
SNOWFLAKE
CAKE WHEAT FLOUR

Chemistry of Cooking

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Contents

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

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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. Image sources: 1)apple by PublicDomainPictures b) watermelon-summer by jill111 c) man by olichel

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.
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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.

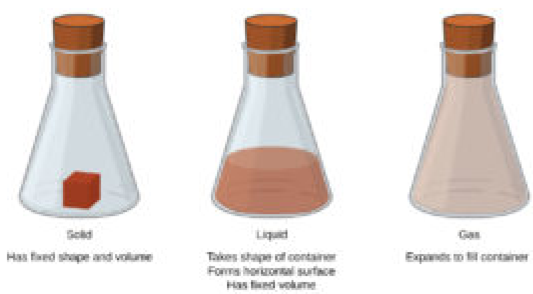

Figure 1 *The three most common states or phases of matter are solid, liquid, and gas.*

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.

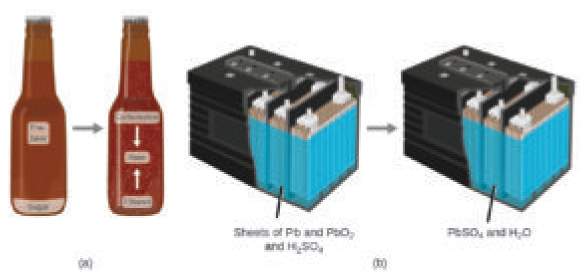

Figure 3. *(a) The mass of beer precursor materials is the same as the mass of beer produced: Sugar has become alcohol and carbonation. (b) The mass of the lead, lead oxide plates, and sulfuric acid that goes into the production of electricity is exactly equal to the mass of lead sulfate and water that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

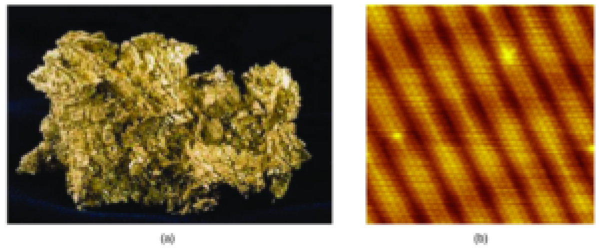

Figure 4. *(a) This photograph shows a gold nugget. (b) A scanning-tunneling microscope (STM) can generate views of the surfaces of solids, such as this image of a gold crystal. Each sphere represents one gold atom. (credit a: modification of work by United States Geological Survey; credit b: modification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

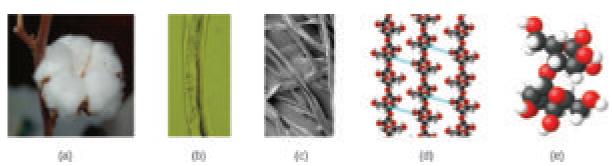

Figure 5. *These images provide an increasingly closer view: (a) a cotton boll, (b) a single cotton fiber viewed under an optical microscope (magnified 40 times), (c) an image of a cotton fiber obtained with an electron microscope (much higher magnification than with the optical microscope); and (d and 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.

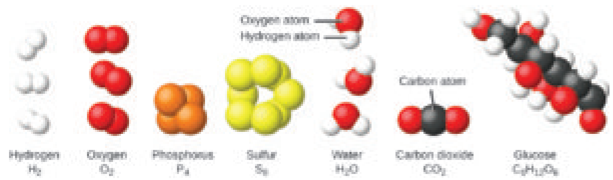

Figure 6. *The elements hydrogen, oxygen, phosphorus, and sulfur form molecules consisting of two or more atoms of the same element. The compounds water, carbon dioxide, and glucose consist of combinations of atoms of different elements.*

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).

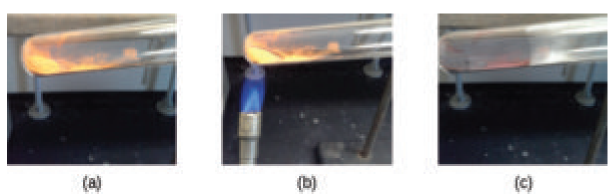

Figure 7. *(a)The compound mercury(II) oxide, (b)when heated, (c) decomposes into silvery droplets of liquid mercury and invisible oxygen gas. (credit: modification of work by Paul Flowers)*

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).

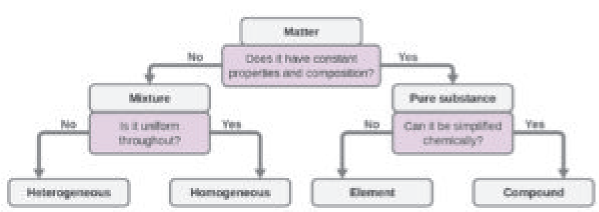

Figure 9. *Depending on its properties, a given substance can be classified as a homogeneous mixture, a heterogeneous mixture, a compound, or an element.*

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).

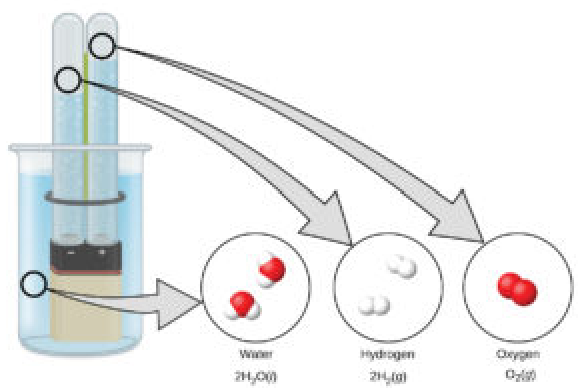

Figure 10. *The decomposition of water is shown at the macroscopic, microscopic, and symbolic levels. The battery provides an electric current (microscopic) that decomposes water. At the macroscopic level, the liquid separates into the gases hydrogen (on the left) and oxygen (on the right). Symbolically, 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.

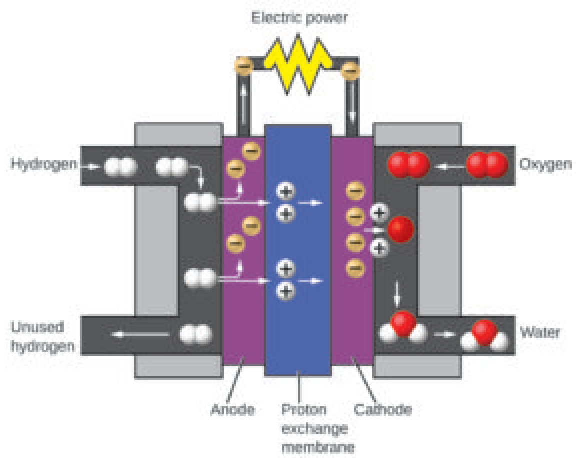

Figure 11 *A fuel cell generates electrical energy from hydrogen and oxygen via an electrochemical process and produces only water as the waste product.*

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.

