mL, determine the mass in kilograms and pounds of the fuel in a full tank.

As an instructor is preparing for an experiment, he requires 225 g phosphoric acid. The only container readily available is a 150-mL Erlenmeyer flask. Is it large enough to contain the acid, whose density is 1.83 g/mL?

To prepare for a laboratory period, a student lab assistant needs 125 g of a compound. A bottle containing 1/4 lb is available. Did the student have enough of the compound?

A chemistry student is 159 cm tall and weighs 45.8 kg. What is her height in inches and weight in pounds?

In a recent Grand Prix, the winner completed the race with an average speed of 229.8 km/h. What was his speed in miles per hour, meters per second, and feet per second?

Solve these problems about lumber dimensions.

(a) To describe to a European how houses are constructed in the US, the dimensions of "two-by-four" lumber must be converted into metric units. The thickness × width × length dimensions are 1.50 in. × 3.50 in. × 8.00 ft in the US. What are the dimensions in cm × cm × m?

(b) This lumber can be used as vertical studs, which are typically placed 16.0 in. apart. What is that distance in centimeters?

The mercury content of a stream was believed to be above the minimum considered safe—1 part per billion (ppb) by weight. An analysis indicated that the concentration was 0.68 parts per billion. What quantity of mercury in grams was present in 15.0 L of the water, the density

of which is 0.998 g/ml? $\left(1 \text{ ppb Hg} = \frac{1 \text{ ng Hg}}{1 \text{ g water}}\right)$

Calculate the density of aluminum if 27.6 cm³ has a mass of 74.6 g.

Osmium is one of the densest elements known. What is its density if 2.72 g has a volume of

0.121 cm³?

Calculate these masses.

(a) What is the mass of 6.00 cm³ of mercury, density = 13.5939 g/cm³?

(b) What is the mass of 25.0 mL octane, density = 0.702 g/cm³?

Calculate these masses.

(a) What is the mass of 4.00 cm³ of sodium, density = 0.97 g/cm³?

(b) What is the mass of 125 mL gaseous chlorine, density = 3.16 g/L?

Calculate these volumes.

(a) What is the volume of 25 g iodine, density = 4.93 g/cm³?

(b) What is the volume of 3.28 g gaseous hydrogen, density = 0.089 g/L?

Calculate these volumes.

(a) What is the volume of 11.3 g graphite, density = 2.25 g/cm³?

(b) What is the volume of 39.657 g bromine, density = 2.928 g/cm³?

Convert the boiling temperature of gold, 2966 °C, into degrees Fahrenheit and kelvin.

Convert the temperature of scalding water, 54 °C, into degrees Fahrenheit and kelvin.

Convert the temperature of the coldest area in a freezer, -10 °F, to degrees Celsius and kelvin.

Convert the temperature of dry ice, -77 °C, into degrees Fahrenheit and kelvin.

Convert the boiling temperature of liquid ammonia, -28.1 °F, into degrees Celsius and kelvin.

The label on a pressurized can of spray disinfectant warns against heating the can above 130 °F. What are the corresponding temperatures on the Celsius and kelvin temperature scales?

The weather in Europe was unusually warm during the summer of 1995. The TV news reported temperatures as high as 45 °C. What was the temperature on the Fahrenheit scale?

Glossary

dimensional analysis

(also, factor-label method) versatile mathematical approach that can be applied to computations ranging from simple unit conversions to more complex, multi-step calculations involving several different quantities

Fahrenheit

unit of temperature; water freezes at 32 °F and boils at 212 °F on this scale

unit conversion factor

ratio of equivalent quantities expressed with different units; used to convert from one unit to a different unit

Attribution

Energy Basics

Chemical changes and their accompanying changes in energy are important parts of our everyday world (Figure 1). The macronutrients in food (proteins, fats, and carbohydrates) undergo metabolic reactions that provide the energy to keep our bodies functioning. We burn a variety of fuels (gasoline, natural gas, coal) to produce energy for transportation, heating, and the generation of electricity. Industrial chemical reactions use enormous amounts of energy to produce raw materials (such as iron and aluminum). Energy is then used to manufacture those raw materials into useful products, such as cars, skyscrapers, and bridges.



Figure 1 The energy involved in chemical changes is important to our daily lives: (a) A cheeseburger for lunch provides the energy you need to get through the rest of the day; (b) the combustion of gasoline provides the energy that moves your car (and you) between home, work, and school; and (c) coke, a processed form of coal, provides the energy needed to convert iron ore into iron, which is essential for making many of the products we use daily. (credit a: modification of work by “Pink Sherbet Photography”/Flickr; credit b: modification of work by Jeffery Turner; credit)

Over 90% of the energy we use comes originally from the sun. Every day, the sun provides the earth with almost 10,000 times the amount of energy necessary to meet all of the world’s energy needs for that day. Our challenge is to find ways to convert and store incoming solar energy so that it can be used in reactions or chemical processes that are both convenient and nonpolluting. Plants and many bacteria capture solar energy through photosynthesis. We release the energy stored in plants when we burn wood or plant products such as ethanol. We also use this energy to fuel our bodies by eating food that comes directly from plants or from animals that got their energy by eating plants. Burning coal and petroleum also releases stored solar energy: These fuels are fossilized plant and animal matter.

This chapter will introduce the basic ideas of an important area of science concerned with the amount of heat absorbed or released during chemical and physical changes—an area called thermochemistry. The concepts introduced in this chapter are widely used in almost all scientific and technical fields. Food scientists use them to determine the energy content of foods. Biologists study the energetics of living organisms, such as the metabolic combustion of sugar into carbon dioxide and water. The oil, gas, and transportation industries, renewable energy providers, and many others endeavor to find better methods to produce energy for our commercial and personal needs. Engineers strive to improve energy efficiency, find better ways to heat and cool our homes, refrigerate our food and drinks, and meet the energy and cooling needs of computers and electronics, among other applications. Understanding thermochemical principles is essential for chemists, physicists, biologists, geologists, every type of engineer, and just about anyone who studies or does any kind of science.

Energy

Energy can be defined as the capacity to supply heat or do work. One type of work (w) is the process of causing matter to move against an opposing force. For example, we do work when we inflate a bicycle tire—we move matter (the air in the pump) against the opposing force of the air already in the tire.

Like matter, energy comes in different types. One scheme classifies energy into two types: potential energy, the energy an object has because of its relative position, composition, or condition, and kinetic energy, the energy that an object possesses because of its motion. Water at the top of a waterfall or dam has potential energy because of its position; when it flows downward through generators, it has kinetic energy that can be used to do work and produce electricity in a hydroelectric plant (Figure 2). A battery has potential energy because the chemicals within it can produce electricity that can do work.



Figure 2 (a) Water that is higher in elevation, for example, at the top of Victoria Falls, has a higher potential energy than water at a lower elevation. As the water falls, some of its potential energy is converted into kinetic energy. (b) If the water flows through generators at the bottom of a dam, such as the Hoover Dam shown here, its kinetic energy is converted into electrical energy. (credit a: modification of work by Steve Jurvetson; credit b: modification of work by “curimedia”/Wikimedia commons)

Energy can be converted from one form into another, but all of the energy present before a change occurs always exists in some form after the change is completed. This observation is expressed in the law of conservation of energy: during a chemical or physical change, energy can be neither created nor destroyed, although it can be changed in form. (This is also one version of the first law of thermodynamics, as you will learn later.)

When one substance is converted into another, there is always an associated conversion of one form of energy into another. Heat is usually released or absorbed, but sometimes the conversion involves light, electrical energy, or some other form of energy. For example, chemical energy (a type of potential energy) is stored in the molecules that compose gasoline. When gasoline is combusted within the cylinders of a car’s engine, the rapidly expanding gaseous products of this chemical reaction generate mechanical energy (a type of kinetic energy) when they move the cylinders’ pistons.

According to the law of conservation of matter (seen in an earlier chapter), there is no detectable change in the total amount of matter during a chemical change. When chemical reactions occur, the energy changes are relatively modest and the mass changes are too small to measure, so the laws of conservation of matter and energy hold well. However, in nuclear reactions, the energy changes are much larger (by factors of a million or so), the mass changes are measurable, and matter-energy conversions are significant. This will be examined in more detail in a later chapter on nuclear chemistry. To encompass both chemical and nuclear changes, we combine these laws into one statement: The total quantity of matter and energy in the universe is fixed.

Thermal Energy, Temperature, and Heat

Thermal energy is kinetic energy associated with the random motion of atoms and molecules. Temperature is a quantitative measure of “hot” or “cold.” When the atoms and molecules in an object are moving or vibrating quickly, they have a higher average kinetic energy (KE), and we say that the object is “hot.” When

the atoms and molecules are moving slowly, they have lower KE, and we say that the object is “cold” (Figure 3). Assuming that no chemical reaction or phase change (such as melting or vaporizing) occurs, increasing the amount of thermal energy in a sample of matter will cause its temperature to increase. And, assuming that no chemical reaction or phase change (such as condensation or freezing) occurs, decreasing the amount of thermal energy in a sample of matter will cause its temperature to decrease.

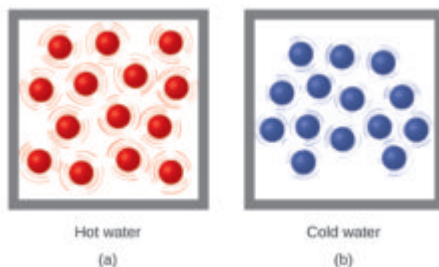


Figure 3 a) The molecules in a sample of hot water move more rapidly than (b) those in a sample of cold water.

Click on this [interactive simulation](#) to view the effects of temperature on molecular motion.

Most substances expand as their temperature increases and contract as their temperature decreases. This property can be used to measure temperature changes, as shown in Figure 4. The operation of many thermometers depends on the expansion and contraction of substances in response to temperature changes.

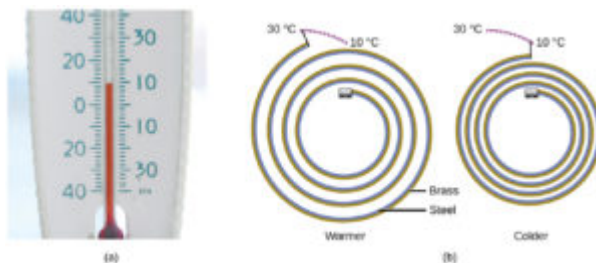


Figure 4 (a) In an alcohol or mercury thermometer, the liquid (dyed red for visibility) expands when heated and contracts when cooled, much more so than the glass tube that contains the liquid. (b) In a bimetallic thermometer, two different metals (such as brass and steel) form a two-layered strip. When heated or cooled, one of the metals (brass) expands or contracts more than the other metal (steel), causing the strip to coil or uncoil. Both types of thermometers have a calibrated scale that indicates the temperature. (credit a: modification of work by “dwstucke”/Flickr)

The following [demonstration](#) allows one to view the effects of heating and cooling a coiled bimetallic strip.

Heat (q) is the transfer of thermal energy between two bodies at different temperatures. Heat flow (a redundant term, but one commonly used) increases the thermal energy of one body and decreases the thermal energy of the other. Suppose we initially have a high temperature (and high thermal energy) substance (H) and a low temperature (and low thermal energy) substance (L). The atoms and molecules in H have a higher average KE than those in L. If we place substance H in contact with substance L, the thermal energy will flow spontaneously from substance H to substance L. The temperature of substance H will decrease, as will the average KE of its molecules; the temperature of substance L will increase, along with the average KE of its molecules. Heat flow will continue until the two substances are at the same temperature (Figure 5).

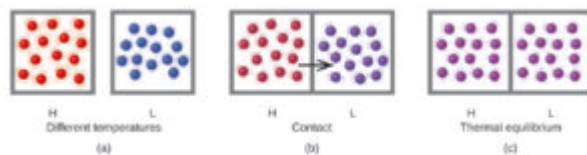


Figure 5 (a) Substances H and L are initially at different temperatures, and their atoms have different average kinetic energies. (b) When they are put into contact with each other, collisions between the molecules result in the transfer of kinetic (thermal) energy from the hotter to the cooler matter. (c) The two objects reach “thermal equilibrium” when both substances are at the same temperature, and their molecules have the same average kinetic energy.

Click on the [PhET simulation](#) to explore energy forms and changes. Visit the Energy Systems tab to create combinations of energy sources, transformation methods, and outputs. Click on Energy Symbols to visualize the transfer of energy.

Matter undergoing chemical reactions and physical changes can release or absorb heat. A change that releases heat is called an exothermic process. For example, the combustion reaction that occurs when using an oxyacetylene torch is an exothermic process—this process also releases energy in the form of light as evidenced by the torch’s flame (Figure 6). A reaction or change that absorbs heat is an endothermic process. A cold pack used to treat muscle strains provides an example of an endothermic process. When the substances in the cold pack (water and a salt like ammonium nitrate) are brought together, the resulting process absorbs heat, leading to the sensation of cold.



Figure 6 (a) An oxyacetylene torch produces heat by the combustion of acetylene in oxygen. The energy released by this exothermic reaction heats and then melts the metal being cut. The sparks are tiny bits of the molten metal flying away. (b) A cold pack uses an endothermic process to create the sensation of cold. (credit a: modification of work by “Skatebiker”/Wikimedia commons)

Historically, energy was measured in units of calories (cal). A calorie is the amount of energy required to raise one gram of water by 1 degree C (1 kelvin). However, this quantity depends on the atmospheric pressure and the starting temperature of the water. The ease of measurement of energy changes in calories has meant that the calorie is still frequently used. The Calorie (with a capital C), or large calorie, commonly used in quantifying food energy content, is a kilocalorie. The SI unit of heat, work, and energy is the joule. A joule (J) is defined as the amount of energy used when a force of 1 newton moves an object 1 meter. It is named in honor of the English physicist James Prescott Joule. One joule is equivalent to 1 kg m²/s², which is also called 1 newton-meter. A kilojoule (kJ) is 1000 joules. To standardize its definition, 1 calorie has been set to equal 4.184 joules.

We now introduce two concepts useful in describing heat flow and temperature change. The heat capacity (C) of a body of matter is the quantity of heat (q) it absorbs or releases when it experiences a temperature change (ΔT) of 1 degree Celsius (or equivalently, 1 kelvin):

$$C = \frac{q}{\Delta T}$$

Heat capacity is determined by both the type and amount of substance that absorbs or releases heat. It is therefore an extensive property—its value is proportional to the amount of the substance.

For example, consider the heat capacities of two cast iron frying pans. The heat capacity of the large pan is five times greater than that of the small pan because, although both are made of the same material, the mass of the large pan is five times greater than the mass of the small pan. More mass means more atoms are present in the larger pan, so it takes more energy to make all of those atoms vibrate faster. The heat capacity of the small cast iron frying pan is found by observing that it takes 18,150 J of energy to raise the temperature of the pan by 50.0 °C:

$$C_{\text{small pan}} = \frac{18,140 \text{ J}}{50.0^\circ\text{C}} = 363 \text{ J}/^\circ\text{C}$$

The larger cast iron frying pan, while made of the same substance, requires 90,700 J of energy to raise its temperature by 50.0 °C. The larger pan has a (proportionally) larger heat capacity because the larger amount of material requires a (proportionally) larger amount of energy to yield the same temperature change:

$$C_{\text{large pan}} = \frac{90,700 \text{ J}}{50.0^\circ\text{C}} = 1814 \text{ J}/^\circ\text{C}$$

The specific heat capacity (c) of a substance, commonly called its “specific heat,” is the quantity of heat required to raise the temperature of 1 gram of a substance by 1 degree Celsius (or 1 kelvin):

$$c = \frac{q}{m\Delta T}$$

Specific heat capacity depends only on the kind of substance absorbing or releasing heat. It is an intensive property—the type, but not the amount, of the substance is all that matters. For example, the small cast iron frying pan has a mass of 808 g. The specific heat of iron (the material used to make the pan) is therefore:

$$C_{\text{iron}} = \frac{18,140 \text{ J}}{(808 \text{ g})(50.0^\circ\text{C})} = 0.449 \text{ J}/\text{g}^\circ\text{C}$$

The large frying pan has a mass of 4040 g. Using the data for this pan, we can also calculate the specific heat of iron:

$$C_{\text{iron}} = \frac{90,700 \text{ J}}{(4040 \text{ g})(50.0^\circ\text{C})} = 0.449 \text{ J}/\text{g}^\circ\text{C}$$

Although the large pan is more massive than the small pan, since both are made of the same material, they both yield the same value for specific heat (for the material of construction, iron). Note that specific heat is measured in units of energy per temperature per mass and is an intensive property, being derived from a ratio of two extensive properties (heat and mass). The molar heat capacity, also an intensive property, is the heat capacity per mole of a particular substance and has units of J/mol °C (Figure 7).



Figure 7 Due to its larger mass, a large frying pan has a larger heat capacity than a small frying pan. Because they are made of the same material, both frying pans have the same specific heat. (credit: Mark Blaser)

Liquid water has a relatively high specific heat (about 4.2 J/g °C); most metals have much lower specific

heats (usually less than 1 J/g °C). The specific heat of a substance varies somewhat with temperature. However, this variation is usually small enough that we will treat specific heat as constant over the range of temperatures that will be considered in this chapter. Specific heats of some common substances are listed in Table 1.

Table 1: Specific Heats of Common Substances at 25 °C and 1 bar		
Substance	Symbol (state)	Specific Heat (J/g °C)
helium	He(<i>g</i>)	5.193
water	H ₂ O(<i>l</i>)	4.184
ethanol	C ₂ H ₆ O(<i>l</i>)	2.376
ice	H ₂ O(<i>s</i>)	2.093 (at -10 °C)
water vapor	H ₂ O(<i>g</i>)	1.864
nitrogen	N ₂ (<i>g</i>)	1.040
air		1.007
oxygen	O ₂ (<i>g</i>)	0.918
aluminum	Al(<i>s</i>)	0.897
carbon dioxide	CO ₂ (<i>g</i>)	0.853
argon	Ar(<i>g</i>)	0.522
iron	Fe(<i>s</i>)	0.449
copper	Cu(<i>s</i>)	0.385
lead	Pb(<i>s</i>)	0.130
gold	Au(<i>s</i>)	0.129
silicon	Si(<i>s</i>)	0.712

If we know the mass of a substance and its specific heat, we can determine the amount of heat, q , entering or leaving the substance by measuring the temperature change before and after the heat is gained or lost:

$$q = (\text{specific heat}) \times (\text{mass of substance}) \times (\text{temperature change})$$

$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$$

In this equation, c is the specific heat of the substance, m is its mass, and ΔT (which is read “delta T”) is the temperature change, $T_{\text{final}} - T_{\text{initial}}$. If a substance gains thermal energy, its temperature increases, its final temperature is higher than its initial temperature, $T_{\text{final}} - T_{\text{initial}}$ has a positive value, and the value of q is positive. If a substance loses thermal energy, its temperature decreases, the final temperature is lower than the initial temperature, $T_{\text{final}} - T_{\text{initial}}$ has a negative value, and the value of q is negative.



Example 1 Measuring Heat

A flask containing 8.0×10^2 g of water is heated, and the temperature of the water increases from 21 °C to 85 °C. How much heat did the water absorb?

Solution

To answer this question, consider these factors:

- the specific heat of the substance being heated (in this case, water)
- the amount of substance being heated (in this case, 800 g)
- the magnitude of the temperature change (in this case, from 21 °C to 85 °C).

The specific heat of water is 4.184 J/g °C, so to heat 1 g of water by 1 °C requires 4.184 J. We note that since 4.184 J is required to heat 1 g of water by 1 °C, we will need 800 times as much to heat 800 g of water by 1 °C. Finally, we observe that since 4.184 J are required to heat 1 g of water by 1 °C, we will need 64 times as much to heat it by 64 °C (that is, from 21 °C to 85 °C).

This can be summarized using the equation:

$$\begin{aligned}q &= c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}}) \\&= (4.184 \text{ J/g}^\circ\text{C}) \times (800 \text{ g}) \times (85 - 21)^\circ\text{C} \\&= (4.184 \text{ J/g}^\circ\text{C}) \times (800 \text{ g}) \times (64)^\circ\text{C} \\&= 210,000 \text{ J} (=210 \text{ kJ})\end{aligned}$$

Because the temperature increased, the water absorbed heat and q is positive.

Check Your Learning

How much heat, in joules, must be added to a 5.00×10^2 -g iron skillet to increase its temperature from 25 °C to 250 °C? The specific heat of iron is 0.451 J/g °C.

ANSWER:

$$5.05 \times 10^4 \text{ J}$$

Note that the relationship between heat, specific heat, mass, and temperature change can be used to determine any of these quantities (not just heat) if the other three are known or can be deduced.

Example 2: Determining Other Quantities

A piece of unknown metal weighs 348 g. When the metal piece absorbs 6.64 kJ of heat, its temperature increases from 22.4 °C to 43.6 °C. Determine the specific heat of this metal (which might provide a clue to its identity).

Solution

Since mass, heat, and temperature change are known for this metal, we can determine its specific heat using the relationship:

$$q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$$

Substituting the known values:

$$6640 \text{ J} = c \times (348 \text{ g}) \times (43.6 - 22.4)^\circ\text{C}$$

Solving:

$$c = \frac{6640 \text{ J}}{(348 \text{ g}) \times (21.2^\circ\text{C})} = 0.900 \text{ J/g}^\circ\text{C}$$

Comparing this value with the values in Table 1, this value matches the specific heat of aluminum, which suggests that the unknown metal may be aluminum.

Check Your Learning

A piece of unknown metal weighs 217 g. When the metal piece absorbs 1.43 kJ of heat, its temperature increases from 24.5 °C to 39.1 °C. Determine the specific heat of this metal, and predict its identity.

ANSWER:

$c = 0.45 \text{ J/g } ^\circ\text{C}$; the metal is likely to be iron

Note: SOLAR THERMAL ENERGY POWER PLANTS

The sunlight that reaches the earth contains thousands of times more energy than we presently capture. Solar thermal systems provide one possible solution to the problem of converting energy from the sun into energy we can use. Large-scale solar thermal plants have different design specifics, but all concentrate sunlight to heat some substance; the heat “stored” in that substance is then converted into electricity.

The Solana Generating Station in Arizona’s Sonora Desert produces 280 megawatts of electrical power. It uses parabolic mirrors that focus sunlight on pipes filled with a heat transfer fluid (HTF) (Figure 8). The HTF then does two things: It turns water into steam, which spins turbines, which in turn produces electricity, and it melts and heats a mixture of salts, which functions as a thermal energy storage system. After the sun goes down, the molten salt mixture can then release enough of its stored heat to produce steam to run the turbines for 6 hours. Molten salts are used because they possess a number of beneficial properties, including high heat capacities and thermal conductivities.

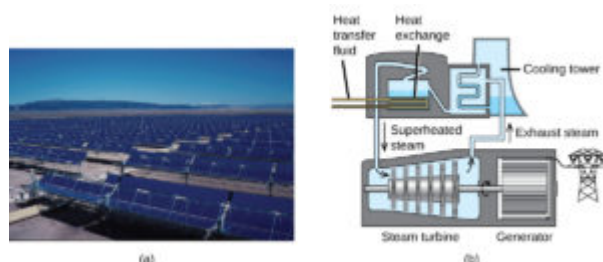


Figure 8 This solar thermal plant uses parabolic trough mirrors to concentrate sunlight. (credit a: modification of work by Bureau of Land Management)

The 377-megawatt Ivanpah Solar Generating System, located in the Mojave Desert in California, is the largest solar thermal power plant in the world (Figure 9). Its 170,000 mirrors focus huge amounts of sunlight on three water-filled towers, producing steam at over $538 \text{ }^\circ\text{C}$ that drives electricity-producing turbines. It produces enough energy to power 140,000 homes. Water is used as the working fluid because of its large heat capacity and heat of vaporization.



Figure 9 (a) The Ivanpah solar thermal plant uses 170,000 mirrors to concentrate sunlight on water-filled towers. (b) It covers 4000 acres of public land near the Mojave Desert and the California-Nevada border. (credit a: modification of work by Craig Dietrich; credit b: modification of work by “USFWS Pacific Southwest Region”/Flickr)

Key Concepts and Summary

Energy is the capacity to do work (applying a force to move matter). Kinetic energy (KE) is the energy of motion; potential energy is energy due to relative position, composition, or condition. When energy is converted from one form into another, energy is neither created nor destroyed (law of conservation of

energy or first law of thermodynamics).

Matter has thermal energy due to the KE of its molecules and temperature that corresponds to the average KE of its molecules. Heat is energy that is transferred between objects at different temperatures; it flows from a high to a low temperature. Chemical and physical processes can absorb heat (endothermic) or release heat (exothermic). The SI unit of energy, heat, and work is the joule (J).

Specific heat and heat capacity are measures of the energy needed to change the temperature of a substance or object. The amount of heat absorbed or released by a substance depends directly on the type of substance, its mass, and the temperature change it undergoes.

Key Equations

- $q = c \times m \times \Delta T = c \times m \times (T_{\text{final}} - T_{\text{initial}})$

Energy Basics Exercises

Why do we use an object's mass, rather than its weight, to indicate the amount of matter it contains

What properties distinguish solids from liquids? Liquids from gases? Solids from gases?

How does a heterogeneous mixture differ from a homogeneous mixture? How are they similar?

How does a homogeneous mixture differ from a pure substance? How are they similar?

How does an element differ from a compound? How are they similar?

How do molecules of elements and molecules of compounds differ? In what ways are they similar?

How does an atom differ from a molecule? In what ways are they similar?

Many of the items you purchase are mixtures of pure compounds. Select three of these commercial products and prepare a list of the ingredients that are pure compounds.

Classify each of the following as an element, a compound, or a mixture:(a) copper(b) water(c) nitrogen(d) sulfur(e) air

(f) sucrose

(g) a substance composed of molecules each of which contains two iodine atoms

(h) gasoline

Classify each of the following as an element, a compound, or a mixture:(a) iron(b) oxygen(c) mercury oxide(d) pancake syrup(e) carbon dioxide

(f) a substance composed of molecules each of which contains one hydrogen atom and one chlorine atom

(g) baking soda

(h) baking powder

A sulfur atom and a sulfur molecule are not identical. What is the difference?

How are the molecules in oxygen gas, the molecules in hydrogen gas, and water molecules similar? How do they differ?

We refer to astronauts in space as weightless, but not without mass. Why?

As we drive an automobile, we don't think about the chemicals consumed and produced.

Prepare a list of the principal chemicals consumed and produced during the operation of an automobile.

Matter is everywhere around us. Make a list by name of fifteen different kinds of matter that you encounter every day. Your list should include (and label at least one example of each) the following: a solid, a liquid, a gas, an element, a compound, a homogenous mixture, a heterogeneous mixture, and a pure substance.

When elemental iron corrodes it combines with oxygen in the air to ultimately form red brown iron(III) oxide which we call rust. (a) If a shiny iron nail with an initial mass of 23.2 g is weighed after being coated in a layer of rust, would you expect the mass to have increased, decreased, or remained the same? Explain. (b) If the mass of the iron nail increases to 24.1 g, what mass of oxygen combined with the iron?

As stated in the text, convincing examples that demonstrate the law of conservation of matter outside of the laboratory are few and far between. Indicate whether the mass would increase, decrease, or stay the same for the following scenarios where chemical reactions take place:(a) Exactly one pound of bread dough is placed in a baking tin. The dough is cooked in an oven at 350 °F releasing a wonderful aroma of freshly baked bread during the cooking process. Is the mass of the baked loaf less than, greater than, or the same as the one pound of original dough? Explain.(b) When magnesium burns in air a white flaky ash of magnesium oxide is produced. Is the mass of magnesium oxide less than, greater than, or the same as the original piece of magnesium? Explain.(c) Antoine Lavoisier, the French scientist credited with first stating the law of conservation of matter, heated a mixture of tin and air in a sealed flask to produce tin oxide. Did the mass of the sealed flask and contents decrease, increase, or remain the same after the heating?

Yeast converts glucose to ethanol and carbon dioxide during anaerobic fermentation as depicted in the simple chemical equation here:glucose \rightarrow ethanol + carbon dioxide(a) If 200.0 g of glucose is fully converted, what will be the total mass of ethanol and carbon dioxide produced?(b) If the fermentation is carried out in an open container, would you expect the mass of the container and contents after fermentation to be less than, greater than, or the same as the mass of the container and contents before fermentation? Explain.(c) If 97.7 g of carbon dioxide is produced, what mass of ethanol is produced?

Glossary

calorie (cal)

unit of heat or other energy; the amount of energy required to raise 1 gram of water by 1 degree Celsius; 1 cal is defined as 4.184 J

endothermic process

chemical reaction or physical change that absorbs heat

energy

capacity to supply heat or do work

exothermic process

chemical reaction or physical change that releases heat

heat (q)

transfer of thermal energy between two bodies

heat capacity (C)

extensive property of a body of matter that represents the quantity of heat required to increase its temperature by 1 degree Celsius (or 1 kelvin)

joule (J)

SI unit of energy; 1 joule is the kinetic energy of an object with a mass of 2 kilograms moving with a velocity of 1 meter per second, $1 \text{ J} = 1 \text{ kg m}^2/\text{s}^2$ and $4.184 \text{ J} = 1 \text{ cal}$

kinetic energy

energy of a moving body, in joules, equal to $\frac{1}{2} mv^2$ (where m = mass and v = velocity)

potential energy

energy of a particle or system of particles derived from relative position, composition, or condition

specific heat capacity (c)

intensive property of a substance that represents the quantity of heat required to raise the temperature of 1 gram of the substance by 1 degree Celsius (or 1 kelvin)

temperature

intensive property of matter that is a quantitative measure of “hotness” and “coldness”

thermal energy

kinetic energy associated with the random motion of atoms and molecules

thermochemistry

study of measuring the amount of heat absorbed or released during a chemical reaction or a physical change

work (w)

energy transfer due to changes in external, macroscopic variables such as pressure and volume; or causing matter to move against an opposing force

Attribution

Calorimetry

One technique we can use to measure the amount of heat involved in a chemical or physical process is known as calorimetry. Calorimetry is used to measure amounts of heat transferred to or from a substance. To do so, the heat is exchanged with a calibrated object (calorimeter). The change in temperature of the measuring part of the calorimeter is converted into the amount of heat (since the previous calibration was used to establish its heat capacity). The measurement of heat transfer using this approach requires the definition of a system (the substance or substances undergoing the chemical or physical change) and its surroundings (the other components of the measurement apparatus that serve to either provide heat to the system or absorb heat from the system). Knowledge of the heat capacity of the surroundings, and careful measurements of the masses of the system and surroundings and their temperatures before and after the process allows one to calculate the heat transferred as described in this section.

A calorimeter is a device used to measure the amount of heat involved in a chemical or physical process. For example, when an exothermic reaction occurs in solution in a calorimeter, the heat produced by the reaction is absorbed by the solution, which increases its temperature. When an endothermic reaction occurs, the heat required is absorbed from the thermal energy of the solution, which decreases its temperature (Figure 1). The temperature change, along with the specific heat and mass of the solution, can then be used to calculate the amount of heat involved in either case.

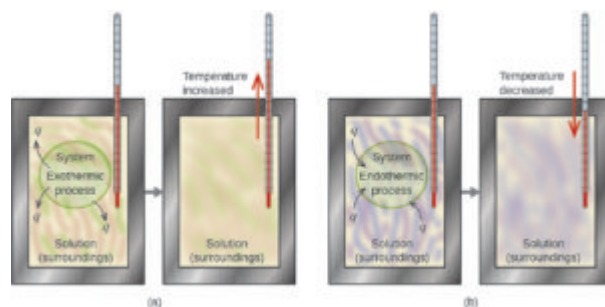


Figure 1 In a calorimetric determination, either (a) an exothermic process occurs and heat, q , is negative, indicating that thermal energy is transferred from the system to its surroundings, or (b) an endothermic process occurs and heat, q , is positive, indicating that thermal energy is transferred from the surroundings to the system.

Scientists use well-insulated calorimeters that all but prevent the transfer of heat between the calorimeter and its environment. This enables the accurate determination of the heat involved in chemical processes, the energy content of foods, and so on. General chemistry students often use simple calorimeters constructed from polystyrene cups (Figure 2). These easy-to-use “coffee cup” calorimeters allow more heat exchange with their surroundings, and therefore produce less accurate energy values.

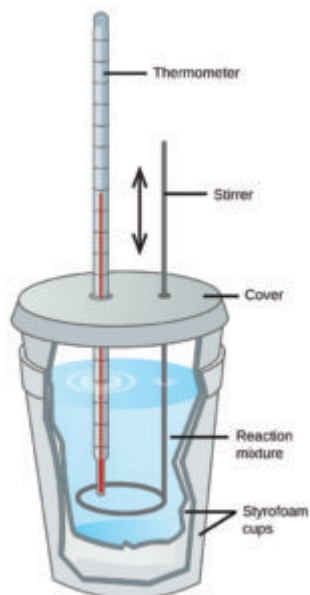


Figure 2 A simple calorimeter can be constructed from two polystyrene cups. A thermometer and stirrer extend through the cover into the reaction mixture.

Commercial solution calorimeters are also available. Relatively inexpensive calorimeters often consist of two thin-walled cups that are nested in a way that minimizes thermal contact during use, along with an insulated cover, handheld stirrer, and simple thermometer. More expensive calorimeters used for industry and research typically have a well-insulated, fully enclosed reaction vessel, motorized stirring mechanism, and a more accurate temperature sensor (Figure 3).

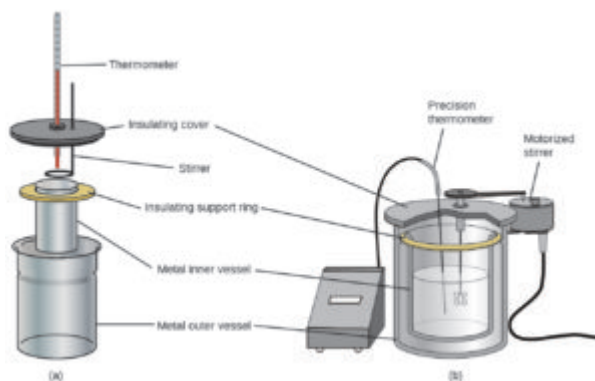


Figure 3 Commercial solution calorimeters range from (a) simple, inexpensive models for student use to (b) expensive, more accurate models for industry and research.

Before we practice calorimetry problems involving chemical reactions, consider a simpler example that illustrates the core idea behind calorimetry. Suppose we initially have a high-temperature substance, such as a hot piece of metal (M), and a low-temperature substance, such as cool water (W). If we place the metal in the water, heat will flow from M to W. The temperature of M will decrease, and the temperature of W will increase, until the two substances have the same temperature—that is, when they reach thermal equilibrium (Figure 4). If this occurs in a calorimeter, ideally all of this heat transfer occurs between the two substances, with no heat gained or lost by either the calorimeter or the calorimeter’s surroundings. Under these ideal circumstances, the net heat change is zero:

$$q_{\text{substance M}} + q_{\text{substance W}} = 0$$

This relationship can be rearranged to show that the heat gained by substance M is equal to the heat lost by substance W:

$$q_{\text{substance M}} = -q_{\text{substance W}}$$

The magnitude of the heat (change) is therefore the same for both substances, and the negative sign merely shows that $q_{\text{substance M}}$ and $q_{\text{substance W}}$ are opposite in direction of heat flow (gain or loss) but does not indicate the arithmetic sign of either q value (that is determined by whether the matter in question gains or loses heat, per definition). In the specific situation described, $q_{\text{substance M}}$ is a negative value and $q_{\text{substance W}}$ is positive, since heat is transferred from M to W.

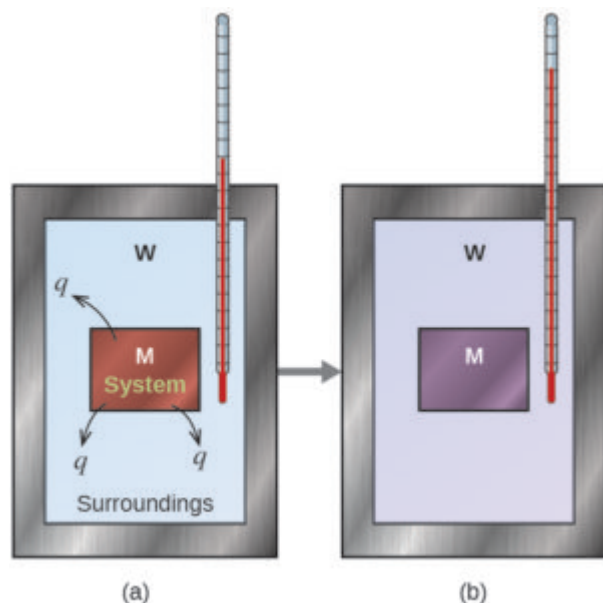


Figure 4 In a simple calorimetry process, (a) heat, q , is transferred from the hot metal, M, to the cool water, W, until (b) both are at the same temperature.

Example 1: Heat Transfer between Substances at Different Temperatures

A 360-g piece of rebar (a steel rod used for reinforcing concrete) is dropped into 425 mL of water at 24.0 °C. The final temperature of the water was measured as 42.7 °C. Calculate the initial temperature of the piece of rebar. Assume the specific heat of steel is approximately the same as that for iron, and that all heat transfer occurs between the rebar and the water (there is no heat exchange with the surroundings).

Solution

The temperature of the water increases from 24.0 °C to 42.7 °C, so the water absorbs heat. That heat came from the piece of rebar, which initially was at a higher temperature. Assuming that all heat transfer was between the rebar and the water, with no heat “lost” to the surroundings, then heat given off by rebar = –heat taken in by water, or:

$$q_{\text{rebar}} = -q_{\text{water}}$$

Since we know how heat is related to other measurable quantities, we have:

$$(c \times m \times \Delta T)_{\text{rebar}} = -(c \times m \times \Delta T)_{\text{water}}$$

Letting f = final and i = initial, in expanded form, this becomes:

$$c_{\text{rebar}} \times m_{\text{rebar}} \times (T_{f,\text{rebar}} - T_{i,\text{rebar}}) = -c_{\text{water}} \times m_{\text{water}} \times (T_{f,\text{water}} - T_{i,\text{water}})$$

The density of water is 1.0 g/mL, so 425 mL of water = 425 g. Noting that the final temperature of both the rebar and water is 42.7 °C, substituting known values yields:

$$(0.449\text{J/g} \cdot ^\circ\text{C})(360\text{g})(42.7^\circ\text{C} - T_{i,\text{rebar}}) = (4.184\text{J/g} \cdot ^\circ\text{C})(425\text{g})(42.7^\circ\text{C} - 24.0^\circ\text{C})$$

$$T_{i,\text{rebar}} = \frac{(4.184\text{J/g}^\circ\text{C})(425\text{g})(42.7^\circ\text{C}-24.0^\circ\text{C})}{(0.449\text{J/g}^\circ\text{C})(360\text{g})} + 42.7^\circ\text{C}$$

Solving this gives $T_{i,\text{rebar}} = 248^\circ\text{C}$, so the initial temperature of the rebar was 248°C .

Check Your Learning

A 248-g piece of copper is dropped into 390 mL of water at 22.6°C . The final temperature of the water was measured as 39.9°C . Calculate the initial temperature of the piece of copper. Assume that all heat transfer occurs between the copper and the water.

Answer:

The initial temperature of the copper was 335.6°C .

Check Your Learning A 248-g piece of copper initially at 314°C is dropped into 390 mL of water initially at 22.6°C . Assuming that all heat transfer occurs between the copper and the water, calculate the final temperature.

Answer:

The final temperature (reached by both copper and water) is 38.8°C .

This method can also be used to determine other quantities, such as the specific heat of an unknown metal.

Example 2: Identifying a Metal by Measuring Specific Heat

A 59.7 g piece of metal that had been submerged in boiling water was quickly transferred into 60.0 mL of water initially at 22.0°C . The final temperature is 28.5°C . Use these data to determine the specific heat of the metal. Use this result to identify the metal.

Solution

Assuming perfect heat transfer, heat given off by metal = -heat taken in by water, or:

$$q_{\text{metal}} = -q_{\text{water}}$$

In expanded form, this is:

$$c_{\text{metal}} \times m_{\text{metal}} \times (T_{f,\text{metal}} - T_{i,\text{metal}}) = -c_{\text{water}} \times m_{\text{water}} \times (T_{f,\text{water}} - T_{i,\text{water}})$$

Noting that since the metal was submerged in boiling water, its initial temperature was 100.0°C ; and that for water, $60.0\text{ mL} = 60.0\text{ g}$; we have:

$$(c_{\text{metal}})(59.7\text{g})(28.5^\circ\text{C}-100.0^\circ\text{C}) = -(4.18\text{J/g}^\circ\text{C})(60.0\text{g})(28.5^\circ\text{C} - 22.0^\circ\text{C})$$

Solving this:

$$c_{\text{metal}} = -(4.184\text{J/g}^\circ\text{C})(60.0\text{g})(6.5^\circ\text{C})(59.7\text{g})(-71.5^\circ\text{C}) = 0.38\text{J/g}^\circ\text{C}$$

Comparing this with values in [link], our experimental specific heat is closest to the value for copper ($0.39\text{ J/g}^\circ\text{C}$), so we identify the metal as copper.

Check Your Learning

A 92.9-g piece of a silver/gray metal is heated to 178.0°C , and then quickly transferred into 75.0 mL of water initially at 24.0°C . After 5 minutes, both the metal and the water have reached the same temperature: 29.7°C . Determine the specific heat and the identity of the metal. (Note: You should find that the specific heat is close to that of two different metals. Explain how you can confidently determine the identity of the metal).

Answer:

$$c_{\text{metal}} = 0.13\text{ J/g}^\circ\text{C}$$

This specific heat is close to that of either gold or lead. It would be difficult to determine which metal this was based solely on the numerical values. However, the observation that the metal is silver/gray in addition to the value for the specific heat indicates that the metal is lead.

When we use calorimetry to determine the heat involved in a chemical reaction, the same principles we have been discussing apply. The amount of heat absorbed by the calorimeter is often small enough that we can neglect it (though not for highly accurate measurements, as discussed later), and the calorimeter minimizes energy exchange with the surroundings. Because energy is neither created nor destroyed during a chemical reaction, there is no overall energy change during the reaction. The heat produced or consumed in the reaction (the “system”), q_{reaction} , plus the heat absorbed or lost by the solution (the “surroundings”), q_{solution} , must add up to zero:

$$q_{\text{reaction}} + q_{\text{solution}} = 0$$

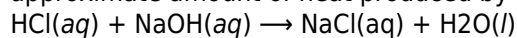
This means that the amount of heat produced or consumed in the reaction equals the amount of heat absorbed or lost by the solution:

$$q_{\text{reaction}} = -q_{\text{solution}}$$

This concept lies at the heart of all calorimetry problems and calculations.

Example 3: Heat Produced by an Exothermic Reaction

When 50.0 mL of 0.10 M HCl(aq) and 50.0 mL of 0.10 M NaOH(aq), both at 22.0 °C, are added to a coffee cup calorimeter, the temperature of the mixture reaches a maximum of 28.9 °C. What is the approximate amount of heat produced by this reaction?



Solution

To visualize what is going on, imagine that you could combine the two solutions so quickly that no reaction took place while they mixed; then after mixing, the reaction took place. At the instant of mixing, you have 100.0 mL of a mixture of HCl and NaOH at 22.0 °C. The HCl and NaOH then react until the solution temperature reaches 28.9 °C.

The heat given off by the reaction is equal to that taken in by the solution. Therefore:

$$q_{\text{reaction}} = -q_{\text{solution}}$$

(It is important to remember that this relationship only holds if the calorimeter does not absorb any heat from the reaction, and there is no heat exchange between the calorimeter and its surroundings.)

Next, we know that the heat absorbed by the solution depends on its specific heat, mass, and temperature change:

$$q_{\text{solution}} = (c \times m \times \Delta T)_{\text{solution}}$$

To proceed with this calculation, we need to make a few more reasonable assumptions or approximations. Since the solution is aqueous, we can proceed as if it were water in terms of its specific heat and mass values. The density of water is approximately 1.0 g/mL, so 100.0 mL has a mass of about 1.0×10^2 g (two significant figures). The specific heat of water is approximately 4.18 J/g °C, so we use that for the specific heat of the solution. Substituting these values gives:

$$q_{\text{solution}} = (4.184\text{J/g } ^\circ\text{C})(1.0 \times 10^2\text{g})(28.9^\circ\text{C} - 22.0^\circ\text{C}) = 2.89 \times 10^3\text{J}$$

Finally, since we are trying to find the heat of the reaction, we have:

$$q_{\text{reaction}} = -q_{\text{solution}} = -2.89 \times 10^3\text{J}$$

The negative sign indicates that the reaction is exothermic. It produces 2.89 kJ of heat.

Check Your Learning

When 100 mL of 0.200 M NaCl(aq) and 100 mL of 0.200 M AgNO₃(aq), both at 21.9 °C, are mixed in a coffee cup calorimeter, the temperature increases to 23.5 °C as solid AgCl forms. How much heat is produced by this precipitation reaction? What assumptions did you make to determine your value?

Answer:

1.34×10^3 J; assume no heat is absorbed by the calorimeter, no heat is exchanged between the

calorimeter and its surroundings, and that the specific heat and mass of the solution are the same as those for water

Thermochemistry of Hand Warmers

When working or playing outdoors on a cold day, you might use a hand warmer to warm your hands (Figure 5). A common reusable hand warmer contains a supersaturated solution of $\text{NaC}_2\text{H}_3\text{O}_2$ (sodium acetate) and a metal disc. Bending the disk creates nucleation sites around which the metastable $\text{NaC}_2\text{H}_3\text{O}_2$ quickly crystallizes (a later chapter on solutions will investigate saturation and supersaturation in more detail).

The process $\text{NaC}_2\text{H}_3\text{O}_2(\text{aq}) \rightarrow \text{NaC}_2\text{H}_3\text{O}_2(\text{s})$

is exothermic, and the heat produced by this process is absorbed by your hands, thereby warming them (at least for a while). If the hand warmer is reheated, the $\text{NaC}_2\text{H}_3\text{O}_2$ redissolves and can be reused.



Figure 5 Chemical hand warmers produce heat that warms your hand on a cold day. In this one, you can see the metal disc that initiates the exothermic precipitation reaction. (credit: modification of work by Science Buddies TV/YouTube)

Another common hand warmer produces heat when it is ripped open, exposing iron and water in the hand warmer to oxygen in the air. One simplified version of this exothermic reaction is $2\text{Fe}(\text{s}) + 3/2\text{O}_2(\text{g}) \rightarrow \text{Fe}_2\text{O}_3(\text{s})$. Salt in the hand warmer catalyzes the reaction, so it produces heat more rapidly; cellulose, vermiculite, and activated carbon help distribute the heat evenly. Other types of hand warmers use lighter fluid (a platinum catalyst helps lighter fluid oxidize exothermically), charcoal (charcoal oxidizes in a special case), or electrical units that produce heat by passing an electrical current from a battery through resistive wires.