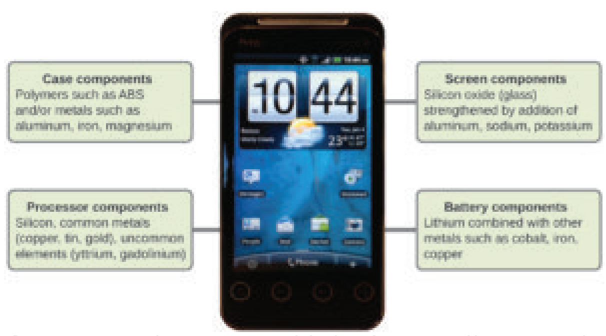

Figure 12 *Almost one-third of naturally occurring elements are used to make a cell phone. (credit: modification of work by John Taylor)*

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 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⟶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

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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.

(a) (b)

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.

(a) (b)

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*

“Atoma”/Wikimedia Commons)

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).

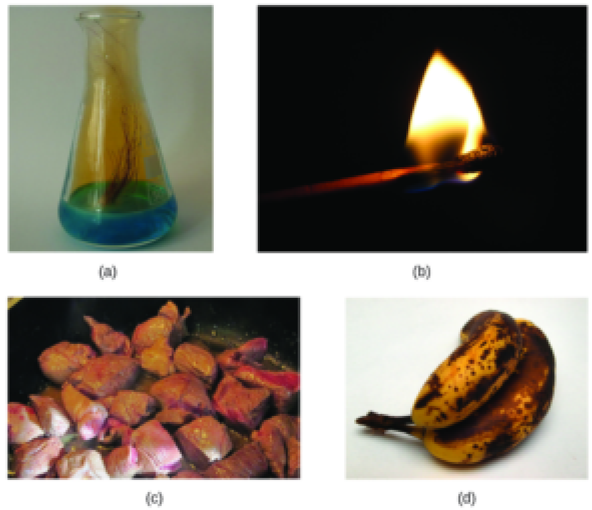


Figure 3 *(a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes 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.

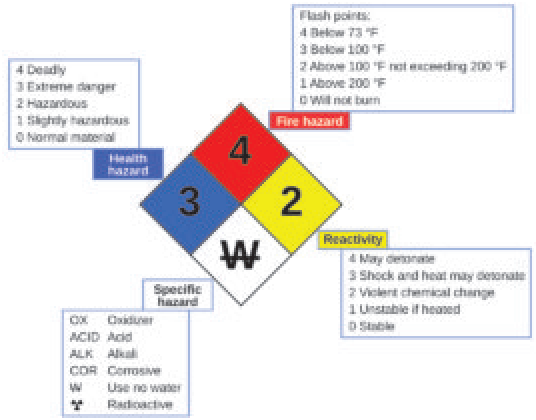


Figure 4 *The National Fire Protection Agency (NFPA) hazard diamond summarizes the major hazards of a chemical substance.*

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.

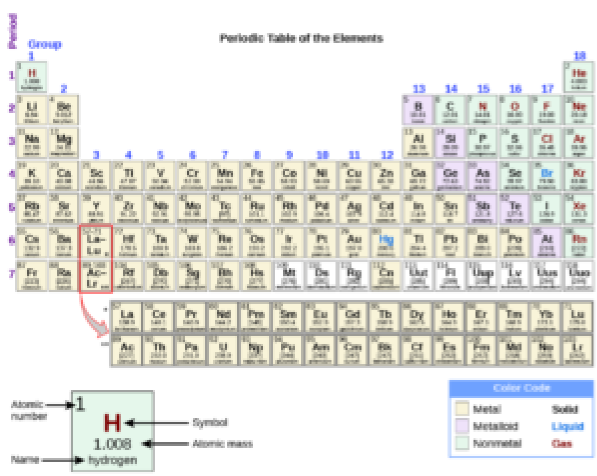


Figure 5 *The periodic table shows how elements may be grouped according to certain similar properties. Note the background color denotes whether an element is a metal, metalloid, or nonmetal, whereas the element symbol color indicates whether it is a solid, liquid, or gas.*

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 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). $\text{density} = \frac{\text{mass}}{\text{volume}} \quad \text{d} = \frac{\text{m}}{\text{v}}$
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

4

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 (µL) = 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).

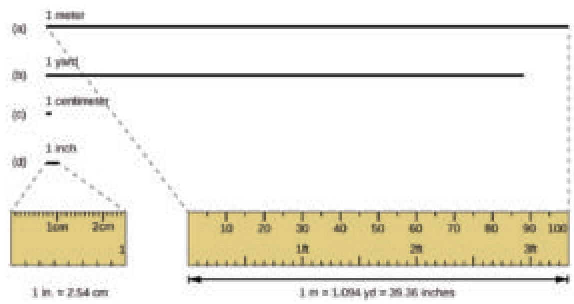


Figure 1 *The relative lengths of 1 m, 1 yd, 1 cm, and 1 in. are shown (not actual size), as well as comparisons of 2.54 cm and 1 in., and of 1 m and 1.094 yd.*

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 **c**ubic **c**entimeter) is often used by health professionals. A cubic centimeter is also called a **milliliter (mL) and is 1/1000 of a liter.**

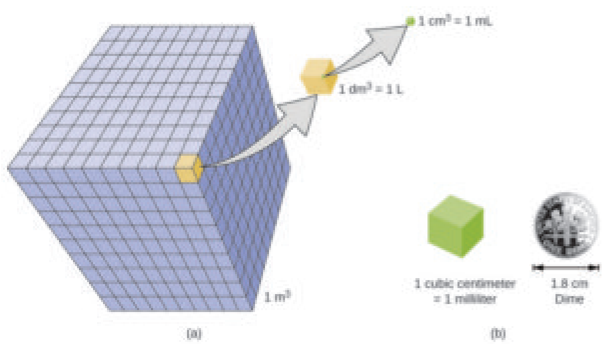


***Figure 3** (a) The relative volumes are shown for cubes of 1 m3, 1 dm3 (1 L), and 1 cm3 (1 mL) (not to scale). (b) The diameter of a dime is compared relative to the edge length of a 1-cm3 (1-mL) cube.(image credit: OpenStax College, CC BY)*

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 as 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.

Table 3

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) 0.60–0.90 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 0.70–0.77 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.

volume of lead cube = 2.00 cm × 2.00 cm × 2.00 cm=8.00 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:

(a) To three decimal places, what is the volume of a cube (cm3) with an edge length of 0.843 cm?
(b) 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/cm^3 (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 (10^6 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 (cm^3 or cc)

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

cubic meter (m^3)

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 ºC

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.