This [link](#) shows the precipitation reaction that occurs when the disk in a chemical hand warmer is flexed.

Example 4: Heat Flow in an Instant Ice Pack

When solid ammonium nitrate dissolves in water, the solution becomes cold. This is the basis for an “instant ice pack” (Figure 6). When 3.21 g of solid NH_4NO_3 dissolves in 50.0 g of water at 24.9 °C in a calorimeter, the temperature decreases to 20.3 °C.

Calculate the value of q for this reaction and explain the meaning of its arithmetic sign. State any assumptions that you made.



Figure 6 An instant cold pack consists of a bag containing solid ammonium nitrate and a second bag of water. When the bag of water is broken, the pack becomes cold because the dissolution of ammonium nitrate is an endothermic process that removes thermal energy from the water. The cold pack then removes thermal energy from your body.

Solution

We assume that the calorimeter prevents heat transfer between the solution and its external environment (including the calorimeter itself), in which case:

$$q_{\text{rxn}} = -q_{\text{soln}}$$

with “rxn” and “soln” used as shorthand for “reaction” and “solution,” respectively.

Assuming also that the specific heat of the solution is the same as that for water, we have:

$$\begin{aligned} q_{\text{rxn}} = -q_{\text{soln}} &= -(c \times m \times \Delta T)_{\text{soln}} \\ &= -[(4.184\text{J/g } ^\circ\text{C}) \times (53.2\text{g}) \times (20.3^\circ\text{C} - 24.9^\circ\text{C})] \\ &= -[(4.184\text{J/g } ^\circ\text{C}) \times (53.2\text{g}) \times (-4.6^\circ\text{C})] \end{aligned}$$

$$+1.0 \times 10^3\text{J} = +1.0\text{kJ}$$

The positive sign for q indicates that the dissolution is an endothermic process.

Check Your Learning

When a 3.00-g sample of KCl was added to 3.00×10^2 g of water in a coffee cup calorimeter, the temperature decreased by 1.05°C . How much heat is involved in the dissolution of the KCl? What assumptions did you make?

Answer:

1.33 kJ; assume that the calorimeter prevents heat transfer between the solution and its external environment (including the calorimeter itself) and that the specific heat of the solution is the same as that for water

If the amount of heat absorbed by a calorimeter is too large to neglect or if we require more accurate results, then we must take into account the heat absorbed both by the solution and by the calorimeter.

The calorimeters described are designed to operate at constant (atmospheric) pressure and are convenient to measure heat flow accompanying processes that occur in solution. A different type of calorimeter that operates at constant volume, colloquially known as a bomb calorimeter, is used to measure the energy produced by reactions that yield large amounts of heat and gaseous products, such as combustion reactions. (The term “bomb” comes from the observation that these reactions can be vigorous enough to

resemble explosions that would damage other calorimeters.) This type of calorimeter consists of a robust steel container (the “bomb”) that contains the reactants and is itself submerged in water (Figure 7). The sample is placed in the bomb, which is then filled with oxygen at high pressure. A small electrical spark is used to ignite the sample. The energy produced by the reaction is trapped in the steel bomb and the surrounding water. The temperature increase is measured and, along with the known heat capacity of the calorimeter, is used to calculate the energy produced by the reaction. Bomb calorimeters require calibration to determine the heat capacity of the calorimeter and ensure accurate results. The calibration is accomplished using a reaction with a known q , such as a measured quantity of benzoic acid ignited by a spark from a nickel fuse wire that is weighed before and after the reaction. The temperature change produced by the known reaction is used to determine the heat capacity of the calorimeter. The calibration is generally performed each time before the calorimeter is used to gather research data.

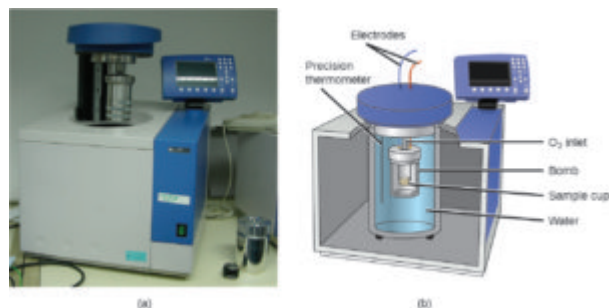


Figure 7 (a) A bomb calorimeter is used to measure heat produced by reactions involving gaseous reactants or products, such as combustion. (b) The reactants are contained in the gas-tight “bomb,” which is submerged in water and surrounded by insulating materials. (credit a: modification of work by “Harbor1”/Wikimedia commons)

Click on this [link](#) to view how a bomb calorimeter is prepared for action. This [site](#) shows calorimetric calculations using sample data.

Example 5: Bomb Calorimetry

When 3.12 g of glucose, $C_6H_{12}O_6$, is burned in a bomb calorimeter, the temperature of the calorimeter increases from 23.8 °C to 35.6 °C. The calorimeter contains 775 g of water, and the bomb itself has a heat capacity of 893 J/°C. How much heat was produced by the combustion of the glucose sample?

Solution

The combustion produces heat that is primarily absorbed by the water and the bomb. (The amounts of heat absorbed by the reaction products and the unreacted excess oxygen are relatively small and dealing with them is beyond the scope of this text. We will neglect them in our calculations.)

The heat produced by the reaction is absorbed by the water and the bomb:

$$\begin{aligned} q_{\text{rxn}} &= -(q_{\text{water}} + q_{\text{bomb}}) \\ &= -[(4.184\text{J/g}^\circ\text{C}) \times (775\text{ g}) \times (35.6^\circ\text{C} - 23.8^\circ\text{C}) + 893\text{J/}^\circ\text{C} \times (35.6^\circ\text{C} - 23.8^\circ\text{C})] \\ &= -(38,300\text{J} + 10,500\text{J}) \\ &= -48,800\text{ J} = -48.8\text{ kJ} \end{aligned}$$

This reaction released 48.7 kJ of heat when 3.12 g of glucose was burned.

Check Your Learning

When 0.963 g of benzene, C_6H_6 , is burned in a bomb calorimeter, the temperature of the calorimeter increases by 8.39 °C. The bomb has a heat capacity of 784 J/°C and is submerged in 925 mL of water. How much heat was produced by the combustion of the glucose sample?

Answer:

39.0 kJ

Since the first one was constructed in 1899, 35 calorimeters have been built to measure the heat produced by a living person.^[1] These whole-body calorimeters of various designs are large enough to hold an individual human being. More recently, whole-room calorimeters allow for relatively normal activities to be performed, and these calorimeters generate data that more closely reflect the real world. These calorimeters are used to measure the metabolism of individuals under different environmental conditions, different dietary regimes, and with different health conditions, such as diabetes. In humans, metabolism is typically measured in Calories per day. A nutritional calorie (Calorie) is the energy unit used to quantify the amount of energy derived from the metabolism of foods; one Calorie is equal to 1000 calories (1 kcal), the amount of energy needed to heat 1 kg of water by 1 °C.

Note: Measuring Nutritional Calories

In your day-to-day life, you may be more familiar with energy being given in Calories, or nutritional calories, which are used to quantify the amount of energy in foods. One calorie (cal) = exactly 4.184 joules, and one Calorie (note the capitalization) = 1000 cal, or 1 kcal. (This is approximately the amount of energy needed to heat 1 kg of water by 1 °C.)

The macronutrients in food are proteins, carbohydrates, and fats or oils. Proteins provide about 4 Calories per gram, carbohydrates also provide about 4 Calories per gram, and fats and oils provide about 9 Calories/g. Nutritional labels on food packages show the caloric content of one serving of the food, as well as the breakdown into Calories from each of the three macronutrients (Figure 8).

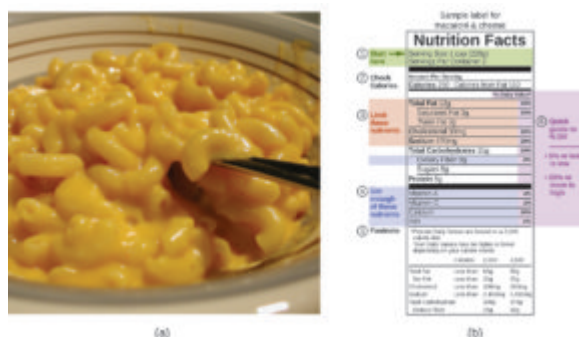


Figure 8 (a) Macaroni and cheese contain energy in the form of the macronutrients in the food. (b) The food's nutritional information is shown on the package label. In the US, the energy content is given in Calories (per serving); the rest of the world usually uses kilojoules. (credit a: modification of work by "Rex Roof"/Flickr)

For the example shown in (b), the total energy per 228-g portion is calculated by:

$$(5 \text{ g protein} \times 4 \text{ Calories/g}) + (31 \text{ g carb} \times 4 \text{ Calories/g}) + (12 \text{ g fat} \times 9 \text{ Calories/g}) = 252 \text{ Calories}$$

So, you can use food labels to count your Calories. But where do the values come from? And how accurate are they? The caloric content of foods can be determined by using bomb calorimetry; that is, by burning the food and measuring the energy it contains. A sample of food is weighed, mixed in a blender, freeze-dried, ground into powder, and formed into a pellet. The pellet is burned inside a bomb calorimeter, and the measured temperature change is converted into energy per gram of food.

Today, the caloric content on food labels is derived using a method called the Atwater system that uses the average caloric content of the different chemical constituents of food, protein, carbohydrate, and fats. The average amounts are those given in the equation and are derived from the various results given by bomb calorimetry of whole foods. The carbohydrate amount is discounted a certain amount for the fiber content, which is indigestible carbohydrate. To determine the energy content of a food, the quantities of carbohydrate, protein, and fat are each multiplied by the average Calories per gram for each and the products summed to obtain the total energy.

Click on this [link](#) to access the US Department of Agriculture (USDA) National Nutrient Database, containing nutritional information on over 8000 foods.

Key Concepts and Summary

Calorimetry is used to measure the amount of thermal energy transferred in a chemical or physical process. This requires careful measurement of the temperature change that occurs during the process and the masses of the system and surroundings. These measured quantities are then used to compute the amount of heat produced or consumed in the process using known mathematical relations.

Calorimeters are designed to minimize energy exchange between the system being studied and its surroundings. They range from simple coffee cup calorimeters used by introductory chemistry students to sophisticated bomb calorimeters used to determine the energy content of food.

Calorimetry Exercises

A 500-mL bottle of water at room temperature and a 2-L bottle of water at the same temperature were placed in a refrigerator. After 30 minutes, the 500-mL bottle of water had cooled to the temperature of the refrigerator. An hour later, the 2-L of water had cooled to the same temperature. When asked which sample of water lost the most heat, one student replied that both bottles lost the same amount of heat because they started at the same temperature and finished at the same temperature. A second student thought that the 2-L bottle of water lost more heat because there was more water. A third student believed that the 500-mL bottle of water lost more heat because it cooled more quickly. A fourth student thought that it was not possible to tell because we do not know the initial temperature and the final temperature of the water. Indicate which of these answers is correct and describe the error in each of the other answers.

Would the amount of heat measured for the reaction in Example 3 be greater, lesser, or remain the same if we used a calorimeter that was a poorer insulator than a coffee cup calorimeter? Explain your answer.

Would the amount of heat absorbed by the dissolution in Example 4 appear greater, lesser, or remain the same if the experimenter used a calorimeter that was a poorer insulator than a coffee cup calorimeter? Explain your answer.

Would the amount of heat absorbed by the dissolution in Example 4 appear greater, lesser, or remain the same if the heat capacity of the calorimeter were taken into account? Explain your answer.

How many milliliters of water at 23 °C with a density of 1.00 g/mL must be mixed with 180 mL (about 6 oz) of coffee at 95 °C so that the resulting combination will have a temperature of 60 °C? Assume that coffee and water have the same density and the same specific heat.

How much will the temperature of a cup (180 g) of coffee at 95 °C be reduced when a 45 g silver spoon (specific heat 0.24 J/g °C) at 25 °C is placed in the coffee and the two are allowed to reach the same temperature? Assume that the coffee has the same density and specific heat as water.

A 45-g aluminum spoon (specific heat 0.88 J/g °C) at 24 °C is placed in 180 mL (180 g) of coffee at 85 °C and the temperature of the two become equal. (a) What is the final temperature when the two become equal? Assume that coffee has the same specific heat as water. (b) The first time a student solved this problem she got an answer of 88 °C. Explain why this is clearly an incorrect answer.

The temperature of the cooling water as it leaves the hot engine of an automobile is 240 °F. After it passes through the radiator it has a temperature of 175 °F. Calculate the amount of heat transferred from the engine to the surroundings by one gallon of water with a specific

heat of 4.184 J/g °C.

A 70.0-g piece of metal at 80.0 °C is placed in 100 g of water at 22.0 °C contained in a calorimeter like that shown in Figure 2. The metal and water come to the same temperature at 24.6 °C. How much heat did the metal give up to the water? What is the specific heat of the metal?

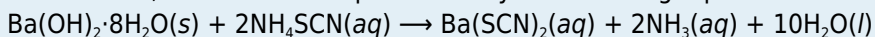
If a reaction produces 1.506 kJ of heat, which is trapped in 30.0 g of water initially at 26.5 °C in a calorimeter like that in Figure 2, what is the resulting temperature of the water?

A 0.500-g sample of KCl is added to 50.0 g of water in a calorimeter (Figure 2). If the temperature decreases by 1.05 °C, what is the approximate amount of heat involved in the dissolution of the KCl, assuming the heat capacity of the resulting solution is 4.18 J/g °C? Is the reaction exothermic or endothermic?

Dissolving 3.0 g of CaCl₂(s) in 150.0 g of water in a calorimeter (Figure 2) at 22.4 °C causes the temperature to rise to 25.8 °C. What is the approximate amount of heat involved in the dissolution, assuming the heat capacity of the resulting solution is 4.18 J/g °C? Is the reaction exothermic or endothermic?

When 50.0 g of 0.200 M NaCl(aq) at 24.1 °C is added to 100.0 g of 0.100 M AgNO₃(aq) at 24.1 °C in a calorimeter, the temperature increases to 25.2 °C as AgCl(s) forms. Assuming the specific heat of the solution and products is 4.20 J/g °C, calculate the approximate amount of heat in joules produced.

The addition of 3.15 g of Ba(OH)₂·8H₂O to a solution of 1.52 g of NH₄SCN in 100 g of water in a calorimeter caused the temperature to fall by 3.1 °C. Assuming the specific heat of the solution and products is 4.20 J/g °C, calculate the approximate amount of heat absorbed by the reaction, which can be represented by the following equation:



The reaction of 50 mL of acid and 50 mL of base described in Example 3 increased the temperature of the solution by 6.9 degrees. How much would the temperature have increased if 100 mL of acid and 100 mL of base had been used in the same calorimeter starting at the same temperature of 22.0 °C? Explain your answer.

If the 3.21 g of NH₄NO₃ in Example 4 were dissolved in 100.0 g of water under the same conditions, how much would the temperature change? Explain your answer.

When 1.0 g of fructose, C₆H₁₂O₆(s), a sugar commonly found in fruits, is burned in oxygen in a bomb calorimeter, the temperature of the calorimeter increases by 1.58 °C. If the heat capacity of the calorimeter and its contents is 9.90 kJ/°C, what is q for this combustion?

When a 0.740-g sample of trinitrotoluene (TNT), C₇H₅N₂O₆, is burned in a bomb calorimeter, the temperature increases from 23.4 °C to 26.9 °C. The heat capacity of the calorimeter is 534 J/°C, and it contains 675 mL of water. How much heat was produced by the combustion of the TNT sample?

One method of generating electricity is by burning coal to heat water, which produces steam that drives an electric generator. To determine the rate at which coal is to be fed into the burner in this type of plant, the heat of combustion per ton of coal must be determined using a bomb calorimeter. When 1.00 g of coal is burned in a bomb calorimeter (Figure 7), the temperature increases by 1.48 °C. If the heat capacity of the calorimeter is 21.6 kJ/°C, determine the heat produced by combustion of a ton of coal (2.000 × 10³ pounds).

The amount of fat recommended for someone with a daily diet of 2000 Calories is 65 g. What percent of the calories in this diet would be supplied by this amount of fat if the average number of Calories for fat is 9.1 Calories/g?

A teaspoon of the carbohydrate sucrose (common sugar) contains 16 Calories (16 kcal). What is the mass of one teaspoon of sucrose if the average number of Calories for carbohydrates is 4.1 Calories/g?

What is the maximum mass of carbohydrate in a 6-oz serving of diet soda that contains less than 1 Calorie per can if the average number of Calories for carbohydrates is 4.1 Calories/g?

A pint of premium ice cream can contain 1100 Calories. What mass of fat, in grams and pounds, must be produced in the body to store an extra 1.1 × 10³ Calories if the average

number of Calories for fat is 9.1 Calories/g?

A serving of a breakfast cereal contains 3 g of protein, 18 g of carbohydrates, and 6 g of fat.

What is the Calorie content of a serving of this cereal if the average number of Calories for fat is 9.1 Calories/g, for carbohydrates is 4.1 Calories/g, and for protein is 4.1 Calories/g?

Which is the least expensive source of energy in kilojoules per dollar: a box of breakfast cereal that weighs 32 ounces and costs \$4.23, or a liter of isooctane (density, 0.6919 g/mL) that costs \$0.45? Compare the nutritional value of the cereal with the heat produced by combustion of the isooctane under standard conditions. A 1.0-ounce serving of the cereal provides 130 Calories.

Glossary

bomb calorimeter

device designed to measure the energy change for processes occurring under conditions of constant volume; commonly used for reactions involving solid and gaseous reactants or products

calorimeter

device used to measure the amount of heat absorbed or released in a chemical or physical process

calorimetry

process of measuring the amount of heat involved in a chemical or physical process

nutritional calorie (Calorie)

unit used for quantifying energy provided by digestion of foods, defined as 1000 cal or 1 kcal

surroundings

all matter other than the system being studied

system

portion of matter undergoing a chemical or physical change being studied

Attribution

Francis D. Reardon et al. "The Snellen human calorimeter revisited, re-engineered and upgraded: Design and performance characteristics." *Medical and Biological Engineering and Computing* 8 (2006)721–28, <http://link.springer.com/article/10.1007/s11517-006-0086-5>. ↵

Atoms, Molecules, and Ions

The Periodic Table

As early chemists worked to purify ores and discovered more elements, they realized that various elements could be grouped together by their similar chemical behaviors. One such grouping includes lithium (Li), sodium (Na), and potassium (K): These elements all are shiny, conduct heat and electricity well, and have similar chemical properties. A second grouping includes calcium (Ca), strontium (Sr), and barium (Ba), which also are shiny, good conductors of heat and electricity, and have chemical properties in common. However, the specific properties of these two groupings are notably different from each other. For example: Li, Na, and K are much more reactive than are Ca, Sr, and Ba; Li, Na, and K form compounds with oxygen in a ratio of two of their atoms to one oxygen atom, whereas Ca, Sr, and Ba form compounds with one of their atoms to one oxygen atom. Fluorine (F), chlorine (Cl), bromine (Br), and iodine (I) also exhibit similar properties to each other, but these properties are drastically different from those of any of the elements above.

Dimitri **Mendeleev** in Russia (1869) and Lothar **Meyer** in Germany (1870) independently recognized that there was a periodic relationship among the properties of the elements known at that time. Both published tables with the elements arranged according to increasing atomic mass. But Mendeleev went one step further than Meyer: He used his table to predict the existence of elements that would have the properties similar to aluminum and silicon, but were yet unknown. The discoveries of gallium (1875) and germanium (1886) provided great support for Mendeleev's work. Although Mendeleev and Meyer had a long dispute over priority, Mendeleev's contributions to the development of the periodic table are now more widely recognized (Figure 1).

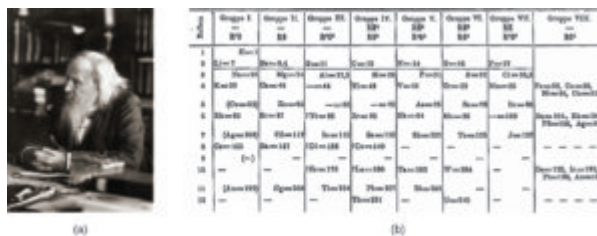


Figure 1 (a) Dimitri Mendeleev is widely credited with creating (b) the first periodic table of the elements. (credit a: modification of work by Serge Lachinov; credit b: modification of work by "Den fjättrade ankan"/Wikimedia Commons)

By the twentieth century, it became apparent that the periodic relationship involved atomic numbers rather than atomic masses. The modern statement of this relationship, the periodic law, is as follows: *the properties of the elements are periodic functions of their atomic numbers*. A modern periodic table arranges the elements in increasing order of their atomic numbers and groups atoms with similar properties in the same vertical column (Figure 2). Each box represents an element and contains its atomic number, symbol, average atomic mass, and (sometimes) name. The elements are arranged in seven horizontal rows, called periods or series, and 18 vertical columns, called groups. Groups are labeled at the top of each column. In the United States, the labels traditionally were numerals with capital letters. However, IUPAC recommends that the numbers 1 through 18 be used, and these labels are more common. For the table to fit on a single page, parts of two of the rows, a total of 14 columns, are usually written

below the main body of the table.

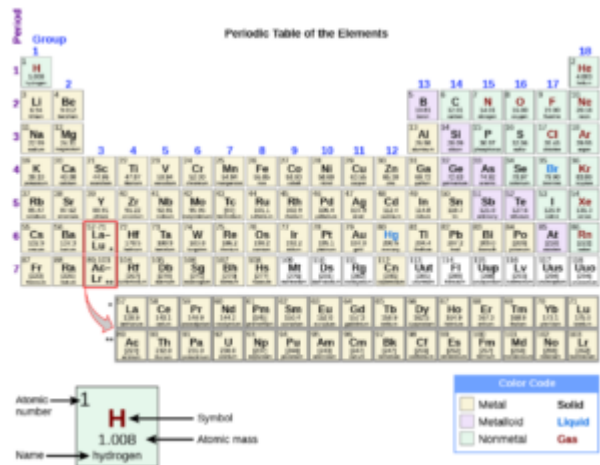


Figure 2 Elements in the periodic table are organized according to their properties.

Many elements differ dramatically in their chemical and physical properties, but some elements are similar in their behaviors. For example, many elements appear shiny, are malleable (able to be deformed without breaking) and ductile (can be drawn into wires), and conduct heat and electricity well. Other elements are not shiny, malleable, or ductile, and are poor conductors of heat and electricity. We can sort the elements into large classes with common properties: metals (elements that are shiny, malleable, good conductors of heat and electricity—shaded yellow); nonmetals (elements that appear dull, poor conductors of heat and electricity—shaded green); and metalloids (elements that conduct heat and electricity moderately well, and possess some properties of metals and some properties of nonmetals—shaded purple).

The elements can also be classified into the main-group elements (or representative elements) in the columns labeled 1, 2, and 13-18; the transition metals in the columns labeled 3-12; and inner transition metals in the two rows at the bottom of the table (the top-row elements are called lanthanides and the bottom-row elements are actinides; Figure 3). The elements can be subdivided further by more specific properties, such as the composition of the compounds they form. For example, the elements in group 1 (the first column) form compounds that consist of one atom of the element and one atom of hydrogen. These elements (except hydrogen) are known as alkali metals, and they all have similar chemical properties. The elements in group 2 (the second column) form compounds consisting of one atom of the element and two atoms of hydrogen: These are called alkaline earth metals, with similar properties among members of that group. Other groups with specific names are the pnictogens (group 15), chalcogens (group 16), halogens (group 17), and the noble gases (group 18, also known as inert gases). The groups can also be referred to by the first element of the group: For example, the chalcogens can be called the oxygen group or oxygen family. Hydrogen is a unique, nonmetallic element with properties similar to both group 1A and group 7A elements. For that reason, hydrogen may be shown at the top of both groups, or by itself.

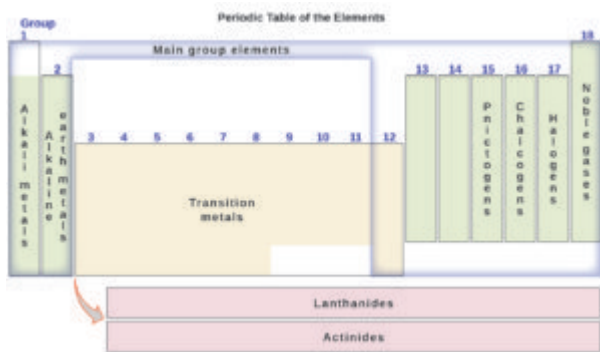


Figure 3 The periodic table organizes elements with similar properties into groups.

Click on this [link](#) for an interactive periodic table, which you can use to explore the properties of the elements (includes podcasts and videos of each element). You may also want to try this [one](#) that shows photos of all the elements.

Example 1: Naming Groups of Elements

Atoms of each of the following elements are essential for life. Give the group name for the following elements:

- (a) chlorine
- (b) calcium
- (c) sodium
- (d) sulfur

Solution:

The family names are as follows:

- (a) halogen
- (b) alkaline earth metal
- (c) alkali metal
- (d) chalcogen

Check Your Learning:

Give the group name for each of the following elements:

- (a) krypton
- (b) selenium
- (c) barium
- (d) lithium

Answer:

(a) noble gas; (b) chalcogen; (c) alkaline earth metal; (d) alkali metal

In studying the periodic table, you might have noticed something about the atomic masses of some of the elements. Element 43 (technetium), element 61 (promethium), and most of the elements with atomic number 84 (polonium) and higher have their atomic mass given in square brackets. This is done for elements that consist entirely of unstable, radioactive isotopes (you will learn more about radioactivity in the nuclear chemistry chapter). An average atomic weight cannot be determined for these elements because their radioisotopes may vary significantly in relative abundance, depending on the source, or may not even exist in nature. The number in square brackets is the atomic mass number (and approximate atomic mass) of the most stable isotope of that element.

Key Concepts and Summary

The discovery of the periodic recurrence of similar properties among the elements led to the formulation of the periodic table, in which the elements are arranged in order of increasing atomic number in rows known as periods and columns known as groups. Elements in the same group of the periodic table have similar chemical properties. Elements can be classified as metals, metalloids, and nonmetals, or as a main-group elements, transition metals, and inner transition metals. Groups are numbered 1–18 from left to right. The elements in group 1 are known as the alkali metals; those in group 2 are the alkaline earth metals; those in 15 are the pnictogens; those in 16 are the chalcogens; those in 17 are the halogens; and those in 18 are the noble gases.

The Periodic Table Exercises

Using the periodic table, classify each of the following elements as a metal or a nonmetal, and

then further classify each as a main-group (representative) element, transition metal, or inner transition metal:

- (a) uranium
- (b) bromine
- (c) strontium
- (d) neon
- (e) gold
- (f) americium
- (g) rhodium
- (h) sulfur
- (i) carbon
- (j) potassium

Using the periodic table, classify each of the following elements as a metal or a nonmetal, and then further classify each as a main-group (representative) element, transition metal, or inner transition metal:

- (a) cobalt
- (b) europium
- (c) iodine
- (d) indium
- (e) lithium
- (f) oxygen
- (h) cadmium
- (i) terbium
- (j) rhenium

Using the periodic table, identify the heaviest member of each of the following groups:

- (a) noble gases
- (b) alkaline earth metals
- (c) alkali metals
- (d) chalcogens

Using the periodic table, identify the heaviest member of each of the following groups:

- (a) alkali metals
- (b) chalcogens
- (c) noble gases
- (d) alkaline earth metals

Use the periodic table to give the name and symbol for each of the following elements:

- (a) the noble gas in the same period as germanium
- (b) the alkaline earth metal in the same period as selenium
- (c) the halogen in the same period as lithium
- (d) the chalcogen in the same period as cadmium

Use the periodic table to give the name and symbol for each of the following elements:

- (a) the halogen in the same period as the alkali metal with 11 protons
- (b) the alkaline earth metal in the same period with the neutral noble gas with 18 electrons
- (c) the noble gas in the same row as an isotope with 30 neutrons and 25 protons
- (d) the noble gas in the same period as gold

Write a symbol for each of the following neutral isotopes. Include the atomic number and mass number for each.

- (a) the alkali metal with 11 protons and a mass number of 23
- (b) the noble gas element with 75 neutrons in its nucleus and 54 electrons in the neutral atom
- (c) the isotope with 33 protons and 40 neutrons in its nucleus
- (d) the alkaline earth metal with 88 electrons and 138 neutrons

Write a symbol for each of the following neutral isotopes. Include the atomic number and mass number for each.

- (a) the chalcogen with a mass number of 125
- (b) the halogen whose longest-lived isotope is radioactive
- (c) the noble gas, used in lighting, with 10 electrons and 10 neutrons
- (d) the lightest alkali metal with three neutrons

Glossary

actinide

inner transition metal in the bottom of the bottom two rows of the periodic table

alkali metal

element in group 1

alkaline earth metal

element in group 2

chalcogen

element in group 16

group

vertical column of the periodic table

halogen

element in group 17

inert gas

(also, noble gas) element in group 18

inner transition metal

(also, lanthanide or actinide) element in the bottom two rows; if in the first row, also called lanthanide, or if in the second row, also called actinide

lanthanide

inner transition metal in the top of the bottom two rows of the periodic table

main-group element

(also, representative element) element in columns 1, 2, and 12-18

metal

element that is shiny, malleable, good conductor of heat and electricity

metalloid

element that conducts heat and electricity moderately well, and possesses some properties of

metals and some properties of nonmetals

noble gas

(also, inert gas) element in group 18

nonmetal

element that appears dull, poor conductor of heat and electricity

period

(also, series) horizontal row of the periodic table

periodic law

properties of the elements are periodic function of their atomic numbers.

periodic table

table of the elements that places elements with similar chemical properties close together

pnictogen

element in group 15

representative element

(also, main-group element) element in columns 1, 2, and 12-18

series

(also, period) horizontal row of the period table

transition metal

element in columns 3-11

Attribution

Atoms, Molecules, and Ions

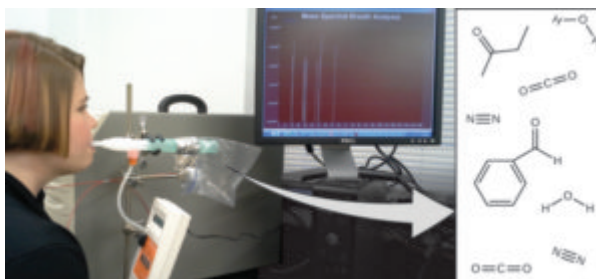


Figure 1 Analysis of molecules in an exhaled breath can provide valuable information, leading to early diagnosis of diseases or detection of environmental exposure to harmful substances. (credit: modification of work by Paul Flowers)

Your overall health and susceptibility to disease depends upon the complex interaction between your genetic makeup and environmental exposure, with the outcome difficult to predict. Early detection of biomarkers, substances that indicate an organism's disease or physiological state, could allow diagnosis and treatment before a condition becomes serious or irreversible. Recent studies have shown that your exhaled breath can contain molecules that may be biomarkers for recent exposure to environmental contaminants or for pathological conditions ranging from asthma to lung cancer. Scientists are working to develop biomarker "fingerprints" that could be used to diagnose a specific disease based on the amounts and identities of certain molecules in a patient's exhaled breath. An essential concept underlying this goal is that of a molecule's identity, which is determined by the numbers and types of atoms it contains, and how they are bonded together. This chapter will describe some of the fundamental chemical principles related to the composition of matter, including those central to the concept of molecular identity.

Attribution

Atomic Structure and Symbolism

The development of modern atomic theory revealed much about the inner structure of atoms. It was learned that an atom contains a very small nucleus composed of positively charged protons and uncharged neutrons, surrounded by a much larger volume of space containing negatively charged electrons. The nucleus contains the majority of an atom's mass because protons and neutrons are much heavier than electrons, whereas electrons occupy almost all of an atom's volume. The diameter of an atom is on the order of 10^{-10} m, whereas the diameter of the nucleus is roughly 10^{-15} m—about 100,000 times smaller. For a perspective about their relative sizes, consider this: If the nucleus were the size of a blueberry, the atom would be about the size of a football stadium (Figure 1).



Figure 1 If an atom could be expanded to the size of a football stadium, the nucleus would be the size of a single blueberry. (credit middle: modification of work by “babyknight”/Wikimedia Commons; credit right: modification of work by Paxson Woelber)

Atoms—and the protons, neutrons, and electrons that compose them—are extremely small. For example, a carbon atom weighs less than 2×10^{-23} g, and an electron has a charge of less than 2×10^{-19} C (coulomb). When describing the properties of tiny objects such as atoms, we use appropriately small units of measure, such as the **atomic mass unit (amu)** and the **fundamental unit of charge (e)**. The amu was originally defined based on hydrogen, the lightest element, then later in terms of oxygen. Since 1961, it has been defined with regard to the most abundant isotope of carbon, atoms of which are assigned masses of exactly 12 amu. (This isotope is known as “carbon-12” as will be discussed later in this module.) Thus, one amu is exactly $\frac{1}{12}$ of the mass of one carbon-12 atom: $1 \text{ amu} = 1.6605 \times 10^{-24} \text{ g}$. (The **Dalton (Da)** and the **unified atomic mass unit (u)** are alternative units that are equivalent to the amu.) The fundamental unit of charge (also called the elementary charge) equals the magnitude of the charge of an electron (e) with $e = 1.602 \times 10^{-19} \text{ C}$.

A proton has a mass of 1.0073 amu and a charge of 1+. A neutron is a slightly heavier particle with a mass 1.0087 amu and a charge of zero; as its name suggests, it is neutral. The electron has a charge of 1– and is a much lighter particle with a mass of about 0.00055 amu (it would take about 1800 electrons to equal the mass of one proton. The properties of these fundamental particles are summarized in Table 1. (An observant student might notice that the sum of an atom's subatomic particles does not equal the atom's actual mass: The total mass of six protons, six neutrons, and six electrons is 12.0993 amu, slightly larger than 12.00 amu. This “missing” mass is known as the mass defect, and you will learn about it in the chapter on nuclear chemistry.)

Table 1

Properties of Subatomic Particles					
Name	Location	Charge (C)	Unit Charge	Mass (amu)	Mass (g)
electron	outside nucleus	-1.602×10^{-19}	1-	0.00055	0.00091×10^{-24}
proton	nucleus	1.602×10^{-19}	1+	1.00727	1.67262×10^{-24}
neutron	nucleus	0	0	1.00866	1.67493×10^{-24}

The number of protons in the nucleus of an atom is its **atomic number (Z)**. This is the defining trait of an element: its value determines the identity of the atom. For example, any atom that contains six protons is the element carbon and has the atomic number 6, regardless of how many neutrons or electrons it may have. A neutral atom must contain the same number of positive and negative charges, so the number of protons equals the number of electrons. Therefore, the atomic number also indicates the number of electrons in an atom. The total number of protons and neutrons in an atom is called its **mass number (A)**. The number of neutrons is therefore the difference between the mass number and the atomic number: $A - Z = \text{number of neutrons}$.

$$\begin{aligned} \text{atomic number (Z)} &= \text{number of protons} \\ \text{atomic mass (A)} &= \text{number of protons} + \text{number of neutrons} \\ A - Z &= \text{number of neutrons} \end{aligned}$$

Atoms are electrically neutral if they contain the same number of positively charged protons and negatively charged electrons. When the numbers of these subatomic particles are *not* equal, the atom is electrically charged and is called an *ion*. The charge of an atom is defined as follows:

$$\text{Atomic charge} = \text{number of protons} - \text{number of electrons}$$

As will be discussed in more detail later in this chapter, atoms (and molecules) typically acquire charge by gaining or losing electrons. An atom that gains one or more electrons will exhibit a negative charge and is called an **anion**. Positively charged atoms called **cations** are formed when an atom loses one or more electrons. For example, a neutral sodium atom ($Z = 11$) has 11 electrons. If this atom loses one electron, it will become a cation with a 1+ charge ($11 - 10 = 1+$). A neutral oxygen atom ($Z = 8$) has eight electrons, and if it gains two electrons it will become an anion with a 2- charge ($8 - 10 = 2-$).

Example 1: Composition of an Atom

Iodine is an essential trace element in our diet; it is needed to produce thyroid hormone. Insufficient iodine in the diet can lead to the development of a goiter, an enlargement of the thyroid gland (Figure 2).

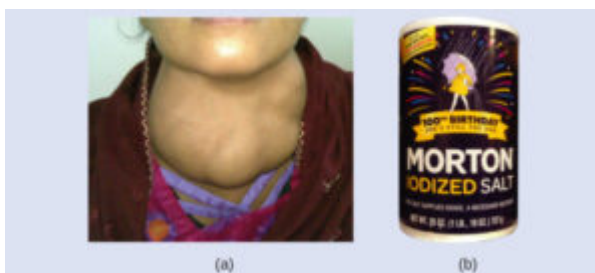


Figure 2 (a) Insufficient iodine in the diet can cause an enlargement of the thyroid gland called a goiter. (b) The addition of small amounts of iodine to salt, which prevents the formation of goiters, has helped eliminate this concern in the US where salt consumption is high. (credit a: modification of work by "Almazi"/Wikimedia Commons; credit b: modification of work by Mike Mozart)

Solution:

The atomic number of iodine (53) tells us that a neutral iodine atom contains 53 protons in its nucleus and 53 electrons outside its nucleus. Because the sum of the numbers of protons and neutrons equals the mass number, 127, the number of neutrons is 74 ($127 - 53 = 74$). Since the iodine is added as a 1- anion, the number of electrons is 54 [$53 - (1-) = 54$].

Check Your Learning:

An atom of platinum has a mass number of 195 and contains 74 electrons. How many protons and neutrons does it contain, and what is its charge?

Answer:

78 protons; 117 neutrons; charge is 4+

Chemical Symbols

A **chemical symbol** is an abbreviation that we use to indicate an element or an atom of an element. For example, the symbol for mercury is Hg (Figure 3). We use the same symbol to indicate one atom of mercury (microscopic domain) or to label a container of many atoms of the element mercury (macroscopic domain).



Figure 3 The symbol Hg represents the element mercury regardless of the amount; it could represent one atom of mercury or a large amount of mercury.

The symbols for several common elements and their atoms are listed in Table 2. Some symbols are derived from the common name of the element; others are abbreviations of the name in another language. Most symbols have one or two letters, but three-letter symbols have been used to describe some elements that have atomic numbers greater than 112. To avoid confusion with other notations, only the first letter of a symbol is capitalized. For example, Co is the symbol for the element cobalt, but CO is the notation for the compound carbon monoxide, which contains atoms of the elements carbon (C) and oxygen (O). All known elements and their symbols are in the [periodic table](#).

Table 2.

Some Common Elements and Their Symbols			
Element	Symbol	Element	Symbol
aluminum	Al	iron	Fe (from <i>ferrum</i>)
bromine	Br	lead	Pb (from <i>plumbum</i>)

Some Common Elements and Their Symbols			
Element	Symbol	Element	Symbol
calcium	Ca	magnesium	Mg
carbon	C	mercury	Hg (from <i>hydrargyrum</i>)
chlorine	Cl	nitrogen	N
chromium	Cr	oxygen	O
cobalt	Co	potassium	K (from <i>kalium</i>)
copper	Cu (from <i>cuprum</i>)	silicon	Si
fluorine	F	silver	Ag (from <i>argentum</i>)
gold	Au (from <i>aurum</i>)	sodium	Na (from <i>natrium</i>)
helium	He	sulfur	S
hydrogen	H	tin	Sn (from <i>stannum</i>)
iodine	I	zinc	Zn

Traditionally, the discoverer (or discoverers) of a new element names the element. However, until the name is recognized by the International Union of Pure and Applied Chemistry (IUPAC), the recommended name of the new element is based on the Latin word(s) for its atomic number. For example, element 106 was called unnilhexium (Unh), element 107 was called unnilseptium (Uns), and element 108 was called unniloctium (Uno) for several years. These elements are now named after scientists (or occasionally locations); for example, element 106 is now known as *seaborgium* (Sg) in honor of Glenn Seaborg, a Nobel Prize winner who was active in the discovery of several heavy elements.

Visit this [site](#) to learn more about IUPAC, the International Union of Pure and Applied Chemistry, and explore its periodic table.

Key Concepts and Summary

An atom consists of a small, positively charged nucleus surrounded by electrons. The nucleus contains protons and neutrons; its diameter is about 100,000 times smaller than that of the atom. The mass of one atom is usually expressed in atomic mass units (amu), which is referred to as the atomic mass. An amu is defined as exactly 1/12 of the mass of a carbon-12 atom and is equal to 1.6605×10^{-24} g.

Protons are relatively heavy particles with a charge of 1+ and a mass of 1.0073 amu. Neutrons are relatively heavy particles with no charge and a mass of 1.0087 amu. Electrons are light particles with a charge of 1– and a mass of 0.00055 amu. The number of protons in the nucleus is called the atomic number (Z) and is the property that defines an atom's elemental identity. The sum of the numbers of protons and neutrons in the nucleus is called the mass number and, expressed in amu, is approximately equal to the mass of the atom. An atom is neutral when it contains equal numbers of electrons and protons.

Atomic Structure and Symbolism Exercises

Write the symbol for each of the following ions:

- the ion with a 1+ charge, atomic number 55, and mass number 133
- the ion with 54 electrons, 53 protons, and 74 neutrons

(c) the ion with atomic number 15, mass number 31, and a 3⁻ charge

(d) the ion with 24 electrons, 30 neutrons, and a 3⁺ charge

Write the symbol for each of the following ions:

(a) the ion with a 3⁺ charge, 28 electrons, and a mass number of 71

(b) the ion with 36 electrons, 35 protons, and 45 neutrons

(c) the ion with 86 electrons, 142 neutrons, and a 4⁺ charge

(d) the ion with a 2⁺ charge, atomic number 38, and mass number 87

Open the [Build an Atom simulation](#) and click on the Atom icon.

(a) Pick any one of the first 10 elements that you would like to build and state its symbol.

(b) Drag protons, neutrons, and electrons onto the atom template to make an atom of your element. State the numbers of protons, neutrons, and electrons in your atom, as well as the net charge and mass number.

(c) Click on "Net Charge" and "Mass Number," check your answers to (b), and correct, if needed.

(d) Predict whether your atom will be stable or unstable. State your reasoning.

(e) Check the "Stable/Unstable" box. Was your answer to (d) correct? If not, first predict what you can do to make a stable atom of your element, and then do it and see if it works. Explain your reasoning.

Open the [Build an Atom simulation](#)

(a) Drag protons, neutrons, and electrons onto the atom template to make a neutral atom of Lithium-6 and give the isotope symbol for this atom.

(b) Now remove one electron to make an ion and give the symbol for the ion you have created.

Glossary

anion

negatively charged atom or molecule (contains more electrons than protons)

atomic mass

average mass of atoms of an element, expressed in amu

atomic mass unit (amu)

(also, unified atomic mass unit, u, or Dalton, Da) unit of mass equal to 1/12 of the mass of a ¹²C atom

atomic number (Z)

number of protons in the nucleus of an atom

cation

positively charged atom or molecule (contains fewer electrons than protons)

chemical symbol

one-, two-, or three-letter abbreviation used to represent an element or its atoms

Dalton (Da)

alternative unit equivalent to the atomic mass unit

fundamental unit of charge

(also called the elementary charge) equals the magnitude of the charge of an electron (e) with $e = 1.602 \times 10^{-19} \text{ C}$

ion

electrically charged atom or molecule (contains unequal numbers of protons and electrons)

mass number (A)

sum of the numbers of neutrons and protons in the nucleus of an atom

unified atomic mass unit (u)

alternative unit equivalent to the atomic mass unit

Attribution

Chemical Formulas

A **molecular formula** is a representation of a molecule that uses chemical symbols to indicate the types of atoms followed by subscripts to show the number of atoms of each type in the molecule. (A subscript is used only when more than one atom of a given type is present.) Molecular formulas are also used as abbreviations for the names of compounds.

The **structural formula** for a compound gives the same information as its molecular formula (the types and numbers of atoms in the molecule) but also shows how the atoms are connected in the molecule. The structural formula for methane contains symbols for one C atom and four H atoms, indicating the number of atoms in the molecule (Figure 1). The lines represent bonds that hold the atoms together. (A chemical bond is an attraction between atoms or ions that holds them together in a molecule or a crystal.) We will discuss chemical bonds and see how to predict the arrangement of atoms in a molecule later. For now, simply know that the lines are an indication of how the atoms are connected in a molecule. A ball-and-stick model shows the geometric arrangement of the atoms with atomic sizes not to scale, and a space-filling model shows the relative sizes of the atoms.

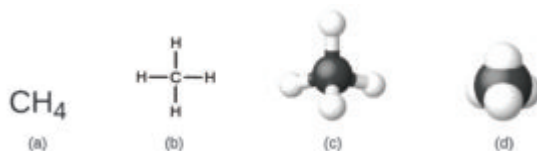


Figure 1 A methane molecule can be represented as (a) a molecular formula, (b) a structural formula, (c) a ball-and-stick model, and (d) a space-filling model. Carbon and hydrogen atoms are represented by black and white spheres, respectively.

Although many elements consist of discrete, individual atoms, some exist as molecules made up of two or more atoms of the element chemically bonded together. For example, most samples of the elements hydrogen, oxygen, and nitrogen are composed of molecules that contain two atoms each (called diatomic molecules) and thus have the molecular formulas H_2 , O_2 , and N_2 , respectively. Other elements commonly found as diatomic molecules are fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2). The most common form of the element sulfur is composed of molecules that consist of eight atoms of sulfur; its molecular formula is S_8 (Figure 2).

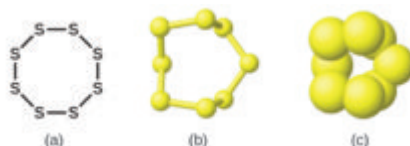


Figure 2 A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S_8 . It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres.

It is important to note that a subscript following a symbol and a number in front of a symbol do not represent the same thing; for example, H_2 and $2H$ represent distinctly different species. H_2 is a molecular

formula; it represents a diatomic molecule of hydrogen, consisting of two atoms of the element that are chemically bonded together. The expression 2H , on the other hand, indicates two separate hydrogen atoms that are not combined as a unit. The expression 2H_2 represents two molecules of diatomic hydrogen (Figure 3).



Figure 3 The symbols H , 2H , H_2 , and 2H_2 represent very different entities.

Compounds are formed when two or more elements chemically combine, resulting in the formation of bonds. For example, hydrogen and oxygen can react to form water, and sodium and chlorine can react to form table salt. We sometimes describe the composition of these compounds with an **empirical formula**, which indicates the types of atoms present and *the simplest whole-number ratio of the number of atoms (or ions) in the compound*. For example, titanium dioxide (used as pigment in white paint and in the thick, white, blocking type of sunscreen) has an empirical formula of TiO_2 . This identifies the elements titanium (Ti) and oxygen (O) as the constituents of titanium dioxide, and indicates the presence of twice as many atoms of the element oxygen as atoms of the element titanium (Figure 4).

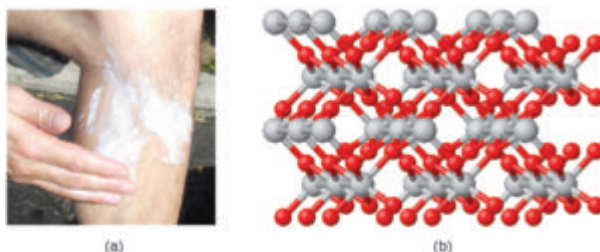


Figure 4 (a) The white compound titanium dioxide provides effective protection from the sun. (b) A crystal of titanium dioxide, TiO_2 , contains titanium and oxygen in a ratio of 1 to 2. The titanium atoms are gray and the oxygen atoms are red. (credit a: modification of work by "osseous"/Flickr)

As discussed previously, we can describe a compound with a molecular formula, in which the subscripts indicate the actual numbers of atoms of each element in a molecule of the compound. In many cases, the molecular formula of a substance is derived from experimental determination of both its empirical formula and its molecular mass (the sum of atomic masses for all atoms composing the molecule). For example, it can be determined experimentally that benzene contains two elements, carbon (C) and hydrogen (H), and that for every carbon atom in benzene, there is one hydrogen atom. Thus, the empirical formula is CH. An experimental determination of the molecular mass reveals that a molecule of benzene contains six carbon atoms and six hydrogen atoms, so the molecular formula for benzene is C_6H_6 (Figure 5).

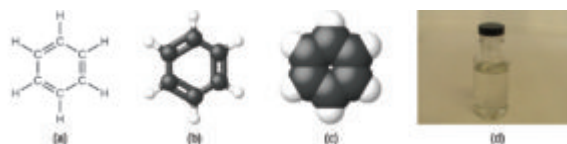


Figure 5 Benzene, C_6H_6 , is produced during oil refining and has many industrial uses. A benzene molecule can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. (d) Benzene is a clear liquid. (credit d: modification of work by Sahar Atwa)

If we know a compound's formula, we can easily determine the empirical formula. (This is somewhat of an academic exercise; the reverse chronology is generally followed in actual practice.) For example, the molecular formula for acetic acid, the component that gives vinegar its sharp taste, is $\text{C}_2\text{H}_4\text{O}_2$. This formula

indicates that a molecule of acetic acid (Figure 6) contains two carbon atoms, four hydrogen atoms, and two oxygen atoms. The ratio of atoms is 2:4:2. Dividing by the lowest common denominator (2) gives the simplest, whole-number ratio of atoms, 1:2:1, so the empirical formula is CH_2O . Note that a molecular formula is always a whole-number multiple of an empirical formula.

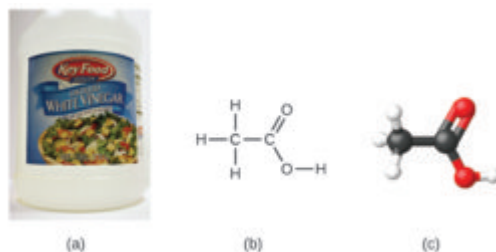


Figure 6 (a) Vinegar contains acetic acid, $\text{C}_2\text{H}_4\text{O}_2$, which has an empirical formula of CH_2O . It can be represented as (b) a structural formula and (c) as a ball-and-stick model. (credit a: modification of work by "HomeSpot HQ"/Flickr)

Empirical and Molecular Formulas

Molecules of glucose (blood sugar) contain 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms. What are the molecular and empirical formulas of glucose?

Solution

The molecular formula is $\text{C}_6\text{H}_{12}\text{O}_6$ because one molecule actually contains 6 C, 12 H, and 6 O atoms. The simplest whole-number ratio of C to H to O atoms in glucose is 1:2:1, so the empirical formula is CH_2O .

Check Your Learning

A molecule of metaldehyde (a pesticide used for snails and slugs) contains 8 carbon atoms, 16 hydrogen atoms, and 4 oxygen atoms. What are the molecular and empirical formulas of metaldehyde?

ANSWER:

Molecular formula, $\text{C}_8\text{H}_{16}\text{O}_4$; empirical formula, $\text{C}_2\text{H}_4\text{O}$

You can explore [molecule building](#) using an online simulation.

Note: LEE CRONIN

What is it that chemists do? According to Lee Cronin (Figure 7), chemists make very complicated molecules by "chopping up" small molecules and "reverse engineering" them. He wonders if we could "make a really cool universal chemistry set" by what he calls "app-ing" chemistry. Could we "app" chemistry?

In a 2012 TED talk, Lee describes one fascinating possibility: combining a collection of chemical "inks" with a 3D printer capable of fabricating a reaction apparatus (tiny test tubes, beakers, and the like) to fashion a "universal toolkit of chemistry." This toolkit could be used to create custom-tailored drugs to fight a new superbug or to "print" medicine personally configured to your genetic makeup, environment, and health situation. Says Cronin, "What Apple did for music, I'd like to do for the

discovery and distribution of prescription drugs.”^[1] View his [full talk](#) at the TED website.



Figure 7 Chemist Lee Cronin has been named one of the UK's 10 most inspirational scientists. The youngest chair at the University of Glasgow, Lee runs a large research group, collaborates with many scientists worldwide, has published over 250 papers in top scientific journals, and has given more than 150 invited talks. His research focuses on complex chemical systems and their potential to transform technology, but also branches into nanoscience, solar fuels, synthetic biology, and even artificial life and evolution. (credit: image courtesy of Lee Cronin)

It is important to be aware that it may be possible for the same atoms to be arranged in different ways: Compounds with the same molecular formula may have different atom-to-atom bonding and therefore different structures. For example, could there be another compound with the same formula as acetic acid, $C_2H_4O_2$? And if so, what would be the structure of its molecules?

If you predict that another compound with the formula $C_2H_4O_2$ could exist, then you demonstrated good chemical insight and are correct. Two C atoms, four H atoms, and two O atoms can also be arranged to form a methyl formate, which is used in manufacturing, as an insecticide, and for quick-drying finishes. Methyl formate molecules have one of the oxygen atoms between the two carbon atoms, differing from the arrangement in acetic acid molecules. Acetic acid and methyl formate are examples of isomers—compounds with the same chemical formula but different molecular structures (Figure 8). Note that this small difference in the arrangement of the atoms has a major effect on their respective chemical properties. You would certainly not want to use a solution of methyl formate as a substitute for a solution of acetic acid (vinegar) when you make salad dressing.

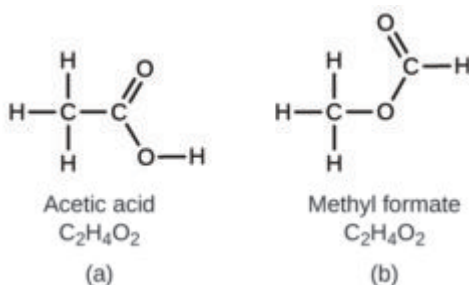


Figure 8 Molecules of (a) acetic acid and methyl formate (b) are structural isomers; they have the same formula

($C_2H_4O_2$) but different structures (and therefore different chemical properties).

Many types of isomers exist (Figure 9). Acetic acid and methyl formate are structural isomers, compounds in which the molecules differ in how the atoms are connected to each other. There are also various types of spatial isomers, in which the relative orientations of the atoms in space can be different. For example, the compound carvone (found in caraway seeds, spearmint, and mandarin orange peels) consists of two isomers that are mirror images of each other. *S*-(+)-carvone smells like caraway, and *R*-(-)-carvone smells like spearmint.

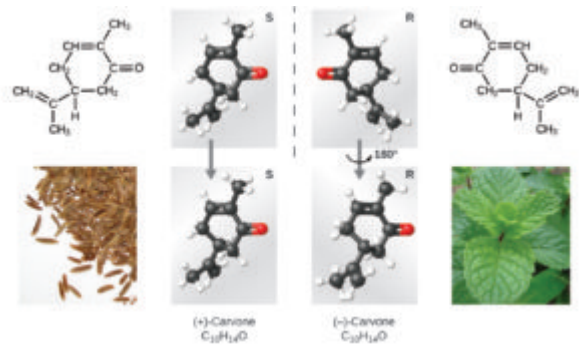


Figure 9 Molecules of carvone are spatial isomers; they only differ in the relative orientations of the atoms in space. (credit bottom left: modification of work by “Miansari66”/Wikimedia Commons; credit bottom right: modification of work by Forest & Kim Starr)

Select this [link](#) to view an explanation of isomers, spatial isomers, and why they have different smells (select the video titled “Mirror Molecule: Carvone”).

Key Concepts and Summary

A molecular formula uses chemical symbols and subscripts to indicate the exact numbers of different atoms in a molecule or compound. An empirical formula gives the simplest, whole-number ratio of atoms in a compound. A structural formula indicates the bonding arrangement of the atoms in the molecule. Ball-and-stick and space-filling models show the geometric arrangement of atoms in a molecule. Isomers are compounds with the same molecular formula but different arrangements of atoms.

Chemical Formulas Exercises

Explain why the symbol for an atom of the element oxygen and the formula for a molecule of oxygen differ.

Explain why the symbol for the element sulfur and the formula for a molecule of sulfur differ.

Write the molecular and empirical formulas of the following compounds:

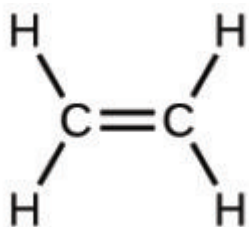
(a)



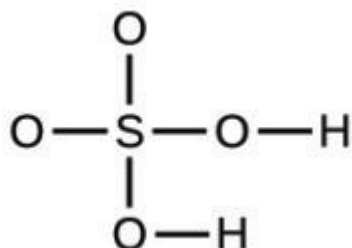
(b)



(c)

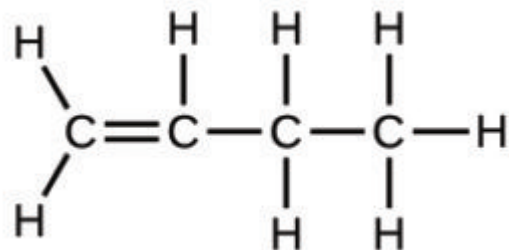


(d)

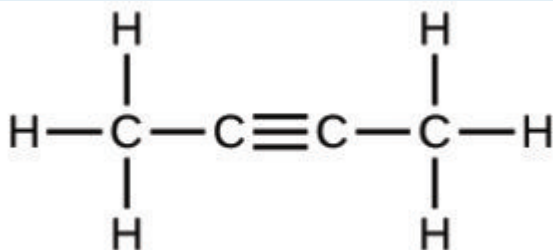


Write the molecular and empirical formulas of the following compounds:

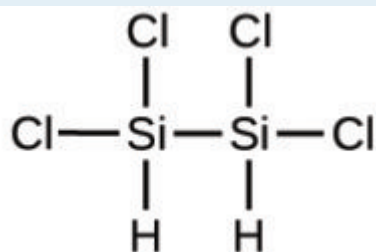
(a)



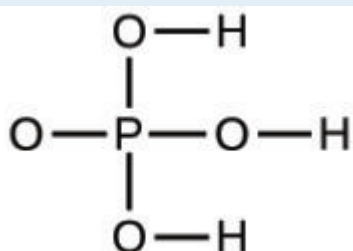
(b)



(c)



(d)



Determine the empirical formulas for the following compounds:

(a) caffeine, $\text{C}_8\text{H}_{10}\text{N}_4\text{O}_2$

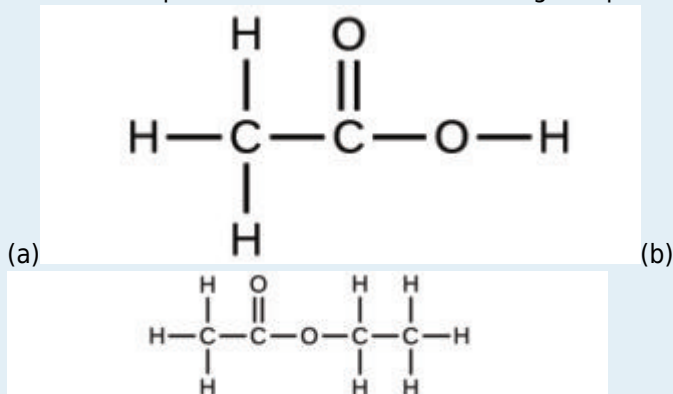
(b) fructose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$

- (c) hydrogen peroxide, H_2O_2
- (d) glucose, $\text{C}_6\text{H}_{12}\text{O}_6$
- (e) ascorbic acid (vitamin C), $\text{C}_6\text{H}_8\text{O}_7$

Determine the empirical formulas for the following compounds:

- (a) acetic acid, $\text{C}_2\text{H}_4\text{O}_2$
- (b) citric acid, $\text{C}_6\text{H}_8\text{O}_7$
- (c) hydrazine, N_2H_4
- (d) nicotine, $\text{C}_{10}\text{H}_{14}\text{N}_2$
- (e) butane, C_4H_{10}

Write the empirical formulas for the following compounds:



Open the [Build a Molecule simulation](#) and select the “Larger Molecules” tab. Select an appropriate atoms “Kit” to build a molecule with two carbon and six hydrogen atoms. Drag atoms into the space above the “Kit” to make a molecule. A name will appear when you have made an actual molecule that exists (even if it is not the one you want). You can use the scissors tool to separate atoms if you would like to change the connections. Click on “3D” to see the molecule, and look at both the space-filling and ball-and-stick possibilities.

- (a) Draw the structural formula of this molecule and state its name.
- (b) Can you arrange these atoms in any way to make a different compound?

Use the [Build a Molecule simulation](#) to repeat the previous exercise, but build a molecule with two carbons, six hydrogens, and one oxygen.

- (a) Draw the structural formula of this molecule and state its name.
- (b) Can you arrange these atoms to make a different molecule? If so, draw its structural formula and state its name.
- (c) How are the molecules drawn in (a) and (b) the same? How do they differ? What are they called (the type of relationship between these molecules, not their names).

Use the [Build a Molecule simulation](#) to repeat exercise 8, but build a molecule with three carbons, seven hydrogens, and one chlorine.

- (a) Draw the structural formula of this molecule and state its name.
- (b) Can you arrange these atoms to make a different molecule? If so, draw its structural formula and state its name.
- (c) How are the molecules drawn in (a) and (b) the same? How do they differ? What are they called (the type of relationship between these molecules, not their names)?

Glossary

empirical formula

formula showing the composition of a compound given as the simplest whole-number ratio of atoms

isomers

compounds with the same chemical formula but different structures

molecular formula

formula indicating the composition of a molecule of a compound and giving the actual number of atoms of each element in a molecule of the compound.

spatial isomers

compounds in which the relative orientations of the atoms in space differ

structural formula

shows the atoms in a molecule and how they are connected

structural isomer

one of two substances that have the same molecular formula but different physical and chemical properties because their atoms are bonded differently

Attribution

Lee Cronin, "Print Your Own Medicine," Talk presented at TED Global 2012, Edinburgh, Scotland, June 2012. [↩](#)

Molecular and Ionic Compounds

In ordinary chemical reactions, the nucleus of each atom (and thus the identity of the element) remains unchanged. Electrons, however, can be added to atoms by transfer from other atoms, lost by transfer to other atoms, or shared with other atoms. The transfer and sharing of electrons among atoms govern the chemistry of the elements. During the formation of some compounds, atoms gain or lose electrons, and form electrically charged particles called ions (Figure 1).

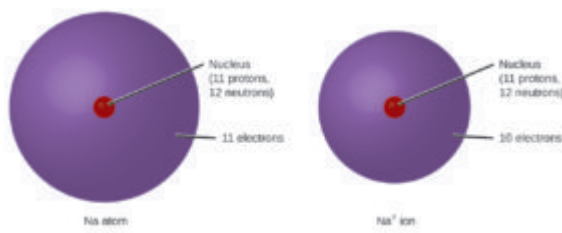


Figure 1 (a) A sodium atom (Na) has equal numbers of protons and electrons (11) and is uncharged. (b) A sodium cation (Na⁺) has lost an electron, so it has one more proton (11) than electrons (10), giving it an overall positive charge, signified by a superscripted plus sign.

You can use the periodic table to predict whether an atom will form an anion or a cation, and you can often predict the charge of the resulting ion. Atoms of many main-group metals lose enough electrons to leave them with the same number of electrons as an atom of the preceding noble gas. To illustrate, an atom of an alkali metal (group 1) loses one electron and forms a cation with a 1+ charge; an alkaline earth metal (group 2) loses two electrons and forms a cation with a 2+ charge, and so on. For example, a neutral calcium atom, with 20 protons and 20 electrons, readily loses two electrons. This results in a cation with 20 protons, 18 electrons, and a 2+ charge. It has the same number of electrons as atoms of the preceding noble gas, argon, and is symbolized Ca²⁺. The name of a metal ion is the same as the name of the metal atom from which it forms, so Ca²⁺ is called a calcium ion.

When atoms of nonmetal elements form ions, they generally gain enough electrons to give them the same number of electrons as an atom of the next noble gas in the periodic table. Atoms of group 17 gain one electron and form anions with a 1− charge; atoms of group 16 gain two electrons and form ions with a 2− charge, and so on. For example, the neutral bromine atom, with 35 protons and 35 electrons, can gain one electron to provide it with 36 electrons. This results in an anion with 35 protons, 36 electrons, and a 1− charge. It has the same number of electrons as atoms of the next noble gas, krypton, and is symbolized Br[−]. (A discussion of the theory supporting the favored status of noble gas electron numbers reflected in these predictive rules for ion formation is provided in a later chapter of this text.)

Note the usefulness of the periodic table in predicting likely ion formation and charge (Figure). Moving from the far left to the right on the periodic table, main-group elements tend to form cations with a charge equal to the group number. That is, group 1 elements form 1+ ions; group 2 elements form 2+ ions, and so on. Moving from the far right to the left on the periodic table, elements often form anions with a negative charge equal to the number of groups moved left from the noble gases. For example, group 17 elements (one group left of the noble gases) form 1− ions; group 16 elements (two groups left) form 2− ions, and so

on. This trend can be used as a guide in many cases, but its predictive value decreases when moving toward the center of the periodic table. In fact, transition metals and some other metals often exhibit variable charges that are not predictable by their location in the table. For example, copper can form ions with a 1+ or 2+ charge, and iron can form ions with a 2+ or 3+ charge.

Figure 2 Some elements exhibit a regular pattern of ionic charge when they form ions.

Example 1: Composition of Ions

An ion found in some compounds used as antiperspirants contains 13 protons and 10 electrons. What is its symbol?

Solution:

Because the number of protons remains unchanged when an atom forms an ion, the atomic number of the element must be 13. Knowing this lets us use the periodic table to identify the element as Al (aluminum). The Al atom has lost three electrons and thus has three more positive charges (13) than it has electrons (10). This is the aluminum cation, Al^{3+} .

Check Your Learning:

Give the symbol and name for the ion with 34 protons and 36 electrons.

Answer:

Se^{2-} , the selenide ion

Example 2: Formation of Ions

Magnesium and nitrogen react to form an ionic compound. Predict which forms an anion, which forms a cation, and the charges of each ion. Write the symbol for each ion and name them.

Solution:

Magnesium's position in the periodic table (group 2) tells us that it is a metal. Metals form positive ions (cations). A magnesium atom must lose two electrons to have the same number electrons as an atom of the previous noble gas, neon. Thus, a magnesium atom will form a cation with two fewer electrons than protons and a charge of 2+. The symbol for the ion is Mg^{2+} , and it is called a magnesium ion.

Nitrogen's position in the periodic table (group 15) reveals that it is a nonmetal. Nonmetals form negative ions (anions). A nitrogen atom must gain three electrons to have the same number of electrons as an atom of the following noble gas, neon. Thus, a nitrogen atom will form an anion with

three more electrons than protons and a charge of 3⁻. The symbol for the ion is N³⁻, and it is called a nitride ion.

Check Your Learning:

Aluminum and carbon react to form an ionic compound. Predict which forms an anion, which forms a cation, and the charges of each ion. Write the symbol for each ion and name them.

Answer:

Al will form a cation with a charge of 3+: Al³⁺, an aluminum ion. Carbon will form an anion with a charge of 4-: C⁴⁻, a carbide ion.

The ions that we have discussed so far are called monatomic ions, that is, they are ions formed from only one atom. We also find many polyatomic ions. These ions, which act as discrete units, are electrically charged molecules (a group of bonded atoms with an overall charge). Some of the more important polyatomic ions are listed in Table 1. Oxyanions are polyatomic ions that contain one or more oxygen atoms. At this point in your study of chemistry, you should memorize the names, formulas, and charges of the most common polyatomic ions. Because you will use them repeatedly, they will soon become familiar.

Charge	Name	Formula
1+	ammonium	NH ₄ ⁺
1-	acetate	C ₂ H ₃ O ₂ ⁻
1-	cyanide	CN ⁻
1-	hydroxide	OH ⁻
1-	nitrate	NO ₃ ⁻
1-	nitrite	NO ₂ ⁻
1-	perchlorate	ClO ₄ ⁻
1-	chlorate	ClO ₃ ⁻
1-	chlorite	ClO ₂ ⁻
1-	hypochlorite	ClO ⁻
1-	permanganate	MnO ₄ ⁻
1-	hydrogen carbonate, or bicarbonate	HCO ₃ ⁻
2-	carbonate	CO ₃ ²⁻
2-	peroxide	O ₂ ²⁻

Table 1		
Common Polyatomic Ions		
Charge	Name	Formula
1-	hydrogen sulfate, or bisulfate	HSO_4^-
2-	sulfate	SO_4^{2-}
2-	sulfite	SO_3^{2-}
1-	dihydrogen phosphate	H_2PO_4^-
2-	hydrogen phosphate	HPO_4^{2-}
3-	phosphate	PO_4^{3-}

Note that there is a system for naming some polyatomic ions; *-ate* and *-ite* are suffixes designating polyatomic ions containing more or fewer oxygen atoms. *Per-* (short for “hyper”) and *hypo-* (meaning “under”) are prefixes meaning more oxygen atoms than *-ate* and fewer oxygen atoms than *-ite*, respectively. For example, perchlorate is ClO_4^- , chlorate is ClO_3^- , chlorite is ClO_2^- , and hypochlorite is ClO^- . Unfortunately, the number of oxygen atoms corresponding to a given suffix or prefix is not consistent; for example, nitrate is NO_3^- while sulfate is SO_4^{2-} . This will be covered in more detail in the next module on nomenclature.

The nature of the attractive forces that hold atoms or ions together within a compound is the basis for classifying chemical bonding. When electrons are transferred and ions form, ionic bonds result. Ionic bonds are electrostatic forces of attraction, that is, the attractive forces experienced between objects of opposite electrical charge (in this case, cations and anions). When electrons are “shared” and molecules form, covalent bonds result. Covalent bonds are the attractive forces between the positively charged nuclei of the bonded atoms and one or more pairs of electrons that are located between the atoms. Compounds are classified as ionic or molecular (covalent) on the basis of the bonds present in them.

Ionic Compounds

When an element composed of atoms that readily lose electrons (a metal) reacts with an element composed of atoms that readily gain electrons (a nonmetal), a transfer of electrons usually occurs, producing ions. The compound formed by this transfer is stabilized by the electrostatic attractions (ionic bonds) between the ions of opposite charge present in the compound. For example, when each sodium atom in a sample of sodium metal (group 1) gives up one electron to form a sodium cation, Na^+ , and each chlorine atom in a sample of chlorine gas (group 17) accepts one electron to form a chloride anion, Cl^- , the resulting compound, NaCl , is composed of sodium ions and chloride ions in the ratio of one Na^+ ion for each Cl^- ion. Similarly, each calcium atom (group 2) can give up two electrons and transfer one to each of two chlorine atoms to form CaCl_2 , which is composed of Ca^{2+} and Cl^- ions in the ratio of one Ca^{2+} ion to two Cl^- ions.

A compound that contains ions and is held together by ionic bonds is called an ionic compound. The periodic table can help us recognize many of the compounds that are ionic: When a metal is combined with one or more nonmetals, the compound is usually ionic. This guideline works well for predicting ionic compound formation for most of the compounds typically encountered in an introductory chemistry course. However, it is not always true (for example, aluminum chloride, AlCl_3 , is not ionic).

You can often recognize ionic compounds because of their properties. Ionic compounds are solids that typically melt at high temperatures and boil at even higher temperatures. For example, sodium chloride melts at 801 °C and boils at 1413 °C. (As a comparison, the molecular compound water melts at 0 °C and boils at 100 °C.) In solid form, an ionic compound is not electrically conductive because its ions are unable to flow (“electricity” is the flow of charged particles). When molten, however, it can conduct electricity because its ions are able to move freely through the liquid (Figure 3).

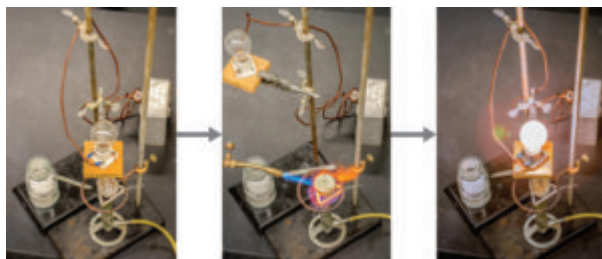


Figure 3 Sodium chloride melts at 801 °C and conducts electricity when molten. (credit: modification of work by Mark Blaser and Matt Evans)

Watch this [video](#) to see a mixture of salts melt and conduct electricity.

In every ionic compound, the total number of positive charges of the cations equals the total number of negative charges of the anions. Thus, ionic compounds are electrically neutral overall, even though they contain positive and negative ions. We can use this observation to help us write the formula of an ionic compound. The formula of an ionic compound must have a ratio of ions such that the numbers of positive and negative charges are equal.

Example 3: Predicting the Formula of an Ionic Compound

The gemstone sapphire (Figure 4) is mostly a compound of aluminum and oxygen that contains aluminum cations, Al^{3+} , and oxygen anions, O^{2-} . What is the formula of this compound?



Figure 4 Although pure aluminum oxide is colorless, trace amounts of iron and titanium give blue sapphire its characteristic color. (credit: modification of work by Stanislav Doronenko)

Solution:

Because the ionic compound must be electrically neutral, it must have the same number of positive and negative charges. Two aluminum ions, each with a charge of 3+, would give us six positive charges, and three oxide ions, each with a charge of 2–, would give us six negative charges. The formula would be Al_2O_3 .

Check Your Learning:

Predict the formula of the ionic compound formed between the sodium cation, Na^+ , and the sulfide anion, S^{2-} .

Answer:



Many ionic compounds contain polyatomic ions (Table 1) as the cation, the anion, or both. As with simple ionic compounds, these compounds must also be electrically neutral, so their formulas can be predicted by treating the polyatomic ions as discrete units. We use parentheses in a formula to indicate a group of atoms that behave as a unit. For example, the formula for calcium phosphate, one of the minerals in our bones, is $\text{Ca}_3(\text{PO}_4)_2$. This formula indicates that there are three calcium ions (Ca^{2+}) for every two phosphate (PO_4^{3-}) groups. The PO_4^{3-} groups are discrete units, each consisting of one phosphorus atom and four oxygen atoms, and having an overall charge of 3-. The compound is electrically neutral, and its formula shows a total count of three Ca, two P, and eight O atoms.

Example 4: Predicting the Formula of a Compound with a Polyatomic Anion

Baking powder contains calcium dihydrogen phosphate, an ionic compound composed of the ions Ca^{2+} and H_2PO_4^- . What is the formula of this compound?

Solution:

The positive and negative charges must balance, and this ionic compound must be electrically neutral. Thus, we must have two negative charges to balance the 2+ charge of the calcium ion. This requires a ratio of one Ca^{2+} ion to two H_2PO_4^- ions. We designate this by enclosing the formula for the dihydrogen phosphate ion in parentheses and adding a subscript 2. The formula is $\text{Ca}(\text{H}_2\text{PO}_4)_2$.

Check Your Learning:

Predict the formula of the ionic compound formed between the lithium ion and the peroxide ion, O_2^{2-} (Hint: Use the periodic table to predict the sign and the charge on the lithium ion.)

Answer:



Because an ionic compound is not made up of single, discrete molecules, it may not be properly symbolized using a molecular formula. Instead, ionic compounds must be symbolized by a formula indicating the *relative numbers* of its constituent cations. For compounds containing only monatomic ions (such as NaCl) and for many compounds containing polyatomic ions (such as CaSO_4), these formulas are just the empirical formulas introduced earlier in this chapter. However, the formulas for some ionic compounds containing polyatomic ions are not empirical formulas. For example, the ionic compound sodium oxalate is comprised of Na^+ and $\text{C}_2\text{O}_4^{2-}$ ions combined in a 2:1 ratio, and its formula is written as $\text{Na}_2\text{C}_2\text{O}_4$. The subscripts in this formula are not the smallest-possible whole numbers, as each can be divided by 2 to yield the empirical formula, NaCO_2 . This is not the accepted formula for sodium oxalate, however, as it does not accurately represent the compound's polyatomic anion, $\text{C}_2\text{O}_4^{2-}$.

Molecular Compounds

Many compounds do not contain ions but instead consist solely of discrete, neutral molecules. These molecular compounds (covalent compounds) result when atoms share, rather than transfer (gain or lose), electrons. Covalent bonding is an important and extensive concept in chemistry, and it will be treated in considerable detail in a later chapter of this text. We can often identify molecular compounds on the basis of their physical properties. Under normal conditions, molecular compounds often exist as gases, low-boiling liquids, and low-melting solids, although many important exceptions exist.

Whereas ionic compounds are usually formed when a metal and a nonmetal combine, covalent compounds are usually formed by a combination of nonmetals. Thus, the periodic table can help us recognize many of the compounds that are covalent. While we can use the positions of a compound's elements in the periodic table to predict whether it is ionic or covalent at this point in our study of chemistry, you should be aware that this is a very simplistic approach that does not account for a number of interesting exceptions. Shades of gray exist between ionic and molecular compounds, and you'll learn more about those later.

Example 5: Predicting the Type of Bonding in Compounds

Predict whether the following compounds are ionic or molecular:

- (a) KI, the compound used as a source of iodine in table salt
- (b) H_2O_2 , the bleach and disinfectant hydrogen peroxide
- (c) CHCl_3 , the anesthetic chloroform
- (d) Li_2CO_3 , a source of lithium in antidepressants

Solution:

- (a) Potassium (group 1) is a metal, and iodine (group 17) is a nonmetal; KI is predicted to be ionic.
- (b) Hydrogen (group 1) is a nonmetal, and oxygen (group 16) is a nonmetal; H_2O_2 is predicted to be molecular.
- (c) Carbon (group 14) is a nonmetal, hydrogen (group 1) is a nonmetal, and chlorine (group 17) is a nonmetal; CHCl_3 is predicted to be molecular.
- (d) Lithium (group 1A) is a metal, and carbonate is a polyatomic ion; Li_2CO_3 is predicted to be ionic.

Check Your Learning:

Using the periodic table, predict whether the following compounds are ionic or covalent:

- (a) SO_2
- (b) CaF_2
- (c) N_2H_4
- (d) $\text{Al}_2(\text{SO}_4)_3$

Answer:

- (a) molecular; (b) ionic; (c) molecular; (d) ionic

Key Concepts and Summary

Metals (particularly those in groups 1 and 2) tend to lose the number of electrons that would leave them with the same number of electrons as in the preceding noble gas in the periodic table. By this means, a positively charged ion is formed. Similarly, nonmetals (especially those in groups 16 and 17, and, to a lesser extent, those in Group 15) can gain the number of electrons needed to provide atoms with the same number of electrons as in the next noble gas in the periodic table. Thus, nonmetals tend to form negative ions. Positively charged ions are called cations, and negatively charged ions are called anions. Ions can be either monatomic (containing only one atom) or polyatomic (containing more than one atom).

Compounds that contain ions are called ionic compounds. Ionic compounds generally form from metals and nonmetals. Compounds that do not contain ions, but instead consist of atoms bonded tightly together in

molecules (uncharged groups of atoms that behave as a single unit), are called covalent compounds. Covalent compounds usually form from two nonmetals.

Molecular and Ionic Compounds Exercises

Using the periodic table, predict whether the following chlorides are ionic or covalent: KCl, NCl_3 , ICl, MgCl_2 , PCl_5 , and CCl_4 .

Using the periodic table, predict whether the following chlorides are ionic or covalent: SiCl_4 , PCl_3 , CaCl_2 , CsCl, CuCl_2 , and CrCl_3 .

For each of the following compounds, state whether it is ionic or covalent. If it is ionic, write the symbols for the ions involved:

- (a) NF_3
- (b) BaO,
- (c) $(\text{NH}_4)_2\text{CO}_3$
- (d) $\text{Sr}(\text{H}_2\text{PO}_4)_2$
- (e) IBr
- (f) Na_2O

For each of the following compounds, state whether it is ionic or covalent, and if it is ionic, write the symbols for the ions involved:

- (a) KClO_4
- (b) $\text{MgC}_2\text{H}_3\text{O}_2$
- (c) H_2S
- (d) Ag_2S
- (e) N_2Cl_4
- (f) $\text{Co}(\text{NO}_3)_2$

For each of the following pairs of ions, write the symbol for the formula of the compound they will form:

- (a) Ca^{2+} , S^{2-}
- (b) NH_4^+ , SO_4^{2-}
- (c) Al^{3+} , Br^-
- (d) Na^+ , HPO_4^{2-}
- (e) Mg^{2+} , PO_4^{3-}

For each of the following pairs of ions, write the symbol for the formula of the compound they will form:

- (a) K^+ , O^{2-}
- (b) NH_4^+ , PO_4^{3-}
- (c) Al^{3+} , O^{2-}
- (d) Na^+ , CO_3^{2-}
- (e) Ba^{2+} , PO_4^{3-}

Glossary

covalent bond

attractive force between the nuclei of a molecule's atoms and pairs of electrons between the atoms

covalent compound

(also, molecular compound) composed of molecules formed by atoms of two or more different elements

ionic bond

electrostatic forces of attraction between the oppositely charged ions of an ionic compound

ionic compound

compound composed of cations and anions combined in ratios, yielding an electrically neutral substance

molecular compound

(also, covalent compound) composed of molecules formed by atoms of two or more different elements

monatomic ion

ion composed of a single atom

oxyanion

polyatomic anion composed of a central atom bonded to oxygen atoms

polyatomic ion

ion composed of more than one atom

Attribution

Cooking Techniques: Thickening and Concentrating Flavors

Viscosity

Viscous means “sticky” and the term *viscosity* refers to the way in which the chocolate flows. Chocolate comes in various viscosities, and the confectioner chooses the one that is most appropriate to his or her needs. The amount of cocoa butter in the chocolate is largely responsible for the viscosity level. Emulsifiers like lecithin can help thin out melted chocolate, so it flows evenly and smoothly. Because it is less expensive than cocoa butter at thinning chocolate, it can be used to help lower the cost of chocolate.

Molded pieces such as Easter eggs require a chocolate of less viscosity. That is, the chocolate should be somewhat runny so it is easier to flow into the moulds. This is also the case for coating cookies and most cakes, where a thin, attractive and protective coating is all that is needed. A somewhat thicker chocolate is advisable for things such as ganache and flavoring of creams and fillings. Where enrobers (machines to dip chocolate centers) are used, the chocolate may also be thinner to ensure that there is an adequate coat of couverture.

Viscosity varies between manufacturers, and a given type of chocolate made by one manufacturer may be available in more than one viscosity. Bakers sometimes alter the viscosity depending on the product. A vegetable oil is sometimes used to thin chocolate for coating certain squares. This makes it easier to cut afterwards.

Chips, Chunks, and Other Baking Products

Content and quality of chocolate chips and chunks vary from one manufacturer to another. This chocolate is developed to be more heat stable for use in cookies and other baking where you want the chips and chunks to stay whole. Ratios of chocolate liquor, sugar, and cocoa butter differ. All these variables affect the flavor.

Chips and chunks may be pure chocolate or have another fat substituted for the cocoa butter. Some high-quality chips have up to 65% chocolate liquor, but in practice, liquor content over 40% tends to smear in baking, so high ratios defeat the purpose.

Many manufacturers package their chips or chunks by count (ct) size. This refers to how many pieces there are in 1 kg of the product. As the count size number increases, the size of the chip gets smaller. With this information, you can choose the best size of chip for the product you are producing.

Other chocolate products available are chocolate sprinkles or “hail,” used as a decoration; chocolate curls, rolls, or decorative shapes for use on cakes and pastries; and chocolate sticks or “batons,” which are often baked inside croissants.

Attribution

Thickening Agents

Learning Objectives

- Identify and describe thickening agents used in the food service industry
- Describe the production and properties of thickening agents
- Describe the function of thickening agents in baking

Two types of thickening agents are recognized: starches and gums. Most thickening agents are of vegetable origin; the only exception is gelatin. All the starches are products of the land; some of the gums are of marine origin.

Bakers use thickening agents primarily to:

- Make fillings easier to handle and bake
- Firm up products to enable them to be served easily
- Provide a glossy “skin” to improve finish and reduce drying

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Types of Thickening Agents

Cornstarch

Cornstarch is the most common thickening agent used in the industry. It is mixed with water or juice and boiled to make fillings and to give a glossy semi-clear finish to products. Commercial cornstarch is made by soaking maize in water containing sulphur dioxide. The soaking softens the corn and the sulphur dioxide prevents possible fermentation. It is then crushed and passed to water tanks where the germ floats off. The mass is then ground fine and, still in a semi-fluid state, passed through silk screens to remove the skin particles. After filtration, the product, which is almost 100% starch, is dried.

Cornstarch in cold water is insoluble, granular, and will settle out if left standing. However, when cornstarch is cooked in water, the starch granules absorb water, swell, and rupture, forming a translucent thickened mixture. This phenomenon is called *gelatinization*. Gelatinization usually begins at about 60°C (140°F), reaching completion at the boiling point.

The commonly used ingredients in a starch recipe affect the rate of gelatinization of the starch. Sugar, added in a high ratio to the starch, will inhibit the granular swelling. The starch gelatinization will not be completed even after prolonged cooking at normal temperature. The result is a filling of thin consistency, dull color, and a cereal taste. Withhold some of the sugar from the cooking step in such cases, and add it after gelatinization of the starch has been completed.

Other ingredients such as egg, fat, and dry milk solids have a similar effect. Fruits with high acidity such as rhubarb will also inhibit starch setting. Cook the starch paste first and add the fruit afterward.

In cooking a filling, about 1.5 kg (3 1/3 lb.) of sugar should be cooked with the water or juice for every 500 g (18 oz.) of starch used as a thickener. Approximately 100 g (4 oz.) of starch is used to thicken 1 L of water or fruit juice. The higher the acidity of the fruit juice, the more thickener required to hold the gel. Regular cornstarch thickens well but makes a cloudy solution. Another kind of cornstarch, waxy maize starch, makes a more fluid mix of great clarity.

Pre-gelatinized Starches

Pre-gelatinized starches are mixed with sugar and then added to the water or juice. They thicken the filling in the presence of sugar and water without heating. This is due to the starch being precooked and not requiring heat to enable it to absorb and gelatinize. There are several brands of these starches on the market (e.g., Clear Jel), and they all vary in absorption properties. For best results, follow the manufacturer's guidelines. Do not put pre-gelatinized starch directly into water, as it will form lumps immediately.

Note: If fruit fillings are made with these pre-cooked starches, there is a potential for breakdown if the fillings are kept. Enzymes in the uncooked fruit may "attack" the starch and destroy some of the gelatinized structure. For example, if you are making a week's supply of pie filling from fresh rhubarb, use a regular cooked formula.

Arrowroot

Arrowroot is a highly nutritious farinaceous starch obtained from the roots and tubers of various West Indian plants. It is used in the preparation of delicate soups, sauces, puddings, and custards.

Agar-Agar

Agar-agar is a jelly-like substance extracted from red seaweed found off the coasts of Japan, California, and Sri Lanka. It is available in strips or slabs and in powder form. Agar-agar only dissolves in hot water and is colorless. Use it at 1% to make a firm gel. It has a melting point much higher than gelatin and its jelling power is eight times greater. It is used in pie fillings and to some extent in the stiffening of jams. It is a permitted ingredient in some dairy products, including ice cream at 0.5%. One of its largest uses is in the production of materials such as piping jelly and marshmallow.

Algin (Sodium Alginate)

Extracted from kelp, this gum dissolves in cold water and a 1% concentration to give a firm gel. It has the disadvantage of not working well in the presence of acidic fruits. It is popular in uncooked icings because it works well in the cold state and holds a lot of moisture. It reduces stickiness and prevents recrystallization.

Carrageenan or Irish Moss

Carrageenan is another marine gum extracted from red seaweed. It is used as a thickening agent in various products, from icing stabilizers to whipping cream, at an allowable rate of 0.1% to 0.5%.

Gelatin

Gelatin is a glutinous substance made from the bones, connective tissues, and skins of animals. The calcium is removed and the remaining substance is soaked in cold water. Then it is heated to 40°C to 60°C (105°F 140°F). The partially evaporated liquid is defatted and coagulated on glass plates and then poured into moulds. When solid, the blocks of gelatin are cut into thin layers and dried on wire netting.

Gelatin is available in sheets of leaf gelatin, powders, granules, or flakes. Use it at a 1% ratio. Like some of the other gelling agents, acidity adversely affects its gelling capacity.

The quality of gelatin often varies because of different methods of processing and manufacturing. For this reason, many bakers prefer leaf gelatin because of its reliable strength.

Gum Arabic or Acacia

This gum is obtained from various kinds of trees and is soluble in hot or cold water. Solutions of gum arabic are used in the bakery for glazing various kinds of goods, particularly marzipan fruits.

Gum Tragacanth

This gum is obtained from several species of *Astragalus*, low-growing shrubs found in Western Asia. It can be purchased in flakes or powdered form. Gum tragacanth was once used to make gum paste and gum paste wedding ornaments, but due to high labour costs and a prohibitive price for the product, its use nowadays is uncommon.

Pectin

Pectin is a mucilaginous substance (gummy substance extracted from plants), occurring naturally in pears, apples, quince, oranges, and other citrus fruits. It is used as the gelling agent in traditional jams and jellies.

Attribution

Coagulation

Coagulation is defined as the transformation of proteins from a liquid state to a solid form. Once proteins are coagulated, they cannot be returned to their liquid state. Coagulation often begins around 38°C (100°F), and the process is complete between 71°C and 82°C (160°F and 180°F). Within the baking process, the natural structures of the ingredients are altered irreversibly by a series of physical, chemical, and biochemical interactions. The three main types of protein that cause coagulation in the bakeshop are outlined below.

Egg proteins

Eggs contain many different proteins. The white, or **albumen**, contains approximately 40 different proteins, the most predominant being ovalbumin (54%) and ovotransferrin (12%). The yolk contains mostly lipids (fats), but also lipoproteins. These different proteins will all coagulate when heated, but do so at different temperatures. The separated white of an egg coagulates between 60°C and 65°C (140°F and 149°F) and the yolk between 62°C and 70°C (144°F and 158°F), which is why you can cook an egg and have a fully set white and a still runny yolk. These temperatures are raised when eggs are mixed into other liquids. For example, the coagulation and thickening of an egg, milk, and sugar mixture, as in custard, will take place between 80°C and 85°C (176°F and 185°F) and will start to curdle at 88°C to 90°C (190°F and 194°F).

Dairy and soy proteins

Casein, a semi-solid substance formed by the coagulation of milk, is obtained and used primarily in cheese. Rennet, derived from the stomach linings of cattle, sheep, and goats, is used to coagulate, or thicken, milk during the cheese-making process. Plant-based rennet is also available. Chymosin (also called rennin) is the **enzyme** used to produce rennet, and is responsible for curdling the milk, which will then separate into solids (curds) and liquid (whey).

Milk and milk products will also coagulate when treated with an acid, such as citric acid (lemon juice) or vinegar, used in the preparation of fresh ricotta, and tartaric acid, used in the preparation of mascarpone, or will naturally curdle when sour as lactic acid develops in the milk. In some cases, as in the production of yogurt or crème fraîche, acid-causing bacteria are added to the milk product to cause the coagulation. Similarly, **tofu** is made from soybean milk that has been coagulated with the use of either salt, acid, or enzyme-based coagulants.

Flour proteins (gluten)

Two main proteins are found in wheat flour: glutenin and gliadin (smaller quantities are also found in other grains). During mixing and in contact with liquid, these two form into a stretchable substance called **gluten**. The coagulation of gluten is what happens when bread bakes; that is, it is the firming or hardening of these gluten proteins, usually caused by heat, which solidify to form a firm structure.

Attribution

Gelatinization

Hydrocolloids

A **hydrocolloid** is a substance that forms a gel in contact with water. There are two main categories:

- Thermo-reversible gel: A gel that melts upon reheating and sets upon cooling. Examples are gelatin and agar agar.
- Thermo-irreversible gel: A gel that does not melt upon reheating. Examples are cornstarch and pectin.

Excessive heating, however, may cause evaporation of the water and shrinkage of the gel. Hydrocolloids do not hydrate (or dissolve) instantly, and that hydration is associated with swelling, which easily causes lumping. It is therefore necessary to disperse hydrocolloids in water. Classically, this has always been done with cornstarch, where a portion of the liquid from the recipe is mixed to form a “slurry” before being added to the cooking liquid. This can also be done with an immersion blender or a conventional blender, or by mixing the hydrocolloid with a helping agent such as sugar, oil, or alcohol prior to dispersion in water.

Starches

Starch **gelatinization** is the process where starch and water are subjected to heat, causing the starch granules to swell. As a result, the water is gradually absorbed in an irreversible manner. This gives the system a viscous and transparent texture. The result of the reaction is a gel, which is used in sauces, puddings, creams, and other food products, providing a pleasing texture. Starch-based gels are thermo-irreversible, meaning that they do not melt upon heating (unlike gelatin, which we will discuss later). Excessive heating, however, may cause evaporation of the water and shrinkage of the gel.

The most common examples of starch gelatinization are found in sauce and pasta preparations and baked goods.