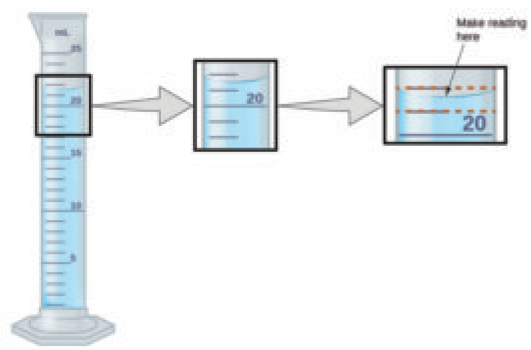
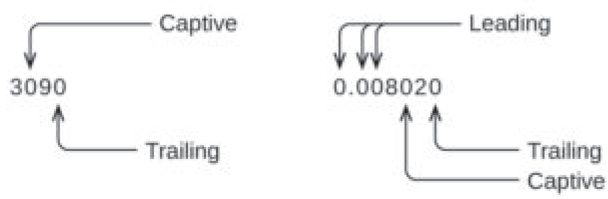


Figure 1 *To measure the volume of liquid in this graduated cylinder, you must mentally subdivide the distance between the 21 and 22 mL marks into tenths of a milliliter, and then make a reading (estimate) at the bottom of the meniscus.*

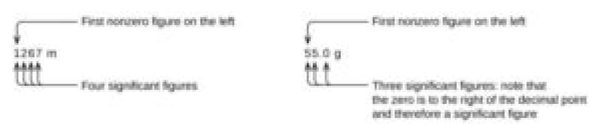
Refer to the illustration in Figure 1. The bottom of the meniscus in this case clearly lies between the 21 and 22 markings, meaning the liquid volume is *certainly* greater than 21 mL but less than 22 mL. The meniscus appears to be a bit closer to the 22-mL mark than to the 21-mL mark, and so a reasonable 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.



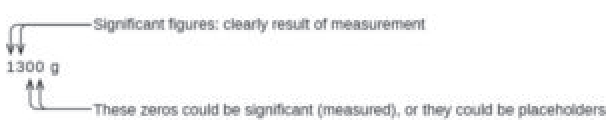
Starting with the first nonzero digit on the left, count this digit and all remaining digits to the right. This is the number of significant figures in the measurement unless the last digit is a trailing zero lying to the left of the decimal point.



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 g + 4.383 g = 5.3853 g
Answer is 5.385 g (round to the thousandths place; three decimal places)
(b) 486 g – 421.23 g = 64.77 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) 0.623 cm x 6.6 cm^2 → result is 4.1 cm^2 (round to two significant figures)
four significant figures x two significant figures → two significant figures answer
(b) $\frac{421.23\,\mathrm{g}}{486\,\mathrm{mL}}$ = 0.86728…g/mL→ result is 0.867 g/mL (round to three significant figures)
$\frac{\text{five significant figures}}{\text{three significant figures}}$ → 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 cm^2
(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:

$V = l \times w \times d$
= 13.44 dm × 5.920 dm × 2.54 dm

= 202.09459...dm^3 (value from calculator)

= 202 dm^3, or 202 L (answer rounded to three significant figures)

Check Your Learning:

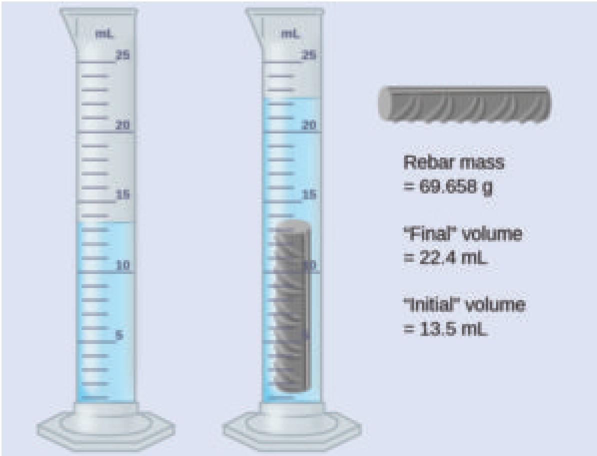
What is the density of a liquid with a mass of 31.1415 g and a volume of 30.13 cm^3?

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:

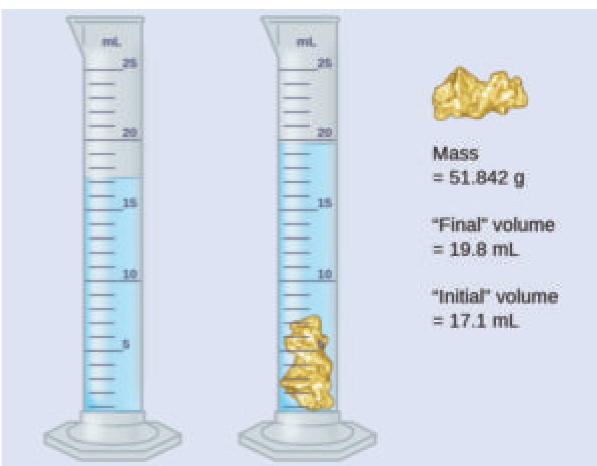
$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{69.658\text{ g}}{8.9\text{ cm}^3} = 7.8\text{g/cm}^3$$

(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.

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 × 95 × 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?

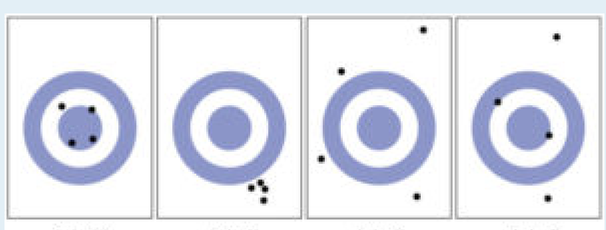


Figure 2 *(a) These arrows are close to both the bull's eye and one another, so they are both accurate and precise. (b) These arrows are close to one another but not on target, so they are precise but not accurate. (c) These arrows are neither on target nor close to one another, so they are neither accurate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

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

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 oz = 125 g × 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 oz = $125\,\text{g} \times \frac{1\,\text{oz}}{28.349\,\text{g}}$
= $\left(\frac{125}{28.349}\right)\,\text{oz}$
= 4.41 oz (three significant figures)

Check Your Learning:

Convert a volume of 9.345 qt to liters.

Answer:

8.844 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:

2.956×10^{-2}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.

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

(b) 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,

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

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

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:

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

The equation relating the temperature scales is then:

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

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

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

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

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

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/°C. Following the same approach, the equations for converting between the kelvin and Celsius temperature scales are derived to be:

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

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

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.

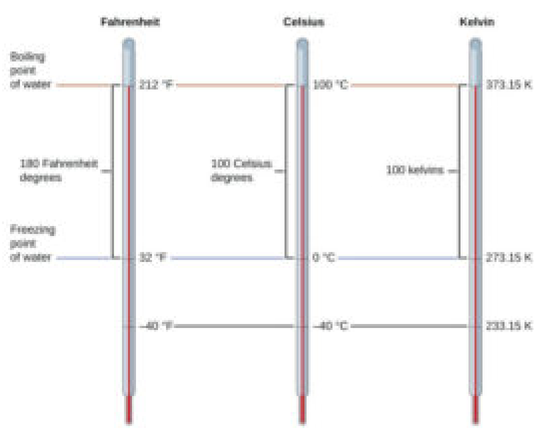


Figure 1 *The Fahrenheit, Celsius, and kelvin temperature scales are compared.*

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 = °C + 273.15 = 37.0 = 273.2 = 310.2 K

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

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:

$^\circ\mathrm{C} = \frac{5}{9}\left(^\circ\mathrm{F} - 32\right) = \frac{5}{9}\left(450 - 32\right) = \frac{5}{9} \times 418 = 232\,^\circ\mathrm{C}$ → set oven to 230 ºC (two significant figures)

K = ºC + 273.15 = 230 +273 =503 K → 5.0 × 10^2K (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

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

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 μL 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 μm
(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 μg = ________ 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 4¼ 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 mi^2, to square kilometers
(d) the volume of 1 gill (exactly 4 oz) to milliliters
(e) the estimated volume of the oceans, 330,000,000 mi^3, 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 cm^2
(d) the displacement volume of an automobile engine, 161 $in.^3$, to liters
(e) the estimated mass of the atmosphere, 5.6 t 10^{15} 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 cm^3has 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 cm^3?

Calculate these masses.

(a) What is the mass of 6.00 cm^3 of mercury, density = 13.5939 g/cm^3?

(b) What is the mass of 25.0 mL octane, density = 0.702 g/cm^3?

Calculate these masses.

(a) What is the mass of 4.00 cm^3 of sodium, density = 0.97 g/cm?

(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/cm^3?

(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/cm^3?

(b) What is the volume of 39.657 g bromine, density = 2.928 g/cm^3?

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.

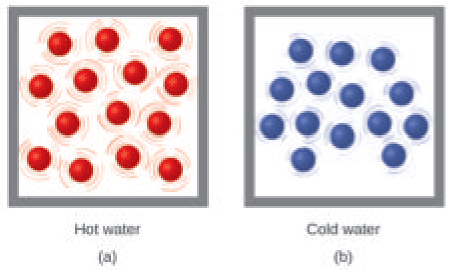


Figure 3 *a) The molecules in a sample of hot water move more rapidly than (b) those in a sample of cold water.*

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.

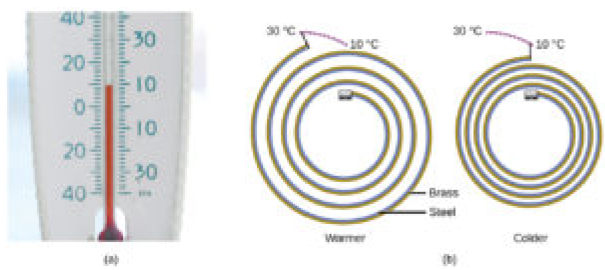


Figure 4 *(a) In an alcohol or mercury thermometer, the liquid (dyed red for visibility) expands when heated and contracts when cooled, much more so than the glass tube that contains the liquid. (b) In a bimetallic thermometer, two different metals (such as brass and steel) form a two-layered strip. 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).

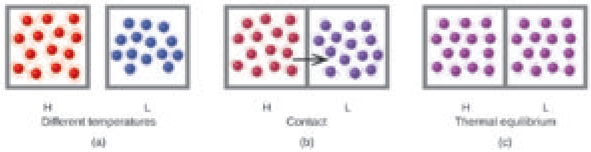


Figure 5 *(a) Substances H and L are initially at different temperatures, and their atoms have different average kinetic energies. (b) When they are put into contact with each other, collisions between the molecules result in the transfer of kinetic (thermal) energy from the hotter to the cooler matter. (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 m^2/s^2, 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:

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

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:

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

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):

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

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:

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

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:

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

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	$He(g)$	5.193
water	$H_2O(l)$	4.184
ethanol	$C_2H_6O(l)$	2.376
ice	$H_2O(s)$	2.093 (at −10 °C)
water vapor	$H_2O(g)$	1.864
nitrogen	$N_2(g)$	1.040
air		1.007
oxygen	$O_2(g)$	0.918
aluminum	$Al(s)$	0.897
carbon dioxide	$CO_2(g)$	0.853
argon	$Ar(g)$	0.522
iron	$Fe(s)$	0.449
copper	$Cu(s)$	0.385
lead	$Pb(s)$	0.130
gold	$Au(s)$	0.129
silicon	$Si(s)$	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 = (specific heat) × (mass of substance) × (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 ×× 102 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:
$q = c \times m \times \Delta T = c \times m \times (T_{final} - T_{initial})$
$= (4.184\ J/\cancel{g}°C) \times (800\ \cancel{g}) \times (85 - 20)°C$
$= (4.184\ J/\cancel{g}\cancel{°C}) \times (800\ \cancel{g}) \times (65)\cancel{°C}$
$=210{,}000\ J\ (=210\ kJ)$
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:

5.05×10^4 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:
$q = c \times m \times \Delta T = c \times m \times (T_{final} - T_{initial})$
Substituting the known values:
$6640\ J = c \times (348\ g) \times (43.6 - 22.4)°C$
Solving:
$c = \frac{6640\ J}{(348\ g)\times(21.2°C)} = 0.900\ J/g°C$
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 J/g °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.

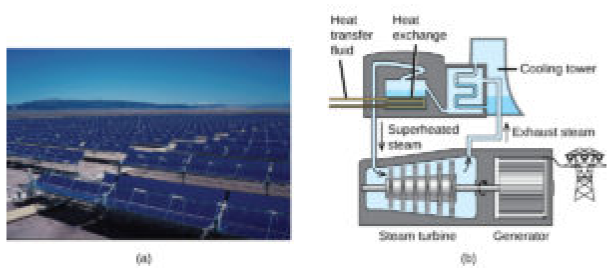


Figure 8 *This solar thermal plant uses parabolic trough mirrors to concentrate sunlight. (credit a: modification of work by Bureau of Land Management)*

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 °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 ⟶ 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 J = 1 kg m^2/s and 4.184 J = 1 cal

kinetic energy

energy of a moving body, in joules, equal to 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.