- In sauces, starches are added to liquids, usually while heating. The starch will absorb liquid and swell, resulting in the liquid becoming thicker. The type of starch determines the final product. Some starches will remain cloudy when cooked; others will remain clear.
- Pasta is made mostly of semolina wheat (durum wheat flour), which contains high amounts of starch. When pasta is cooked in boiling water, the starch in the pasta swells as it absorbs water, and as a result the texture of the pasta softens.

Starch molecules make up the majority of most baked goods, so starch is an important part of the structure. Although starches by themselves generally can't support the shape of the baked items, they do give bulk to the structure. Starches develop a softer structure when baked than proteins do. The softness of the crumb of baked bread is due largely to the starch. The more protein structure there is, the chewier the bread.

Starches can be fairly straightforward extracts of plants, such as **cornstarch**, **tapioca**, or **arrowroot**, but there are also **modified starches** and **pre-gelatinized starches** available that have specific uses. See Table 1 for a list of different thickening and binding agents and their characteristics.

Table 1 - Common starches and gels used in the bakeshop

Starch or Gel	Ratio	Preparation	Characteristics and Uses
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Cornstarch	20-40 g starch thickens 1 L liquid	A slurry (mixture of cornstarch and water) is mixed and added to a simmering liquid while whisking until it dissolves and the liquid thickens; or Cornstarch mixed with sugar, and cold liquid added Thickened mixture simmered until no starch taste remains	Used to thicken sauces when a clear glossy texture is desired, such as dessert sauces and in Asian-inspired dishes Translucent, thickens further as it cools; forms a “sliceable” gel Sensitive to extended heat exposure, so products become thin if held at heat for long periods of time
Waxy maize, waxy rice	Dissolved in cold water 20-40 g starch thickens 1 L liquid	Added to hot liquid while whisking until it dissolves and the liquid thickens	Used in desserts and dessert sauces Clear, does not thicken further as it cools Does not gel at cool temperatures, good for cold sauces Quite stable at extreme temperatures (heat and freezing)
Modified starches	Dissolved in cold water 20-40 g starch thickens 1 L liquid	Added to hot liquid while whisking until it dissolves and the liquid thickens	Modified starches are often used in commercially processed foods and convenience products Modified to improve specific characteristics (e.g., stability or texture under extreme conditions; heat and freezing) Translucent, thickens further as it cools
Pre-gelatinized starches	Powder, dissolved in cold liquid 20-40 g starch thickens 1 L liquid	Added to liquid at any temperature	Used when thickening liquids that might lose color or flavor during cooking Become viscous without the need for additional cooking Translucent, fairly clear, shiny, does NOT gel when cold
Arrowroot	Powder, dissolved in cold liquid 20-40 g starch thickens 1 L liquid	Added to hot liquid while whisking until it dissolves and the liquid thickens	Derived from cassava root Used in Asian cuisines Very clear; possesses a gooey texture Translucent, shiny, very light gel when cold
Gelatin	15-30 g gelatin sets 1 L liquid	Powder or sheets (leaves) dissolved in cold water Added to cold or simmering liquid Activates with heat, sets when cold	Derived from collagens in bones and meats of animals Used in aspic, glazes, cold sauces, and desserts Clear, firm texture Dissolves when reheated, thickens when cold

Agar agar	15-30g agar agar sets 1 L liquid	Powder dissolved in cold water Added to cold or simmering liquid Activates with heat, sets when cold	Extracted from seaweed Used in Asian desserts and molecular gastronomy cooking Used in place of gelatin in vegetarian dishes Clear firm texture Does not thin when reheated, thickens more when cold
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Gelling agents

Gelatin is a water-soluble protein extracted from animal tissue and used as a gelling agent, a thickener, an emulsifier, a whipping agent, a stabilizer, and a substance that imparts a smooth mouth feel to foods. It is thermo-reversible, meaning the setting properties or action can be reversed by heating. Gelatin is available in two forms: powder and sheet (leaf).

Gelatin is often used to stabilize whipped cream and mousses; confectionery, such as gummy bears and marshmallows; desserts including pannacotta; commercial products like Jell-O; "lite" or low-fat versions of foods including some margarines; and dairy products such as yogurt and ice cream. Gelatin is also used in hard and soft gel capsules for the pharmaceutical industry.

Agar agar is an extract from red algae and is often used to stabilize emulsions or foams and to thicken or gel liquids. It is thermo-reversible and heat resistant.

It is typically hydrated in boiling liquids and is stable across a wide range of acidity levels. It begins to gel once it cools to around 40°C (100°F) and will not melt until it reaches 85°C (185°F).

Pectin

Pectin is taken from citrus and other tree fruits (apples, pears, etc.). Pectin is found in many different foods such as jam, milk-based beverages, jellies, sweets, and fruit juices. Pectin is also used in molecular gastronomy mainly as a gelling agent, thickener, and stabilizer.

There are a variety of types of pectin that react differently according to the ingredients used. Low-methoxyl pectin (which is activated with the use of calcium for gelling) and high-methoxyl pectin that requires sugar for thickening are the two most common types used in cooking. High-methoxyl pectin is what is traditionally used to make jams and jellies. Low-methoxyl pectin is often used in modern cuisine due to the thermo-irreversible gel that it forms and its good reaction to calcium. Its natural capability to emulsify and gel creates stable preparations.

Increasingly, cooks, bakers, and pastry chefs are turning to many different gels, chemicals, and other substances used in commercial food processing as new ingredients to modify liquids or other foods. These will be outlined in detail in the section on molecular gastronomy.

Attribution

Crystallization

Many factors can influence **crystallization** in food. Controlling the crystallization process can affect whether a particular product is spreadable, or whether it will feel gritty or smooth in the mouth. In some cases, crystals are something you try to develop; in others, they are something you try to avoid.

It is important to know the characteristics and quality of the crystals in different food. Butter, margarine, ice cream, sugar, and chocolate all contain different types of crystals, although they all contain fat crystals. For example, ice cream has fat crystals, ice crystals, and sometimes lactose crystals.

The fact that sugar solidifies into crystals is extremely important in candy making. There are basically two categories of candies: crystalline (candies that contain crystals in their finished form, such as fudge and fondant); and non-crystalline (candies that do not contain crystals, such as lollipops, taffy, and caramels). Recipe ingredients and procedures for non-crystalline candies are specifically designed to prevent the formation of sugar crystals because they give the resulting candy a grainy texture. One way to prevent the crystallization of sucrose in candy is to make sure that there are other types of sugar—usually fructose and glucose—to get in the way and slow down or inhibit the process. Acids can also be added to “invert” the sugar, and to prevent or slow down crystallization. Fats added to certain confectionary items will have a similar effect.

When boiling sugar for any application, the formation of crystals is generally not desired. These are some of the things that can promote crystal growth:

- Pot and utensils that are not clean
- Sugar with impurities in it (A scoop used in the flour bin, and then used for sugar, may have enough particles on it to promote crystallization.)
- Water with a high mineral content (“hard water”)
- Too much stirring (agitation) during the boiling phase

Crystallization may be prevented by adding an interferent, such as acid (lemon, vinegar, tartaric, etc.) or glucose or corn syrup, during the boiling procedure.

As mentioned above, ice cream can have ice and fat crystals that co-exist along with other structural elements (emulsion, air cells, and hydrocolloid stabilizers such as locust bean gum) that make up the “body” of the ice cream. Some of these components crystallize either partially or completely. The bottom line is that the nature of the crystalline phase in the food will determine the quality, appearance, texture, feel in the mouth, and stability of the product. The texture of ice cream is derived, in part, from the large number of small ice crystals. These small ice crystals provide a smooth texture with excellent melt-down and cooling properties. When these ice crystals grow larger during storage (recrystallization), the product becomes coarse and less enjoyable. Similar concerns apply to sugar crystals in fondant and frostings, and to fat crystals in chocolate, butter, and margarine.

Control of crystallization in fats is important in many food products, including chocolate, margarine, butter, and shortening. In these products, the aim is to produce the appropriate number, size, and distribution of crystals in the correct shape because the crystalline phase plays such a large role in appearance, texture, spreadability, and flavor release. Thus, understanding the processes that control crystallization is critical to controlling quality in these products.

To control crystallization in foods, certain factors must be controlled:

- Number and size of crystals
- Crystal distribution
- Proper polymorph (crystal shape)

Crystallization is important in working with chocolate. The **tempering** process, sometimes called pre-crystallization, is an important step that is used for decorative and moulding purposes, and is a major contributor to the mouth feel and enjoyment of chocolate. Tempering is a process that encourages the cocoa butter in the chocolate to harden into a specific crystalline pattern, which maintains the sheen and texture for a long time.

When chocolate isn't tempered properly it can have a number of problems. For example, it may not ever set up hard at room temperature; it may become hard, but look dull and blotchy; the internal texture may be spongy rather than crisp; and it can be susceptible to fat **bloom**, meaning the fats will migrate to the surface and make whitish streaks and blotches.

Attribution

Non-traditional thickeners

Non-traditional thickeners

In addition to traditional starches, there are new ways to thicken sauces and to change the texture of liquids. Some of these thickening agents work without heating and are simply blended with the cold liquid, such as modified starch or xanthan gum. These allow the creation of sauces and other liquids with a fresh, uncooked taste.

Foams, froths, and bubbles

Liquids can be stabilized with gelatin, **lecithin**, and other ingredients, and then used to create foams by whipping or using a special dispenser charged with nitrogen gas. A well-made foam adds an additional flavor dimension to the dish without adding bulk, and an interesting texture as the foam dissolves in the mouth (Figure 1).



Figure 1. "Dinner in the Dark 21-Dessert" by Esther Little is licensed under CC BY SA 2.0

Espuma

Espuma is the Spanish term for froth or foam, and it is created with the use of a siphon (ISO) bottle. This is a specific term, since culinary foams may be attained through other means.

Espuma from a siphon creates foam without the use of an emulsifying agent such as egg. As a result, it offers an unadulterated flavor of the ingredients used. It also introduces much more air into a preparation compared to other culinary aerating processes.

Espuma is created mainly with liquid that has air incorporated in it to create froth. But solid ingredients can be used too; these can be liquefied by cooking, puréeing, and extracting natural juices. It should be noted, though, that the best flavors to work with are those that are naturally diluted. Otherwise, the *espuma* tends to lose its flavor as air is introduced into it.

Stabilizers may be used alongside the liquids to help retain their shape longer; however, this is not always necessary. Prepared liquids can also be stored in a siphon bottle and kept for use. The pressure from the bottle will push out the aerated liquid, producing the *espuma*.

Foam

Foam is created by trapping air within a solid or liquid substance. Although culinary foams are most recently associated with molecular gastronomy, they are part of many culinary preparations that date back to even earlier times. Mousse, soufflé, whipped cream, and froth in cappuccino are just some examples of common foams. Common examples of “set” foams are bread, pancakes, and muffins.

Foam does not rely on pressure to encase air bubbles into a substance. Like espuma, foam may also be created with the help of a surfactant and gelling or thickening agents to help it hold shape. The production of a culinary foam starts with a liquid or a solid that has been puréed. The thickening or gelling agent is then diluted into this to form a solution. Once dissolved, the solution is whipped to introduce air into it.

The process of whipping is done until the foam has reached the desired stiffness. Note that certain ingredients may break down if they are whipped for too long, especially without the presence of a stabilizing agent.

Gels

Turning a liquid, such as a vegetable juice or raspberry purée, into a solid not only gives it a different texture but also allows the food to be cut into many shapes, enabling different visual presentations (Figure 2). Regular gelatin can be used as well as other gelling agents, such as agar agar, which is derived from red algae.



Figure 2. “Papayagelee” by hedonistin is licensed under CC BY NC 2.0

Brittle gels

Gelling agents are often associated with jelly-like textures, which may range from soft to firm. However, certain gels produced by specific agents may not fit this description.

Rather than forming an elastic or pliable substance, brittle gels may also be formed. These are gels that are firm in nature yet fragile at the same time. This characteristic is caused by the formation of a gel network that is weak and susceptible to breaking. This property allows brittle gels to crumble in the mouth and create a melt-in-the-mouth feeling. As a result, new sensations and textures are experienced while dining. At the same time, tastes within a dish are also enhanced due to the flavour release caused by the gel breakdown. Brittle gels are made by diluting the gelling agent into a liquid substance such as water, milk, or a stock. This mixture is left to set to attain a gelled end product. It should be noted that the concentration of gelling agents used, as well as the amount of liquid, both affect gelation.

Agar agar is a common agent used to create brittle gels. However, when combined with sugar it tends to create a more elastic substance. Low-acyl gellan gum, locust bean gum, and carrageenan also create brittle gels.

Fluid gels

A fluid gel is a cross between a sauce, gel, and purée. It is a controlled liquid that has properties of all three preparations. A fluid gel displays viscosity and fluidity at the same time, being thick yet still spreadable.

Fluid gels behave as solids when undisturbed, and flow when exposed to sufficient agitation. They are used in many culinary dishes where fluids need to be controlled, and they provide a rich, creamy texture.

A fluid gel is created using a base liquid that can come from many different sources. The base liquid is commonly extracted from fruits and vegetables, taken from stocks, or even puréed from certain ingredients. The longer the substance is exposed to stress, and the more intense the outside stress, the more fluidity is gained. More fluidity causes a finer consistency in the gel.

Fluid gels can be served either hot or cold, as many of the gelling agents used for such preparations are stable at high temperatures.

Drying and powdering

Drying a food intensifies its flavour and, of course, changes its texture. Eating a piece of apple that has been cooked and then dehydrated until crisp is very different from eating a fresh fruit slice. If the dehydrated food is powdered, it becomes yet another flavour and texture experience.

When maltodextrin (or tapioca maltodextrin) is mixed with fat, it changes to a powder. Because maltodextrin dissolves in water, peanut butter (or olive oil) that has been changed to a powder changes back to an oil in the mouth.

Freezing

In molecular gastronomy, liquid nitrogen is often used to freeze products or to create a frozen item without the use of a freezer.

Liquid nitrogen is the element nitrogen in a liquefied state. It is a clear, colourless liquid with a temperature of -196°C (-321°F). It is classified as a cryogenic fluid, which causes rapid freezing when it comes into contact with living tissues.

The extremely cold temperatures provided by this liquefied gas are most often used in modern cuisine to produce frozen foams and ice cream. After freezing food, nitrogen boils away, creating a thick nitrogen fog that may also add to the aesthetic features of a dish.

Given the extreme temperature of liquid nitrogen, it must be handled with care. Mishandling may cause serious burns to the skin. Nitrogen must be stored in special flasks and handled only by trained people. Aprons, gloves, and other specially designed safety gear should be used when handling liquid nitrogen.

Used mainly in the form of a coolant for molecular gastronomy, liquid nitrogen is not ingested. It is poured directly onto the food that needs to be cooled, causing it to freeze. Any remaining nitrogen evaporates, although sufficient time must be provided to allow the liquefied gas to be eliminated and for the dish to warm up to the point that it will not cause damage during consumption.

Spherification

Spherification is a modern cuisine technique that involves creating semi-solid spheres with thin membranes out of liquids. Spheres can be made in various sizes and of various firmnesses, such as the “caviar” shown in Figure 3. The result is a burst-in-the-mouth effect, achieved with the liquid. Both flavour and texture are enhanced with this culinary technique.

There are two versions of the spherification process: direct and reverse.

In direct spherification, a flavoured liquid (containing either sodium alginate, gellan gum, or carrageenan) is dripped into a water bath that is mixed with calcium (either calcium chloride or calcium lactate). The outer layer is induced by calcium to form a thin gel layer, leaving a liquid centre. In this version, the spheres are easily breakable and should be consumed immediately.

Calcium chloride and sodium alginate are the two basic components used for this technique. Calcium chloride is a type of salt used in cheese making, and sodium alginate is taken from seaweed. The sodium alginate is used to gel the chosen liquid by dissolving it directly into the fluid. This causes the liquid to become sticky, and proper dissolving must be done by mixing. The liquid is then left to set to eliminate any bubbles.

Once ready, a bath is prepared with calcium chloride and water. The liquid is then dripped into the bath using a spoon or syringe depending on the desired sphere size. The gel forms a membrane encasing the liquid when it comes into contact with the calcium chloride. Once set, the spheres are then removed and rinsed with water to remove any excess calcium chloride.

In reverse spherification, a calcium-containing liquid (or ingredients mixed with a soluble calcium salt) is dripped into a setting bath containing sodium alginate. Surface tension causes the drop to become spherical. A skin of calcium alginate immediately forms around the top. Unlike in the direct version, the gelling stops and does not continue into the liquid orb. This results in thicker shells so the products do not have to be consumed immediately.



Figure 3. “White chocolate spaghetti with raspberry sauce and chocolate martini caviar” by ayngelina is licensed under CC BY NC-ND 2.0

Videos on spherification:

Direct: <https://www.youtube.com/watch?v=BeRMBv95gLk>

Reverse: <https://www.youtube.com/watch?v=JPNo79U77yl>

Specialty ingredients used in molecular gastronomy

There are a number of different ingredients used in molecular gastronomy as gelling, thickening, or emulsifying agents. Many of these are available in specialty food stores or can be ordered online.

Algin

Another name for sodium alginate, algin is a natural gelling agent taken from the cell walls of certain brown seaweed species.

Calcium chloride

Calcium chloride, also known as CaCl_2 , is a compound of chlorine and calcium that is a by-product of sodium bicarbonate (baking soda) manufacturing. At room temperature it is a solid salt, which is easily dissolved in water.

This is very salty and is often used for preservation, pickling, cheese production, and adding taste without increasing the amount of sodium. It is also used in molecular gastronomy in the spherification technique (see above) for the production of ravioli, spheres, pearls, and caviar (Figure 3).

Calcium lactate

Calcium lactate is a calcium salt resulting from the fermentation of lactic acid and calcium. It is a white crystalline powder when solid and is highly soluble in cold liquids. It is commonly used as a calcium fortifier in various food products including beverages and supplements.

Calcium lactate is also used to regulate acidity levels in cheese and baking powder, as a food thickener, and as a preservative for fresh fruits. In molecular gastronomy, it is most commonly used for basic spherification and reverse spherification due to the lack of bitterness in the finished products.

Like calcium chloride, calcium lactate is used alongside sodium alginate. In regular spherification, it is used in the bath. It is also used as a thickener in reverse spherification.

Carob bean gum

Carob bean gum is another name for locust bean gum. It is often used to stabilize, texturize, thicken, and gel liquids in the area of modern cuisine, although it has been a popular thickener and stabilizer for many years.

Carrageenan

Carrageenan refers to any linear sulfated polysaccharide taken from the extracts of red algae. This seaweed derivative is classified mainly as iota, kappa, and lambda. It is a common ingredient in many foods.

There are a number of purposes that it serves, including binding, thickening, stabilizing, gelling, and emulsifying. Carrageenan can be found in ice cream, salad dressings, cheese, puddings, and many more foods. It is often used with dairy products because of its good interaction with milk proteins. Carrageenan also works well with other common kitchen ingredients and offers a smooth texture and taste that blends well and does not affect flavour.

More often than not, carrageenan is found in powder form, which is hydrated in liquid before being used. For best results, carrageenan powder should be sprinkled in cold liquid and blended well to dissolve, although it may also be melted directly in hot liquids.

Citric acid

Classified as a weak organic acid, citric acid is a naturally occurring preservative that can be found in citrus fruits. Produced as a result of the fermentation of sugar, it has a tart to bitter taste and is usually in powder form when sold commercially. It is used mainly as a preservative and acidulant, and it is a common food additive in a wide range of foods such as candies and soda. Other than extending shelf life by adjusting the acidity or pH of food, it can also help enhance flavours. It works especially well with other fruits, providing a fresh taste.

In modern cooking, citric acid is often used as an emulsifier to keep fats and liquids from separating. It is also a common component in spherification, where it may be used as an acid buffer.

Gellan gum

Gellan gum is a water-soluble, high-molecular-weight polysaccharide gum that is produced through the fermentation of carbohydrates in algae by the bacterium *Pseudomonas elodea*. This fermented carbohydrate is purified with isopropyl alcohol, then dried and milled to produce a powder.

Gellan gum is used as a stabilizer, emulsifier, thickener, and gelling agent in cooking. Aspics and terrines are only some of the dishes that use gellan. It comes in both high-acyl and low-acyl forms. High-acyl gellan gum produces a flexible elastic gel, while low-acyl gellan gum will give way to a more brittle gel.

Like many other hydrocolloids, gellan gum is used with liquids. The powder is normally dispersed in the chosen liquid to dissolve it. Once dissolved, the solution is then heated to facilitate liquid absorption and gelling by the hydrocolloid. A temperature between 85°C and 95°C (185°F and 203°F) will start the

dissolution process. Gelling will begin upon cooling around 10°C and 80°C (50°F and 176°F).

Gellan gum creates a thermo-irreversible gel and can withstand high heat without reversing in form. This makes it ideal for the creation of warm gels.

Guar gum

Guar gum, or guaran, is a carbohydrate. This galactomannan is taken from the seeds of the guar plant by dehusking, milling, and screening. The end product is a pale, off-white, loose powder. It is most commonly used as a thickening agent and stabilizer for sauces and dressings in the food industry. Baked goods such as bread may also use guar gum to increase the amount of soluble fibre. At the same time, it also aids with moisture retention in bread and other baked items.

Being a derivative of a legume, guar gum is considered to be vegan and a good alternative to starches. In modern cuisine, guar gum is used for the creation of foams from acidic liquids, for fluid gels, and for stabilizing foams.

Guar gum must first be dissolved in cold liquid. The higher the percentage of guar gum used, the more viscous the liquid will become. Dosage may also vary according to the ingredients used as well as desired results and temperature.

Iota carrageenan

Iota carrageenan is a hydrocolloid taken from red seaweed (*Eucheuma denticulatum*). It is one of three varieties of carrageenan and is used mainly as a thickening or gelling agent.

Gels produced from iota carrageenan are soft and flexible, especially when used with calcium salts. It produces a clear gel that exhibits little syneresis. Iota is a fast-setting gel that is thermo-reversible and remains stable through freezing and thawing. In modern cuisine it is used to create hot foams as well as custards and jellies with a creamy texture.

Like most other hydrocolloids, iota carrageenan must first be dispersed and hydrated in liquid before use. Unlike lambda carrageenan, it is best dispersed in cold liquid. Once hydrated, the solution must be heated to about 70°C (158°F) with shear to facilitate dissolution. Gelling will happen between 40°C and 70°C (104°F and 158°F) depending on the number of calcium ions present.

Kappa carrageenan

Kappa carrageenan is another type of red seaweed extract taken specifically from *Kappaphycus alvarezii*. Like other types of carrageenan, it is used as a gelling, thickening, and stabilizing agent. When mixed with water, kappa carrageenan creates a strong and firm solid gel that may be brittle in texture.

This particular variety of carrageenan blends well with milk and other dairy products. Since it is taken from seaweed, it is considered to be vegan and is an alternative to traditional gelling agents such as gelatin.

Kappa carrageenan is used in various cooking preparations including hot and cold gels, jelly toppings, cakes, breads, and pastries. When used in molecular gastronomy preparations and other dishes, kappa carrageenan should be dissolved in cold liquid.

Once dispersed, the solution must be heated between 40°C and 70°C (104°F and 158°F). Gelling will begin between 30°C and 60°C (86°F and 140°F). Kappa carrageenan is a thermo-reversible gel and will stay stable up to 70°C (158°F). Temperatures beyond this will cause the gel to melt and become liquid once again.

Locust bean gum

Locust bean gum, also known as LBG and carob bean gum, is a vegetable gum derived from Mediterranean-region carob tree seeds. This hydrocolloid is used to stabilize, texturize, thicken, and gel liquids in modern cuisine, although it has been a popular thickener and stabilizer for many years.

It has a neutral taste that does not affect the flavour of food that it is combined with. It also provides a creamy mouth feel and has reduced syneresis when used alongside pectin or carrageenan for dairy and fruit applications. The neutral behaviour of this hydrocolloid makes it ideal for use with a wide range of ingredients.

To use locust bean gum, it must be dissolved in liquid. It is soluble with both hot and cold liquids.

Maltodextrin

Maltodextrin is a sweet polysaccharide that is produced from starch, corn, wheat, tapioca, or potato through partial hydrolysis and spray drying. This modified food starch is a white powder that has the capacity to absorb and hold water as well as oil. It is an ideal additive since it has fewer calories than sugar and is easily absorbed and digested by the body in the form of glucose.

Coming from a natural source, it ranges from nearly flavourless to fairly sweet without any odour. Maltodextrin is a common ingredient in processed foods such as soda and candies. In molecular gastronomy, it can be used both as a thickener and a stabilizer for sauces and dressings, for encapsulation, and as a sweetener. In many cases, it is also used as an aroma carrier due to its capacity to absorb oil. It is also often used to make powders or pastes out of fat.

Sodium alginate

Sodium alginate, which is also called algin, is a natural gelling agent taken from the cell walls of certain brown seaweed species. This salt is obtained by drying the seaweed, followed by cleaning, boiling, gelling, and pulverizing it. A light yellow powder is produced from the process. When dissolved in liquids, sodium alginate acts as a thickener, creating a viscous fluid. Conversely, when it is used with calcium it forms a gel through a cold process.

In molecular gastronomy, sodium alginate is most commonly used as a texturizing agent. Foams and sauces may be created with it. It is also used in spherification for the creation of pearls, raviolis, mock caviar, marbles, and spheres. Sodium alginate can be used directly by dissolving it into the liquid that needs to be gelled, as in the case of basic spherification. It may also be used inversely by adding it directly to a bath, as in the case of reverse spherification.

This versatile product is soluble in both hot and cold liquids, and gels made with it will set at any temperature.

Soy lecithin

Soy lecithin, also called just lecithin, is a natural emulsifier that comes from fatty substances found in plant tissues. It is derived from soybeans either mechanically or chemically, and is a by-product of soybean oil creation. The end product is a light brown powder that has low water solubility.

As an emulsifier, it works to blend immiscible ingredients together, such as oil and water, giving way to stable preparations. It can be whisked directly into the liquid of choice.

Soy lecithin is also used in creating foams, airs, mousses, and other aerated dishes that are long lasting and full of flavour. It is used in pastries, confections, and chocolate to enhance dough and increase moisture tolerance.

As with most ingredients, dosage and concentration for soy lecithin will depend on the ingredients used, specific properties desired in the resulting preparation, as well as other conditions.

Tapioca maltodextrin

Tapioca maltodextrin is a form of maltodextrin made from tapioca starch. It is a common ingredient in molecular gastronomy because it can be used both as a thickener and stabilizer for sauces and dressings, for encapsulation, and as a sweetener. In many cases it is also used as an aroma carrier due to its capacity to absorb oil. It is often used to make powders or pastes out of fat.

Xanthan gum

Xanthan gum is a food additive used as a thickening agent. It is produced through the fermentation of glucose. As a gluten-free additive it can be used as a substitute in cooking and baking.

As a thickener, when used in low dosages, xanthan gum produces a weak gel with high viscosity that is shear reversible with a high pourability. It also displays excellent stabilizing abilities that allow for particle suspension.

Moreover, xanthan gum mixes well with other flavours without masking them and provides an improved mouth feel to preparations. The presence of bubbles within the thickened liquids often makes way for light and creamy textures. It is used in the production of emulsions, suspensions, raviolis, and foams.

Being a hydrocolloid, xanthan gum must be hydrated before use. High versatility allows it to be dissolved over a wide range of temperatures, acid, and alcohol levels. Once set, xanthan gum may lose some of its effectiveness when exposed to heat.

Attribution

Sauces

Sauces enhance desserts by both their flavor and their appearance, just as savory sauces enhance meats, fish, and vegetables. Crème anglaise, chocolate sauce, caramel sauce, and the many fruit sauces and coulis are the most versatile. One or another of these sauces will complement nearly every dessert.

Examples of dessert sauces

Caramel sauce: A proper caramel flavor is a delicate balance between sweetness and bitterness. As sugar cooks and begins to change color, a flavor change will occur. The darker the sugar, the more bitter it will become. Depending on the application for the finished caramel, it can be made mild or strong. At this point, a liquid is added. This liquid will serve several roles: it will stop the cooking process, it can add richness and flavor, and it will soften the sauce. The fluidity of the finished sauce will depend on the amount of liquid added to it, and the temperature it is served at. Dairy products, such as cream, milk, or butter, will add richness; use water for a clear sauce; use fruit purées to add different flavor elements.

Chocolate sauce: Sometimes called fudge sauce, chocolate sauce is generally made from cream (or milk), butter, and chocolate, and can be served hot or cold. The proportion of each of the ingredients will affect the thickness of the final product.

Compote: French for “mixture,” a **compote** is cooked fruit served in its own cooking liquid, usually a sugar syrup. Compotes can be made with fresh, frozen, or dried fruits, and served hot or cold.

Coulis: French for “strained liquid,” a **coulis** is most often an uncooked, strained purée. Flavors remain pure, and the colors bright. One of the drawbacks of using a coulis is that it may separate quickly when used as a plating sauce. It’s best to use à la minute.

Crème anglaise: French for “English custard,” **crème anglaise** is a rich, pourable custard sauce that can be served hot or cold over cake, fruits, or other desserts. Made with eggs, sugar, and milk or cream, it is stirred over heat until it thickens into a light sauce. However, it’s a delicate operation: too much heat turns it into scrambled eggs! It should not get above 85°C (185°F) during the cooking process. Vanilla is the classic flavoring, but coffee, spices, chocolate, or liqueurs can be added. With additional yolks and heavy cream, it becomes the “custard” used for French ice cream. With additional yolks, gelatin, whipped cream, and flavoring, it becomes Bavarian cream.

Curd: A **curd** is creamy and fruit based, with citrus and berry flavors being the most popular. Made from fruit juices, eggs, butter, and sugar cooked in a process similar to crème anglaise, curds can be thick, pourable sauces or spreads.

Fruit butter: **Fruit butter** is a spread made from whole fruits, cooked, reduced, and puréed (if you don’t want any chunks in it) until very thick. It does not contain any butter; the term refers to the consistency.

Fruit sauce: A fruit sauce is a fruit purée, cooked and thickened with a starch. It is normally served cold.

Hard sauce: This traditional sauce for Christmas pudding, or any steamed pudding, is made by combining butter, sugar, and flavorings, often liqueurs. It is normally piped into shapes and chilled, then placed on the warm dessert just before serving.

Sabayon: **Sabayon** is a mixture of egg yolks, flavoring, and sugar beaten over simmering water until thick, then beaten until cool. It is traditionally flavored with sweet white wine or liquor, then served over fresh fruit and grilled (when it is called a gratin). The Italian version of this is called a **zabaglione** and is

flavored with Madeira wine.

Whipped cream: This very popular dessert topping can be served plain, sweetened, or flavored. **Crème chantilly**, a classic version of this, is a combination of whipped cream, sugar, and vanilla.

Applying dessert sauces

Except in the case of some home-style or frozen desserts, sauces are usually not ladled over the dessert because doing so would mar the appearance. Instead, the sauce is applied in a decorative fashion to the plate rather than the dessert. Many different styles of plate saucing are available.

Pouring a pool of sauce onto the plate is known as *flooding*. Although plate flooding often looks old-fashioned today, it can still be a useful technique for many desserts. Flooded plates can be made more attractive by applying a contrasting sauce and then blending or feathering the two sauces decoratively with a pick or the end of a knife. For this technique to work, the two sauces should be at about the same fluidity or consistency.

Rather than flooding the entire plate, it may be more appropriate for some desserts to apply a smaller pool of sauce to the plate, as this avoids overwhelming the dessert with too much sauce.

A variation of the flooding technique is outlining, where a design is piped onto the plate with chocolate and allowed to set. The spaces can then be flooded with colorful sauces.

A squeeze bottle is useful for making dots, lines, curves, and streaks of sauce in many patterns. Or just a spoon is needed to drizzle random patterns of sauce onto a plate. Another technique for saucing is applying a small amount of sauce and streaking it with a brush, an offset spatula, or the back of a spoon.

Sauces are a great way to highlight flavors. Choose ones that will create balance on the plate, not just for color, but with all the components. A tart berry sauce will complement a rich cheesecake or chocolate dessert because sourness (acid) will cut through fat, making it taste lighter than it is. A sweet sauce served with a sweet dessert will have the overall effect of hiding flavors in both. Hold back on sweetness in order to intensify other flavors.

Many modern presentations may have a minimal amount of sauce. Sometimes this is done just for aesthetic reasons and not for how it will complement the dessert. Think of the dish and the balance of the components. This is the most important factor: flavor first, presentation second.

Attribution

Low-temperature and sous-vidé

Sous-vidé cooking is about immersing a food item in a precisely controlled water bath, where the temperature of the water is the same as the target temperature of the food being cooked. Food is placed in a food-grade plastic bag and vacuum-sealed before going into the water bath. Temperatures will vary depending on desired end result. This allows the water in the bath to transfer heat into the food while preventing the water from coming into direct contact with it. This means the water does not chemically interact with the food: the flavors of the food remain stronger, because the water is unable to dissolve or carry away any compounds in the food (Figure 1).



Figure 1. "Img_0081" by Derek is licensed under CC BY-SA-ND 2.0

Sous-vidé fruits and vegetables

Cooking vegetable and fruits sous-vidé is a great way to tenderize them without losing as many of the vitamins and minerals that are normally lost through blanching or steaming. Fruits can also be infused with liquid when cooked at lower temperatures by adding liquid to the bag. Sous-vidé helps preserve the nutrients present in fruits and vegetables by not cooking them above the temperatures that cause the cell walls to fully break down. This allows them to tenderize without losing all their structure. The bag also helps to catch any nutrients that do come out of the vegetable.

While time and temperature do not factor into safety for fruits and vegetables, they do have a unique effect on their structure. There are two components in fruits and vegetables that make them crisp: pectin and starch. Pectin, which is a gelling agent commonly used in jams and jellies for structure, breaks down at 83°C (183°F) at a slower rate than the starch cells do. In many cases this allows for more tender fruits and vegetables that have a unique texture to them.

Custards

The term *custard* spans so many possible ingredients and techniques that it is most useful to think of a custard as simply a particular texture and mouth feel. Custards have been made for centuries by lightly cooking a blend of eggs, milk, and heavy cream, but modernist chefs have invented myriad ways to make custards.

Using the sous-vidé method to prepare crème anglaise, curds, ice cream bases, custard bases, sabayons, and dulce de leche is possible. The technique offers greater consistency and more control over the texture, which can range from airy, typical of a sabayon, to dense, as in a posset. For custards, eggs will be properly cooked at 82°C (180°F), so if the water bath is set to this temperature, no overcooking can happen. The

one constant among custards is the use of plenty of fat, which not only provides that distinctive mouth feel but also makes custard an excellent carrier of fat-soluble flavors and aromas. Lighter varieties of custard, prepared sous-vide style and cooled, can be aerated in a whipping siphon into smooth, creamy foams.

Other applications for vacuum-seal processes

Fruit compression

Vacuum-compressing fruits and vegetables is a popular modern technique that can give many plant foods an attractive, translucent appearance (as shown in the watermelon in Figure 2) and a pleasant, surprising texture. This technique exploits the ability of a vacuum chamber to reduce surrounding pressure, which causes air and moisture within the plant tissue to rapidly expand and rupture the structures within the food. When the surrounding pressure is restored to a normal level, the labyrinth of air-filled spaces collapses. As a result, light tends to pass through the food rather than being scattered and diffused, which is why vacuum-compressed plant foods appear translucent. Causing the porous structure of a plant food to collapse also imparts a somewhat dense, toothsome texture that can give a familiar ingredient, such as watermelon, an entirely new appeal.



Figure 5. “WD-50 (7th Course)” by Peter Dillon is licensed under CC BY 2.0

Infusions

When adding liquids, the vacuum-seal process creates a rapid infusion—especially with more porous foods (such as adding spices to cream or herbs to melon). This can add flavor and texture in a shorter time than traditional infusions.

Attribution

Understanding Ingredients: Flour

Introduction -- Understanding Ingredients

Ingredients play an important role in baking. Not only do they provide the structure and flavour of all of the products produced in the bakery or pastry shop, their composition and how they react and behave in relation to each other are critical factors in understanding the science of baking. This is perhaps most evident when it comes to adapting formulas and recipes to accommodate additional or replacement ingredients while still seeking a similar outcome to the original recipe.

In this book, we look at each of the main categories of baking ingredients, listed below, and then explore their composition and role in the baking process. In addition to these categories, we will discuss the role that salt and water play in the baking process.

The main categories of baking ingredients are:

- Grains and flours
- Sweeteners
- Fats oils
- Leavening agents
- Eggs
- Dairy products
- Chocolate and other cocoa products
- Nuts and seeds
- Thickening agents
- Spices and other flavourings
- Fruit

Note: For most measurements used in the open textbook series, both S.I. (metric) and U.S./imperial values are given. The exception is nutritional information, which is always portrayed using metric values in both Canada and the United States.

Attribution

The History of Wheat Flour

Archaeologists who did excavations in the region of the lake dwellers of Switzerland found grains of wheat, millet, and rye 10,000 years old. The Romans perfected the rotary mill for turning wheat into flour. By the time of Christ, Rome had more than 300 bakeries, and Roman legions introduced wheat throughout their empire. Improved milling processes were needed because even when wheat was milled twice and bolted (sifted) through silk gauze, the result was still a yellowish flour of uneven texture and flecked with germ and bran.

In the second half of the 19th century, there were great changes in the flour milling process. An American inventor, Edmund LaCroix, improved the process with a purifier to separate the **middlings** (bran, germ, and other coarse particles) from the particles that form smooth-textured white flour. In recent years, the demand for whole grain milling has increased because whole grain food products have proved to be more nutritious than products made from white flour. (More information on whole grain and artisan milling is provided later in this section.)

In Canada, large-scale wheat growing didn't occur until after the Prairies were settled in the 1800s. Hard wheat, such as Red Fife, Marquis, and Selkirk, earned Canada a position as the granary for Britain and many other European countries. Today, most of the wheat grown in Western Canada is the hard Red Spring variety. Soft wheats, such as soft red and soft white, are primarily grown in Quebec and Ontario. Many of the original wheat growers have passed on their farms to the next generations, while others branched out to organic farming and milling. One of these farms, [Nunweiler's](#), has a heritage that goes back to the early 1900s when the original wheat in Canada, Red Fife and Marquis, was grown on this farm.

Today, the major wheat growing areas of North America are in the central part of the continent, in the Great Plains of the United States and the Canadian Prairies. From Nebraska south, winter wheat can be grown, while to the north through Saskatchewan spring wheat dominates. Many American states and some Canadian provinces grow both kinds. In fact, there are very few states that don't grow some wheat. Kansas, the site of the [American Institute of Baking](#), could be said to be at the heart of the U.S. wheat growing area, while Saskatchewan is the Canadian counterpart.

Attribution

Milling of Wheat

Milling of wheat is the process that turns whole grains into flours. The overall aims of the miller are to produce:

- A consistent product
- A range of flours suitable for a variety of functions
- Flours with predictable performance

The very first mill operation is analyzing the grain, which determines criteria such as the **gluten** content and amylase activity. It is at this point that decisions about blending are made.

Following analysis, milling may be divided into three stages:

- Cleaning and conditioning – ridding the grain of all impurities and readying it for milling
- Crushing or breaking – breaking down the grain in successive stages to release its component parts
- Reduction – progressive rollings and siftings to refine the flour and separate it into various categories, called *streams*

Cleaning

Wheat received at the mill contains weeds, seeds, chaff, and other foreign material. Strong drafts of air from the aspirator remove lighter impurities. The **disc separator** removes barley, oats, and other foreign materials. From there, the wheat goes to the scourers in which it is driven vigorously against perforated steel casings by metal beaters. In this way, much of the dirt lodged in the crease of the wheat berry is removed and carried away by a strong blast of air. Then the magnetic separator removes any iron or steel.

At this point, the wheat is moistened. Machines known as *whizzers* take off the surface moisture. The wheat is then **tempered**, or allowed to lie in bins for a short time while still damp, to toughen the bran coat, thus making possible a complete separation of the bran from the flour-producing portion of the wheat berry. After tempering, the wheat is warmed to a uniform temperature before the crushing process starts.

Crushing or Breaking

The objectives at this stage are twofold:

- Separate as much bran and germ as possible from the endosperm
- Maximize the flour from the resulting endosperm

Household grain mills create flour in one step — grain in one end, flour out the other — but the commercial mill breaks the grain down in a succession of very gradual steps, ensuring that little bran and germ are mixed with any endosperm.

Although the process is referred to as *crushing*, flour mills crack rather than crush the wheat with large steel rollers. The rollers at the beginning of the milling system are corrugated and break the wheat into coarse particles. The grain passes through screens of increasing fineness. Air currents draw off impurities from the **middlings**. *Middlings* is the name given to coarse fragments of endosperm, somewhere between the size of semolina and flour. Middlings occur after the “break” of the grain.

Bran and germ are sifted out, and the coarse particles are rolled, sifted, and purified again. This separation of germ and bran from the endosperm is an important goal of the miller. It is done to improve dough-

making characteristics and colour. As well, the germ contains oil and can affect keeping qualities of the flour.

Reduction

In the reduction stage, the coarser particles go through a series of fine rollers and sieves. After the first crushing, the wheat is separated into five or six streams. This is accomplished by means of machines called **plansifters** that contain sieves, stacked vertically, with meshes of various sizes. The finest mesh is as fine as the finished flour, and some flour is created at an early stage of reduction.

Next, each of the divisions or streams passes through cleaning machines, known as *purifiers*, a series of sieves arranged horizontally and slightly angled. An upcurrent draught of air assists in eliminating dust. The product is crushed a little more, and each of the resulting streams is again divided into numerous portions by means of sifting. The final crushings are made by perfectly smooth steel rollers that reduce the middlings into flour. The flour is then bleached and put into bulk storage. From bulk storage, the flour is **enriched** (thiamine, niacin, riboflavin, and iron are added), and either bagged for home and bakery use or made ready for bulk delivery.

Extraction Rates

The extraction rate is a figure representing the percentage of flour produced from a given quantity of grain. For example, if 82 kg of flour is produced from 100 kg of grain, the extraction rate is 82% ($82 \div 100 \times 100$). Extraction rates vary depending on the type of flour produced. A whole grain flour, which contains all of the germ, bran, and endosperm, can have an extraction rate of close to 100%, while white all-purpose flours generally have extraction rates of around 70%. Since many of the nutrients are found in the germ and bran, flours with a higher extraction rate have a higher nutritional value.

Attribution

Flour Streams and Types of Wheat Flour

Modern milling procedures produce many different **flour streams** (approximately 25) that vary in quality and chemical analysis. These are combined into four basic streams of edible flour, with four other streams going to feed.

- **Top patent flour:** This stream is composed of only the purest and most highly refined streams from the mill. It is low in ash and is approximately 50% of the flour extracted. The term *ash* indicates the mineral content (e.g., phosphorus) of the flour. When flour is burned, all that is left is the burned mineral elements that constitute ash.
- **Second patent flour:** This flour is composed of streams with an intermediate degree of refinement. It has an average ash content of approximately 0.45% and represents about 35% of the total flour.
- **First clear flour:** This stream contains the balance of the flour that possesses baking properties, and is high in ash and protein content. It is usually about 15% of the total flour.
- **Second clear flour:** This grade contains the poorest flour streams. It is very high in ash (approximately 0.75%), and has little or no baking quality. It is about 2% of the total flour.
- **Feed streams:** The balance of the streams from the mill are classed as feed. Feeds are marketed as bran, **wheat shorts**, flour middlings, and wheat germ.

Within the streams of edible flours, there are a number of different types of flour used in food preparation. Each has different characteristics, and with those come different uses, as described below.

All-Purpose Flour

General purpose or home use flours are usually a blend of hard spring wheats that are lower in protein (gluten) content than bread flours. They are top patent flours and contain sufficient protein to make good yeast breads, yet not too much for good quick breads, cakes, and cookies.

Note: A word about gluten quality as opposed to gluten quantity: The fact that a particular flour contains a high quantity of protein, say 13% to 15%, does not necessarily mean that it is of high quality. It may contain too much ash or too much damaged starch to warrant this classification. High quality is more important in many bread applications than high quantity. All-purpose flour is an example of a high-quality flour, with a protein content of about 12%.

Graham Flour

A U.S. patented flour, graham flour is a combination of whole wheat flour (slightly coarser), with added bran and other constituents of the wheat kernel.

Bread Flour

Bread flour is milled from blends of hard spring and hard winter wheats. They average about 13% protein and are slightly granular to the touch. This type of flour is sold chiefly to bakers because it makes excellent bread with bakery equipment, but has too much protein for home use. It is also called strong flour or hard flour and is second patent flour.

For example, the specification sheet on bread flour produced by a Canadian miller might include the following information:

Ingredients: Wheat flour, amylase, ascorbic acid, niacin, iron, thiamine mononitrate, riboflavin, azodicarbonamide, folic acid.

Moisture: 14.2%

Ash: 0.54%

Protein (5.7 x N) 13.00%

Along with this information there is microbiological data and an allergen declaration. (Note that the formula in parentheses beside "Protein" is simply the laboratory's way of deriving the protein figure from the nitrogen content.)

Cake Flour

Cake flour is milled from soft winter wheats. The protein content is about 7% and the granulation is so uniform and fine that the flour feels satiny. An exception is a high-protein cake flour formulated especially for fruited pound cakes (to prevent the fruit from sinking).

Clear Flour

Clear flour comes from the part of the wheat berry just under the outer covering. Comparing it to first patent flour is like comparing cream to skim milk. It is dark in colour and has a very high gluten content. It is used in rye and other breads requiring extra strength.

Gluten Flour

Gluten flour is made from wheat flour by removing a large part of the starch. It contains no more than 10% moisture and no more than 44% starch.

Pastry Flour

Pastry flour is made from either hard or soft wheat, but more often from soft. It is fairly low in protein and is finely milled, but not so fine as cake flour. It is unsuitable for yeast breads but ideal for cakes, pastries, cookies, and quick breads.

Self-Rising Flour

Self-rising flour has leavening and salt added to it in controlled amounts at the mill.

Wheat Germ Flour

Wheat germ flour consists entirely of the little germ or embryo part of the wheat separated from the rest of the kernel and flattened into flakes. This flour should be refrigerated.

Whole Wheat Flour

Whole wheat flour contains all the natural parts of the wheat kernel up to 95% of the total weight of the wheat. It contains more protein than all-purpose flour and produces heavier products because of the bran particles.

Whole Wheat Pastry Flour

Whole wheat pastry flour is milled from the entire kernel of soft wheat, is low in gluten, and is suitable for pastry, cakes, and cookies.

Hovis Flour

Most of the germ goes away with the shorts and only a small fraction of the total quantity can be recovered in a fairly pure form. At the mill, a special process developed in England to improve its keeping qualities

and flavour cooks this fraction. It is then combined with white flour to make Hovis flour, which produces a loaf that, though small for its weight, has a rich, distinctive flavour.