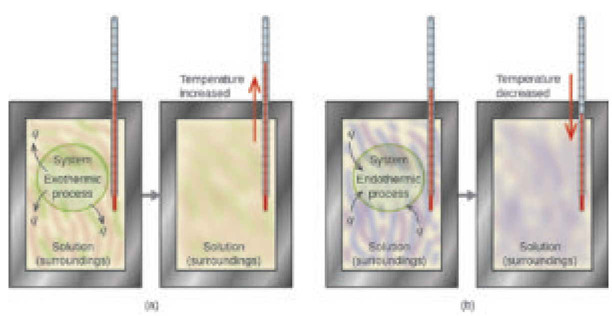


Figure 1 *In a calorimetric determination, either (a) an exothermic process occurs and heat, q, is negative, indicating that thermal energy is transferred from the system to its surroundings, or (b) an endothermic process occurs and heat, q, is positive, indicating that thermal energy is transferred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

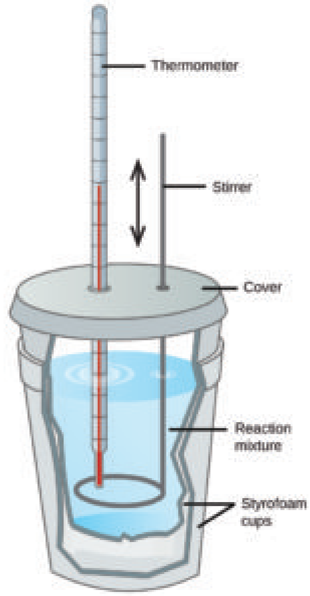


Figure 2 *A simple calorimeter can be constructed from two polystyrene cups. A thermometer and stirrer extend through the cover into the reaction mixture.*

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).

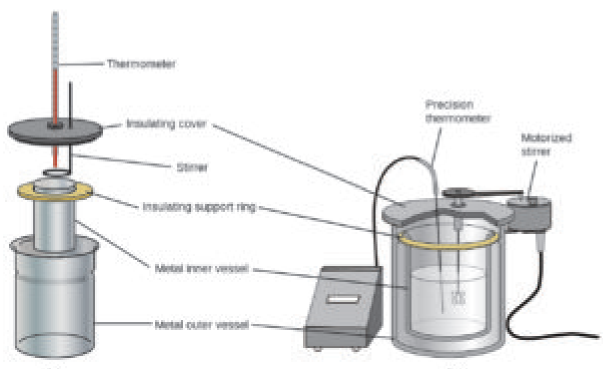


Figure 3 *Commercial solution calorimeters range from (a) simple, inexpensive models for student use to (b) expensive, more accurate models for industry and research.*

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:

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

The magnitude of the heat (change) is therefore the same for both substances, and the negative sign merely shows that $q_{substance\ M}$ and $q_{substance\ W}$ 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, $q_{substance\ M}$ is a negative value and $q_{substance\ W}$ is positive, since heat is transferred from M to W.

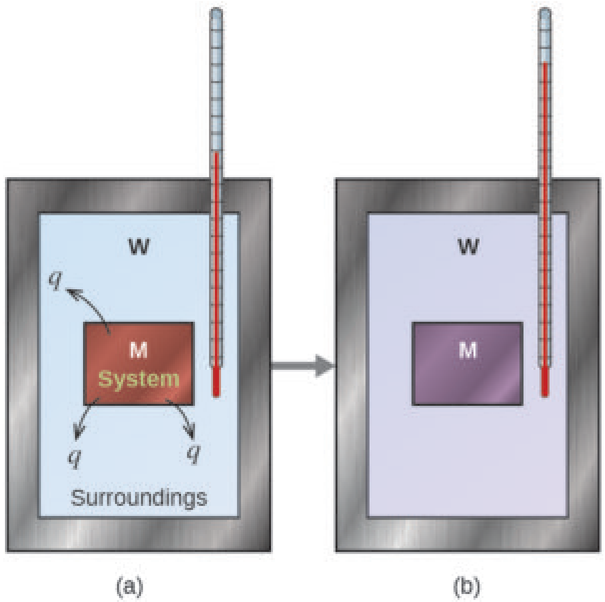


Figure 4 *In a simple calorimetry process, (a) heat, q, is transferred from the hot metal, M, to the cool water, W, until (b) both are at the same temperature.*

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:

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

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

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

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

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

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:

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

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

Solving this gives $T_{\text{i,rebar}}$= 248 °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:
$c_{\text{metal}} \times m_{\text{metal}} \times (T_{\text{f,metal}} - T_{\text{i, metal}}) = -c_{\text{water}} \times m_{\text{water}} \times (T_{\text{f,water}} - \text{T}_{\text{i,water}})$
Noting that since the metal was submerged in boiling water, its initial temperature was 100.0 °C; and that for water, 60.0 mL = 60.0 g; we have:
$(c_{\text{metal}})(59.7\text{g})(28.5°\text{C}-100.0°\text{C}) = -(4.18\text{J/g °C})(60.0\text{g})(28.5°\text{C} - 22.0°\text{C})$
Solving this:
$c_{\text{metal}} = -(4.184\text{J/g °C})(60.0\text{g})(6.5°\text{C})(59.7\text{g})(-71.5°\text{C}) = 0.38\text{J/g °C}$
Comparing this with values in [link], our experimental specific heat is closest to the value for copper (0.39 J/g °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_{metal}= 0.13 J/g °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:

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

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

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

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?
HCl(*aq*) + NaOH(*aq*) ⟶ NaCl(aq) + H2O(*l*)

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:
$q_{reaction} = -q_{solution}$
(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:
$q_{solution} = (c \times m \times \Delta T)_{solution}$
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:
$q_{solution} = (4.184 \text{J/g °C})(1.0 \times 10^2 \text{g})(28.9\text{°C} - 22.0\text{°C}) = 2.89 \times 10^3 \text{J}$
Finally, since we are trying to find the heat of the reaction, we have:
$q_{reaction} = -q_{solution} = -2.89 \times 10^3 \text{J}$
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 $AgNO_3$(*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 NaC2H3O2 (sodium acetate) and a metal disc. Bending the disk creates nucleation sites around which the metastable NaC2H3O2 quickly crystallizes (a later chapter on solutions will investigate saturation and supersaturation in more detail).
The process NaC2H3O2(aq)⟶NaC2H3O2(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 NaC2H3O2 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 $2Fe(s) + 3/2O_2(g) \longrightarrow Fe_2O_3(s)$. 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 NH4NO3 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:

$q_{rxn} = -q_{soln}$
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:

$q_{rxn} = -q_{soln} = -(c \times m \times \Delta T)_{soln}$
$= -[(4.184 J/g\ °C) \times (53.2 g) \times (20.3 °C - 24.9 °C)]$
$= -[(4.184 J/g\ °C) \times (53.2 g) \times (-4.6 °C)]$
$+1.0 \times 10^3 J = +1.0 kJ$
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.

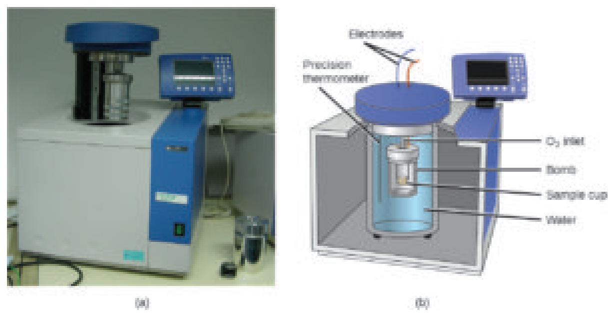


Figure 7 *(a) A bomb calorimeter is used to measure heat produced by reactions involving gaseous reactants or products, such as combustion. (b) The reactants are contained in the gas-tight “bomb,” which is submerged in water and surrounded by insulating materials. (credit a: modification of work by “Harbor1”/Wikimedia 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:
$q_{rxn} = -(q_{water} + q_{bomb})$
$= -[(4.184 J/g°C) \times (775\ g) \times (35.6°C - 23.8°C) + 893 J/°C \times (35.6°C - 23.8°C)]$
$= -(38{,}300 J + 10{,}500 J)$
$= -48{,}800\ J = -48.8\ kJ$
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, C6H6, 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).