Triticale Flour

The world's first new grain, triticale is a hybrid of wheat and rye. It combines the best qualities of both grains. It is now grown commercially in Manitoba.

Semolina

Semolina is the granular product consisting of small fragments of the endosperm of the durum wheat kernel. (The equivalent particles from other hard wheat are called *farina*.) The commonest form of semolina available commercially is the breakfast cereal Cream of Wheat.

No-Time Flour

The primary goal of all bakers has been to reduce production time and keep costs to a minimum without losing quality, flavour, or structure. After extensive research, millers have succeeded in eliminating bulk fermentation for both sponge and straight dough methods. No-time flour is flour with additives such as ascorbic acid, bromate, and cysteine. It saves the baker time and labour, and reduces floor space requirements. The baker can use his or her own formulas with only minor adjustments.

Blending Flours

Blending of flours is done at the mill, and such is the sophistication of the analysis and testing of flours (test baking, etc.) that when problems occur it is generally the fault of the baker and not the product. Today the millers and their chemists ensure that bakers receive the high grade of flour that they need to produce marketable products for a quality-conscious consumer. Due to the vagaries of the weather and its effect on growing conditions, the quality of the grain that comes into the mill is hardly ever constant. For example, if damp weather occurs at harvest time, the grain may start to sprout and will cause what is known as *damaged starch*. Through analysis and adjustments in grain handling and blending, the miller is able to furnish a fairly constant product.

Bakers do blend flours, however. A portion of soft flour may be blended with the bread flour to reduce the toughness of a Danish pastry or sweet dough, for example. Gluten flour is commonly used in multigrain bread to boost the aeration.

Attribution

Flour Terms and Treatments

In addition to types of flour, you may come across various other terms when purchasing flour. These include some terms that refer to the processing and treatment of the flour, and others outlining some of the additives that may be added during the milling and refining process.

Bleached

Bleaching and maturing agents are added to whiten and improve the baking quality quickly, making it possible to market the freshest flour. Even fine wheat flours vary in colour from yellow to cream when freshly milled. At this stage, the flour produces doughs that are usually sticky and do not handle well. Flour improves with age under proper storage conditions up to one year, both in colour and quality.

Because storing flour is expensive, toward the close of the 19th century, millers began to treat freshly milled flour with oxidizing agents to bleach it and give it the handling characteristics of naturally aged flour. Under the category of maturing agents are included materials such as chlorine dioxide, chlorine gas plus a small amount of nitrosyl chloride, ammonium persulfate, and ascorbic acid. No change occurs in the nutritional value of the flour when these agents are present.

There are two classes of material used to bleach flour. A common one, an organic peroxide, reacts with the yellow pigment only, and has no effect on gluten quality. Chlorine dioxide, the most widely used agent in North America, neutralizes the yellow pigment and improves the gluten quality. It does, however, destroy the tocopherols (vitamin E complex).

Enriched

Iron and three of the most necessary B vitamins (thiamin, riboflavin, and niacin), which are partially removed during milling, are returned to white flour by a process known as enrichment. No change occurs in taste, colour, texture, baking quality, or caloric value of the flour.

Pre-sifted

During the milling process, flour is sifted many times through micro-fine silk. This procedure is known as pre-sifting. The mesh size used for sifting varies from flour to flour. There are more holes per square inch for cake flour than, for example, bread flour, so that a cup of cake flour has significantly more minute particles than does a cup of bread flour, is liable to be denser, and weigh slightly more. Sifted flour yields more volume in baked bread than does unsifted flour, simply because of the increased volume of air.

Attribution

Flour Additives

A number of additives may be found in commercial flours, from agents used as dough conditioners, to others that aid in the fermentation process. Why use so many additives? Many of these products are complementary – that is, they work more effectively together and the end product is as close to “ideal” as possible. Nevertheless, in some countries the number of additives allowed in flour are limited. For instance, in Germany, ascorbic acid remains the only permitted additive.

Some of the additives that are commonly added to flour include those described below.

Bromate

Until the early 1990s, bromate was added to flour because it greatly sped up the oxidation or aging of flour. Millers in Canada stopped using it after health concerns raised by the U.S. Food and Drug Administration (FDA). In the United States, bromate is allowed in some states but banned in others (e.g., California).

Azodicarbonamide (ADA)

Approved in the United States since 1962, but banned in Europe, ADA falls under the food additives permitted in Canada. ADA is a fast-acting flour treatment resulting in a cohesive, dry dough that tolerates high water absorption. It is not a bleach, but because it helps produce bread with a finer texture it gives an apparently whiter crumb. It does not destroy any vitamins in the dough.

Bakers who want to know if their flours contain ADA or other chemical additives can request this information from their flour suppliers.

L-Cysteine

An amino acid, L-cysteine speeds up reactions within the dough, thus reducing or almost eliminating bulk fermentation time. In effect, it gives the baker a “no-time” dough. It improves dough elasticity and gas retention.

Ascorbic Acid

Ascorbic acid was first used as a bread improver in 1932, after it was noticed that old lemon juice added to dough gave better results because it improved gas retention and loaf volume. Essentially vitamin C (ascorbic acid) has the advantage of being safe even if too much is added to the dough, as the heat of baking destroys the vitamin component. The addition of ascorbic acid consistent with artisan bread requirements is now routine for certain flours milled in North America.

Calcium Peroxide

Calcium peroxide (not to be confused with the peroxide used for bleaching flour) is another dough-maturing agent.

Glycerides

Glycerides are multi-purpose additives used in both cake mixes and yeast doughs. They are also known as surfactants, which is a contraction for “surface-acting agents.” In bread doughs, the main function of glycerides is as a crumb-softening agent, thus retarding bread staling. Glycerides also have some dough-

strengthening properties.

Sodium Stearoyl Lactylate

Approved for use in the United States since 1961, this additive improves gas retention, shortens proofing time, increases loaf volume, and works as an anti-staling agent.

Attribution

Whole Grain and Artisan Milling

Whole grain and artisan milling is the type of milling that was practiced before the consumer market demanded smooth white flours that are refined and have chemical additives to expedite aging of flours. Artisan milling produces flours that are less refined and better suited to traditional breads, but also contain little to no additives and have higher nutritional content. For that reason, demand for these types of flour is on the rise.

Artisan millers (also known as micro millers) process many non-stream grains, including spelt, kamut, buckwheat, and other non-gluten grains and pulses. This offers bakers opportunities to work with different grains and expand their businesses. Artisan flours are readily available directly from millers or through a distributor. Knowing the origin of the grains and the quality of the ingredients in baking is important for artisan bakers.

Whole grain flours are on the increase as consumers become more aware of their benefits. Whole grain flour, as the name suggests, is made from whole grains.

Many artisan millers purchase their grains directly from growers. This method of purchasing establishes trustworthy working relationships with the grain growers and promotes transparency in grain growing and food safety practices. Grain growers that sell their grains to artisan millers apply conventional or organic growing practices. Grain growers and millers have to go through vigorous processes to obtain the **certified organic** certification for their grains or products, which guarantees that no chemical additives have been used.

How organic grain is processed varies. Stone milling and [impact hammer milling](#) methods are typical when minimal refined whole grain flour is preferred. Information on several American artisan millers that produce various whole grain flours can be found at [Firebird Mills](#); [Hayden Flour Mills](#); and [Baker Miller Chicago](#). Organic flours have gained popularity in the baking industry. As consumers become more aware of them, we see the demand swinging back toward whole grain and artisan milling as a preference.

Attribution

Flour in Baking

Flour forms the foundation for bread, cakes, and pastries. It may be described as the skeleton, which supports the other ingredients in a baked product. This applies to both yeast and chemically leavened products.

The strength of flour is represented in protein (gluten) quality and quantity. This varies greatly from flour to flour. The quality of the protein indicates the strength and stability of the flour, and the result in bread making depends on the method used to develop the gluten by proper handling during the fermentation. Gluten is a rubber-like substance that is formed by mixing flour with water. Before it is mixed it contains two proteins. In wheat, these two proteins are **gliadin** and **glutenin**. Although we use the terms *protein* and *gluten* interchangeably, gluten only develops once the flour is moistened and mixed. The protein in the flour becomes gluten.

Hard spring wheat flours are considered the best for bread making as they have a larger percentage of good quality gluten than soft wheat flours. It is not an uncommon practice for mills to blend hard spring wheat with hard winter wheat for the purpose of producing flour that combines the qualities of both. Good bread flour should have about 13% gluten.

Storing Flour

Flour should be kept in a dry, well-ventilated storeroom at a fairly uniform temperature. A temperature of about 21°C (70°F) with a relative humidity of 60% is considered ideal. Flour should never be stored in a damp place. Moist storerooms with temperatures greater than 23°C (74°F) are conducive to mould growth, bacterial development, and rapid deterioration of the flour. A well-ventilated storage room is necessary because flour absorbs and retains odors. For this reason, flour should not be stored in the same place as onions, garlic, coffee, or cheese, all of which give off strong odors.

Flour Tests

Wheat that is milled and blended with modern milling methods produce flours that have a fairly uniform quality all year round and, if purchased from a reliable mill, they should not require any testing for quality. The teacher, student, and professional baker, however, should be familiar with qualitative differences in flours and should know the most common testing methods.

Flours are mainly tested for:

- Color
- Absorption
- Gluten strength
- Baking quality

Other tests, done in a laboratory, are done for:

- Albumen
- Starch
- Sugar
- Dextrin
- Mineral and fat content

Color

The color of the flour has a direct bearing on baked bread, providing that fermentation has been carried out properly. The addition of other ingredients to the dough, such as **brown sugar**, malt, molasses, salt, and colored margarine, also affects the color of bread.

To test the color of the flour, place a small quantity on a smooth glass, and with a spatula, work until a firm smooth mass about 5 cm (2 in.) square is formed. The thickness should be about 2 cm (4/5 in.) at the back of the plate to a thin film at the front. The test should be made in comparison with a flour of known grade and quality, both flours being worked side by side on the same glass. A creamy white color indicates a hard flour of good gluten quality. A dark or greyish color indicates a poor grade of flour or the presence of dirt. Bran specks indicate a low grade of flour.

After making a color comparison of the dry samples, dip the glass on an angle into clean water and allow to partially dry. Variations in color and the presence of bran specks are more easily identified in the damp samples.

Absorption

Flours are tested for absorption because different flours absorb different amounts of water and therefore make doughs of different consistencies. The absorption ability of a flour is usually between 55% and 65%. To determine the absorption factor, place a small quantity of flour (100 g/4 oz.) in a bowl. Add water gradually from a beaker containing a known amount of water. As the water is added, mix with a spoon until the dough reaches the desired consistency. You can knead the dough by hand for final mixing and determination of consistency. Weigh the unused water. Divide the weight of the water used by the weight of the flour used. The result is the absorption ability in percentage. For example:

Weight of flour used 100 g (4 oz.)

Weight of water used 60 g (2.7 oz.)

Therefore absorption = $6/10$ or 60%

Prolonged storage in a dry place results in a natural moisture loss in flour and has a noticeable effect on the dough. For example, a sack of flour that originally weighed 40 kg (88 lb.) with a moisture content of 14% may be reduced to 39 kg (86 lb.) during storage. This means that 1 kg (2 lb.) of water is lost and must be made up when mixing. The moisture content of the wheat used to make the flour is also important from an economic standpoint.

Hard wheat flour absorbs more liquid than soft flour. Good hard wheat flour should feel somewhat granular when rubbed between the thumb and fingers. A soft, smooth feeling indicates a soft wheat flour or a blend of soft and hard wheat flour. Another indicator is that hard wheat flour retains its form when pressed in the hollow of the hand and falls apart readily when touched. Soft wheat flour tends to remain lumped together after pressure.

Gluten Strength

The gluten test is done to find the variation of gluten quality and quantity in different kinds of flour. Hard flour has more gluten of better quality than soft flour. The gluten strength and quality of two different kinds of hard flour may also vary with the weather conditions and the place where the wheat is grown. The difference may be measured exactly by laboratory tests, or roughly assessed by the variation of gluten balls made from different kinds of hard flours.

For example, to test the gluten in hard flour and all-purpose flour, mix 250 g (9 oz.) of each in separate mixing bowls with enough water to make each dough stiff. Mix and develop each dough until smooth. Let the dough rest for about 10 minutes. Wash each dough separately while kneading it under a stream of cold water until the water runs clean and all the starch is washed out. (Keep a flour sieve in the sink to prevent dough pieces from being washed down the drain.) What remains will be crude gluten. Shape the crude gluten into round balls, then place them on a paper-lined baking pan and bake at 215°C (420°F) for about

one hour. The gluten ball made from the hard flour will be larger than the one made from all-purpose flour. This illustrates the ability of hard flour to produce a greater volume because of its higher gluten content.

Ash Content

Ash or mineral content of flour is used as another measurement of quality. Earlier in the chapter, we talked about extraction rates as an indicator of how much of the grain has been refined. Ash content refers to the amount of ash that would be left over if you were to burn 100 g of flour. A higher ash content indicates that the flour contains more of the germ, bran, and outer endosperm. Lower ash content means that the flour is more highly refined (i.e., a lower extraction rate).

Baking Quality

The final and conclusive test of any flour is the kind of bread that can be made from it. The baking test enables the baker to check on the completed loaf that can be expected from any given flour. Good volume is related to good quality gluten; poor volume to young or green flour. Flour that lacks stability or power to hold during the entire fermentation may result in small, flat bread. Flour of this type may sometimes respond to an increase in the amount of yeast. More yeast shortens the fermentation time and keeps the dough in better condition during the pan fermentation period.

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Rye Flour

Rye is a hardy cereal grass cultivated for its grain. Its use by humans can be traced back over 2,000 years. Once a staple food in Scandinavia and Eastern Europe, rye declined in popularity as wheat became more available through world trade. A crop well suited to northern climates, rye is grown on the Canadian Prairies and in the northern states such as the Dakotas and Wisconsin.

Rye flour is the only flour other than wheat that can be used without blending (with wheat flour) to make yeast-raised breads. Nutritionally, it is a grain comparable in value to wheat. In some cases, for example, its lysine content (an amino acid), is even biologically superior.

The brown grain is cleaned, tempered, and milled much like wheat grain. One difference is that the rye endosperm is soft and breaks down into flour much quicker than wheat. As a result, it does not yield semolina, so purifiers are seldom used. The bran is separated from the flour by the break roller, and the flour is further rolled and sifted while being graded into chop, meal, light flour, medium flour, and dark flour:

- Chop: This is the miller's name for the coarse stock after grinding in a break roller mill.
- Meal: Like chop, meal is made of 100% extraction obtained by grinding the entire rye kernel.
- Light rye flour: This is obtained from the centre of the rye kernel and is low in protein and high in starch content. It can be compared to white bread flour and is used to make light rye breads.
- Medium rye flour: This is straight flour and consists of all the kernels after the bran and shorts have been removed. It is light grey in colour, has an ash content of 1%, and is used for a variety of sourdough breads.
- Dark rye flour: This is comparable to first clear wheat flour. It has an ash content of 2% and a protein content of 16%. It is used primarily for heavier types of rye bread.

The lighter rye flours are generally bleached, usually with a chlorine treatment. The purpose of bleaching is to lighten the colour, since there is no improvement on the gluten capability of the flour.

Extraction of Rye Flour

The grade of extraction of rye flour is of great importance to the yield of the dough and the creation of a particular flavour in the baked bread. Table 1 shows the percentage of the dry substances of rye flour by grade of extraction.

Substance	Grade of Extraction	
	70%	85%
Ash	0.8%	1.4%
Fat	1.2%	1.7%
Protein	8.1%	9.6%
Sugar	6.5%	7.5%

Starch	72.5%	65.1%
Crude fibre	0.5%	1.3%
Pentosans	5.2%	7.6%
Undefinable	5.2%	5.8%

Note that ash, fibre, and **pentosans** are higher in the 85% extraction rate flour, and starch is lower. Pentosans are gummy carbohydrates that tend to swell when moistened and, in baking, help to give the rye loaf its cohesiveness and structure. The pentosan level in rye flour is greater than that of wheat flour and is of more significance for successful rye bread baking.

Rye flours differ from wheat flours in the type of gluten that they contain. Although some dark rye flours can have a gluten content as high as 16%, this is only gliadin. The glutenin, which forms the elasticity in dough is absent, and therefore doughs made only with rye flour will not hold the gas produced by the yeast during fermentation. This results in a small and compact loaf of bread.

Starch and pentosans are far more important to the quality of the dough yield than gluten. Starch is the chief component of the flour responsible for the structure of the loaf. Its bread-making ability hinges on the age of the flour and the acidity. While rye flour does not have to be aged as much as wheat flour, it has both a “best after” and a “best before” date. Three weeks after milling is considered to be good.

When the rye flour is freshly milled, the starch gelatinizes (sets) quickly at a temperature at which amylases are still very active. As a result, bread made from fresh flour may be sticky and very moist. At the other extreme, as the starch gets older, it gelatinizes less readily, the enzymes cannot do their work, and the loaf may split and crack. A certain amount of starch breakdown must occur for the dough to be able to swell.

The moisture content of rye flour should be between 13% and 14%. The less water in the flour, the better its storage ability. Rye should be stored under similar conditions to wheat flour.

Differences between Rye and Wheat

Here is a short list of the differences between rye and wheat:

- Rye is more easily pulverized.
- Rye does not yield semolina.
- Gluten content in rye is not a significant dough-making factor.
- Starch is more important for bread making in rye flour than in wheat flour.
- The pentosan level in rye flour is higher and more important for bread making.
- Rye flour has greater water binding capability than wheat flour, due to its starch and pentosan content.

In summary, both wheat and rye have a long history in providing the “staff of life.” They are both highly nutritious. North American mills have state-of-the-art technology that compensates for crop differences, thus ensuring that the baker has a reliable and predictable raw material. Flour comes in a great variety of types, specially formulated so that the baker can choose according to product and customer taste.

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Other Grains and Flours

Several other types of grains are commonly used in baking. In particular, corn and oats feature predominantly in certain types of baking (quick breads and cookies respectively, for instance) but increasingly rice flour is being used in baked goods, particularly for people with gluten sensitivities or intolerances. The trend to whole grains and the influence of different ethnic cultures has also meant the increase in the use of other grains and pulses for flours used in breads and baking in general.

Corn

Corn is one of the most widely used grains in the world, and not only for baking. Corn is used in breads and cereals, but also to produce sugars (such as dextrose and corn syrup), starch, plastics, adhesives, fuel (ethanol), and alcohol (bourbon and other whisky). It is produced from the **maize** plant (the preferred scientific and formal name of the plant that we call *corn* in North America). There are different varieties of corn, some of which are soft and sweet (corn you use for eating fresh or for cooking) and some of which are starchy and are generally dried to use for baking, animal feed, and popcorn.

Varieties Used in Baking

- Cornmeal has a sandy texture and is ground to fine, medium, and coarse consistencies. It has most of the husk and germ removed, and is used in recipes from the American South (e.g., cornbread) and can be used to add texture to other types of breads and pastry.
- Stone-ground cornmeal has a texture not unlike whole wheat flour, as it contains some of the husk and germ. Stone ground cornmeal has more **nutrients**, but it is also more perishable. In baking, it acts more like cake flour due to the lack of gluten.
- Corn flour in North America is very finely ground cornmeal that has had the husk and germ removed. It has a very soft powdery texture. In the U.K. and Australia, corn flour refers to cornstarch.
- Cornstarch is the starch extracted from the maize kernel. It is primarily used as a thickener in baking and other cooking. Cornstarch has a very fine powdery consistency, and can be dissolved easily in water. As a thickening agent, it requires heat to set, and will produce products with a shiny, clear consistency.
- Blue cornmeal has a light blue or violet colour and is produced from whole kernels of blue corn. It is most similar to stone-ground cornmeal and has a slightly sweet flavour.

Rice

Rice is another of the world's most widely used cereal crops and forms the staple for much of the world's diet. Because rice is not grown in Canada, it is not regulated by the Canadian Grain Commission.

Varieties Used in Baking

- Rice flour is prepared from finely ground rice that has had the husks removed. It has a fine, slightly sandy texture, and provides crispness while remaining tender due to its lack of gluten. For this reason, many gluten-free breads are based on rice flours or blends that contain rice flour.
- Short grain or pearl rice is also used in the pastry shop to produce rice pudding and other desserts.

Oats

Oats are widely used for animal feed and food production, as well as for making breads, cookies, and dessert toppings. Oats add texture to baked goods and desserts.

Varieties Used in Baking

- Bakers will most often encounter rolled oats, which are produced by pressing the de-husked whole kernels through rollers.
- Oat bran and oat flour are produced by grinding the oat kernels and separating out the bran and endosperm.
- Whole grain oat flour is produced by grinding the whole kernel but leaving the ground flour intact.
- Steel-cut oats are more commonly used in cooking and making breakfast cereals, and are the chopped oat kernels.

Other Grains and Pulses

A wide range of additional flours and grains that are used in ethnic cooking and baking are becoming more and more widely available in Canada. These may be produced from grains (such as kamut, spelt, and quinoa), pulses (such as lentils and chickpeas), and other crops (such as buckwheat) that have a grain-like consistency when dried. Increasingly, with allergies and intolerances on the rise, these flours are being used in bakeshops as alternatives to wheat-based products for customers with special dietary needs.

Attribution

Understanding Ingredients: Fat

Understanding Fats and Oils

Fats and oils are organic compounds that, like carbohydrates, are composed of the elements carbon (C), hydrogen (H), and oxygen (O), arranged to form molecules. There are many types of fats and oils and a number of terms and concepts associated with them, which are detailed further here.

Lipids

In baking, lipids are generally a synonym for fats. Baking books may talk about the “lipid content of eggs,” for example.

Triglycerides

Triglycerides is another chemical name for the most common type of fats found in the body, indicating that they are usually made up of three (tri) fatty acids and one molecule of glycerol (glycerine is another name) as shown in Figure 3. (The mono and diglycerides that are used as emulsifiers have one and two fatty acids respectively.)



Figure 1 Composition of fats (triglycerides)

Fatty Acids

Each kind of fat or oil has a different combination of fatty acids. The nature of the fatty acid will determine the consistency of the fat or oil. For example, stearic acid is the major fatty acid in beef fat, and linoleic acid is dominant in seed oils. Fatty acids are defined as short, medium, or long chain, depending on the number of atoms in the molecule.

The reason that some fat melts gradually is that as the temperature rises, each fatty acid will, in turn, soften, as its melting point is reached. Fats that melt all of a sudden mean that the fatty acids are of the same or similar type and have melting points within a narrow range. An example of such a fat is coconut fat: one second it is solid, the next, liquid.

Table 1 shows the characteristics of three fatty acids.

Table 1: Characteristics of Fatty Acids		
Type of Fatty Acid	Melting Point	Physical State (at room temperature)
Stearic	69°C (157°F)	Solid
Oleic	16°C (61°F)	Liquid
Linoleic	-12°C (9°F)	Liquid

Rancid

Rancid is a term used to indicate that fat has spoiled. The fat takes on an unpleasant flavor when exposed to air and heat. Unsalted butter, for example, will go rancid quickly if left outside the refrigerator, especially in warm climates.

Oxidation/Antioxidants

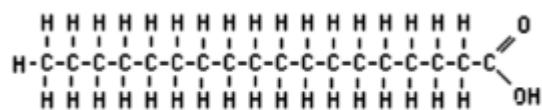
Oxidation (exposure to air) causes rancidity in fats over time. This is made worse by combination with certain metals, such as copper. This is why doughnuts are never fried in copper pans!

Some oils contain natural antioxidants, such as tocopherols (vitamin E is one kind), but these are often destroyed during the processing. As a result, manufacturers add synthetic antioxidants to retard rancidity. BHA and BHT are synthetic antioxidants commonly used by fat manufacturers.

Saturated/Unsaturated

Saturated and unsaturated refer to the extent to which the carbon atoms in the molecule of fatty acid are linked or bonded (saturated) to hydrogen atoms. One system of fatty acid classification is based on the number of double bonds.

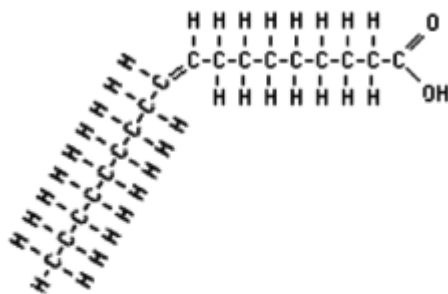
- 0 double bonds: saturated fatty acids. Stearic acid is a typical long-chain saturated fatty acid (Figure 2).^[1]



Stearic acid, a saturated fatty acid

Figure 2 Stearic Acid

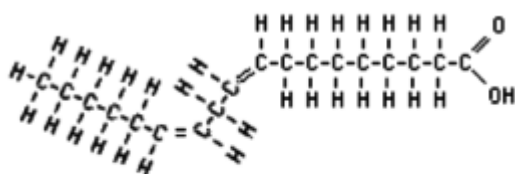
- 1 double bond: monounsaturated fatty acids. Oleic acid is a typical monounsaturated fatty acid (Figure 3).^[2]



Oleic acid, a monounsaturated fatty acid.
Note that the double bond is *cis*; this is the common natural configuration.

Figure 3 Oleic Acid

- 2 or more double bonds: polyunsaturated fatty acids. Linoleic acid is a typical polyunsaturated fatty acid (Figure 4).^[3]



Linoleic acid, a polyunsaturated fatty acid.
Both double bonds are *cis*.

Figure 4 *Linoleic Acid*

Saturated fat is a type of fat found in food. For many years, there has been a concern that saturated fats may lead to an increased risk of heart disease; however, there have been studies to the contrary and the literature is far from conclusive. The general assumption is that the less saturated fat the better as far as health is concerned. For the fat manufacturer, however, low saturated fat levels make it difficult to produce oils that will stand up to the high temperatures necessary for processes such as deep-frying. Hydrogenation has been technology's solution. Hydrogenation will be discussed later in the chapter.

Saturated fat is found in many foods:

- Animal foods (like beef, chicken, lamb, pork, and veal)
- Coconut, palm, and palm kernel oils
- Dairy products (like butter, cheese, and whole milk)
- Lard
- Shortening

Unsaturated fat is also in the foods you eat. Replacing saturated and trans fats (see below) with unsaturated fats has been shown to help lower cholesterol levels and may reduce the risk of heart disease. Unsaturated fat is also a source of omega-3 and omega-6 fatty acids, which are generally referred to as "healthy" fats. Choose foods with unsaturated fat as part of a balanced diet using the U.S. Department of Health and Human Service's [Dietary Guidelines](#).

Even though unsaturated fat is a "good fat," having too much in your diet may lead to having too many calories, which can increase your risk of developing [obesity](#), [type 2 diabetes](#), [heart disease](#), and certain types of [cancer](#).

There are two main types of unsaturated fats:

- Monounsaturated fat, which can be found in:
 - Avocados
 - Nuts and seeds (like cashews, pecans, almonds, and peanuts)
 - Vegetable oils (like canola, olive, peanut, safflower, sesame, and sunflower)
- Polyunsaturated fat, which can be found in:
 - Fatty fish (like herring, mackerel, salmon, trout and smelt)
 - Fish oils
 - Nuts and seeds (like cashews, pecans, almonds and peanuts)
 - Vegetable oils (like canola, corn, flaxseed, soybean and sunflower)

Hydrogenation

Simply put, hydrogenation is a process of adding hydrogen gas to alter the melting point of the oil or fat. The injected hydrogen bonds with the available carbon, which changes liquid oil into solid fat. This is practical, in that it makes fats versatile. Think of the different temperature conditions within a bakery during which fat must be workable; think of the different climatic conditions encountered in bakeries.

Trans Fat

Trans fat is made from a chemical process known as “partial hydrogenation.” This is when liquid oil is made into a solid fat. Like saturated fat, trans fat has been shown to raise LDL or “bad” cholesterol levels, which may in turn increase your risk for heart disease. Unlike saturated fat, trans fat also lowers HDL or “good” cholesterol. A low level of HDL-cholesterol is also a risk factor for heart disease.

Until recently, most of the trans fat found in a typical American diet came from:

- Fried foods (like doughnuts)
- baked goods including cakes, pie crusts, biscuits, frozen pizza, cookies, and crackers
- stick margarine and other spreads

The US Food and Drug Administration (FDA) specifically prescribe what information must be displayed on a label. The *trans fat* content of food is one piece of core nutrition information that is required to be declared in a nutrition facts table. More information on a nutrition facts table and labeling details can be found in <http://www.fda.gov/food/ingredientspackaginglabeling/labelingnutrition/ucm274590.htm>

Emulsification (Emulsified Shortenings)

Emulsification is the process by which normally unmixable ingredients (such as oil and water) can be combined into a stable substance. Emulsifiers are substances that can aid in this process. There are natural emulsifiers such as lecithin, found in egg yolks. Emulsifiers are generally made up of monoglycerides and diglycerides and have been added to many hydrogenated fats, improving the fat’s ability to:

- Develop a uniformly fine structure
- Absorb a high percentage of sugar
- Hold in suspension a high percentage of liquid

Emulsified shortenings are ideal for cakes and icings, but they are not suitable for deep-frying.

Stability

Stability refers to the ability of a shortening to have an extended shelf life. It refers especially to deep-frying fats, where a smoke point (see below) of 220°C to 230°C (428°F to 446°F) indicates a fat of high stability.

Smoke Point

The smoke point is the temperature reached when fat first starts to smoke. The smoke point will decline over time as the fat breaks down (see below).

Fat Breakdown

The technical term for fat breakdown is hydrolysis, which is the chemical reaction of a substance with water. In this process, fatty acids are separated from their glycerol molecules and accumulate over time in the fat. When their concentration reaches a certain point, the fat takes on an unpleasant taste, and continued use of the fat will yield a nasty flavor. The moisture, which is at the root of this problem, comes from the product being fried. This is why it is a good reason to turn off the fryer or turn it to “standby” between batches of frying foods such as doughnuts. Another cause of fat breakdown is excessive flour on the product or particles breaking off the product.

Attribution

Figure 2. Stearic Acid. Retrieved from http://library.med.utah.edu/NetBiochem/FattyAcids/3_3.html ←

Figure 3 Oleic Acid Retrieved from: http://library.med.utah.edu/NetBiochem/FattyAcids/3_3.html ←

Figure 4 Linoleic Acid Retrieved from: http://library.med.utah.edu/NetBiochem/FattyAcids/3_3.html ←

Sources of Bakery Fats and Oils

Edible fats and oils are obtained from both animal and vegetable sources. Animal sources include:

- Beef
- Pork
- Sheep
- Fish

In North America, the first two are the prime sources.

Vegetable sources include canola, coconut, corn, cotton, olive, palm fruit and palm kernel, peanut, soya bean, safflower, and sunflower.

Refining of Fats and Oils

The major steps in refining fats and oils are as follows:

- Free fatty acids are neutralized and treated with an alkali.
- Color is removed.
- The fat is hydrogenated.
- The fat is deodorized.
- The fat is chilled and beaten to make it softer and whiter. This is done by a votator (a machine that cools and kneads liquid margarine).
- Fat is stored to facilitate the correct crystallization (tempering).

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Major Fats and Oils Used in Bakeries

Lard

Lard is obtained from the fatty tissues of pigs, with a water content of 12% to 18%. Due to dietary concerns, lard has gradually lost much of its former popularity. It is still extensively used, however, for:

- Yeast dough additions
- Pie pastry
- Pan greasing

Lard has a good plastic range, which enables it to be worked in a pie dough at fairly low temperatures (try the same thing with butter!). It has a fibrous texture and does not cream well. It is therefore not suitable for cake making. Some grades of lard also have a distinctive flavor, which is another reason it is unsuitable for cake making.

Butter

Butter is made from sweet, neutralized, or ripened creams pasteurized and standardized to a fat content of 30% to 40%. When cream is churned or overwhipped, the fat particles separate from the watery liquid known as buttermilk. The separated fat is washed and kneaded in a water wheel to give it plasticity and consistency. Color is added during this process to make it look richer, and salt is added to improve its keeping quality.

In Canada, the following regulations apply to butter:

- Minimum 80% milk fat by weight
- Permitted ingredients: milk solids, salt, air or inert gas, permitted food color, permitted bacterial culture
- The grade and grade name for butter and butter products is Canada 1.

Sweet (or unsalted) butter is made from a cream that has a very low acid content and no salt is added to it. It is used in some baking products like French butter cream, where butter should be the only fat used in the recipe. Keep sweet butter in the refrigerator.

From the standpoint of flavor, butter is the most desirable fat used in baking. Its main drawback is its relatively high cost. It has moderate but satisfactory shortening and creaming qualities. When used in cake mixing, additional time, up to five minutes more, should be allowed in the creaming stage to give maximum volume. Adding an emulsifier (about 2% based on flour weight) will also help in cake success, as butter has a poor plastic range of 18°C to 20°C (64°F to 68°F).

Butter and butter products may also be designated as “whipped” where they have had air or inert gas uniformly incorporated into them as a result of whipping. Whipped butter may contain up to 1% added edible casein or edible caseinates.

Butter and butter products may also be designated as “cultured” where they have been produced from cream to which a permitted bacterial culture has been added.

Margarine

Margarines are made primarily from vegetable oils (to some extent hydrogenated) with a small fraction of

milk powder and bacterial culture to give a butter-like flavor. Margarines are very versatile and include:

- General purpose margarine with a low melting point, suitable for blending in dough and general baking
- Cake margarine with excellent creaming qualities
- Roll-in margarine, which is plastic and suitable for Danish pastries
- Puff pastry roll-in, which is the most waxy and has the highest melting point

Margarine may be obtained white, but is generally colored. Margarine has a fat content ranging from 80% to 85%, with the balance pretty much the same as butter.

Oil content claims on margarine

The claim that margarine contains a certain percentage of a specific oil in advertisements should always be based on the percentage of oil by weight of the total product. All the oils used in making the margarine should be named. For example, if a margarine is made from a mixture of corn oil, cottonseed oil, and soybean oil, it would be considered misleading to refer only to the corn oil content in an advertisement for the margarine. On the other hand, the mixture of oils could be correctly referred to as vegetable oils.

It used to be that you could only buy margarines in solid form full of saturated and trans fat. The majority of today's margarines come in tubs, are soft and spreadable, and are non-hydrogenated, which means they have low levels of saturated and trans fat. Great care must be taken when attempting to substitute spreadable margarine for solid margarine in recipes.

Shortenings

Since the invention of [hydrogenated vegetable oil](#) in the early 20th century, *shortening* has come almost exclusively to mean hydrogenated vegetable oil. Vegetable shortening shares many properties with lard: both are semi-solid fats with a higher [smoke point](#) than butter and margarine. They contain less water and are thus less prone to splattering, making them safer for frying. Lard and shortening have a higher fat content (close to 100%) compared to about 80% for butter and margarine. Cake margarines and shortenings tend to contain a bit higher percentage of [monoglycerides](#) than margarines. Such "high-ratio shortenings" blend better with hydrophilic (attracts water) ingredients such as starches and sugar.

Health concerns and reformulation

Early in this century, vegetable shortening became the subject of some health concerns due to its traditional formulation from partially hydrogenated vegetable oils that contain trans fats, which have been linked to a number of [adverse health effects](#). Consequently, a low trans-fat variant of Crisco brand shortening was introduced in 2004. In January 2007, all Crisco products were reformulated to contain less than one gram of trans fat per serving, and the separately marketed trans-fat free version introduced in 2004 was consequently discontinued. Since 2006, many other brands of shortening have also been reformulated to remove trans fats. Non-hydrogenated vegetable shortening can be made from [palm oil](#).

Hydrogenated vegetable shortenings

Hydrogenated shortenings are the biggest group of fats used in the commercial baking industry. They feature the following characteristics:

- They are made from much the same oils as margarine.
- They are versatile fats with good creaming ability.
- Their hydrogenation differs according to the specific use for which the fat is designed.
- They are 100% fat - no water.
- They keep well for six to nine months.

Variations on these shortenings are: emulsified vegetable shortenings, roll-in pastry shortenings, and deep-frying fats.

Emulsified vegetable shortenings

Emulsified vegetable shortenings are also termed *high-ratio fats*. The added emulsifiers (mono- and diglycerides) increase fat dispersion and give added fineness to the baked product. They are ideal for high-ratio cakes, where relatively large amounts of sugar and liquid are incorporated. The result is a cake:

- Fine in texture
- Light in weight and of excellent volume
- Superior in moisture retention (good shelf life)
- Tender to eat

This is also the fat of choice for many white cake icings.

Roll-in pastry shortenings

This type of shortening is also called special pastry shortening (SPS). These fats have a semi-waxy consistency and offer:

- Large plastic range
- Excellent extensibility
- Excellent lifting ability

They are primarily used in puff pastry and Danish pastry products where lamination is required. They come in various specialized forms, with varying qualities and melting points. It is all a matter of compromise between cost, palatability, and leavening power. A roll-in that does not have “palate cling” may have a melting point too low to guarantee maximum lift in a puff pastry product.

Deep-Frying Fats

Deep-frying fats are special hydrogenated fats that have the following features:

- High smoke point of up to 250°C (480°F)
- High heat stability and resistance to fat breakdown
- No undesirable flavor on finished products
- No greasiness when cold

These fats contain an anti-foaming agent.

Vegetable Oils

Vegetable oil is an acceptable common name for an oil that contains more than one type of vegetable oil. Generally, when such a vegetable oil blend is used as an ingredient in another food, it may be listed in the ingredients as “vegetable oil.”

There are two exceptions: if the vegetable oils are ingredients of a cooking oil, salad oil, or table oil, the oils must be specifically named in the ingredient list (e.g., canola oil, corn oil, safflower oil), and using the general term *vegetable oil* is not acceptable. As well, if any of the oils are coconut oil, palm oil, palm kernel oil, peanut oil, or cocoa butter, the oils must be specifically named in the ingredient list.

When two or more vegetable oils are present and one or more of them has been modified or hydrogenated, the common name on the principal display panel and in the list of ingredients must include the word “modified” or “hydrogenated,” as appropriate (e.g., modified vegetable oil, hydrogenated vegetable oil, modified palm kernel oil).

Vegetable oils are used in:

- Chemically leavened batters (e.g., muffin mixes)
- Dough additives (to replace the fat)
- Short sponges (to replace the butter or fat)

Coconut Fat

Coconut fat is often used to stabilize butter creams as it has a very small plastic range. It has a quite low melting point and its hardness is due to other factors. It can be modified to melt at different temperatures, generally between 32°C and 36°C (90°F and 96°F).

The Importance of Melting Points

As mentioned above, all fats become oils and vice versa, depending on temperature. Physically, fats consist of minute solid fat particles enclosing a microscopic liquid oil fraction. The consistency of fat is very important to the baker. It is very difficult to work with butter (relatively low melting point) in hot weather, for example. At the other extreme, fats with a very high melting point are not very palatable, since they tend to stick to the palate. Fat manufacturers have therefore attempted to customize fats to accommodate the various needs of the baker.

Fats with a melting range between 40°C and 44°C (104°F and 112°F) are considered to be a good compromise between convenience in handling and palatability. New techniques allow fats with quite high melting points without unpleasant palate-cling.

Table 1 shows the melting points of some fats.

Table 1 Melting points of typical fats.	
Type of Fat	Melting Point
Coconut fat	32.5°C-34.5°C (90.5°F-4.1°F)
Regular margarine	34°C (93°F)
Butter	38°C (100°F)
Regular shortenings	44°C-47°C (111°F-116°F)
Roll-in shortenings	40°C-50°C (104°F-122°F)
Roll-in margarine	44°C-54°C (111°F-130°F)

Blending

It is probably safe to say that most fats are combinations or blends of different oils and/or fats.

They may be all vegetable sources. They may be combined vegetable and animal sources. A typical ratio is 90% vegetable source to 10% animal (this is **not** a hard and fast rule). Formerly, blends of vegetable and animal oils and fats were termed *compound fats*. Nowadays, this term, if used at all, may refer also to combinations of purely vegetable origin.

Attribution

Functions of Fat in Baking

The following summarize the various functions of fat in baking.

Tenderizing Agents

Used in sufficient quantity, fats tend to “shorten” the gluten strands in flour; hence their name: shortenings. Traditionally, the best example of such fat was lard.

Creaming Ability

This refers to the extent to which fat, when beaten with a paddle, will build up a structure of air pockets. This aeration, or creaming ability, is especially important for cake baking; the better the creaming ability, the lighter the cake.

Plastic Range

Plastic range relates to the temperature at which the fatty acid component melts and over which shortening will stay workable and will “stretch” without either cracking (too cold) or softening (too warm). A fat that stays “plastic” over a temperature range of 4°C to 32°C (39°F to 90°F) would be rated as excellent. A dough made with such a fat could be taken from the walk-in cooler to the bench in a hot bakeshop and handled interchangeably. Butter, on the other hand, does not have a good plastic range; it is almost too hard to work at 10°C (50°F) and too soft at 27°C (80°F).

Lubrication

In dough making, the fat portion makes it easier for the gluten network to expand. The dough is also easier to mix and to handle. This characteristic is known as *lubrication*.

Moistening Ability

Whether in dough or in a cake batter, fat retards drying out. For this purpose, a 100% fat shortening will be superior to either butter or margarine.

Nutrition

As one of the three major food categories, fats provide a very concentrated source of energy. They contain many of the fatty acids essential for health.

Attribution

Understanding Ingredients: Sugar

Sugar Chemistry (ADD US)

Chemically, sugar consists of carbon (C), oxygen (O), and hydrogen (H) atoms, and is classified as a **carbohydrate**. There are three main groups of sugars, classified according to the way the atoms are arranged together in the molecular structure. These groups are the following:

- **Monosaccharides** or simple sugars. **Dextrose (glucose)** is the major monosaccharide. Others are **levulose** or **fructose** (found in **honey** and many fruits), and galactose, which is a milk sugar. Such sugars do not readily crystallize. (*Mono* means *one*, indicating that the sugar consists of only one molecule.)
- **Disaccharides** or complex sugars. Sucrose (common sugar) is the primary example of a disaccharide. Maltose, found in cereals, and **lactose**, found in milk, are others.
- **Polysaccharides**. Examples are starches, dextrans, and cellulose.

Bakers are not concerned with polysaccharides but rather with the monosaccharides and disaccharides. The latter two both sweeten, but they cannot be used interchangeably because they have different effects on the end product. These differences are touched on later in the book.

Sugar Names

It is helpful to understand some of the conventions of the names of different sugars. Note that sugar names often end in “ose”: sucrose, dextrose, maltose, lactose, etc. **Sucrose** is the chemical name for sugar that comes from the cane and beet sugar plants.

Note that glucose is the chemical name for a particular type of sugar. What is sometimes confusing is that glucose occurs naturally, as a sugar molecule in substances such as honey, but it is also produced industrially from the maize plant (corn).

The Canadian Food and Drug Regulations (FDR) govern the following definitions:

- **Sugars:** All monosaccharides and disaccharides. Used for nutrition labelling purposes.
- **Sweetening agent:** Any food for which a standard is provided in Division 18 of the Food and Drug Regulation, or any combination of these. Includes sugar (sucrose), sugar syrups, and molasses derived from sugar cane or sugar beet, dextrose, glucose and syrups, honey and lactose. Excludes sweeteners considered to be food additives.
- **Sweetening ingredient:** Any sugar, **invert sugar**, honey, dextrose, glucose, or glucose solids, or any combination of these in dry or liquid form. Designed for sweetening fruits, vegetables, and their products and substitutes.
- **Maple syrup:** The syrup obtained by the concentration of maple sap or by the dilution or solution of a maple product, other than maple sap, in potable water.
- **Sweetener:** Any food additive listed as a sweetener. Includes both sugar alcohols and high intensity-sweeteners such as acesulfame-potassium, aspartame, and sucralose.
- **Sugar alcohols:** Food additives that may be used as sweeteners. Includes isomalt, lactitol, maltitol, maltitol syrup, mannitol, sorbitol, sorbitol syrup, xylitol, and erythritol.

Attribution

Sugar Refining

While some refining usually occurs at source, most occurs in the recipient country. The raw sugar that arrives at the ports is not legally edible, being full of impurities.

At the refinery, the raw brown sugar goes through many stages:

- Washing and boiling
- Filtering to remove impurities
- Evaporation to the desired crystal size under vacuum to avoid caramelization
- Centrifuging, in which the fluid is spun off leaving the crystals
- Drying in a rotating drum with hot air
- Packaging in various sizes, depending on the intended market

Sugar beet undergoes identical steps after the initial processing, which involves:

- Slicing the beets and extracting the sugar with hot water
- Removing impurities
- Filtration
- Concentration in evaporators

From here, the process is identical to the final steps in cane processing. See Figure 2 which illustrates the process.

Some of the sugar passes through a machine that presses the moist sugar into cubes and wraps and packages them; still other sugar is made into icing sugar. The sugar refining process is completely mechanical, and machine operators' hands never touch the sugar.

Brown and yellow sugars are produced only in cane sugar refineries. When sugar syrup flows from the centrifuge machine, it passes through further filtration and purification stages and is re-boiled in vacuum pans such as the two illustrated in Figure 2. The sugar crystals are then centrifuged but not washed, so the sugar crystals still retain some of the syrup that gives the product its special flavour and colour.

During the whole refining process almost 100 scientific checks for quality control are made, while workers in research laboratories at the refineries constantly carry out experiments to improve the refining process and the final product. Sugar is carefully checked at the mills and is guaranteed to have a high purity. Government standards both in the United States and Canada require a purity of at least 99.5% sucrose.

Are animal ingredients included in white sugar?

Bone char — often referred to as natural carbon — is widely used by the sugar industry as a decolourizing filter, which allows the sugar cane to achieve its desirable white colour. Other types of filters involve granular carbon or an ion-exchange system rather than bone char.

Bone char is made from the bones of cattle, and it is heavily regulated by the European Union and the USDA. Only countries that are deemed BSE-free can sell the bones of their cattle for this process.

Bone char is also used in other types of sugar. Brown sugar is created by adding molasses to refined sugar, so companies that use bone char in the production of their regular sugar also use it in the production of their brown sugar. Confectioner's sugar — refined sugar mixed with cornstarch — made by these

companies also involves the use of bone char. Fructose may, but does not typically, involve a bone-char filter.

Bone char is not used at the sugar beet factory in Taber, Alberta, or in Montreal's cane refinery. Bone char is used only at the Vancouver cane refinery. All products under the Lantic trademark are free of bone char. For the products under the Rogers trademark, all Taber sugar beet products are also free of bone char. In order to differentiate the Rogers Taber beet products from the Vancouver cane products, you can verify the inked-jet code printed on the product. Products with the code starting with the number "22" are from Taber, Alberta, while products with the code starting with the number "10" are from Vancouver.