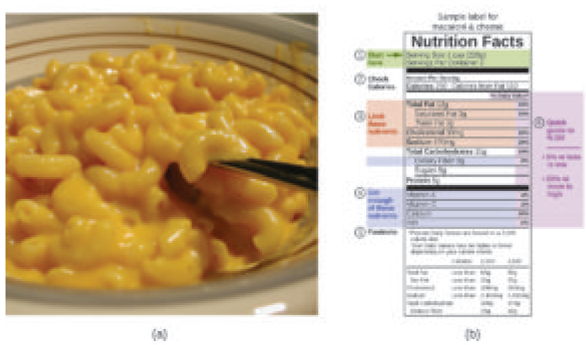


Figure 8 *(a) Macaroni and cheese contain energy in the form of the macronutrients in the food. (b) The food's nutritional information is shown on the package label. In the US, the energy content is given in Calories (per serving); the rest of the world usually uses kilojoules. (credit a: modification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

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

(5 g protein × 4 Calories/g) + (31 g carb × 4Calories/g) + (12 g fat × 9 Calories/g) = 252Calories
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 $AgNO_3$(*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:

$Ba(OH)_2{\cdot}8H_2O(s) + 2NH_4SCN(aq) \longrightarrow Ba(SCN)_2(aq) + 2NH_3(aq) + 10H_2O(l)$

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 NH_4NO_3 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, $C_6H_{12}O_6$(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), $C_7H_5N_2O_6$, 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 (2.000×10^3 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 1.1×10^3 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

9

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).

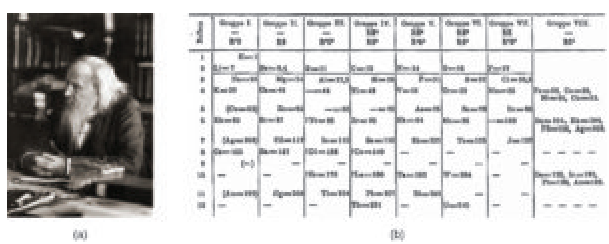


Figure 1 *(a) Dimitri Mendeleev is widely credited with creating (b) the first periodic table of the elements. (credit a: modification of work by Serge Lachinov; credit b: modification of work by "Den fjättrade ankan"/Wikimedia Commons)*

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.

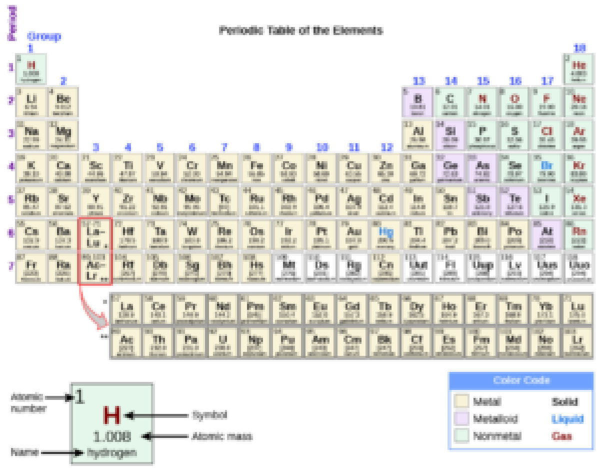


Figure 2 *Elements in the periodic table are organized according to their properties.*

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.

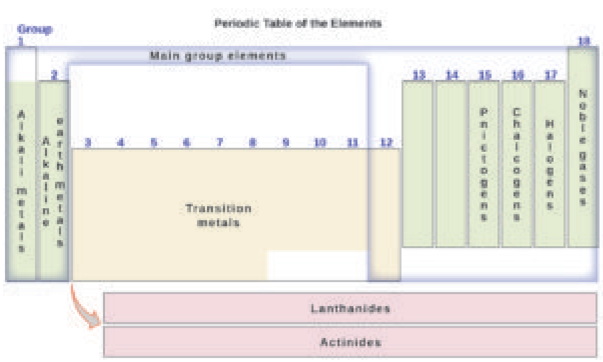


Figure 3 *The periodic table organizes elements with similar properties into groups.*

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 and 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

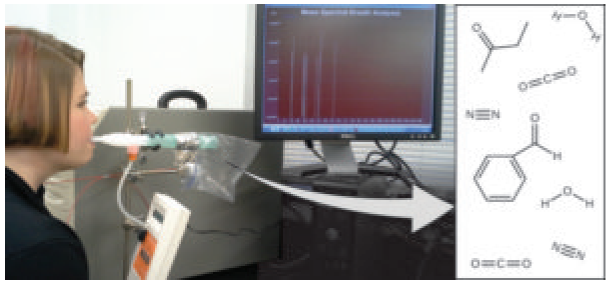


Figure 1 *Analysis of molecules in an exhaled breath can provide valuable information, leading to early diagnosis of diseases or detection of environmental exposure to harmful substances. (credit: modification of work by Paul Flowers)*

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

11

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 amu = 1.6605×10^{-24} 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 = 1.602×10^{-19} 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 = number of neutrons.

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

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).

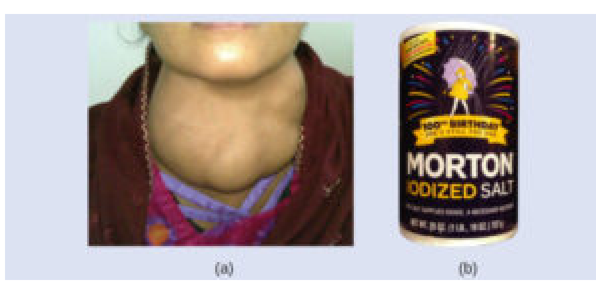


Figure 2 *(a) Insufficient iodine in the diet can cause an enlargement of the thyroid gland called a goiter. (b) The addition of small amounts of iodine to salt, which prevents the formation of goiters, has helped eliminate this concern in the US where salt consumption is high. (credit a: modification 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 *ferrum*)
bromine	Br	lead	Pb (from *plumbum*)

Some Common Elements and Their Symbols			
Element	**Symbol**	**Element**	**Symbol**
calcium	Ca	magnesium	Mg
carbon	C	mercury	Hg (from *hydrargyrum*)
chlorine	Cl	nitrogen	N
chromium	Cr	oxygen	O
cobalt	Co	potassium	K (from *kalium*)
copper	Cu (from *cuprum*)	silicon	Si
fluorine	F	silver	Ag (from*argentum*)
gold	Au (from *aurum*)	sodium	Na (from *natrium*)
helium	He	sulfur	S
hydrogen	H	tin	Sn (from *stannum*)
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:
(a) the ion with a 1+ charge, atomic number 55, and mass number 133
(b) 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 112112 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 ×× 10−19 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

12

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.