If you want to avoid all refined sugars, there are alternatives such as sucanat and turbinado sugar, which are not filtered with bone char. Additionally, beet sugar — though normally refined — never involves the use of bone char.

Attribution

The Application of Sugar

Sugar is the third most used ingredient in the bakeshop. Sugar has several functions in baking. The most recognized purpose is, of course, to sweeten food, but there are many other reasons sugar is used in cooking and baking:

- It can be used for browning effect, both **caramelization** and the **Maillard reaction**, on everything from breads to cookies to cakes. Browning gives a pleasant colour and flavour to the finished product. Caramelization results from the action of heat on sugars. At high temperatures, the chemical changes associated with melting sugars result in a deep brown colour and new flavours. The Maillard reaction results from chemical interactions between sugars and proteins at high heat. An amino group from a protein combines with a reducing sugar to produce a brown colour in a variety of foods (e.g., brewed coffee, fried foods, and breads).
- It acts as the most important tenderizing agent in all baked goods, and one of the factors responsible for the spread in cookies. It helps delay the formation of gluten, which is essential for maintaining a soft or tender product.
- It makes an important contribution to the way we perceive the texture of food. For example, adding sugar to ice cream provides body and texture, which is perceived as smoothness. This addition helps prevent lactose crystallization and thus reduces sugar crystal formation that otherwise causes a grainy texture sometimes associated with frozen dairy products.
- It preserves food when used in sufficient quantity.
- In baking, it increases the effectiveness of yeast by providing an immediate and more usable source of nourishment for the yeast's growth. This hastens the leavening process by producing more carbon dioxide, which allows the dough to rise at a quicker and more consistent rate.

Just as there are many functions of sugar in the bakeshop, there are different uses for the various types of sugar as well:

- Fine granulated sugar is most used by bakers. It generally dissolves easily in mixes and is pure enough for sugar crafters to boil for “pulled” sugar decorations.
- Coarse granulated sugar may be used for a topping on sugar cookies, puff pastry, and Danish pastries as it doesn't liquefy or caramelize so readily. In some European countries, an extra coarse sugar (called *hail* — a literal translation) is used for this purpose.
- Icing or powdered sugar is used in icings and fillings and in sifted form as a top decoration on many baked goods.
- Brown or yellow sugars are used where their unique flavour is important, or in bakeries where an old-fashioned or rustic image is projected. Brown sugar can usually be substituted for white sugar without technical problems in sugar/batter mixes such as cakes and muffins, and in bread dough.

Attribution

Agave

Agave has gained popularity in the food industry due to some of its nutritional properties. The agave nectar is obtained from the sap of the heart of the agave plant, a desert succulent, which is also used to produce tequila. The syrup/sugar production process of agave is similar to that of sugar. See more about the nutritional properties and application of agave in the chapter Special Diets, Allergies, Intolerances, Emerging Issues, and Trends in the open textbook *Nutrition and Labelling for the Canadian Baker*.

A video on the production of agave syrup is available here:
<https://www.youtube.com/watch?v=DZ2moGQL6n8>
Attribution

Glucose/Dextrose

The sugar known as glucose has two origins:

- In a natural form in most fruits
- In a processed form from corn (corn syrup)

In baking, we usually refer to industrially made glucose. It is made from corn and the resulting product, a thick syrup, is then adjusted to a uniform viscosity or consistency. The particular form of the syrup is defined by what is known as the *dextrose equivalent*, or DE for short. Corn syrup is the most familiar form of glucose.

In plant baking, high-fructose corn syrup (HFCS) is the major sweetening agent in bread and buns. It consists of roughly half fructose and half dextrose. Dextrose (chemically identical to glucose) is available in crystalline form and has certain advantages over sucrose:

- It is easily fermentable.
- It contributes to browning in bread and bun making.
- In crystalline form, it is often used in doughnut sugars as it is more inclined to stay dry and non-greasy.
- It is **hygroscopic** and valued as a moisture-retaining ingredient.
- It retards crystallization in syrups, candies, and fondant.

Corn syrup is made from the starch of maize (corn) and contains varying amounts of glucose and maltose, depending on the processing methods. Corn syrup is used in foods to soften texture, add volume, prevent crystallization of sugar, and enhance flavor.

Glucose/dextrose has a sweetening level of approximately three-quarters that of sugar. Table 1 shows the amount of corn syrup or HFCS needed to replace sugar in a formula.

Table1 Replacement factor for Corn Syrup and High-Fructose Corn Syrup		
Type of Sugar	Solids	Replacement Factor
Granulated sugar	100%	1.0
Regular corn syrup	80%	1.25
High-fructose corn syrup	71%	1.41

Glucose, HFCS, and corn syrup are not appropriate substitutions for sucrose in all bakery products. Certain types of cakes, such as white layer cakes, will brown too much if glucose or HFCS is used in place of sugar.
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Honey

Honey is a natural food, essentially an invert sugar. Bees gather nectar and, through the enzyme invertase, change it into honey. Honey varies in composition and flavor depending on the source of the nectar. The average composition of honey is about 40% levulose, 35% dextrose, and 15% water, with the remainder being ash, waxes, and gum.

Blended honey is a mixture of pure honey and manufactured invert sugar, or a blend of different types of honey mixed together to produce a good consistency, color, and aroma. Dehydrated honey is available in a granular form.

Store honey in a tightly covered container in a dry place and at room temperature because it is hygroscopic, meaning it absorbs and retains moisture. Refrigeration or freezing won't harm the color or flavor but it may hasten granulation. Liquid honey crystallizes during storage and is re-liquefied by warming in a double boiler not exceeding a temperature of 58°C (136°F).

Honey is used in baking:

- As a sweetener
- To add unique flavor
- In gingerbread and special cookies where a certain moistness is characteristic of the product
- To improve keeping qualities

There are several types of honey available:

- Comb honey is "packed by the bees" directly from the hive.
- Liquid honey is extracted from the comb and strained. It is the type used by most bakers.
- Creamed honey has a certain amount of crystallized honey added to liquid honey to give body to the final product.
- Chunk honey consists of pieces of comb honey as well as liquid.
- Granulated honey has been crystallized.

In the United States, honey categories are based on color, from white to dark amber. Honey from orange blossom is an example of white honey. Clover honey is an amber honey, and sage and buckwheat honeys are dark amber honeys.

Attribution

Malt

Malt is the name given to a sweetening agent made primarily from barley. The enzymes from the germ of the seeds become active, changing much of the starch into maltose, a complex sugar. Maltose has a distinct flavor and is used for making yeast products such as bread and rolls. Malt is considered to be relatively nutritious compared to other sweeteners.

Malt is available as:

- Flour
- Malt syrup
- Malt extract
- Dried malt

The flour is not recommended since it can lead to problems if not scaled precisely. Malt syrup is inconvenient to work with, as it is sticky, heavy, and bulky. Dried malt is the most practical, though it must be kept protected from humidity.

There are two distinct types of malt:

- Diastatic malt flour is dried at low temperature, thus retaining the activity of the diastatic enzymes.
- Non-diastatic malt flour is darker in color. It is treated at high temperature, which kills the enzymes, and the result is non-diastatic malt.

Crushing malted grain in water produces malt syrup. This dissolves the maltose and soluble enzymes. The liquid is concentrated, producing the syrup. If the process is continued, a dry crystallized product called dried malt syrup is obtained.

Malt syrup has a peculiar flavor, which many people find desirable. It is used in candy, malted milk, and many other products. The alcoholic beverage industry is the largest consumer of malt by far, but considerable quantities are used in syrup and dried malt syrup, both of which are divided into diastatic and non-diastatic malt.

Both diastatic and non-diastatic malts add sweetness, color, and flavor to baked products. Both are valuable since they contain malt sugar, which is fermented by the yeast in the later stages of fermentation. Other sugars such as glucose and levulose are used up rapidly by fermenting yeast in the early stages of fermentation.

Diastatic malt is made with various levels of active enzymes. Malt with medium diastatic activity is recommended. Normally, bread bakers will find sufficient enzymes in well-balanced flour from a good mill, so it is unnecessary to use diastatic malt.

When using dry diastatic malt, about the same weight should be used as liquid regular diastatic malt. Adjustment is made at the factory insofar as the enzyme level is increased in the dry product to compensate. Since the dry type contains about 20% less moisture than the liquid type, add water to make up the difference if dry diastatic malt is substituted for malt syrup.

The main uses of malt in the bakery are to:

- Add nutritive value, as it is rich in vitamins and essential amino acids
- Lengthen shelf life through its ability to attract moisture

- Help fermentation by strengthening the gluten and feeding the yeast
- Make products more appealing through browning of the crust
- Add unique flavor to products when used in sufficient quantity

Table 1 shows the suggested use levels for malt.

Table 1 Recommended level of malt for various baked goods	
Product	Percentage of Flour Weight
White pan bread	0.5-1.5
Sweet goods	1.5-3.0
French/Italian bread	0.5-2.0
Whole wheat bread	5.0-9.0
Pretzels	1.5-6.0
Hard rolls	3.0-5.5

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Maple Syrup (ADD US)

Canada is responsible for 84% of the world's maple syrup production, with the United States being responsible for the remaining 16%. Maple syrup is made by boiling and evaporating the sap of the sugar maple tree. Because sap is only 2% or 3% sugar, it takes almost 40 liters of sap to make 1 liter of syrup. This makes maple syrup a very expensive sweetener. It is prized for its unique flavor and sweet aroma. Don't confuse maple-flavored pancake or table syrup with real maple syrup. Table syrup is made from inexpensive glucose or corn syrup, with added caramel coloring and maple flavoring.

Maple syrup in Canada has two categories:

Canada Grade A, which has four color/flavor classes

- (i) golden, delicate taste
- (ii) amber, rich taste
- (iii) dark, robust taste
- (iv) very dark, strong taste

Canada Processing Grade, which has no color descriptors (any maple syrup that possesses minimal food quality defects but still meets all government regulatory standards for food quality and safety for human consumption)

This definition and grading system gives consumers more consistent and relevant information about the varieties, and helps them make informed choices when choosing maple syrup.

Darker maple syrups are better for baking as they have a more robust flavor. Using maple sugar is also a good way to impart flavor. Maple sugar is what remains after the sap of the sugar maple is boiled for longer than is needed to create maple syrup. Once almost all the water has been boiled off, all that is left is a solid sugar. It can be used to flavor some maple products and as an alternative to cane sugar.

For a video on maple syrup production, see: <https://www.youtube.com/watch?v=OFIj4pMYpTQ>

Attribution

Sugar Substitutes (ADD US)

In Canada, food additives such as sugar substitutes, which cover both artificial sweeteners and intense sweeteners obtained from natural sources, are subject to rigorous controls under the [Food and Drugs Act](#) and [Regulations](#). New food additives (or new uses of permitted food additives) are permitted only once a safety assessment has been conducted and regulatory amendments have been enacted.

Several sugar substitutes have been approved for use in Canada. These include acesulfame-potassium, aspartame, polydextrose, saccharin, stevia, sucralose, thaumatin, and sugar alcohols (polyols) like sorbitol, isomalt, lactitol, maltitol, mannitol, and xylitol. Please see the [Health Canada website](#) for more information on sugar substitutes.

Bakers must be careful when replacing sugar (sucrose) with these sugar substitutes in recipes. Even though the sweetness comparison levels may be similar (or less), it is generally not possible to do straight 1-for-1 substitution. Sugar (sucrose) plays many roles in a recipe:

- It is a bulking agent.
- It absorbs moisture.
- It is a tenderizer.
- It adds moisture and extends shelf life.
- It adds colour (caramelization).

Sugar substitutes may not work in a recipe in the same way.

More information on sugar substitutes and their relative sweetness can be found here:

<http://www.sugar-and-sweetener-guide.com/sweetener-values.html>

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Understanding Ingredients: Leavening Agents

Catalysis

Learning Objectives

By the end of this section, you will be able to:

- Explain the function of a catalyst in terms of reaction mechanisms and potential energy diagrams
- List examples of catalysis in natural and industrial processes

We have seen that the rate of many reactions can be accelerated by catalysts. A catalyst speeds up the rate of a reaction by lowering the activation energy; in addition, the catalyst is regenerated in the process. Several reactions that are thermodynamically favorable in the absence of a catalyst only occur at a reasonable rate when a catalyst is present. One such reaction is catalytic hydrogenation, the process by which hydrogen is added across an alkene C=C bond to afford the saturated alkane product. A comparison of the reaction coordinate diagrams (also known as energy diagrams) for catalyzed and uncatalyzed alkene hydrogenation is shown in Figure 1.

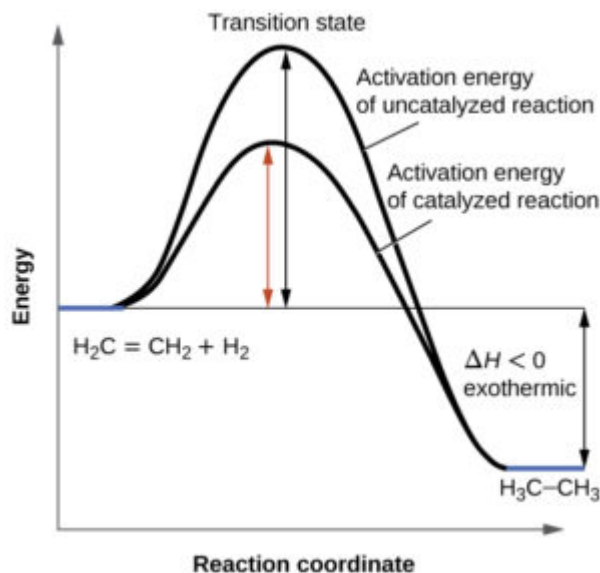


Figure 1 This graph compares the reaction coordinates for catalyzed and uncatalyzed alkene hydrogenation.

Catalysts function by providing an alternate reaction mechanism that has a lower activation energy than would be found in the absence of the catalyst. In some cases, the catalyzed mechanism may include additional steps, as depicted in the reaction diagrams shown in Figure 2. This lower activation energy results in an increase in rate as described by the Arrhenius equation. Note that a catalyst decreases the activation energy for both the forward and the reverse reactions and hence accelerates both the forward and the reverse reactions. Consequently, the presence of a catalyst will permit a system to reach equilibrium more quickly, but it has no effect on the position of the equilibrium as reflected in the value of its equilibrium constant (see the later chapter on chemical equilibrium).

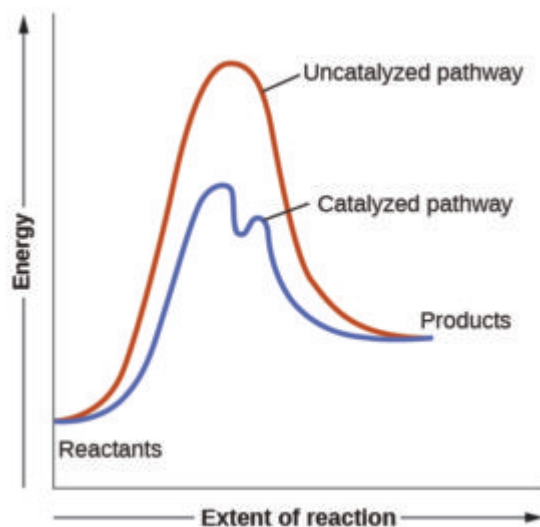
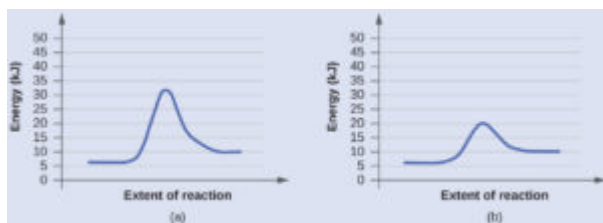


Figure 2 This potential energy diagram shows the effect of a catalyst on the activation energy. The catalyst provides a different reaction path with a lower activation energy. As shown, the catalyzed pathway involves a two-step mechanism (note the presence of two transition states) and an intermediate species (represented by the valley between the two transition states).

Example 1: Using Reaction Diagrams to Compare Catalyzed Reactions

The two reaction diagrams here represent the same reaction: one without a catalyst and one with a catalyst. Identify which diagram suggests the presence of a catalyst, and determine the activation energy for the catalyzed reaction:



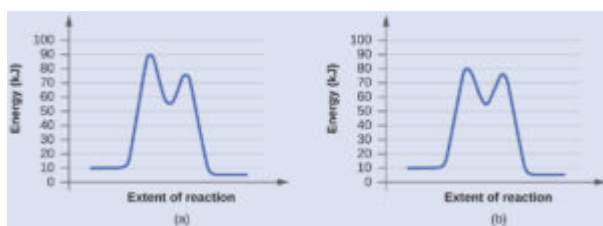
Solution

A catalyst does not affect the energy of reactant or product, so those aspects of the diagrams can be ignored; they are, as we would expect, identical in that respect. There is, however, a noticeable difference in the transition state, which is distinctly lower in diagram (b) than it is in (a). This indicates the use of a catalyst in diagram (b). The activation energy is the difference between the energy of the starting reagents and the transition state—a maximum on the reaction coordinate diagram. The reagents are at 6 kJ and the transition state is at 20 kJ, so the activation energy can be calculated as follows:

$$E_a = 20\text{kJ} - 6\text{kJ} = 14\text{kJ}$$

Check Your Learning

Determine which of the two diagrams here (both for the same reaction) involves a catalyst, and identify the activation energy for the catalyzed reaction:



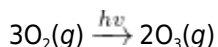
Answer:

Diagram (b) is a catalyzed reaction with an activation energy of about 70 kJ.

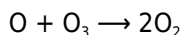
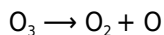
Homogeneous Catalysts

A homogeneous catalyst is present in the same phase as the reactants. It interacts with a reactant to form an intermediate substance, which then decomposes or reacts with another reactant in one or more steps to regenerate the original catalyst and form product.

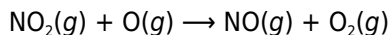
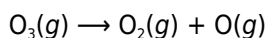
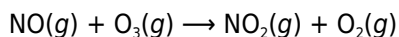
As an important illustration of homogeneous catalysis, consider the earth's ozone layer. Ozone in the upper atmosphere, which protects the earth from ultraviolet radiation, is formed when oxygen molecules absorb ultraviolet light and undergo the reaction:



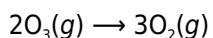
Ozone is a relatively unstable molecule that decomposes to yield diatomic oxygen by the reverse of this equation. This decomposition reaction is consistent with the following mechanism:



The presence of nitric oxide, NO, influences the rate of decomposition of ozone. Nitric oxide acts as a catalyst in the following mechanism:



The overall chemical change for the catalyzed mechanism is the same as:



The nitric oxide reacts and is regenerated in these reactions. It is not permanently used up; thus, it acts as a catalyst. The rate of decomposition of ozone is greater in the presence of nitric oxide because of the catalytic activity of NO. Certain compounds that contain chlorine also catalyze the decomposition of ozone.

Note: Mario J. Molina

The 1995 Nobel Prize in Chemistry was shared by Paul J. Crutzen, Mario J. Molina (Figure 3), and F. Sherwood Rowland "for their work in atmospheric chemistry, particularly concerning the formation and decomposition of ozone."^[1] Molina, a Mexican citizen, carried out the majority of his work at the

Massachusetts Institute of Technology (MIT).

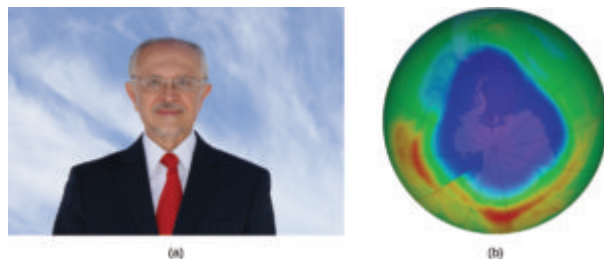
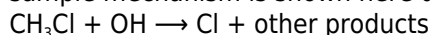
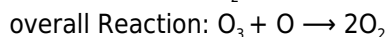
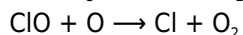
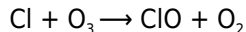


Figure 3 (a) Mexican chemist Mario Molina (1943 –) shared the Nobel Prize in Chemistry in 1995 for his research on (b) the Antarctic ozone hole. (credit a: courtesy of Mario Molina; credit b: modification of work by NASA)

In 1974, Molina and Rowland published a paper in the journal *Nature* (one of the major peer-reviewed publications in the field of science) detailing the threat of chlorofluorocarbon gases to the stability of the ozone layer in earth's upper atmosphere. The ozone layer protects earth from solar radiation by absorbing ultraviolet light. As chemical reactions deplete the amount of ozone in the upper atmosphere, a measurable "hole" forms above Antarctica, and an increase in the amount of solar ultraviolet radiation—strongly linked to the prevalence of skin cancers—reaches earth's surface. The work of Molina and Rowland was instrumental in the adoption of the Montreal Protocol, an international treaty signed in 1987 that successfully began phasing out production of chemicals linked to ozone destruction. Molina and Rowland demonstrated that chlorine atoms from human-made chemicals can catalyze ozone destruction in a process similar to that by which NO accelerates the depletion of ozone. Chlorine atoms are generated when chlorocarbons or chlorofluorocarbons—once widely used as refrigerants and propellants—are photochemically decomposed by ultraviolet light or react with hydroxyl radicals. A sample mechanism is shown here using methyl chloride:



Chlorine radicals break down ozone and are regenerated by the following catalytic cycle:



A single monatomic chlorine can break down thousands of ozone molecules. Luckily, the majority of atmospheric chlorine exists as the catalytically inactive forms Cl_2 and ClONO_2 .

Since receiving his portion of the Nobel Prize, Molina has continued his work in atmospheric chemistry at MIT.

Note: Glucose-6-Phosphate Dehydrogenase Deficiency

Enzymes in the human body act as catalysts for important chemical reactions in cellular metabolism. As such, a deficiency of a particular enzyme can translate to a life-threatening disease. G6PD (glucose-6-phosphate dehydrogenase) deficiency, a genetic condition that results in a shortage of the enzyme glucose-6-phosphate dehydrogenase, is the most common enzyme deficiency in humans. This enzyme, shown in Figure 4, is the rate-limiting enzyme for the metabolic pathway that supplies NADPH to cells (Figure 5).

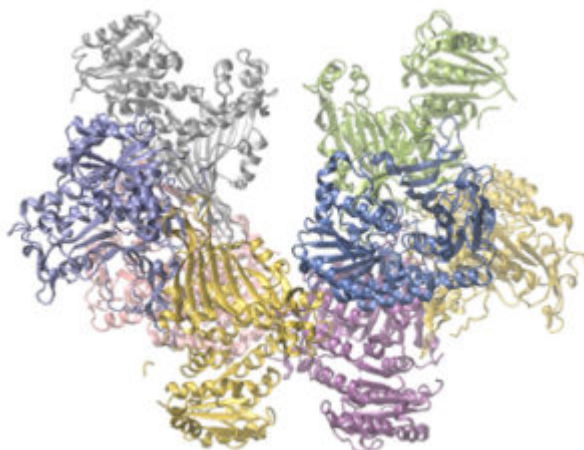


Figure 4 Glucose-6-phosphate dehydrogenase is a rate-limiting enzyme for the metabolic pathway that supplies NADPH to cells.

A disruption in this pathway can lead to reduced glutathione in red blood cells; once all glutathione is consumed, enzymes and other proteins such as hemoglobin are susceptible to damage. For example, hemoglobin can be metabolized to bilirubin, which leads to jaundice, a condition that can become severe. People who suffer from G6PD deficiency must avoid certain foods and medicines containing chemicals that can trigger damage their glutathione-deficient red blood cells.

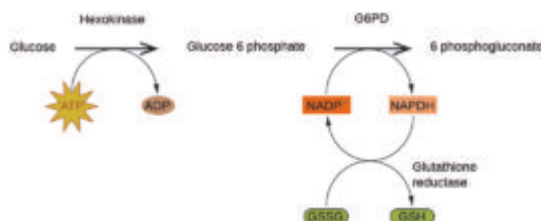


Figure 5 In the mechanism for the pentose phosphate pathway, G6PD catalyzes the reaction that regulates NADPH, a co-enzyme that regulates glutathione, an antioxidant that protects red blood cells and other cells from oxidative damage.

Heterogeneous Catalysts

A heterogeneous catalyst is a catalyst that is present in a different phase (usually a solid) than the reactants. Such catalysts generally function by furnishing an active surface upon which a reaction can occur. Gas and liquid phase reactions catalyzed by heterogeneous catalysts occur on the surface of the catalyst rather than within the gas or liquid phase.

Heterogeneous catalysis has at least four steps:

- Adsorption of the reactant onto the surface of the catalyst
- Activation of the adsorbed reactant
- Reaction of the adsorbed reactant
- Diffusion of the product from the surface into the gas or liquid phase (desorption).

Any one of these steps may be slow and thus may serve as the rate determining step. In general, however, in the presence of the catalyst, the overall rate of the reaction is faster than it would be if the reactants were in the gas or liquid phase.

Figure 6 illustrates the steps that chemists believe to occur in the reaction of compounds containing a carbon-carbon double bond with hydrogen on a nickel catalyst. Nickel is the catalyst used in the

hydrogenation of polyunsaturated fats and oils (which contain several carbon-carbon double bonds) to produce saturated fats and oils (which contain only carbon-carbon single bonds).

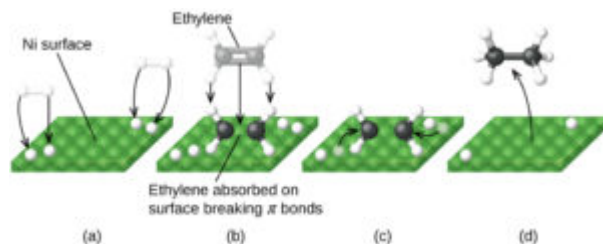


Figure 6 There are four steps in the catalysis of the reaction $C_2H_4 + H_2 \rightarrow C_2H_6$ by nickel. (a) Hydrogen is adsorbed on the surface, breaking the H-H bonds and forming Ni-H bonds. (b) Ethylene is adsorbed on the surface, breaking the π -bond and forming Ni-C bonds. (c) Atoms diffuse across the surface and form new C-H bonds when they collide. (d) C_2H_6 molecules escape from the nickel surface, since they are not strongly attracted to nickel.

Other significant industrial processes that involve the use of heterogeneous catalysts include the preparation of sulfuric acid, the preparation of ammonia, the oxidation of ammonia to nitric acid, and the synthesis of methanol, CH_3OH . Heterogeneous catalysts are also used in the catalytic converters found on most gasoline-powered automobiles (Figure 7).

Note: Automobile Catalytic Converters

Scientists developed catalytic converters to reduce the amount of toxic emissions produced by burning gasoline in internal combustion engines. Catalytic converters take advantage of all five factors that affect the speed of chemical reactions to ensure that exhaust emissions are as safe as possible.

By utilizing a carefully selected blend of catalytically active metals, it is possible to effect complete combustion of all carbon-containing compounds to carbon dioxide while also reducing the output of nitrogen oxides. This is particularly impressive when we consider that one step involves adding more oxygen to the molecule and the other involves removing the oxygen (Figure 7).

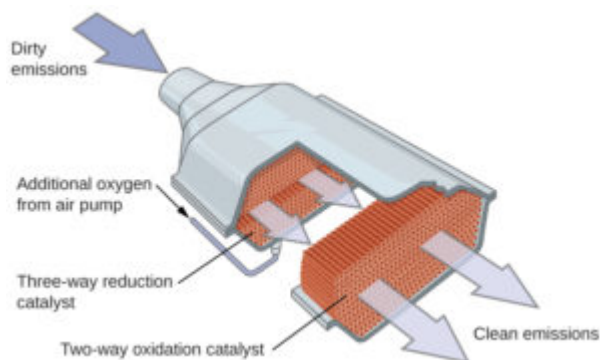
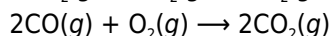
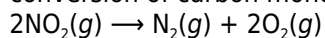
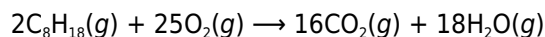


Figure 7 A catalytic converter allows for the combustion of all carbon-containing compounds to carbon dioxide, while at the same time reducing the output of nitrogen oxide and other pollutants in emissions from gasoline-burning engines.

Most modern, three-way catalytic converters possess a surface impregnated with a platinum-rhodium catalyst, which catalyzes the conversion nitric oxide into dinitrogen and oxygen as well as the conversion of carbon monoxide and hydrocarbons such as octane into carbon dioxide and water vapor:





In order to be as efficient as possible, most catalytic converters are preheated by an electric heater. This ensures that the metals in the catalyst are fully active even before the automobile exhaust is hot enough to maintain appropriate reaction temperatures.

The University of California at Davis' "ChemWiki" provides a [thorough explanation](#) of how catalytic converters work.

Note: Enzyme Structure and Function

The study of enzymes is an important interconnection between biology and chemistry. Enzymes are usually proteins (polypeptides) that help to control the rate of chemical reactions between biologically important compounds, particularly those that are involved in cellular metabolism. Different classes of enzymes perform a variety of functions, as shown in Table 1.

Table 1 Classes of Enzymes and Their Functions

Class	Function
oxidoreductases	redox reactions
transferases	transfer of functional groups
hydrolases	hydrolysis reactions
lyases	group elimination to form double bonds
isomerases	isomerization
ligases	bond formation with ATP hydrolysis

Enzyme molecules possess an active site, a part of the molecule with a shape that allows it to bond to a specific substrate (a reactant molecule), forming an enzyme-substrate complex as a reaction intermediate. There are two models that attempt to explain how this active site works. The most simplistic model is referred to as the lock-and-key hypothesis, which suggests that the molecular shapes of the active site and substrate are complementary, fitting together like a key in a lock. The induced fit hypothesis, on the other hand, suggests that the enzyme molecule is flexible and changes shape to accommodate a bond with the substrate. This is not to suggest that an enzyme's active site is completely malleable, however. Both the lock-and-key model and the induced fit model account for the fact that enzymes can only bind with specific substrates, since in general a particular enzyme only catalyzes a particular reaction (Figure 8).

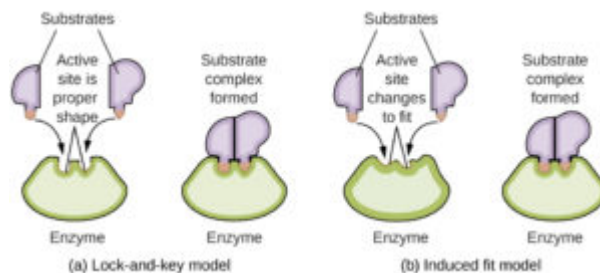


Figure 8 (a) According to the lock-and-key model, the shape of an enzyme's active site is a perfect fit for the substrate. (b) According to the induced fit model, the active site is somewhat flexible, and can change shape in

order to bond with the substrate.

The [Royal Society of Chemistry](#) provides an excellent introduction to enzymes for students and teachers.

Key Concepts and Summary

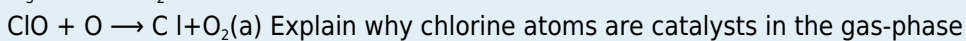
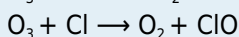
Catalysts affect the rate of a chemical reaction by altering its mechanism to provide a lower activation energy. Catalysts can be homogenous (in the same phase as the reactants) or heterogeneous (a different phase than the reactants).

Catalysis Exercises

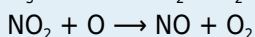
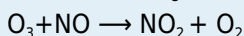
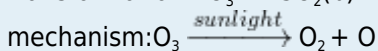
Account for the increase in reaction rate brought about by a catalyst.

Compare the functions of homogeneous and heterogeneous catalysts.

Consider this scenario and answer the following questions: Chlorine atoms resulting from decomposition of chlorofluoromethanes, such as CCl_2F_2 , catalyze the decomposition of ozone in the atmosphere. One simplified mechanism for the decomposition is:

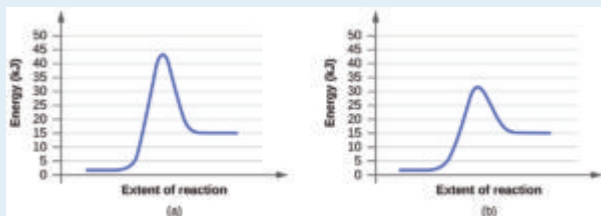
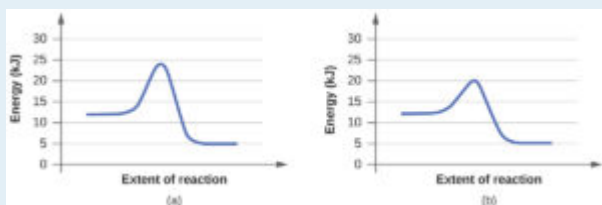


(a) Explain why chlorine atoms are catalysts in the gas-phase transformation: $2\text{O}_3 \rightarrow 3\text{O}_2$



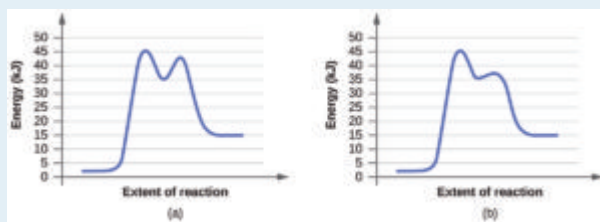
Is NO a catalyst for the decomposition? Explain your answer.

For each of the following pairs of reaction diagrams, identify which of the pair is catalyzed: (a)

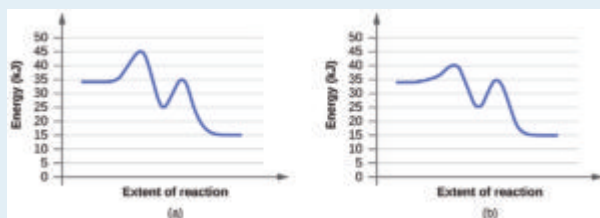


(b)

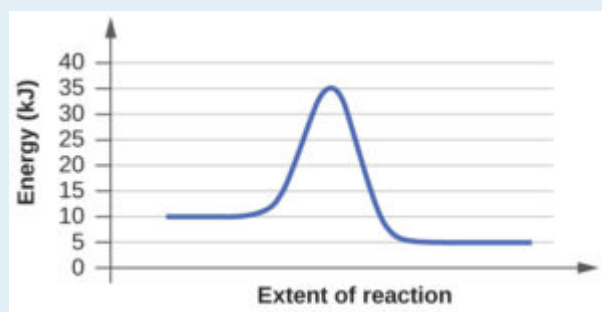
For each of the following pairs of reaction diagrams, identify which of the pairs is catalyzed: (a)



(b)

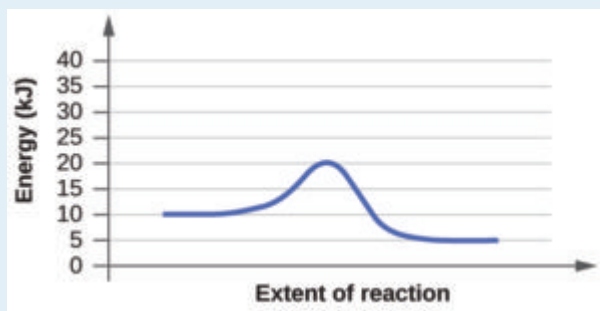


For each of the following reaction diagrams, estimate the activation energy (E_a) of the reaction:

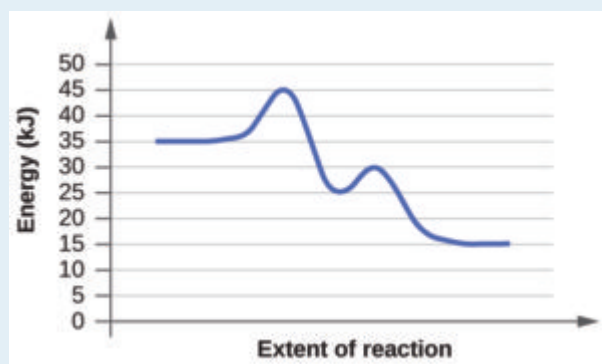


(a)

(b)

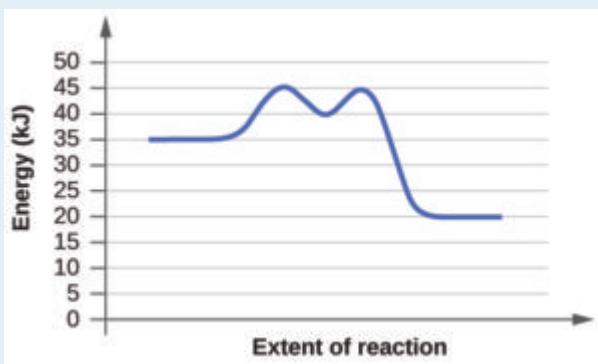


For each of the following reaction diagrams, estimate the activation energy (E_a) of the reaction:



(a)

(b)



Based on the diagrams in Exercise 6, which of the reactions has the fastest rate? Which has the slowest rate?

Based on the diagrams in Exercise 7, which of the reactions has the fastest rate? Which has the slowest rate?

Glossary

heterogeneous catalyst

catalyst present in a different phase from the reactants, furnishing a surface at which a reaction can occur

homogeneous catalyst

catalyst present in the same phase as the reactants

Attribution

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- 1 "The Nobel Prize in Chemistry 1995," Nobel Prize.org, accessed February 18, 2015, http://www.nobelprize.org/nobel_prizes/chemistry/laureates/1995/. ↵

Introduction to Leavening Agents

The word *leavening* in the baking trade is used to describe the source of gas that makes a dough or batter expand in the presence of moisture and heat. Leavening agents are available in different forms, from yeast (the organic leavener) to chemical, mechanical, and physical leaveners. Bakers choose the appropriate type of leavening based on the product they are making.

Attribution

Yeast

Yeast is a microscopic unicellular fungus that multiplies by budding, and under suitable conditions, causes fermentation. Cultivated yeast is widely used in the baking and distilling industries. History tells us that the early Chaldeans, Egyptians, Greeks, and Romans made leavened bread from fermented doughs. This kind of fermentation, however, was not always reliable and easy to control. It was Louis Pasteur, a French scientist who lived in the 19th century, who laid the foundation for the modern commercial production of yeast as we know it today through his research and discoveries regarding the cause and prevention of disease.

Types of Yeast

There are several types of yeast.

Wild Yeast

Wild yeast spores are found floating on dust particles in the air, in flour, on the outside of fruits, etc. Wild yeasts form spores faster than cultivated yeasts, but they are inconsistent and are not satisfactory for controlled fermentation purposes.

Compressed Yeast

Compressed yeast is made by cultivating a select variety, which is known by experiment to produce a yeast that is hardy, consistent, and produces a fermentation with strong enzymatic action. These plants are carefully isolated in a sterile environment free of any other type of yeast and cultivated on a plate containing nutrient agar or gelatin. Wort, a combination of sterilized and purified molasses or malt, nitrogenous matter, and mineral salts is used to supply the food that the growing yeast plants need to make up the bulk of compressed yeast.

After growing to maturity in the fermentation tank, the yeast is separated from the used food or wort by means of centrifugal machines. The yeast is then cooled, filtered, pressed, cut, wrapped, and refrigerated. It is marketed in 454 g (1 lb.) blocks, or in large 20 kg (45 lb.) bags for wholesale bakeries.

Figure 1 illustrates the process of cultivating compressed yeast, and Table 1 summarizes its composition.

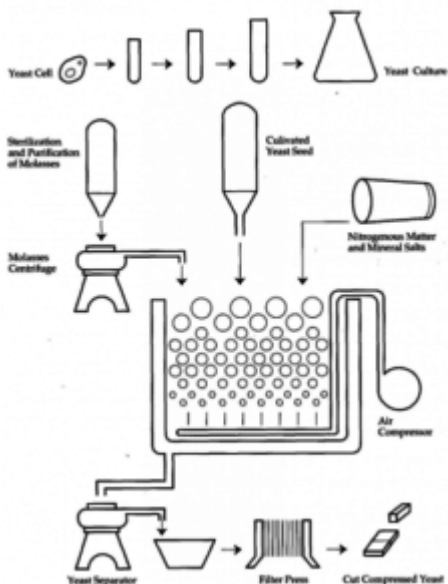


Figure 1 Cultivating compressed yeast

Table 1 Average composition of fresh (compressed) yeast	
Component	Percentage
Water	68% to 73%
Protein	12% to 14%
Fat	0.6% to 0.8%
Carbohydrate	9% to 11%
Mineral Matter	1.7% to 2%

Active Dry Yeast

Active dry yeast is made from a different strain than compressed yeast. The manufacturing process is the same except that the cultivated yeast is mixed with starch or other absorbents and dehydrated. Its production began after World War II, and it was used mainly by the armed forces, homemakers, and in areas where fresh yeast was not readily available.

Even though it is a dry product, it is alive and should be refrigerated below 7°C (45°F) in a closed container for best results. It has a moisture content of about 7%. Storage without refrigeration is satisfactory only for a limited period of time. If no refrigeration is available, the yeast should be kept unopened in a cool, dry place. It should be allowed to warm up to room temperature slowly before being used. Dry yeast must be hydrated for about 15 minutes in water at least four times its weight at a temperature between 42°C and 44°C (108°F and 112°F). The temperature should never be lower than 30°C (86°F), and dry yeast should never be used before it is completely dissolved.

It takes about 550 g (20 oz.) of dry yeast to replace 1 kg (2.2 lb.) of compressed yeast, and for each kilogram of dry yeast used, an additional kilogram of water should be added to the mix. This product is hardly, if ever, used by bakers, having been superseded by instant yeast (see below).

Instant Dry Yeast

Unlike instant active dry yeast that must be dissolved in warm water for proper rehydration and activation,

instant dry yeast can be added to the dough directly, either by:

- Mixing it with the flour before the water is added
- Adding it after all the ingredients have been mixed for one minute

This yeast can be reconstituted. Some manufacturers call for adding it to five times its weight of water at a temperature of 32°C to 38°C (90°F to 100°F). Most formulas suggest a 1:3 ratio when replacing compressed yeast with instant dry. Others vary slightly, with some having a 1:4 ratio. In rich Danish dough, it takes about 400 g (14 oz.), and in bread dough about 250 g to 300 g (9 oz. to 11 oz.) of instant dry yeast to replace 1 kg (2.2 lb.) of compressed yeast. As well, a little extra water is needed to make up for the moisture in compressed yeast. Precise instructions are included with the package; basically, it amounts to the difference between the weight of compressed yeast that would have been used and the amount of dry yeast used.

Instant dry yeast has a moisture content of about 5% and is packed in vacuum pouches. It has a shelf life of about one year at room temperature without any noticeable change in its gassing activity. After the seal is broken, the content turns into a granular powder, which should be refrigerated and used by its best-before date, as noted on the packaging.

Instant dry yeast is especially useful in areas where compressed yeast is not available. However, in any situation, it is practical to use and has the advantages of taking up less space and having a longer shelf life than compressed yeast.

Cream Yeast

Creamy yeast is a soft slurry-type yeast that is used only in large commercial bakeries and is pumped into the dough.

Yeast Food

Yeast food is used in bread production to condition the dough and speed up the fermentation process. It consists of a blend of mineral salts such as calcium salt or ammonium salt and potassium iodate. It has a tightening effect on the gluten and is especially beneficial in dough where soft water is used. The addition of yeast food improves the general appearance and tasting quality of bread. The retail baker does not use it much.

Attribution

The Functions of Yeast

Yeast has two primary functions in fermentation:

- To convert sugar into carbon dioxide gas, which lifts and aerates the dough
- To mellow and condition the gluten of the dough so that it will absorb the increasing gases evenly and hold them at the same time

In baked products, yeast increases the volume and improves the flavor, texture, grain, color, and eating quality. When yeast, water, and flour are mixed together under the right conditions, all the food required for fermentation is present as there is enough soluble protein to build new cells and enough sugar to feed them.

Activity within the yeast cells starts when enzymes in the yeast change complex sugar into invert sugar. The invert sugar is, in turn, absorbed within the yeast cell and converted into carbon dioxide gas and alcohol. Other enzymes in the yeast and flour convert soluble starch into malt sugar, which is converted again by other enzymes into fermentable sugar so that aeration goes on from this continuous production of carbon dioxide.

Proper Handling of Yeast

Compressed yeast ages and weakens gradually even when stored in the refrigerator. Fresh yeast feels moist and firm, and breaks evenly without crumbling. It has a fruity, fresh smell, which changes to a sticky mass with a cheesy odor. It is not always easy to recognize whether or not yeast has lost enough of its strength to affect the fermentation and the eventual outcome of the baked bread, but its working quality definitely depends on the storage conditions, temperature, humidity, and age.

The optimum storage temperature for yeast is -1°C (30°F). At this temperature it is still completely effective for up to two months. Yeast does not freeze at this temperature.

Other guidelines for storing yeast include:

- Rotating it properly and using the older stock first
- Avoiding overheating by spacing it on the shelves in the refrigerator

Yeast needs to breathe, since it is a living fungus. The process is continuous, proceeding slowly in the refrigerator and rapidly at the higher temperature in the shop. When respiration occurs without food, the yeast cells starve, weaken, and gradually die.

Yeast that has been frozen and thawed does not keep and should be used immediately. Freezing temperatures weaken yeast, and thawed yeast cannot be refrozen successfully.

Attribution

Using Yeast in Baking

Many bakers add compressed yeast directly to their dough. A more traditional way to use yeast is to dissolve it in lukewarm water before adding it to the dough. The water should never be higher than 50°C (122°F) because heat destroys yeast cells. In general, salt should not come into direct contact with yeast, as salt dehydrates the yeast. (Table 1 indicates the reaction of yeast at various temperatures.)

It is best to add the dissolved yeast to the flour when the dough is ready for mixing. In this way, the flour is used as a buffer. (Buffers are ingredients that separate or insulate ingredients, which if in too close contact, might start to react prematurely.) In sponges where little or no salt is used, yeast buds quickly and fermentation of the sponge is rapid.

Temperature	Reaction
15°C -20°C (60°F -68°F)	slow reaction
26°C -29°C (80°F -85°F)	normal reaction
32°C -38°C (90°F -100°F)	fast reaction
59°C (138°F)	terminal death point

Never leave compressed yeast out for more than a few minutes. Remove only the amount needed from the refrigerator. Yeast lying around on workbenches at room temperature quickly deteriorates and gives poor results. One solution used by some bakeries to eliminate steps to the fridge is to have a small portable cooler in which to keep the yeast on the bench until it is needed. Yeast must be kept wrapped at all times because if it is exposed to air the edges and the corners will turn brown. This condition is known as air-burn.