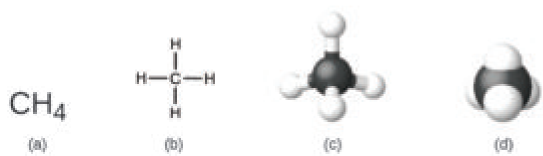


Figure 1 *A methane molecule can be represented as (a) a molecular formula, (b) a structural formula, (c) a ball-and-stick model, and (d) a space-filling model. Carbon and hydrogen atoms are represented by black and white spheres, respectively.*

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).

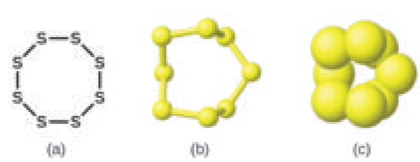


Figure 2 *A molecule of sulfur is composed of eight sulfur atoms and is therefore written as S8. It can be represented as (a) a structural formula, (b) a ball-and-stick model, and (c) a space-filling model. Sulfur atoms are represented by yellow spheres.*

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).

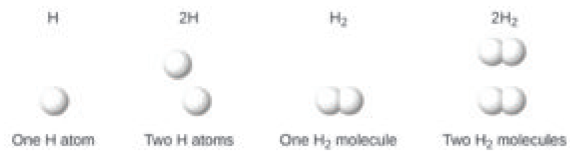


Figure 3 *The symbols H, 2H, H_2, and $2H_2$ represent very different entities.*

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).

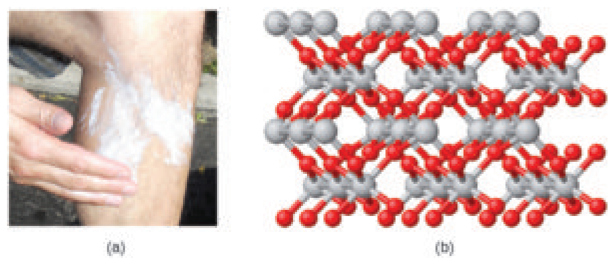


Figure 4 *(a) The white compound titanium dioxide provides effective protection from the sun. (b) A crystal of titanium dioxide, TiO2, contains titanium and oxygen in a ratio of 1 to 2. The titanium atoms are gray and the oxygen atoms are red. (credit a: modification of work by "osseous"/Flickr)*

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 $C_2H_4O_2$. 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.

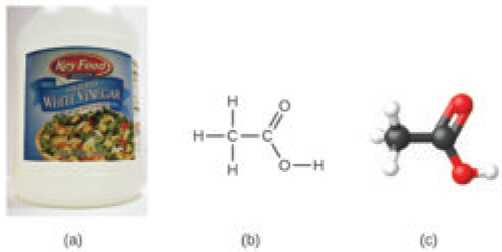


Figure 6 *(a) Vinegar contains acetic acid, C2H4O2, which has an empirical formula of CH2O. It can be represented as (b) a structural formula and (c) as a ball-and-stick model. (credit a: modification of work by "HomeSpot HQ"/Flickr)*

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 C6H12O6 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, $C_8H_{16}O_4$; empirical formula, C_2H_4O

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.

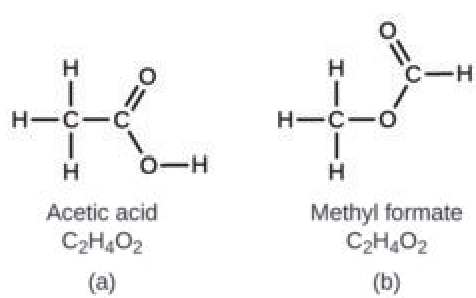


Figure 8 *Molecules of (a) acetic acid and methyl formate (b) are structural isomers; they have the same formula*

($C_2H_4O_2$) 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.

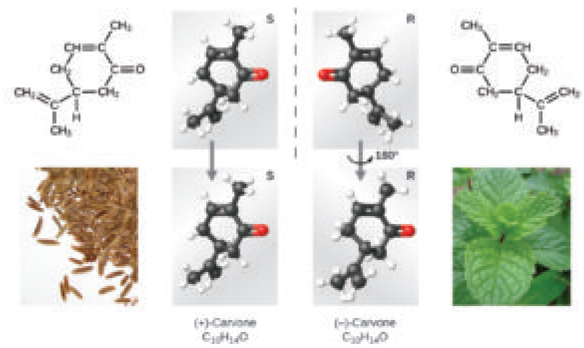


***Figure 9** Molecules of carvone are spatial isomers; they only differ in the relative orientations of the atoms in space. (credit bottom left: modification of work by "Miansari66"/Wikimedia Commons; credit bottom right: modification of work by Forest & Kim Starr)*

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

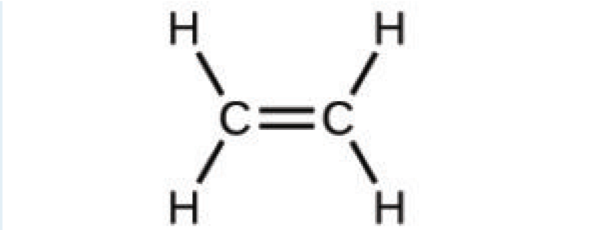
Explain why the symbol for an atom of the element oxygen and the formula for a molecule of oxygen differ.
Explain why the symbol for the element sulfur and the formula for a molecule of sulfur differ.

Write the molecular and empirical formulas of the following compounds:

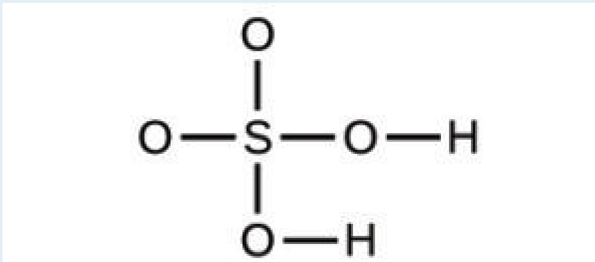
(a)

(b)

(c)

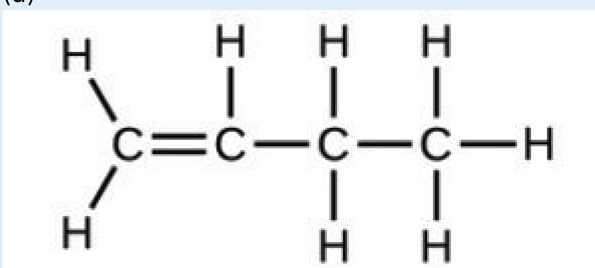


(d)