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Introduction to Acid- bases

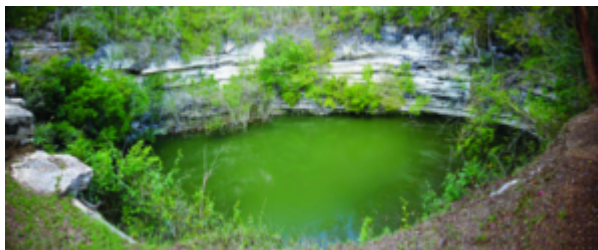


Figure 1 Sinkholes such as this are the result of reactions between acidic groundwaters and basic rock formations, like limestone. (credit: modification of work by Emil Kehnel)

In our bodies, in our homes, and in our industrial society, acids and bases play key roles. Proteins, enzymes, blood, genetic material, and other components of living matter contain both acids and bases. We seem to like the sour taste of acids; we add them to soft drinks, salad dressings, and spices. Many foods, including citrus fruits and some vegetables, contain acids. Cleaners in our homes contain acids or bases. Acids and bases play important roles in the chemical industry. Currently, approximately 36 million metric tons of sulfuric acid are produced annually in the United States alone. Huge quantities of ammonia (8 million tons), urea (10 million tons), and phosphoric acid (10 million tons) are also produced annually.

This chapter will illustrate the chemistry of acid-base reactions and equilibria, and provide you with tools for quantifying the concentrations of acids and bases in solutions.

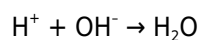
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Brønsted-Lowry Acids and Bases

Acids and bases have been known for a long time. When *Robert Boyle* characterized them in 1680, he noted that acids dissolve many substances, change the color of certain natural dyes (for example, they change litmus from blue to red), and lose these characteristic properties after coming into contact with alkalis (bases). In the eighteenth century, it was recognized that acids have a sour taste, react with limestone to liberate a gaseous substance (now known to be CO_2), and interact with alkalis to form neutral substances. In 1815, *Humphry Davy* contributed greatly to the development of the modern acid-base concept by demonstrating that hydrogen is the essential constituent of acids. Around that same time, *Joseph Louis Gay-Lussac* concluded that acids are substances that can neutralize bases and that these two classes of substances can be defined only in terms of each other. The significance of hydrogen was reemphasized in 1884 when *Carl Axel Arrhenius* defined an acid as a compound that dissolves in water to yield hydrogen cations (now recognized to be hydronium ions) and a base as a compound that dissolves in water to yield hydroxide anions.

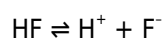
In an earlier chapter on chemical reactions, we defined acids and bases as Arrhenius did: We identified an acid as a compound that dissolves in water to yield hydronium ions (H_3O^+) and a base as a compound that dissolves in water to yield hydroxide ions (OH^-). This definition is not wrong; it is simply limited. Later, we extended the definition of an acid or a base using the more general definition proposed in 1923 by the Danish chemist Johannes Brønsted and the English chemist Thomas Lowry. Their definition centers on the proton, H^+ . A proton is what remains when a normal hydrogen atom, ${}^1_1\text{H}$, loses an electron. A compound that donates a proton to another compound is called a *Brønsted-Lowry acid*, and a compound that accepts a proton is called a *Brønsted-Lowry base*. An acid-base reaction is the transfer of a proton from a proton donor (acid) to a proton acceptor (base). In a subsequent chapter of this text we will introduce the most general model of acid-base behavior introduced by the American chemist G. N. Lewis.

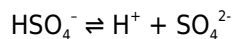
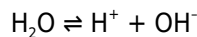
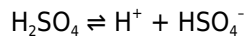
Acids may be compounds such as HCl or H_2SO_4 , organic acids like acetic acid (CH_3COOH) or ascorbic acid (vitamin C), or H_2O . Anions (such as HSO_4^- , H_2PO_4^- , HS^- , and HCO_3^-) and cations (such as H_3O^+ , NH_4^+ , and $[\text{Al}(\text{H}_2\text{O})_6]^{3+}$) may also act as acids. Bases fall into the same three categories. Bases may be neutral molecules (such as H_2O , NH_3 , and CH_3NH_2), anions (such as OH^- , HS^- , HCO_3^- , CO_3^{2-} , F^- , and PO_4^{3-}), or cations (such as $[\text{Al}(\text{H}_2\text{O})_5\text{OH}]^{2+}$). The most familiar bases are ionic compounds such as NaOH and $\text{Ca}(\text{OH})_2$, which contain the hydroxide ion, OH^- . The hydroxide ion in these compounds accepts a proton from acids to form water:



We call the product that remains after an acid donates a proton the *conjugate base* of the acid. This species is a base because it can accept a proton (to re-form the acid):

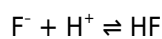
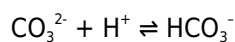
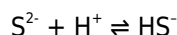
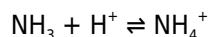
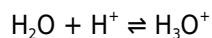
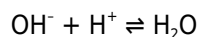
acid \rightleftharpoons proton + conjugate base



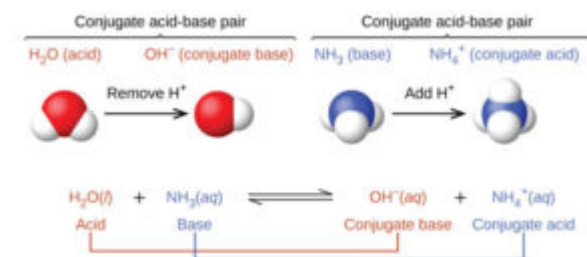


We call the product that results when a base accepts a proton the base's *conjugate acid*. This species is an acid because it can give up a proton (and thus re-form the base):

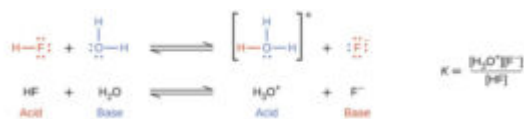
base + proton \rightleftharpoons conjugate acid



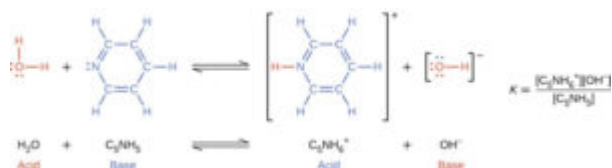
In these two sets of equations, the behaviors of acids as proton donors and bases as proton acceptors are represented in isolation. In reality, all acid-base reactions involve the *transfer* of protons between acids and bases. For example, consider the acid-base reaction that takes place when ammonia is dissolved in water. A water molecule (functioning as an acid) transfers a proton to an ammonia molecule (functioning as a base), yielding the conjugate base of water, OH^- , and the conjugate acid of ammonia, NH_4^+ :



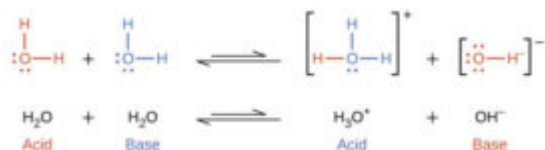
The reaction between a Brønsted-Lowry acid and water is called *acid ionization*. For example, when hydrogen fluoride dissolves in water and ionizes, protons are transferred from hydrogen fluoride molecules to water molecules, yielding hydronium ions and fluoride ions:



When we add a base to water, a *base ionization* reaction occurs in which protons are transferred from water molecules to base molecules. For example, adding ammonia to water yields hydroxide ions and ammonium ions:

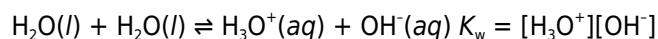


Notice that both these ionization reactions are represented as equilibrium processes. The relative extent to which these acid and base ionization reactions proceed is an important topic treated in a later section of this chapter. In the preceding paragraphs we saw that water can function as either an acid or a base, depending on the nature of the solute dissolved in it. In fact, in pure water or in any aqueous solution, water acts both as an acid and a base. A very small fraction of water molecules donate protons to other water molecules to form hydronium ions and hydroxide ions:



This type of reaction, in which a substance ionizes when one molecule of the substance reacts with another molecule of the same substance, is referred to as *autoionization*.

Pure water undergoes autoionization to a very slight extent. Only about two out of every 10^9 molecules in a sample of pure water are ionized at 25 °C. The equilibrium constant for the ionization of water is called the ion-product constant for water (K_w):



The slight ionization of pure water is reflected in the small value of the equilibrium constant; at 25 °C, K_w has a value of 1.0×10^{-14} . The process is endothermic, and so the extent of ionization and the resulting concentrations of hydronium ion and hydroxide ion increase with temperature. For example, at 100 °C, the value for K_w is about 5.1×10^{-13} , roughly 100-times larger than the value at 25 °C.

Example 1: Ion Concentrations in Pure Water

What are the hydronium ion concentration and the hydroxide ion concentration in pure water at 25 °C?

Solution:

The autoionization of water yields the same number of hydronium and hydroxide ions. Therefore, in pure water, $[\text{H}_3\text{O}^+] = [\text{OH}^-]$. At 25 °C:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = [\text{H}_3\text{O}^+]^2 = [\text{OH}^-]^2 = 1.0 \times 10^{-14}$$

So:

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = \sqrt{1.0 \times 10^{-14}} = 1.0 \times 10^{-7} M$$

The hydronium ion concentration and the hydroxide ion concentration are the same, and we find that both equal $1.0 \times 10^{-7} M$.

Check Your Learning:

The ion product of water at 80 °C is 2.4×10^{-13} . What are the concentrations of hydronium and hydroxide ions in pure water at 80 °C?

Answer:

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 4.9 \times 10^{-7} M$$

It is important to realize that the autoionization equilibrium for water is established in all aqueous solutions.

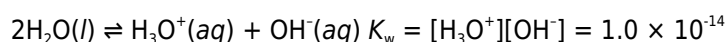
Adding an acid or base to water will not change the position of the equilibrium. Example 2 demonstrates the quantitative aspects of this relation between hydronium and hydroxide ion concentrations.

Example 2: The Inverse Proportionality of $[H_3O^+]$ and $[OH^-]$

A solution of carbon dioxide in water has a hydronium ion concentration of $2.0 \times 10^{-6} M$. What is the concentration of hydroxide ion at $25^\circ C$?

Solution:

We know the value of the ion-product constant for water at $25^\circ C$:



Thus, we can calculate the missing equilibrium concentration.

Rearrangement of the K_w expression yields that $[OH^-]$ is directly proportional to the inverse of $[H_3O^+]$:

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{2.0 \times 10^{-6}} = 5.0 \times 10^{-9}$$

The hydroxide ion concentration in water is reduced to $5.0 \times 10^{-9} M$ as the hydrogen ion concentration increases to $2.0 \times 10^{-6} M$. This is expected from Le Châtelier's principle; the autoionization reaction shifts to the left to reduce the stress of the increased hydronium ion concentration and the $[OH^-]$ is reduced relative to that in pure water.

A check of these concentrations confirms that our arithmetic is correct:

$$K_w = [H_3O^+][OH^-] = (2.0 \times 10^{-6})(5.0 \times 10^{-9}) = 1.0 \times 10^{-14}$$

Check Your Learning:

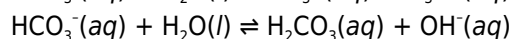
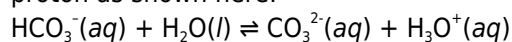
What is the hydronium ion concentration in an aqueous solution with a hydroxide ion concentration of $0.001 M$ at $25^\circ C$?

Answer:

$$[H_3O^+] = 1 \times 10^{-11} M$$

Amphiprotic Species

Like water, many molecules and ions may either gain or lose a proton under the appropriate conditions. Such species are said to be *amphiprotic*. Another term used to describe such species is *amphoteric*, which is a more general term for a species that may act either as an acid or a base by any definition (not just the Brønsted-Lowry one). Consider for example the bicarbonate ion, which may either donate or accept a proton as shown here:

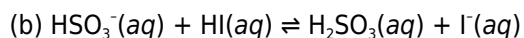
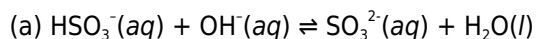


Example 3: Representing the Acid-Base Behavior of an Amphoteric Substance

Write separate equations representing the reaction of HSO_3^-

- (a) as an acid with OH^-
- (b) as a base with HI

Solution:



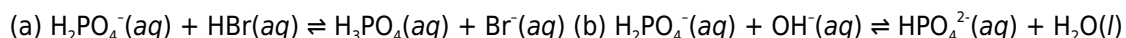
Check Your Learning:

Write separate equations representing the reaction of H_2PO_4^-

(a) as a base with HBr

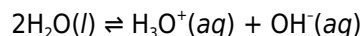
(b) as an acid with OH^-

Answer:



Key Concepts and Summary

A compound that can donate a proton (a hydrogen ion) to another compound is called a Brønsted-Lowry acid. The compound that accepts the proton is called a Brønsted-Lowry base. The species remaining after a Brønsted-Lowry acid has lost a proton is the conjugate base of the acid. The species formed when a Brønsted-Lowry base gains a proton is the conjugate acid of the base. Thus, an acid-base reaction occurs when a proton is transferred from an acid to a base, with formation of the conjugate base of the reactant acid and formation of the conjugate acid of the reactant base. Amphiprotic species can act as both proton donors and proton acceptors. Water is the most important amphiprotic species. It can form both the hydronium ion, H_3O^+ , and the hydroxide ion, OH^- when it undergoes autoionization:



The ion product of water, K_w is the equilibrium constant for the autoionization reaction:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 2.0 \times 10^{-14} \text{ at } 25^\circ$$

Key Equations

- $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ (at } 25^\circ\text{C)}$

Brønsted-Lowry Acids and Bases Exercises

Write equations that show NH_3 as both a conjugate acid and a conjugate base.

Write equations that show H_2PO_4^- acting both as an acid and as a base.

Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry acid:

(a) H_3O^+

- (b) HCl
- (c) NH₃
- (d) CH₃CO₂H
- (e) NH₄⁺
- (f) HSO₄⁻

Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry acid:

- (a) HNO₃
- (b) PH₄⁺
- (c) H₂S
- (d) CH₃CH₂COOH
- (e) H₂PO₄⁻
- (f) HS⁻

Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry base:

- (a) H₂O
- (b) OH⁻
- (c) NH₃
- (d) CN⁻
- (e) S²⁻
- (f) H₂PO₄⁻

Show by suitable net ionic equations that each of the following species can act as a Brønsted-Lowry base:

- (a) HS⁻
- (b) PO₄³⁻
- (c) NH₂⁻
- (d) C₂H₅OH
- (e) O²⁻
- (f) H₂PO₄⁻

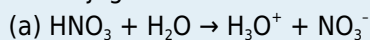
What is the conjugate acid of each of the following? What is the conjugate base of each?

- (a) OH⁻
- (b) H₂O
- (c) HCO₃⁻
- (d) NH₃
- (e) HSO₄⁻
- (f) H₂O₂
- (g) HS⁻
- (h) H₅N₂⁺

What is the conjugate acid of each of the following? What is the conjugate base of each?

- (a) H₂S
- (b) H₂PO₄⁻
- (c) PH₃
- (d) HS⁻
- (e) HSO₃⁻
- (f) H₃O₂⁺
- (g) H₄N₂
- (h) CH₃OH

Identify and label the Brønsted-Lowry acid, its conjugate base, the Brønsted-Lowry base, and its conjugate acid in each of the following equations:



- (b) $\text{CN}^- + \text{H}_2\text{O} \rightarrow \text{HCN} + \text{OH}^-$
 (c) $\text{H}_2\text{SO}_4 + \text{Cl}^- \rightarrow \text{HCl} + \text{HSO}_4^-$
 (d) $\text{HSO}_4^- + \text{OH}^- \rightarrow \text{SO}_4^{2-} + \text{H}_2\text{O}$
 (e) $\text{O}^{2-} + \text{H}_2\text{O} \rightarrow 2\text{OH}^-$
 (f) $[\text{Cu}(\text{H}_2\text{O})_3(\text{OH})]^+ + [\text{Al}(\text{H}_2\text{O})_6]^{3+} \rightarrow [\text{Cu}(\text{H}_2\text{O})_4]^{2+} + [\text{Al}(\text{H}_2\text{O})_5(\text{OH})]^{2+}$
 (g) $\text{H}_2\text{S} + \text{NH}_2^- \rightarrow \text{HS}^- + \text{NH}_3$

Identify and label the Brønsted-Lowry acid, its conjugate base, the Brønsted-Lowry base, and its conjugate acid in each of the following equations:

- (a) $\text{NO}_2^- + \text{H}_2\text{O} \rightarrow \text{HNO}_2 + \text{OH}^-$
 (b) $\text{HBr} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Br}^-$
 (c) $\text{HS}^- + \text{H}_2\text{O} \rightarrow \text{H}_2\text{S} + \text{OH}^-$
 (d) $\text{H}_2\text{PO}_4^- + \text{OH}^- \rightarrow \text{HPO}_4^{2-} + \text{H}_2\text{O}$
 (e) $\text{H}_2\text{PO}_4^- + \text{HCl} \rightarrow \text{H}_3\text{PO}_4 + \text{Cl}^-$
 (f) $[\text{Fe}(\text{H}_2\text{O})_5(\text{OH})]^{2+} + [\text{Al}(\text{H}_2\text{O})_6]^{3+} \rightarrow [\text{Fe}(\text{H}_2\text{O})_6]^{3+} + [\text{Al}(\text{H}_2\text{O})_5(\text{OH})]^{2+}$
 (g) $\text{CH}_3\text{OH} + \text{H}^+ \rightarrow \text{CH}_3\text{O}^+ + \text{H}_2$

What are amphiprotic species? Illustrate with suitable equations.

State which of the following species are amphiprotic and write chemical equations illustrating the amphiprotic character of these species:

- (a) H_2O
 (b) H_2PO_4^-
 (c) S^{2-}
 (d) CO_3^{2-}
 (e) HSO_4^-

State which of the following species are amphiprotic and write chemical equations illustrating the amphiprotic character of these species.

- (a) NH_3
 (b) HPO_4^-
 (c) Br^-
 (d) NH_4^+
 (e) ASO_4^{3-}

Is the self ionization of water endothermic or exothermic? The ionization constant for water (K_w) is 2.9×10^{-14} at 40 °C and 9.6×10^{-14} at 60 °C.

Glossary

acid ionization

reaction involving the transfer of a proton from an acid to water, yielding hydronium ions and the conjugate base of the acid

amphiprotic

species that may either gain or lose a proton in a reaction

amphoteric

species that can act as either an acid or a base

autoionization

reaction between identical species yielding ionic products; for water, this reaction involves transfer of protons to yield hydronium and hydroxide ions

base ionization

reaction involving the transfer of a proton from water to a base, yielding hydroxide ions and the conjugate acid of the base

Brønsted-Lowry acid

proton donor

Brønsted-Lowry base

proton acceptor

conjugate acid

substance formed when a base gains a proton

conjugate base

substance formed when an acid loses a proton

ion-product constant for water (K_w)

equilibrium constant for the autoionization of water

Attribution

pH and pOH

As discussed earlier, hydronium and hydroxide ions are present both in pure water and in all aqueous solutions, and their concentrations are inversely proportional as determined by the ion product of water (K_w). The concentrations of these ions in a solution are often critical determinants of the solution's properties and the chemical behaviors of its other solutes, and specific vocabulary has been developed to describe these concentrations in relative terms. A solution is neutral if it contains equal concentrations of hydronium and hydroxide ions; acidic if it contains a greater concentration of hydronium ions than hydroxide ions; and basic if it contains a lesser concentration of hydronium ions than hydroxide ions.

A common means of expressing quantities, the values of which may span many orders of magnitude, is to use a logarithmic scale. One such scale that is very popular for chemical concentrations and equilibrium constants is based on the p-function, defined as shown where "X" is the quantity of interest and "log" is the base-10 logarithm:

$$pX = -\log X$$

The pH of a solution is therefore defined as shown here, where $[H_3O^+]$ is the molar concentration of hydronium ion in the solution:

$$pH = -\log [H_3O^+]$$

Rearranging this equation to isolate the hydronium ion molarity yields the equivalent expression:

$$[H_3O^+] = 10^{-pH}$$

Likewise, the hydroxide ion molarity may be expressed as a p-function, or pOH:

$$pOH = -\log [OH^-]$$

or

$$[OH^-] = 10^{-pOH}$$

Finally, the relation between these two ion concentration expressed as p-functions is easily derived from the K_w expression:

$$K_w = [H_3O^+] [OH^-]$$

$$-\log K_w = -\log([H_3O^+] [OH^-]) = -\log [H_3O^+] + -\log [OH^-]$$

$$pK_w = pH + pOH$$

At 25 °C, the value of K_w is 1.0×10^{-14} , and so:

$$14.00 = \text{pH} + \text{pOH}$$

As was shown in the [previous section](#), the hydronium ion molarity in pure water (or any neutral solution) is $1.0 \times 10^{-7} \text{ M}$ at 25 °C. The pH and pOH of a neutral solution at this temperature are therefore:

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (1.0 \times 10^{-7}) = 7.00$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log (1.0 \times 10^{-7}) = 7.00$$

And so, *at this temperature*, acidic solutions are those with hydronium ion molarities greater than $1.0 \times 10^{-7} \text{ M}$ and hydroxide ion molarities less than $1.0 \times 10^{-7} \text{ M}$ (corresponding to pH values less than 7.00 and pOH values greater than 7.00). Basic solutions are those with hydronium ion molarities less than $1.0 \times 10^{-7} \text{ M}$ and hydroxide ion molarities greater than $1.0 \times 10^{-7} \text{ M}$ (corresponding to pH values greater than 7.00 and pOH values less than 7.00).

Since the autoionization constant K_w is temperature dependent, these correlations between pH values and the acidic/neutral/basic adjectives will be different at temperatures other than 25 °C. For example, the “Check Your Learning” exercise in the [previous section](#) showed the hydronium molarity of pure water at 80 °C is $4.9 \times 10^{-7} \text{ M}$, which corresponds to pH and pOH values of:

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (4.9 \times 10^{-7}) = 6.31$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log (4.9 \times 10^{-7}) = 6.31$$

At this temperature, then, neutral solutions exhibit $\text{pH} = \text{pOH} = 6.31$, acidic solutions exhibit pH less than 6.31 and pOH greater than 6.31, whereas basic solutions exhibit pH greater than 6.31 and pOH less than 6.31. This distinction can be important when studying certain processes that occur at nonstandard temperatures, such as enzyme reactions in warm-blooded organisms. Unless otherwise noted, references to pH values are presumed to be those at standard temperature (25 °C) (Table 1).

Summary of Relations for Acidic, Basic and Neutral Solutions		
Classification	Relative Ion Concentrations	pH at 25 °C
acidic	$[\text{H}_3\text{O}^+] > [\text{OH}^-]$	$\text{pH} < 7$
neutral	$[\text{H}_3\text{O}^+] = [\text{OH}^-]$	$\text{pH} = 7$
basic	$[\text{H}_3\text{O}^+] < [\text{OH}^-]$	$\text{pH} > 7$

Figure 1 shows the relationships between $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, pH, and pOH, and gives values for these properties at standard temperatures for some common substances.

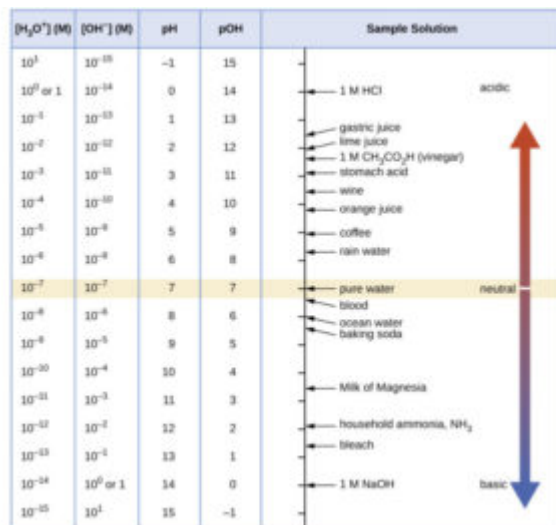


Figure 1 The pH and pOH scales represent concentrations of $[H_3O^+]$ and OH^- , respectively. The pH and pOH values of some common substances at standard temperature (25 °C) are shown in this chart.

Example 1: Calculation of pH from $[H_3O^+]$

What is the pH of stomach acid, a solution of HCl with a hydronium ion concentration of $1.2 \times 10^{-3}M$?

Solution:

$$\begin{aligned} \text{pH} &= -\log [H_3O^+] \\ &= -\log(1.2 \times 10^{-3}) \\ &= -(-2.92) = 2.92 \end{aligned}$$

(The use of logarithms is explained in Appendix B. Recall that, as we have done here, when taking the log of a value, keep as many decimal places in the result as there are significant figures in the value.)

Check Your Learning:

Water exposed to air contains carbonic acid, H_2CO_3 , due to the reaction between carbon dioxide and water: $CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$. Air-saturated water has a hydronium ion concentration caused by the dissolved CO_2 of $2.0 \times 10^{-6}M$, about 20-times larger than that of pure water. Calculate the pH of the solution at 25 °C.

Answer:

5.70

Example 2: Calculation of Hydronium Ion Concentration from pH

Calculate the hydronium ion concentration of blood, the pH of which is 7.3 (slightly alkaline).

Solution:

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = 7.3$$

$$\log[\text{H}_3\text{O}^+] = -7.3$$

$$[\text{H}_3\text{O}^+] = 10^{-7.3} \text{ or } [\text{H}_3\text{O}^+] = \text{antilog of } -7.3$$

$$[\text{H}_3\text{O}^+] = 5 \times 10^{-8} \text{ M}$$

(On a calculator take the antilog, or the “inverse” log, of -7.3 , or calculate $10^{-7.3}$.)

Check Your Learning:

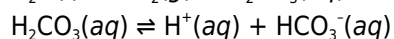
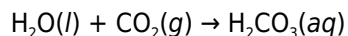
Calculate the hydronium ion concentration of a solution with a pH of -1.07 .

Answer:

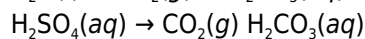
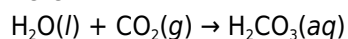
12 M

Note: ENVIRONMENTAL SCIENCE

Normal rainwater has a pH between 5 and 6 due to the presence of dissolved CO_2 which forms carbonic acid:



Acid rain is rainwater that has a pH of less than 5, due to a variety of nonmetal oxides, including CO_2 , SO_2 , SO_3 , NO , and NO_2 being dissolved in the water and reacting with it to form not only carbonic acid, but sulfuric acid and nitric acid. The formation and subsequent ionization of sulfuric acid are shown here:



Carbon dioxide is naturally present in the atmosphere because we and most other organisms produce it as a waste product of metabolism. Carbon dioxide is also formed when fires release carbon stored in vegetation or when we burn wood or fossil fuels. Sulfur trioxide in the atmosphere is naturally produced by volcanic activity, but it also stems from burning fossil fuels, which have traces of sulfur, and from the process of “roasting” ores of metal sulfides in metal-refining processes. Oxides of nitrogen are formed in internal combustion engines where the high temperatures make it possible for the nitrogen and oxygen in air to chemically combine.

Acid rain is a particular problem in industrial areas where the products of combustion and smelting are released into the air without being stripped of sulfur and nitrogen oxides. In North America and Europe until the 1980s, it was responsible for the destruction of forests and freshwater lakes, when the acidity of the rain actually killed trees, damaged soil, and made lakes uninhabitable for all but the most acid-tolerant species. Acid rain also corrodes statuary and building facades that are made of marble and limestone (Figure 2). Regulations limiting the amount of sulfur and nitrogen oxides that can be released into the atmosphere by industry and automobiles have reduced the severity of acid damage to both natural and manmade environments in North America and Europe. It is now a growing problem in industrial areas of China and India.

For further information on acid rain, visit this [website](#) hosted by the US Environmental Protection Agency.



Figure 2 (a) Acid rain makes trees more susceptible to drought and insect infestation, and depletes nutrients in the soil. (b) It also corrodes statues that are carved from marble or limestone. (credit a: modification of work by Chris M Morris; credit b: modification of work by "Eden, Janine and Jim"/Flickr)

Example 3: Calculation of pOH

What are the pOH and the pH of a 0.0125-M solution of potassium hydroxide, KOH?

Solution Potassium hydroxide is a highly soluble ionic compound and completely dissociates when dissolved in dilute solution, yielding $[\text{OH}^-] = 0.0125 \text{ M}$:

$$\begin{aligned} \text{pOH} &= -\log[\text{OH}^-] = -\log 0.0125 \\ &= -(-1.903) = 1.903 \end{aligned}$$

The pH can be found from the pOH:

$$\text{pH} + \text{pOH} = 14.00$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 1.903 = 12.10$$

Check Your Learning:

The hydronium ion concentration of vinegar is approximately $4 \times 10^{-3} \text{ M}$. What are the corresponding values of pOH and pH?

Answer:

$$\text{pOH} = 11.6, \text{pH} = 2.4$$

The acidity of a solution is typically assessed experimentally by measurement of its pH. The pOH of a solution is not usually measured, as it is easily calculated from an experimentally determined pH value. The pH of a solution can be directly measured using a pH meter (Figure 3).

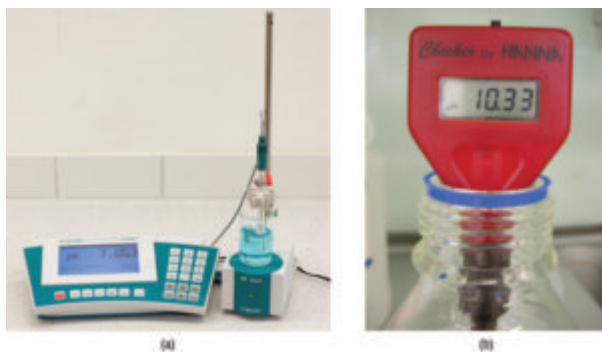


Figure 3 (a) A research-grade pH meter used in a laboratory can have a resolution of 0.001 pH units, an accuracy of ± 0.002 pH units, and may cost in excess of \$1000. (b) A portable pH meter has lower resolution (0.01 pH units), lower accuracy (± 0.2 pH units), and a far lower

price tag. (credit b: modification of work by Jacopo Werther)

The pH of a solution may also be visually estimated using colored indicators (Figure 4).

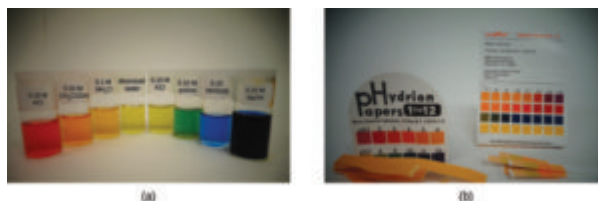


Figure 4 (a) A universal indicator assumes a different color in solutions of different pH values. Thus, it can be added to a solution to determine the pH of the solution. The eight vials each contain a universal indicator and 0.1-M solutions of progressively weaker acids: HCl (pH = 1), CH₃CO₂H (pH = 3), and NH₄Cl (pH = 5), deionized water, a neutral substance (pH = 7); and 0.1-M solutions of the progressively stronger bases: KCl (pH = 7), aniline, C₆H₅NH₂ (pH = 9), NH₃ (pH = 11), and NaOH (pH = 13). (b) pH paper contains a mixture of indicators that give different colors in solutions of differing pH values. (credit: modification of work by Sahar Atwa)

Key Concepts and Summary

The concentration of hydronium ion in a solution of an acid in water is greater than $1.0 \times 10^{-7} M$ at 25°C. The concentration of hydroxide ion in a solution of a base in water is greater than $1.0 \times 10^{-7} M$ at 25°C. The concentration of H₃O⁺ in a solution can be expressed as the pOH of the solution: $\text{pOH} = -\log[\text{OH}^-]$. In pure water, pH = 7.00 and pOH = 7.00

Key Equations

- $\text{pH} = -\log[\text{H}_3\text{O}^+]$
- $\text{pOH} = -\log[\text{OH}^-]$
- $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$
- $[\text{OH}^-] = 10^{-\text{pOH}}$
- $\text{pH} + \text{pOH} = \text{p}K_w = 14.00$ at 25 °C

pH and pOH Exercises

Explain why a sample of pure water at 40 °C is neutral even though $[\text{H}_3\text{O}^+] = 1.7 \times 10^{-7} M$. K_w is 2.9×10^{-14} at 40 °C.

The ionization constant for water (K_w) is 2.9×10^{-14} at 40 °C. Calculate $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, pH, and pOH for pure water at 40 °C.

The ionization constant for water (K_w) is 9.614×10^{-14} at 60 °C. Calculate $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, pH, and pOH for pure water at 60 °C.

Calculate the pH and the pOH of each of the following solutions at 25 °C for which the

substances ionize completely:

(a) 0.200 M HCl

(b) 0.0143 M NaOH

(c) 3.0 M HNO₃

(d) 0.0031 M Ca(OH)₂

Calculate the pH and the pOH of each of the following solutions at 25 °C for which the substances ionize completely:

(a) 0.000259 M HClO₄

(b) 0.21 M NaOH

(c) 0.000071 M Ba(OH)₂

(d) 2.5 M KOH

What are the pH and pOH of a solution of 2.0 M HCl, which ionizes completely?

What are the pH and pOH of a solution of 2.0 M HCl, which ionizes completely?

What are the hydronium and hydroxide ion concentrations in a solution whose pH is 6.52?

Calculate the hydrogen ion concentration and the hydroxide ion concentration in wine from its pH. See Figure 1 for useful information.

Calculate the hydronium ion concentration and the hydroxide ion concentration in lime juice from its pH. See Figure 1 for useful information.

The hydronium ion concentration in a sample of rainwater is found to be 1.7×10^{-6} M at 25 °C.

What is the concentration of hydroxide ions in the rainwater?

The hydroxide ion concentration in household ammonia is 3.2×10^{-3} M at 25 °C. What is the concentration of hydronium ions in the solution?

Glossary

acidic

describes a solution in which $[\text{H}_3\text{O}^+] > [\text{OH}^-]$

basic

describes a solution in which $[\text{H}_3\text{O}^+] < [\text{OH}^-]$

neutral

describes a solution in which $[\text{H}_3\text{O}^+] = [\text{OH}^-]$

pH

logarithmic measure of the concentration of hydronium ions in a solution

pOH

logarithmic measure of the concentration of hydroxide ions in a solution

Attribution

Baking Powder

Baking powder is a dependable, high-quality chemical leavener. To be effective, all baking powders rely on the reaction between one or more acids on sodium bicarbonate to produce carbon dioxide gas. Just as with yeast leavening, the presence of carbon dioxide gas creates air bubbles that cause the product to rise.

There are two main types of baking powders available on the market:

- Continuous or single-action baking powder
- Double- or multiple-action baking powder

The difference between continuous- and double-action baking powders is simply the rate of reaction:

- Continuous-action baking powder uses one acid, which continuously reacts with the soda to release gas steadily throughout the baking process until all the gassing power is spent.
- Double-action baking powder contains two different acids, which react with soda at different stages of the baking process. One acid reacts to give off a small amount of gas at low temperature, and the other major acid reacts at baking temperatures to give off the bulk of the gas.

The Leavening Mechanism of Baking Powder

Before baking, approximately 15% of the CO₂ gas is released in the cold stage. Eighty-five percent of the CO₂ gas is released in the oven starting at approximately 40°C (105°F). Some leavening power is apparently lost in the cold stage, but there is usually still adequate gassing power in the remaining portion.

When the baking powder is activated through moisture and heat, the gas works its way into the many cells created by the mixing or creaming of the batter and starts to expand them. This process comes to a halt when the starch gelatinizes and the cells become rigid. This starts at about 60°C (140°F) and is more or less complete at around 75°C (167°F). After this point, some gas may still be created, but it simply escapes through the porous structure of the product.

Using Baking Powder

For even distribution throughout the batter, baking powder should be sifted with the flour or other dry ingredients. For most cakes, about 5% baking powder to the weight of the flour produces an optimum result. Accurate scaling is important, since a little too much may cause the product to collapse. (Note this is unlike yeast, where an “overdose” will usually simply cause a more rapid rise.)

Attribution

Sodium Bicarbonate

When sodium bicarbonate (baking soda) is moistened and heated, it releases carbon dioxide gas. If it is moistened and heated in the presence of sufficient acid, it will release twice as much gas as if it is moistened and heated without the presence of an acid.

Slightly acidic ingredients provide the mix with some of the necessary acids for the release of carbon dioxide gas. Examples are:

- Honey
- Molasses
- Ginger
- Cocoa
- Bran

For this reason, some of the mixes contain baking powder only while others contain a combination of baking powder and baking soda. If an excessive amount of baking soda is used in a cake batter without the presence of sufficient acid, the normally white cake crumb will have a yellowish-brown color and a strong undesirable smell of soda.

The gas evolves very fast at the beginning of baking when the **pH** level is still on the acidic side (pH of around 5 to 6). Once the soda neutralizes the acid, the dough or batter quickly becomes **alkaline** and the release of gas is reduced. Mixes and doughs leavened with baking soda must be handled without delay, or the release of the gas may be almost exhausted before the product reaches the oven.

The darker color of the crumb found on the bottom half of a cake or muffins is caused by the partial dehydration of the batter that is heated first during baking. In spiced honey cookies and gingerbread, baking soda is used alone to give them quick color during baking and yet keep the products soft.

In chocolate cakes, baking soda is used in conjunction with baking powder to keep the pH at a desirable level. However, it is important to know whether the cocoa powder you are using is natural or treated by the Dutch process. In the Dutch process, some of the acid in the cocoa is already neutralized, and there is less left for the release of gas in the mix. This means more baking powder and less baking soda is used.

Baking soda in a chocolate mix not only counteracts the acid content in the baked cake but also improves the grain and color of the cake. A darker and richer chocolate color is produced if the acid level is sufficient to release all the carbon dioxide gas. On the other hand, the reddish, coarse, open-grained crumb in devil's food cake is the result of using baking soda as the principal leavening agent.

The level of baking soda depends on the nature of the product and on the other ingredients in the formula. Cookies, for example, with high levels of fat and sugar, do not require much, if any, leavening.

Table 1 provides the recommended amounts of baking soda for different products. Note that the percentages appear small compared to the 5% level of baking powder suggested because baking powder contains both an acid agent and a leavening agent.

Table 1 Recommended amounts of baking soda	
Product	Amount of Baking Soda (% of flour weight)
Cookies	0.4-0.6

Cakes	0.5-1.0
Cake doughnuts	0.7-1.0
Pancakes	1.4-2.0

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Ammonium Bicarbonate

Ammonium bicarbonate is a white crystalline powder used in flat, spiced cookies, such as gingerbreads, and in eclair paste. It must be dissolved in the cold liquid portion of the batter. At room temperature, decomposition of CO_2 in the batter is minimal. When heated to approximately 60°C (140°F) decomposition is more noticeable, and at oven temperature, decomposition takes place in a very short time. Ammonium bicarbonate should only be used in low moisture-containing products that are not dense. Providing that these conditions are met, there will be no taste and odor remaining from the ammonium.

Attribution

Water Hardness and pH

Effects on Baking

Most municipal supplies of water contain chlorine, which is used to ensure the purity of the water. Some cities add fluoride to their water supply to stop tooth decay. Neither chlorine nor fluoride is present in large enough quantities to affect dough in any way. In addition, most municipal water is treated to reduce excessive acidity, since this could be corrosive for the water lines. It is therefore unlikely that bakers using municipal water need to be concerned about extremely acidic water.

Soft water is another matter, as it can lead to sticky dough. An addition of yeast food, or a reduction in dough water, will help. Alkaline water tends to tighten the dough and retard fermentation, since enzymes work best in slightly acidic dough.

If there is a possibility of water problems, a sample should be forwarded to a laboratory for a complete analysis.

Attribution

Understanding Ingredients: Dairy Products

Introduction to Dairy Products

Milk and milk products are some of our oldest and best-known natural foods. In baking, milk is used fresh, condensed, powdered, skimmed, or whole. The great bulk, weight, and perishability of fresh milk plus the expense of refrigeration makes it a relatively high-cost ingredient, and for this reason, most modern bakeries use non-fat powdered milk or buttermilk powder.

Over the past 20 years, there has been a trend to lower fat content in dairy products. This reflects the high caloric value of milk fat, and also is compatible with the trend to leaner, healthier nutrition. These “low-fat” products often have the fat replaced with sugars, so care must be taken in substituting these ingredients in a recipe. For bakers, this trend has not meant any great changes in formulas: a 35% milk fat or a 15% cream cheese product usually works equally well in a cheesecake. Some pastry chefs find lowering the richness in pastries and plated desserts can make them more enjoyable, especially after a large meal.

Table 1 provides the nutritional properties of milk products.

Table 1 Nutritional properties of milk products (per 100 g)				
	Whole Milk (3.5% milk fat)	Skim Milk (0.1% milk fat)	Coffee Cream (18% milk fat)	Heavy or Whipping Cream (36% milk fat)
Protein	3.22 g	3.37 g	3 g	2 g
Fat	3.25 g	0.08 g	19 g	37 g
Cholesterol	10 mg	2 mg	66 mg	137 mg
Potassium	143 mg	156 mg	122 mg	75 mg
Calcium	113 mg	125 mg	96 mg	65 mg
Magnesium	10 mg	11 mg	9 mg	7 mg
Sodium	40 mg	42 mg	40 mg	40 mg
Vitamin A (IU)	102 IU	204 IU	656 IU	1470 IU

Note: Besides the elements shown in Table 1, all dairy products contain vitamin B-complex.

IU = International Units, a term used in nutritional measurement
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Milk

Homogenized milk is fresh milk in which the fat particles are so finely divided and emulsified mechanically that the milk fat cannot separate on standing. The milk fat is forced into tiny droplets. As soon as the droplets form, milk proteins and emulsifiers form a protective film around each one, preventing the fat from reuniting. The tiny droplets stay suspended indefinitely, and milk fat no longer separates and rises to the top as a cream layer. In other words, homogenized dairy products are stable emulsions of fat droplets suspended in milk. It is also said that homogenized milk is more readily digestible.

Pasteurization of milk was developed in 1859 by the French chemist Louis Pasteur. One method of pasteurization is to heat milk to above 71°C (160°F), maintain it at this temperature for a set time, then cool it immediately to 10°C (50°F) or lower. This kills all harmful bacteria that carry the potential threat of bovine tuberculosis and fever from cows to humans.

The two main types of pasteurization used today are high-temperature, short-time (HTST, also known as “flash”) and higher-heat, shorter time (HHST). Ultra-high-temperature (UHT) processing is also used.

- High-temperature, short-time (HTST) pasteurization is done by heating milk to 72°C (161°F) for 15 seconds. Milk simply labelled “pasteurized” is usually treated with the HTST method.
- Higher-heat, shorter time (HHST) milk and milk products are pasteurized by applying heat continuously, generally above 100°C (212°F) for such time to extend the shelf life of the product under refrigerated conditions. This type of heat process can be used to produce dairy products with extended shelf life (ESL).
- Ultra-high-temperature (UHT) processing holds the milk at a temperature of 140°C (284°F) for four seconds. During UHT processing, milk is sterilized rather than pasteurized. This process allows milk or juice to be stored several months without refrigeration. The process is achieved by spraying the milk or juice through a nozzle into a chamber that is filled with high-temperature steam under pressure. After the temperature reaches 140°C (284°F) the fluid is cooled instantly in a vacuum chamber and packed in a pre-sterilized, airtight container. Milk labelled UHT has been treated in this way.

For more information on pasteurization, visit the [International Dairy Foods Association](#).

Attribution

Milk Products ADD US

Cream

The usual minimum standard for cream is 10% fat content, though it ranges between 10% and 18%. Cream in this range may be sold as half and half, coffee cream, or table cream.

Whipping cream is about 32% to 36% in milk fat content. Cream with 36% or higher is called heavy cream. This percentage of fat is not a mandated standard; much less than this and the cream simply will not whip. For best whipping results, the cream should be 48 to 60 hours old and be cold. A stabilizer, some sugar, and flavour may be added during whipping. Before adding stabilizer, check the ingredients on the carton; some whipping creams nowadays have added agents such as carrageenan, in which case an additional stabilizer may not be necessary.

Canadian cream definitions are similar to those used in the United States, except for that of “light cream.” In Canada, what the U.S. calls light cream is referred to most commonly as half and half. In Canada, “light cream” is low-fat cream, usually with **5% to 6% fat**. You can make your own light cream by blending milk with half-and-half.

In Quebec, country cream is sold, which contains 15% milk fat. If you are using [a recipe that calls for country cream](#), you may substitute 18% cream.

If you have recipes from the UK, you might see references to *double cream*. This is cream with about 48% milk fat, which is not readily available in Canada, except in some specialty stores. Use whipping cream or heavy cream instead.

Table 1 lists some of the common cream types and their uses.

Name	Minimum Milk Fat	Additional Definition	Main Uses
Whipping cream	32%	Heavy cream has at least 36% milk fat	Whips well, can be piped; custards, cream fillings, confectionary products
Table cream	18%	Coffee cream	Added to coffee, poured over puddings, used in sauces
Half-and-half	10%-12%	Cereal cream	Added to coffee; custards and ice cream mixes
Light cream	5%-10%		Added to coffee

Buttermilk

There are two methods to produce buttermilk:

- Inoculating milk with a specific culture to sour it
- Churning milk and separating the liquid left over from the butter

The second method is where buttermilk gets its name, but today, most of what is commonly called buttermilk is the first type. Buttermilk has a higher acid content than regular milk (pH of 4.6 compared with milk's pH of 6.6).

The fermented dairy product known as *cultured buttermilk* is produced from cow's milk and has a characteristically sour taste caused by lactic acid bacteria. This variant is made using one of two species of bacteria — either *Lactococcus lactis* or *Lactobacillus bulgaricus*, which creates more tartness in certain recipes.

The acid in buttermilk reacts with the sodium bicarbonate (baking soda) to produce carbon dioxide, which acts as the leavening agent.

Sour Cream

Sour cream is made from cream soured by adding lactic acids and thickened naturally or by processing. Milk fat content may vary from 5.5% to 14%. The lactic acid causes the proteins in sour cream to coagulate to a gelled consistency; gums and starches may be added to further thicken it. The added gums and starches also keep the liquid whey in sour cream from separating.

Use sour cream in cheesecakes, coffee cakes, and pastry doughs. Low-fat and fat-free sour cream are available. Low-fat sour cream, which is essentially cultured half-and-half or light cream (and usually contains 7% to 10% milk fat), is often satisfactory as a substitute for regular sour cream in baking. These products are higher in moisture and less rich in flavor than regular sour cream.

Crème Fraîche

Crème fraîche (fresh cream) is a soured cream containing 30% to 45% milk fat and having a pH of around 4.5. It is soured with bacterial culture. Traditionally it is made by setting unpasteurized milk into a pan at room temperature, allowing the cream to rise to the top. After about 12 hours, the cream is skimmed off. During that time, natural bacteria in the unpasteurized milk ripens the cream, turning it into a mildly sour, thickened product.

An effective substitute can be made by adding a small amount of cultured buttermilk or sour cream to whipping cream and allowing it to stand in a warm spot for 10 hours or more before refrigerating. As the cream ripens from the growth of the lactic acid bacteria, it thickens and develops a sour flavour. This product is similar to sour cream, but it has a higher milk fat content.

Milk Substitutes

Milk substitutes are becoming increasingly popular as replacements for straight skim milk powders. Innumerable replacement blends are available to the baker. Their protein contents range from 11% to 40%; some are wet, some are dry-blended. Product types vary from all dairy to mostly cereal. All-dairy blends range from mostly dry skim milk to mostly whey. A popular blend is whey mixed with 40% soy flour solids and a small quantity of sodium hydroxide to neutralize the whey acidity.

Dough consistency may be a little softer if the milk in the replacement blend exceeds 3%, and this could dictate the need to increase dough mixing by at least half a minute. However, absorption and formula changes are seldom necessary when switching from dry milk to a blend, or from a blend to a blend.

For nutritional labelling, or when using a blend in a non-standardized product that must carry an itemized ingredient label, all blend components must be listed in their proper order on the label.

The Canadian Food Inspection Agency defines modified milk ingredients as any of the following in liquid, concentrated, dry, frozen, or reconstituted form:

- Calcium-reduced skim milk
- Casein: This a protein in milk and is used as a binding agent. Caseins are also used in wax to shine fruits and vegetables, as an adhesive, and to fortify bread. Caseins contain common amino acids.
- Caseinate: This protein is derived from skim milk. Bodybuilders sometimes take powder enriched

with calcium caseinate because it releases proteins at an even, measured pace.

- Cultured milk products: These are milk products that have been altered through controlled fermentation, including yogurt, sour cream, and cultured buttermilk.
- Milk serum proteins
- Ultra-filtered milk: The Canadian Food and Drug Regulations define this type of milk as that which “has been subjected to a process in which it is passed over one or more semi-permeable membranes to partially remove water, lactose, minerals, and water-soluble vitamins without altering the whey protein-to-casein ratio and that results in a liquid product.”
- Whey: This is serum by-product created in the manufacture of cheese.
- Whey butter: Typically oily in composition, whey butter is made from cream separated from whey.
- Whey cream: This is cream skimmed from whey, sometimes used as a substitute for sweet cream and butter.
- Any component of milk that has been altered from the form in which it is found in milk.

Milk Powder

Milk powder is available in several different forms: whole milk, skim milk (non-fat dry milk), buttermilk, or whey. They are all processed similarly: the product is first pasteurized, then concentrated with an evaporator, and finally dried (spray or roller dried) to produce powder.

- Whole milk powder must contain no less than 95% milk solids and must not exceed 5% moisture. The milk fat content must be no less than 2.6%. Vitamins A and D may be added and the emulsifying agent lecithin may also be added in an amount not exceeding 0.5%.
- Skim milk powder (non-fat dry milk) must contain no less than 95% milk solids and must not exceed 4% moisture or 1.5% fat.
- Buttermilk powder must contain no less than 95% milk solids and must not exceed 3% moisture or 6% fat.
- Whey powder consists primarily of carbohydrate (lactose), protein (several different whey proteins, mainly lactalbumins and globulins), various minerals, and vitamins. Whey powder is a valuable addition to the functional properties of various foods as well as a source of valuable nutrients because it contains approximately 50% of the nutrients in the original milk.

Table 2 compares the composition of milk and two powdered milk products.

Table 2 Comparison of fresh and powdered milk products (% by weights)			
	Whole Milk	Skim Milk Powder (Non-fat dry milk)	Buttermilk Powder
Milk fat	3.25	0.7	5.0
Protein	3.5	36.0	34.0
Milk sugar (lactose)	4.9	51.0	48.0
Minerals	0.8	8.2	7.9
Water	87.0	3.0	3.0
Calcium	0.12	1.3	1.3

- To make 10 L (22 lb.) of liquid skim milk from skim milk powder, 9.1 L (2.4 gal.) of water and 900 g (2 lb.) of skim milk powder are required.
- To make 10 L (22 lb.) of whole milk from skim milk powder, 8.65 L (2.25 gal) of water, 900 g (2 lb.) of skim milk powder, and 450 g (1 lb.) of butter are needed.

When reconstituting dried milk, add it to the water and whisk in immediately. Delaying this, or adding water

to the milk powder, will usually result in clogging. Water temperature should be around 21°C (70°F).

Evaporated Milk

Sometimes called concentrated milk, this includes evaporated whole, evaporated partly skimmed, and evaporated skim milks, depending on the type of milk used in its production. Canadian standards require 25% milk solids and 7.5% milk fat.

All types of evaporated milk have a darker color than the original milk because at high temperatures a browning reaction occurs between the milk protein and the lactose. After 60% of the water is removed by evaporation, the milk is homogenized, cooled, restandardized, and canned. It is then sterilized by heating for 10 to 15 minutes at 99°C to 120°C (210°F to 248°F). Controlled amounts of disodium phosphate and/or sodium citrate preserve the “salt balance” and prevent coagulation of the milk that might occur at high temperatures and during storage.

Sweetened Condensed Milk

Sweetened condensed milk is a viscous, sweet-colored milk made by condensing milk to one-third of its original volume, which then has sugar added. It contains about 40% sugar, a minimum of 8.5% milk fat, and not less than 28% total milk solids.

Attribution

Milk in bread baking

In the dough stage, milk increases water absorption. Consequently, dough made with milk should come softer from the mixer than dough made with water. Other aspects of milk in yeast doughs include:

- Dough may be mixed more intensively.
- Milk yields dough with a higher pH compared to water dough, and the fermentation will be slower.
- Fermentation tolerance (the ability of the dough to work properly in a range of temperatures) will be slightly improved.
- Bench time will be extended as the dough ferments more slowly at this stage. (Final proof times will be about the same, as by this time the yeast has adjusted to the condition of the dough.)

Bread made with milk will color faster in the oven and allowance should be made for this. If taken out too early after a superficial examination of crust color, it may collapse slightly and be hard to slice. The loaf should be expected to have a darker crust color than bread made without milk.

In the finished product, milk will make bread that has:

- Greater volume (improved capacity to retain gas)
- Darker crust (due to the lactose in the milk)
- Longer shelf life (due partly to the milk fat)
- Finer and more "cottony" grain
- Better slicing due to the finer grain

If skim milk or skim milk powder is used, some of the above benefits will not be so evident (e.g., longer shelf life, which is a result of the fat in the milk).

The type of sugar found in milk, lactose, has little sweetening power and does not ferment, so in dough made with skim milk powder, sugar has to be added or the fermentation will be very slow. While lactose is not fermentable, it caramelizes readily in the oven and produces a healthy crust color. The recommended amount of skim milk powder used in fermented dough is 2% to 8% based on flour, and up 15% in cakes.

Buttermilk and sour milk are used to make variety breads. They have a lower pH and require a shorter fermentation for good results.

Attribution

Yogurt

Yogurt is a thick or semi-solid food made from pasteurized milk fermented by lactic bacteria. The milk coagulates when a sufficient quantity of lactic acid is produced. Yogurt is a rich, versatile food capable of enhancing the flavor and texture of many recipes. It is prepared sweetened or unsweetened, and is used in baking to make yogurt-flavored cream cakes, desserts, and frozen products.

Attribution

Lactose

This milk sugar is a complex sugar (see sugar section). It is available commercially spray-dried and in crystalline form. There are many advantages to using it in various baking applications:

- Because of its low sweetening value compared to sucrose, it can lend texture and create browning while keeping the sweetness level at low values, which many consumers prefer. It can be used to replace sucrose up to a 50% level, or replace it entirely in products like pie pastry.
- Lactose improves dough handling properties and the color of the loaf.
- In pie crusts, it gives good color to top and bottom crusts, more tender crusts, and retards sogginess.
- In machine-dropped cookies, lactose can help the dough release better from the die.
- In cakes and muffins, it gives body without excessive sweetening and improves volume.
- Lactose binds flavors that are normally volatile and thus intensifies or enhances flavor.

Attribution

Cheese

Cheese is a concentrated dairy product made from fluid milk and is defined as the fresh or matured product obtained by draining the whey after coagulation of casein.

Cheese making consists of four steps:

- Curdling of the milk, either by enzyme (rennet) or by lactic curdling (natural process)
- Draining in which the whey (liquid part) is drained from the curd (firm part)
- Pressing, which determines the shape
- Ripening, in which the rind forms and the curd develops flavor

Cheese can be classified, with some exceptions, into five broad categories, as follows. Examples are given of specific cheeses that may be used in baking.

Fresh cheese: High moisture content and no ripening characterize these products. Examples: cottage cheese, baker's cheese, cream cheese, quark, and ricotta.

Soft cheeses: Usually some rind, but with a soft interior. Example: feta.

Semi-soft cheeses: Unripened cheeses of various moisture content. Example: mozzarella.

Firm cheeses: Well-ripened cheese with relatively low moisture content and fairly high fat content.

Examples: Swiss, cheddar, brick.

Hard cheeses: Lengthy aging and very low moisture content. Example: Parmesan.

In baking, cheeses have different functions. Soft cheeses, mixed with other ingredients, are used in fillings for pastries and coffeecakes. They are used for certain European deep-fried goods, such as cannoli. They may also be used, sometimes in combination with a richer cream cheese, for cheesecakes. All the cheeses itemized under fresh cheese (see above) are all more or less interchangeable for these functions. The coarser cheese may be strained first if necessary. The firmer cheeses are used in products like cheese bread, quiches, pizza, and cheese straws.

A brief description of the cheeses most likely to be used by bakers follows.

Dry Curd Cottage Cheese

This is a soft, unripened, acid cheese. Pasteurized skim milk is inoculated with lactic-acid-producing bacteria, and a milk-clotting enzyme (rennet) is added. Following incubation, the milk starts to clot, and it is then cut into cubes. After gentle cooking, the cubes or curds become quite firm. At this point, the whey is drained off, and the curd is washed and cooled with cold water.

Creamed Cottage Cheese

Creamed or dressed cottage cheese consists of dry curd cottage cheese combined with a cream dressing. The milk fat content of the dressing determines whether the final product is "regular" (4% milk fat) or low fat (1% to 2% milk fat).

Baker's Cheese

This is a soft, unripened, uncooked cheese. It is made following exactly the same process as for dry curd cottage cheese, up to and including the point when the milk clot is cut into cubes. This cheese is not cooked to remove the whey from the curd. Rather, the curd is drained through cloth bags or it may be pumped through a curd concentrator. The product is then ready to be packaged. The milk fat content is

generally about 4%.

Quark

Quark (or quarg) is a fresh unripened cheese prepared in a fashion similar to cottage cheese. The mild flavor and smooth texture of quark make it excellent as a topping or filling for a variety of dishes. Quark is similar to baker's cheese, except acid is added to it (it is inoculated with lactic-acid-producing bacteria), and then it is blended with straight cream to produce a smooth spread containing approximately 7% milk fat. Today there are low-fat quarks with lower percentage, and high-fat versions with milk fat adjusted to 18%. Quark cheese can often be used in place of sour cream, cottage cheese, or ricotta cheese.

Cream Cheese

Cream cheese is a soft, unripened, acid cheese. A milk-and-cream mixture is homogenized and pasteurized, cooled to about 27°C (80°F), and inoculated with lactic-acid-producing bacteria. The resulting curd is not cut, but it is stirred until it is smooth, and then heated to about 50°C (122°F) for one hour. The curd is drained through cloth bags or run through a curd concentrator. Regular cream cheese is fairly high fat, but much lighter versions exist now.

Ricotta

Ricotta is a fresh cheese prepared from either milk or whey that has been heated with an acidulating agent added. Traditionally lemon juice or vinegar was used for acidulation, but in commercial production, a bacterial culture is used. The curds are then strained and the ricotta is used for both sweet and savory applications.

Mascarpone

Mascarpone is a rich, fresh cheese that is a relative of both cream cheese and ricotta cheese. Mascarpone is prepared in a similar fashion to ricotta, but using cream instead of whole milk. The cream is acidified (often by the direct addition of tartaric acid) and heated to a temperature of 85°C (185°F), which results in precipitation of the curd. The curd is then separated from the whey by filtration or mechanical means. The cheese is lightly salted and usually whipped. Note that starter culture and rennet are not used in the production of this type of cheese. The high-fat content and smooth texture of mascarpone cheese make it suitable as a substitute for cream or butter. Ingredient applications of mascarpone cheese tend to focus on desserts. The most famous application of mascarpone cheese is in the Italian dessert tiramisu.

Table 1 provides the composition of various types of cheeses.

Table 1 Composition of various cheeses (% by weight)			
	Moisture %	Milk Fat %	Salt %
Dry curd cottage cheese	80	0.4	n/a
Regular creamed cottage cheese	79	4	1
Low fat (1% and 2%) creamed cottage cheese	79	1-2	1
Baker's cheese	79	4	1
Quark	72	5-7	n/a
Quark (high fat)	59	18	n/a
Cream cheese	54 (varies)	17- 37	1
Ricotta	72-75	8-13	n/a

Mascarpone	46	60-75	1
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Attribution

Understanding Ingredients: Eggs

Eggs Grade (ADD US)

Fresh hen eggs are sold by grade in all provinces. All shell eggs that are imported, exported, or shipped from one province to another for commercial sale must be graded. In Canada, it is mandatory to have all eggs graded by the standards set by Agriculture and Agri-Foods Canada (AAFC). The grade name appears on cartons. The grades Canada A and Canada B bear the maple leaf symbol with the grade name inside, and Canada C and Nest Run eggs will have the grade name printed in block text. The grades indicate the quality of the egg and should not be confused with size. Only Canada A are available in different sizes. The average large size egg weighs about 56 g (2 oz.) as indicated in Table 1.

Size	Weight (including shell)
Peewee	Less than 42 g (1.5 oz.)
Small	At least 42 g (1.5 oz.)
Medium	At least 49 g (1.75 oz.)
Large	At least 56 g (2 oz.)
Extra Large	At least 63 g (2.25 oz.)
Jumbo	70 g (2.5 oz.) or more

The Canada grade symbol does not guarantee that the eggs are of Canadian origin, but it does guarantee that the products meet Canadian government standards. Agriculture Canada inspects all egg-processing plants to ensure that the products are wholesome and processed according to sanitary standards. The pasteurization of “packaged” egg product is also monitored.

The criteria for grading eggs are:

- Weight
- Cleanliness
- Soundness and shape of shell
- Shape and relative position of yolk within the egg
- Size of air cell free of abnormalities
- Freedom from dissolved yolk and blood spots

Canada A

Canada A eggs are clean, normal in shape with sound shells, and have the finest interior quality. They are ideal for all uses. The yolks are round and compact and surrounded by very thick, firm albumen. Canada A eggs are a premium quality and in limited supply on the retail market. If eggs are not sold within a limited time, unsold stocks are returned to the supplier. Eggs graded as A must meet the minimum weight for the declared size (see Table 12.) The size designation for Canada A eggs must appear on the label.

Canada B

Canada B eggs have very slight abnormalities. This grade is fine for baking, where appearance is not important. These eggs must weigh at least 49 g (1.75 oz.). There are no size designations on the label for Canada B eggs.

Canada C

Canada C is considered a processing grade and provides a safe outlet for the disposition of cracked eggs. Canada C eggs must be shipped to a federally registered processed egg station and pasteurized as a means of controlling the higher risk of salmonella or other microbial contamination that may be found in cracked eggs.

These eggs are suitable for processing into commercially frozen, liquid, and dried egg products. Sizes are not specified.

Canada Nest Run

Since Canada Nest Run eggs are generally sent for further processing, they are usually not washed, candled (a process discussed later in this chapter), or sized. However, nest run eggs must meet the minimum quality requirements prescribed by the [egg regulations](#). This grade, as with other Canada grades, can only be applied to eggs in a federally registered egg station.

Attribution

Composition and Nutrition

Table 1 Composition of eggs by percent of weight. Traces of sugar and ash are also present.

	Composition of Eggs (%) Whole Egg	Composition of Eggs (%) Yolk	Composition of Eggs (%) White
Moisture	73.0	49.0	86.0
Protein	13.3	16.7	11.6
Lipid	11.5	31.6	0.2

Table 2 Nutritional content of a large egg

	Whole Egg	Yolk	White
Weight	50 g	17 g	33 g
Protein	6 g	3 g	3 g
Fat	5 g	5 g	Trace
Cholesterol	216 mg	216 mg	0
Calcium	25 mg	2 mg	27 mg
Iron	1.0 mg	0.6 mg	Trace
Sodium	63 mg	7 mg	54 mg
Potassium	60 mg	16 mg	47 mg
Vitamin A	96 RE	99 RE	0 RE

Note: B-complex vitamins, not itemized, are well represented in eggs, as are amino acids. RE = retinol equivalent, a term used in nutritional measurement.

Worth noting is the concentration of certain food elements in different parts of the egg. Note for example that all the cholesterol is in the yolk. The yolk is relatively rich in iron and the white is high in calcium.

In practice, when separating large eggs, one estimates the weight of the white as 30 g (1 oz) and the yolk as 20 g (0.7 oz). The color of the shell, which is either a creamy white or brown, is relevant to the breed of the hen, and there is no other basic difference in the content of the egg or the shell.

The color of the yolk depends on the diet of the hens. Bakers have a preference for eggs with dark yolks. Certainly the appearance of cakes made with such eggs is richer. Tests have found that, although eggs with darker yolks tend to produce moister sponge cakes, the cakes are somewhat coarser and less tender.

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Egg Products

A number of egg products besides whole shell eggs are used in the baking and food service industry. By law, all egg products other than shell eggs are pasteurized to protect them against salmonella, and the low temperature at which they are kept inhibits bacterial activity, although under certain conditions they may spoil very rapidly.

The chief categories of egg products available are:

- Liquid eggs (whole eggs and whole eggs with additional yolks)
- Frozen eggs (whole eggs, egg whites, and egg yolks)
- Dried and powdered eggs (whole eggs, egg whites, and meringue powder)

Liquid and Frozen Eggs

Liquid and frozen whole eggs are preferred in large bakeries where cracking and emptying of shells is not economical. They are also one of the most economical ways of purchasing eggs. Liquid and frozen whole eggs are sometimes “fortified” by the addition of egg yolks. Some bakers feel that liquid or frozen eggs don’t yield the same volume in sponge cakes as fresh eggs, and there is a certain bias in favor of shell eggs.

If stored in the freezer at -18°C (0°F) or lower, liquid and frozen eggs will keep for long periods with minimum loss of quality. Thawing should take place in the refrigerator or under cold water without submerging the container. Leaving frozen eggs at room temperature to thaw is a bad practice because the outside layers of egg can reach a temperature favorable to bacteria while the centre is still frozen. Heat should never be used to defrost eggs. Unused portions must be refrigerated and used within 24 hours.

Frozen egg yolks consist of 90% egg yolks and 10% sugar to prevent the yolk from gelling and to avoid separation of the fat.

Spray-Dried Whole Eggs and Egg Whites

Dried eggs are used by some bakers as a convenience and cost saver. As with frozen eggs, some bakers doubt their performance in products such as sponge cakes. But dried eggs produce satisfactory results because of the addition of a carbohydrate to the egg before the drying process, usually corn syrup, which results in foaming comparable to fresh eggs.

Dried whole eggs should be stored unopened in a cool place not over 10°C (50°F), preferably in the refrigerator. They are reconstituted by blending 1 kg (2.2 lb.) of powdered whole egg with 3 kg (6.6 lb) of cold water. The water is added slowly while mixing. Once reconstituted, dried eggs should be used immediately or refrigerated promptly and used within an hour.

In mixes such as muffins and cake doughnuts, dried eggs can be mixed in with the other dry ingredients and do not have to be reconstituted. In layer cake formulas, dried eggs are blended with the other dry ingredients before the fat and some water are added, followed by the balance of liquid in two stages.

Spray-dried egg whites are reconstituted by mixing 1 kg (2.2 lb.) of powdered egg white with 1 kg (2.2 lb.) of cold water, letting it stand for 15 minutes, and then adding 9 kg (20 lb.) of cold water. When used in cake mixes, the powdered egg white is blended with the other dry ingredients, but only 7 L (7 qt.) of cold water is used for every 1 kg (2.2 lb.) of powdered egg white.

Dry Egg Substitutes or Replacements

Egg substitutes are made from sweet cheese, whey, egg whites, dextrose, modified tapioca starch, **sodium** caseinate, and artificial color and flavor. They are cost-cutters and can be used alone or in combination with fresh or dried eggs in cakes, cookies, and fillings. One kg (2.2 lb.) of powder is mixed with 4 kg (9 lb.) of water to replace powdered eggs.

Meringue Powder

While it is not a pure dehydrated egg white, meringue powder is widely used by bakers to make baked Alaska, royal icing, and toppings. It contains vegetable gums and starches to absorb moisture and make it whip better.

Attribution

The Function of Eggs

Eggs are a truly multifunctional ingredient and have many roles to play in the bakeshop. Their versatility means that product formulas may be adjusted once the properties of eggs are understood. For example, in French butter cream, egg whites may be substituted in the summer for whole eggs to give a more stable and bacteria-free product (egg white is alkaline, with pH 8.5). A yolk or two may be worked into a sweet short paste dough to improve its extensibility. Sponge cake formulas can be adjusted, for example, with the addition of egg yolks in jelly rolls to improve rolling up.

If a recipe is changed by replacing some or all of the eggs with water, two factors must be remembered:

Water replacement is about 75% of the egg content, since egg solids constitute about 25% of the egg.

Leavening ability is lessened and must be made up by the addition of chemical leavening.

Other uses of eggs are:

- **Leavening:** They will support many times their own weight of other ingredients through their ability to form a cell structure either alone or in combination with flour. The egg white in particular is capable of forming a large mass of cells by building a fine protein network.
- **Moistening and binding:** The fat in eggs provides a moistening effect, and the proteins present coagulate when heated, binding ingredients together.
- **Thickening:** Eggs are valuable thickeners in the cooking of chiffon pie fillings and custard.
- **Emulsifying:** Lecithin, present in the yolk, is a natural emulsifier and assists in making smooth batters.
- **Customer appeal:** Eggs enhance the appearance of products through their colour and flavour, and they improve texture and grain.
- **Structure:** Eggs bind with other ingredients, primarily flour, creating the supporting structure for other ingredients.
- **Shelf life:** The shelf life of eggs is extended through the fat content of the yolk.
- **Nutrition:** Eggs are a valuable food in every respect. Note, however, that 4% of the lipid in egg yolk is cholesterol, which may be a concern to some people. Developments in poultry feed claim to have reduced or eliminated this cholesterol level.
- **Tenderizing:** The fat in eggs acts like a shortening and improves the tenderness of the baked cake.

Keep these points in mind when using eggs:

- Spots in eggs are due to blood fragments in the ovary. Such eggs are edible and may be used.
- The albumen or egg white is soluble in cold water, congeals at 70°C (158°F), and remains insoluble from then on.
- Cover leftover yolks or whites tightly and refrigerate. Add a little water on top of yolks, or mix in 10% sugar, to prevent crusting. Do not return unused portions to the master container.
- Use clean utensils to dip egg products from their containers.

Attribution

Storing Eggs

Whole eggs are the perfect medium for the development of bacteria and mould. Eggs with an undesirable odor may be high in bacteria or mould. While some of these odors disappear in baking, some will remain and give an off-taste to the product if the odor is concentrated and strong.

Store fresh eggs in the refrigerator in cartons to prevent moisture loss and absorption of odours. If refrigerator space is at a premium, eggs are stable for up to three weeks if kept at a temperature of 13°C to 15°C (55°F to 60°F). Naturally, this must be in a location with invariable conditions.

Food poisoning can result from using eggs held too long before using. Liquid or cracked eggs should be kept under refrigeration at all times.

Whole eggs can be checked for freshness with the candling or salt water method:

- **Candling method:** Hold the egg up to a light in a darkened room or positioned so that the content or condition of the egg may be seen. If the yolk is held firmly by the white when the egg is turned, and the egg is clean and not broken, then the egg is of good quality. Smell or odor is not readily revealed unless the shell is broken.
- **Salt water method:** Add 100 g of salt to 1 L (3.5 oz. to 1 qt.) of water. Allow to dissolve completely. When an egg is placed in this mixture, its level of buoyancy determines the age of the egg. An old egg will float to the surface, while a fresher egg will sink to the bottom.

Attribution

Chocolate

From the Cocoa Bean to the Finished Chocolate

In North America, chocolate manufacturing started in Massachusetts in 1765. Today, in the factory, the beans get cleaned, and magnets take out metallic parts, and then sand, dust, and other impurities are removed. Some starch will be changed into dextrans in the roasting process to improve flavor. Machines break the beans and grind them fine until a flowing liquid is produced, called chocolate liquor. Through hydraulic pressure, cocoa butter is reduced from 55% to approximately 10% to 24% or less, and the residue forms a solid mass called *press cake*.

The press cake is then broken, pulverized, cooled, and sifted to produce commercial cocoa powder. The baking industry uses primarily cocoa powders with a low fat content.

At the factory, chocolate is also subject to an additional refining step called **conching**. Conching has a smoothing effect. The temperature range in this process is between 55°C and 65°C (131°F and 149°F). Sugar interacts with protein to form amino sugars, and the paste loses **acids** and moisture and becomes smoother.

This video explains the chemical reactions related to heat, melting point, and formation of crystal structures in chocolate:
<http://science360.gov/obj/video/27d931d9-c33c-45c6-adac-aa0a42f04ad6/chemistry-chocolate>

Attribution

Chocolate Produced for the Baking Industry

True chocolate contains cocoa butter. The main types of chocolate, in decreasing order of cocoa liquor content, are:

- Unsweetened (bitter) chocolate
- Dark chocolate
- Milk chocolate
- White chocolate

Unsweetened Chocolate

Unsweetened chocolate, also known as bitter chocolate, baking chocolate, or cooking chocolate, is pure cocoa liquor mixed with some form of fat to produce a solid substance. The pure ground, roasted cocoa beans impart a strong, deep chocolate flavor. With the addition of sugar in recipes, however, it is used as the base for cakes, brownies, confections, and cookies.

Dark (Sweet, Semi-Sweet, Bittersweet) Chocolate

Dark chocolate has an ideal balance of cocoa liquor, cocoa butter, and sugar. Thus it has the attractive, rich color and flavor so typical of chocolate, and is also sweet enough to be palatable. It does not contain any milk solids. It can be eaten as is or used in baking. Its flavor does not get lost or overwhelmed, as in many cases when milk chocolate is used. It can be used for fillings, for which more flavorful chocolates with high cocoa percentages ranging from 60% to 99% are often used. *Dark* is synonymous with semi-sweet, and *extra dark* with bittersweet, although the ratio of cocoa butter to solids may vary.

- Sweet chocolate has more sugar, sometimes almost equal to cocoa liquor and butter amounts (45% to 55% range).
- Semi-sweet chocolate is frequently used for cooking. It is a dark chocolate with less sugar than sweet chocolate.
- Bittersweet chocolate has less sugar and more liquor than semi-sweet chocolate, but the two are often interchangeable when baking. Bittersweet and semi-sweet chocolates are sometimes referred to as *couverture* (see below). The higher the percentage of cocoa, the less sweet the chocolate is.

Milk Chocolate

Milk chocolate is solid chocolate made with milk, added in the form of milk powder. Milk chocolate contains a higher percentage of fat (the milk contributes to this) and the melting point is slightly lower. It is used mainly as a flavoring and in the production of candies and moulded pieces.

White Chocolate

The main ingredient in white chocolate is sugar, closely followed by cocoa butter and milk powder. It has no cocoa liquor. It is used mainly as a flavoring in desserts, in the production of candies and, in chunk form in cookies.

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Couverture

The usual term for top quality chocolate is **couverture**. Couverture chocolate is a very high-quality chocolate that contains extra cocoa butter. The higher percentage of cocoa butter, combined with proper tempering, gives the chocolate more sheen, firmer “snap” when broken, and a creamy mellow flavor. Dark, milk, and white chocolate can all be made as couvertures.

The total percentage cited on many brands of chocolate is based on some combination of cocoa butter in relation to cocoa liquor. In order to be labelled as couverture by European Union regulations, the product must contain not less than 35% total dry cocoa solids, including not less than 31% cocoa butter and not less than 2.5% of dry non-fat cocoa solids. Couverture is used by professionals for dipping, coating, moulding, and garnishing.

What the percentages don't tell you is the proportion of cocoa butter to cocoa solids. You can, however, refer to the nutrition label or company information to find the amounts of each. All things being equal, the chocolate with the higher fat content will be the one with more cocoa butter, which contributes to both flavor and mouthfeel. This will also typically be the more expensive chocolate, because cocoa butter is more valuable than cocoa liquor.

But keep in mind that just because two chocolates from different manufacturers have the same percentages, they are not necessarily equal. They could have dramatically differing amounts of cocoa butter and liquor, and dissimilar flavors, and substituting one for the other can have negative effects for your recipe. Determining the amounts of cocoa butter and cocoa liquor will allow you to make informed decisions on chocolate choices.

Attribution

Definitions and Regulations (ADD US)

The legislation for cocoa and chocolate products in Canada is found in Division 4 of the Food and Drug Regulations (FDR), under the [Food and Drugs Act \(FDA\)](#). The [Canadian Food Inspection Agency \(CFIA\)](#) is responsible for administering and enforcing the FDR and FDA. [Here are some of the regulations governing cocoa and chocolate:](#)

- Cocoa butter must be the only fat source. Chocolate sold in Canada cannot contain vegetable fats or oils.
- Chocolate must contain chocolate liquor.
- The only sweetening agents permitted in chocolate in Canada are listed in Division 18 of the Food and Drug Regulations.
- Artificial sweeteners such as aspartame, sucralose, acesulfame potassium, and sugar alcohols (sorbitol, maltitol, etc.) are not permitted.
- Milk and/or milk ingredients are permissible.
- Emulsifying agents are permissible, as are flavors such as vanilla.

Cocoa butter and sugar quantities are not defined in the regulations. Some semi-sweet chocolate may be sweeter than so-called sweet chocolate. And remember that bittersweet chocolate is not, as you might expect, sugarless. Only if the label states “unsweetened,” do you know that there is no sugar added.

Products manufactured or imported into Canada that contain non-permitted ingredients (vegetable fats or oils, artificial sweeteners) cannot legally be called chocolate when sold in Canada. A non-standardized name such as “candy” must be used.

Finally, lecithin, which is the most common emulsifying agent added to chocolate, is approved for use in chocolate in North America and Europe, but Canadian regulations state that no more than 1% can be added during the manufacturing process of chocolate. Emulsifiers like lecithin can help thin out melted chocolate so it flows evenly and smoothly. Because it is less expensive than cocoa butter at thinning chocolate, it can be used to help lower the cost. The lecithin used in chocolate is mainly derived from soy. Both GMO (**genetically modified** organism) and non-GMO soy lecithin are available. Check the manufacturer’s packaging and ingredient listing for the source of soy lecithin in your chocolate.

Attribution

Understanding Ingredients: Spices

Elements of Taste

Essentially there are a handful of elements that compose all of the taste profiles found in the foods we eat. Western definitions of taste conventionally define four major elements of taste:

- Salty
- Sweet
- Sour
- Bitter

Asian cultures have added the following to the list:

- Umami (literally “pleasant savory taste”)
- Spiciness
- Astringency

Foods and recipes that contain a number of these elements in balance are generally those that we think of as tasting good.

Attribution

Introduction to Salt

Historically, salt was a prestigious commodity. “The salt of the earth” describes an outstanding person. The word *salary* comes from the Latin *salaria*, which was the payment made to Roman soldiers for the purchase of salt. In Arabic, the phrase translated as “there is salt between us” expresses the covenant between humans and the divine. Though no longer a valuable commodity in the monetary sense, salt is still valuable in the sense of being crucial to human health.

Salt can be found deposited in Earth’s layers in rock salt deposits. These deposits formed when the water in the oceans that covered Earth many millions of years ago evaporated. The salt was then covered by various types of rocks.

Common salt (sodium chloride) is 40% sodium and 60% chloride. An average adult consumes about 7 kg (15 lb.) per year.

Today, we have three basic methods of obtaining salt from natural sources:

- Mining rock salt

- Extracting salt from salt brines created by pumping water into underground salt deposits

- Evaporating salt water from oceans, seas, and salt lakes

Attribution

Types of Salt

Mined Rock Salt

In some countries, salt is mined from salt beds approximately 150 m to 300 m (490 ft. to 985 ft.) below Earth's surface. Sometimes, impurities such as clay make it impossible to use rock salt without purification. Purification makes it possible to get the desired flavor and color, thus making it edible. Edible salt is highly refined: pure and snow white.

Salt from Salt Brines

Salt can also be mined from natural salt beds by using water to extract the salt in the form of a brine, which saves having to construct a mine. Holes are drilled approximately 20 cm (8 in.) in diameter until the salt deposits are reached. A pipe is then driven into the salt beds and another pipe is driven inside the larger pipe further into the deposits. Pressurized water is forced through the outer pipe into the salt beds, and then pumped back out through the smaller pipe to the refineries. Through separation of the impurities, eventually all water in the brine will evaporate, leaving crystallized salt, which then can be dried, sifted, and graded in different sizes.

Ocean, Sea, and Lake Salt

In some countries, especially those with dry and warm climates, salt is recovered straight from the ocean or salt lakes. The salt water is collected in large shallow ponds (also called *salt gardens*) where, through the heat of the sun, the water slowly evaporates. Moving the salt solution from one pond to another until the salt crystals become clear and the water has evaporated eliminates impurities. The salt is then purified, dried completely, crushed, sifted, and graded.

Attribution

Functions of Salt in Baking

Salt has three major functions in baking. It affects:

- Fermentation
- Dough conditioning
- Flavor

Fermentation

Fermentation is salt's major function:

- Salt slows the rate of fermentation, acting as a healthy check on yeast development.
- Salt prevents the development of any objectionable bacterial action or wild types of fermentation.
- Salt assists in oven browning by controlling the fermentation and therefore lessening the destruction of sugar.
- Salt checks the development of any undesirable or excessive acidity in the dough. It thus protects against undesirable action in the dough and effects the necessary healthy fermentation required to secure a finished product of high quality.

Dough Conditioning

Salt has a binding or strengthening effect on gluten and thereby adds strength to any flour. The additional firmness imparted to the gluten by the salt enables it to hold the water and gas better, and allows the dough to expand without tearing. This influence becomes particularly important when soft water is used for dough mixing and where immature flour must be used. Under both conditions, incorporating a maximum amount of salt will help prevent soft and sticky dough. Although salt has no direct bleaching effect, its action results in a fine-grained loaf of superior texture. This combination of finer grain and thin cell walls gives the crumb of the loaf a whiter appearance.

Flavour

One of the important functions of salt is its ability to improve the taste and flavor of all the foods in which it is used. Salt is one ingredient that makes bread taste so good. Without salt in the dough batch, the resulting bread would be flat and insipid. The extra palatability brought about by the presence of salt is only partly due to the actual taste of the salt itself. Salt has the peculiar ability to intensify the flavor created in bread as a result of yeast action on the other ingredients in the loaf. It brings out the characteristic taste and flavor of bread and, indeed, of all foods. Improved palatability in turn promotes the digestibility of food, so it can be said that salt enhances the nutritive value of bakery products. The lack of salt or too much of it is the first thing noticed when tasting bread. In some bread 2% can produce a decidedly salty taste, while in others the same amount gives a good taste. The difference is often due to the mineralization of the water used in the dough.

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Using Salt in Fermented Doughs

The average amount of salt to use in dough is about 1.75% to 2.25% based on the flour used. Some authorities recommend that the amount of salt used should be based on the actual quantity of water used in making the dough, namely about 30 g per L (1 oz. per qt.) of water.

During the hot summer months, many bakers find it advantageous to use slightly more salt than in the winter as a safeguard against the development of any undesirable changes in the dough fermentation. Salt should never be dissolved in the same water in which yeast is dissolved. It is an antiseptic and dehydrates yeast cells and can even kill part of them, which means that less power is in the dough and a longer fermentation is needed. In bread made by the sponge dough method and in liquid fermentation systems, a small amount of salt included in the first stage strengthens the gluten.

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Storing Salt

Salt is very stable and does not spoil under ordinary conditions. However, it may have a slight tendency to absorb moisture and become somewhat lumpy and hard. Therefore, it is advisable to store it in a clean, cool, and dry place. Inasmuch as salt can absorb odors, the storage room should be free from any odor that might be taken up and carried by the salt.

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Introduction to Spices and Other Flavorings

Food touches all of the senses. We taste, we smell, we see color and shape, we feel texture and temperature, and we hear sounds as we eat.

All of these elements together create a palette with an infinite number of combinations, but the underlying principles that make food taste good are unchanged.

- Variety and diversity in textures and the elements of taste make for interesting food.
- Contrast is as important as harmony; but avoid extremes and imbalance.

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Seasoning and Flavoring

Many ingredients are used to enhance the taste of foods. These ingredients can be used to provide both seasoning and flavoring.

- *Seasoning* means to bring out or intensify the natural flavor of the food without changing it. Seasonings are usually added near the end of the cooking period. The most common seasonings are salt, pepper, and acids (such as lemon juice). When seasonings are used properly, they cannot be tasted; their job is to heighten the flavors of the original ingredients.
- *Flavoring* refers to something that changes or modifies the original flavor of the food. Flavoring can be used to contrast a taste such as adding liqueur to a dessert where both the added flavor and the original flavor are perceptible. Or flavorings can be used to create a unique flavor in which it is difficult to discern what the separate flavorings are. Spice blends used in pumpkin pies are a good example of this.

Knowing how to use seasonings and flavorings skillfully provides cooks and bakers with an arsenal with which they can create limitless flavor combinations.

Flavoring and seasoning ingredients include wines, spirits, fruit zests, extracts, essences, and oils. However, the main seasoning and flavoring ingredients are classified as herbs and spices.

Knowing the difference between herbs and spices is not as important as knowing how to use seasonings and flavorings skillfully. In general, fresh seasonings are added late in the cooking process while dry ones tend to be added earlier. It is good practice to under-season during the cooking process and then add more seasonings (particularly if you are using fresh ones) just before presentation. This is sometimes referred to as “layering.” When baking, it is difficult to add more seasoning at the end, so testing recipes to ensure the proper amount of spice is included is a critical process.

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Herbs

Herbs tend to be the leaves of fragrant plants that do not have a woody stem. Herbs are available fresh or dried, with fresh herbs having a more subtle flavor than dried. You need to add a larger quantity of fresh herbs (up to 50% more) than dry herbs to get the same desired flavor. Conversely, if a recipe calls for a certain amount of fresh herb, you would use about one-half of that amount of dry herb.

The most common fresh herbs are basil, coriander, marjoram, oregano, parsley, rosemary, sage, tarragon, and thyme. Fresh herbs should have a clean, fresh fragrance and be free of wilted or brown leaves. They can be kept for about five days if sealed inside an airtight plastic bag. Fresh herbs are usually added near the completion of the cooking process so flavors are not lost due to heat exposure.

Dried herbs lose their power rather quickly if not properly stored in airtight containers. They can last up to six months if properly stored. Dried herbs are usually added at the start of the cooking process as their flavor takes longer to develop than fresh herbs.

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Spices

Spices are aromatic substances obtained from the dried parts of plants such as the roots, shoots, fruits, bark, and leaves. They are sold as seeds, blends of spices, whole or ground spices, and seasonings. The aromatic substances that give a spice its particular aroma and flavor are the essential oils. The flavor of the essential oil or flavoring compound will vary depending on the quality and freshness of the spice.

The aromas of ground spices are volatile. This means they lose their odor or flavoring when left exposed to the air for extended periods. They should be stored in sealed containers when not in use. Whole beans or unground seeds have a longer shelf life but should also be stored in sealed containers.

Allspice

Allspice is only one spice, yet it has a flavor resembling a blend of cloves, nutmeg, and cinnamon. At harvest time, the mature (but still green) berries from the allspice trees (a small tropical evergreen) are dried in the sun. During drying they turn reddish-brown and become small berries. The berries are about 0.6 cm (1/4 in.) in diameter and contain dark brown seeds.

Allspice is grown principally in Jamaica and to a lesser degree in Mexico. Allspice is available whole or ground. Bakers usually use ground allspice in cakes, cookies, spices, and pies.

Anise

Anise is the small, green-grey fruit or seed of a plant of the parsley family. The plant grows to a height of 45 cm (18 in.) and has fine leaves with clusters of small white flowers. It is native to Mexico and Spain, with the latter being the principal producer. Anise seeds are added to pastries, breads, cookies, and candies.

Caraway

Caraway is the dried fruit or seed of a biennial plant of the parsley family, harvested every second year, primarily in the Netherlands. It is also produced in Poland and Russia. The many-branched, hollow-stemmed herb grows up to 60 cm (24 in.) high and has small white flowers. Caraway is a small crescent-shaped brown seed with a pleasant aroma but somewhat sharp taste. Although it is most familiar in rye bread, caraway is also used in cookies and cakes.

Cardamom

Native to India, Sri Lanka, and Guatemala, cardamom is the fruit or seed of a plant of the ginger family. The three-sided, creamy-white, flavorless pod holds the tiny aromatic, dark brown seeds. It is available in whole and ground (pod removed). Cardamom in ground form flavors Danish pastries and coffee cakes, Christmas baking, and Easter baking such as hot cross buns.

Cinnamon

Cinnamon comes from the bark of an aromatic evergreen tree. It is native to China, Indonesia, and Indochina. Cinnamon may be purchased in ground form or as cinnamon sticks. Ground cinnamon is used in pastries, breads, puddings, cakes, candy, and cookies. Cinnamon sticks are used for preserved fruits and flavoring puddings. Cinnamon sugar is made with approximately 50 g (2 oz.) of cinnamon to 1 kg (2.2 lb.) of granulated sugar.

Cassia

Cassia, sometimes known as Chinese cinnamon, is native to Assam and Myanmar. It is similar to cinnamon but a little darker with a sharper taste. It is considered better for savory rather than sweet foods. It is prized in Germany and some other countries as a flavor in chocolate.

Cloves

Cloves are the dried, unopened buds of a tropical evergreen tree, native to Indonesia. The flavor is characterized by a sweet, pungent spiciness. The nail-shaped whole cloves are mainly used in cooking, but the ground version of this spice heightens the flavor of mincemeat, baked goods, fruit pies, and plum pudding.

Ginger

Ginger is one of the few spices that grow below the ground. It is native to southern Asia but is now imported from Jamaica, India, and Africa. The part of the ginger plant used is obtained from the root. Ground ginger is the most commonly used form in baking — in fruitcakes, cookies, fruit pies, and gingerbread. Candied ginger is used in pastries and confectionery.

Mace

Originating in the East and West Indies, mace is the fleshy growth between the nutmeg shell and outer husk, yellow-orange in color. It is usually sold ground, but sometimes whole mace (blades of mace) is available. Mace is used in pound cakes, breads, puddings, and pastries.

Nutmeg

Nutmeg is the kernel or seed of the nutmeg fruit. The fruit is similar to the peach. The fleshy husk, grooved on one side, splits, releasing the deep-brown aromatic nutmeg. It is available whole or ground. Ground nutmeg is used extensively in custards, cream puddings, spice cakes, gingerbread, and doughnuts.

Poppy Seed

Poppy seed comes from the Netherlands and Asia. The minute, blue-grey, kidney-shaped seeds are so small they seem to be round. Poppy seeds are used in breads and rolls, cakes and cookies, and fillings for pastries.

Sesame or Benne Seed

Sesame or benne seeds are the seeds of the fruit of a tropical annual herb grown in India, China, and Turkey. The seeds are tiny, shiny, and creamy-white with a rich almond-like flavor and aroma. Bakers use sesame seeds in breads, buns, coffee cakes, and cookies.

Vanilla

The Spaniards named vanilla. The word derives from *vaina*, meaning pod. Vanilla is produced from an orchid-type plant native to Central America. The vanilla beans are cured by a complicated process, which helps explain the high cost of genuine vanilla. The cured pods should be black in color and packed in airtight boxes. Imitation vanilla extracts are made from a colorless crystalline synthetic compound called vanillin. Pure vanilla extract is superior to imitation vanilla. Artificial vanilla is more intense than real vanilla by a factor of 3 to 4 and must be used sparingly.

To use vanilla beans, split the pod down the middle to scrape out the seeds. The seeds are the flavoring agents. Alternatively, the split pod can be simmered in the milk or cream used in dessert preparation. Its flavoring power is not spent in one cooking and it can be drained, kept frozen, and reused. A vanilla bean kept in a container of icing sugar imparts the flavor to the sugar, all ready for use in cookies and cakes.

Vanilla extract is volatile at temperatures starting at 138°C (280°F) and is therefore not ideal for flat products such as cookies. It is suitable for cakes, where the interior temperature does not get so high.

Vanilla beans and vanilla extract are used extensively by bakers to flavor a wide range of desserts and other items.

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Flavorings in Baking

Flavors cannot be considered a truly basic ingredient in bakery products but are important in producing the most desirable products. Flavoring materials consist of:

- Extracts or essences
- Emulsions
- Aromas
- Spices

Note: Salt may also be classed as a flavoring material because it intensifies other flavors.

These and others (such as chocolate) enable the baker to produce a wide variety of attractively flavored pastries, cakes, and other bakery products. Flavor extracts, essences, emulsions, and aromas are all solutions of flavor mixed with a solvent, often ethyl alcohol.

The flavors used to make extracts and essences are the extracted essential oils from fruits, herbs, and vegetables, or an imitation of the same. Many fruit flavors are obtained from the natural parts (e.g., rind of lemons and oranges or the exterior fruit pulp of apricots and peaches). In some cases, artificial flavor is added to enhance the taste, and artificial coloring may be added for eye appeal. Both the Canadian and U.S. departments that regulate food restrict these and other additives. The flavors are sometimes encapsulated in corn syrup and emulsifiers. They may also be coated with gum to preserve the flavor compounds and give longer shelf life to the product. Some of the most popular essences are compounded from both natural and artificial sources. These essences have the true taste of the natural flavors.

Aromas are flavors that have an oil extract base. They are usually much more expensive than alcoholic extracts but purer and finer in their aromatic composition. Aromas are used for flavoring delicate creams, sauces, and ice creams.

Emulsions are homogenized mixtures of aromatic oils and water plus a stabilizing agent (e.g., vegetable gum). Emulsions are more concentrated than extracts and are less susceptible to losing their flavor in the oven. They can therefore be used more sparingly.

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