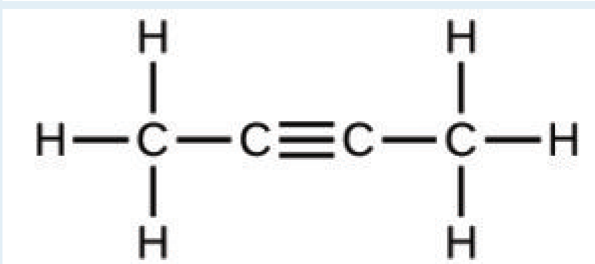


Write the molecular and empirical formulas of the following compounds:

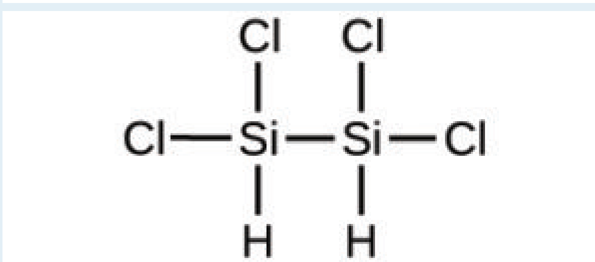
(a)



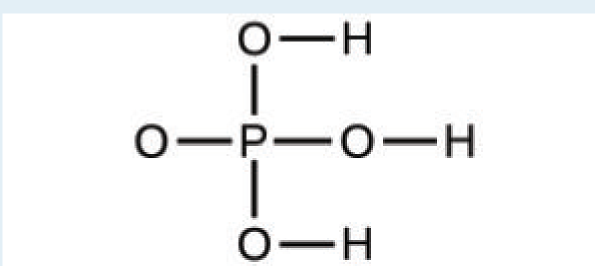
(b)



(c)



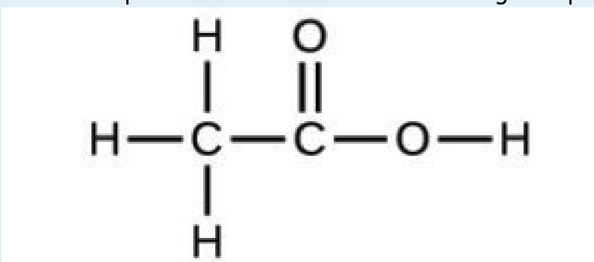
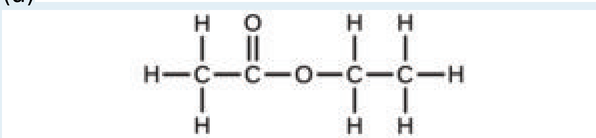
(d)



Determine the empirical formulas for the following compounds:

(a) caffeine, $C_8H_{10}N_4O_2$

(b) fructose, $C_{12}H_{22}O_{11}$

(c) hydrogen peroxide, H_2O_2
(d) glucose, $C_6H_{12}O_6$
(e) ascorbic acid (vitamin C), $C_6H_8O_6$
Determine the empirical formulas for the following compounds:
(a) acetic acid, $C_2H_4O_2$
(b) citric acid, $C_6H_8O_7$
(c) hydrazine, N_2H_4
(d) nicotine, $C_{10}H_{14}N_2$
(e) butane, C_4H_{10}
Write the empirical formulas for the following compounds:
(a)  (b)

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

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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).

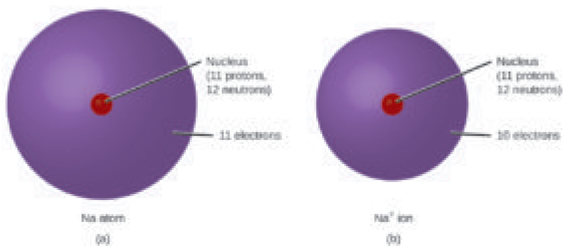

Figure 1 *(a) A sodium atom (Na) has equal numbers of protons and electrons (11) and is uncharged. (b) A sodium cation (Na+) has lost an electron, so it has one more proton (11) than electrons (10), giving it an overall positive charge, signified by a superscripted plus sign.*

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 Ca^{2+}. The name of a metal ion is the same as the name of the metal atom from which it forms, so Ca^{2+} 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.

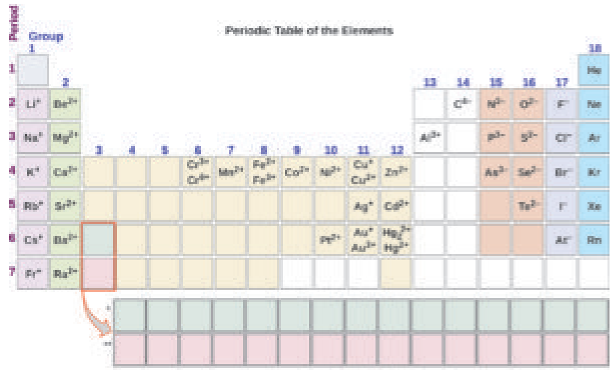

Figure 2 *Some elements exhibit a regular pattern of ionic charge when they form ions.*

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 N^{3-}, 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+: Al^{3+}, an aluminum ion. Carbon will form an anion with a charge of 4−: C^{4-}, 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.

Table 1

Common Polyatomic Ions

Charge	Name	Formula
1+	ammonium	NH_4^+
1−	acetate	$C_2H_3O_2^-$
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).

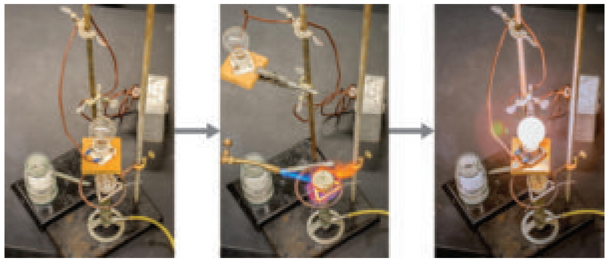

Figure 3 *Sodium chloride melts at 801 °C and conducts electricity when molten. (credit: modification of work by Mark Blaser and Matt Evans)*

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:

Na_2S

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 $Ca_3(PO_4)_2$. 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 $Ca(H_2PO_4)_2$.

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:

Li_2O_2

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 $C_2O_4^{2-}$ ions combined in a 2:1 ratio, and its formula is written as $Na_2C_2O_4$. 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, $C_2O_4^{2-}$.

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) $Al_2(SO_4)_3$

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 charge 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) $(NH_4)_2CO_3$
(d) $Sr(H_2PO_4)_2$
(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) $MgC_2H_3O_2$
(c) H_2S
(d) Ag_2S
(e) N_2Cl_4
(f) $Co(NO_3)_2$
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

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Cooking Techniques: Thickening and Concentrating Flavors

14

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.
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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 jellying 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.
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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.
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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

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

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 40ºC (100ºF) and will not melt until it reaches 85ºC (185ºF).

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.
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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.
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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.