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{{ مكتبة A to Z }}

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CHAPTER

1

THE SCIENCE OF CHEMISTRY

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For one weekend, an ice rink in Tacoma, Washington became a work of art. Thousands of people came to see the amazing collection of ice and lights on display. Huge blocks of ice, each having a mass of about 136 kg, were lit from the inside by lights. The glowing gas in each light made the solid ice shine with color. And as you can see, lights of many different colors were used in the display. In this chapter, you will learn about matter. You will learn about the properties used to describe matter. You will also learn about the changes matter can undergo. Finally, you will learn about classifying matter based on its properties.

START-UP ACTIVITY

SAFETY PRECAUTIONS

Classifying Matter

PROCEDURE

1. Examine the **objects** provided by your teacher.
2. Record in a table observations about each object's individual characteristics.
3. Divide the objects into at least three different categories based on your observations. Be sure that the objects in each category have something in common.

ANALYSIS

1. Describe the basis of your classification for each category you created.
2. Give an example that shows how using these categories makes describing the objects easier.
3. Describe a system of categories that could be used to classify matter. Explain the basis of your categories.

Pre-Reading Questions

1. **Do you think there are "good chemicals" and "bad chemicals"? If so, how do they differ?**
2. **What are some of the classifications of matter?**
3. **What is the difference between a chemical change and a physical change?**

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SECTION 1

What Is Chemistry?

Key Terms

- **chemical**
- **chemical reaction**
- **states of matter**
- **reactant**
- **product**

Objectives

1. **Describe** ways in which chemistry is a part of your daily life.
2. **Describe** the characteristics of three common states of matter.
3. **Describe** physical and chemical changes, and give examples of each.
4. **Identify** the reactants and products in a chemical reaction.
5. **List** four observations that suggest a chemical change has occurred.

Working with the Properties and Changes of Matter

Do you think of chemistry as just another subject to be studied in school? Or maybe you feel it is important only to people working in labs? The effects of chemistry reach far beyond schools and labs. It plays a vital role in your daily life and in the complex workings of your world.

Look at **Figure 1.** Everything you see, including the clothes the students are wearing and the food the students are eating, is made of chemicals. The students themselves are made of chemicals! Even things you cannot see, such as air, are made up of chemicals.

Chemistry is concerned with the properties of chemicals and with the changes chemicals can undergo. A **chemical** is any substance that has a definite composition—it's always made of the same stuff no matter where the chemical comes from. Some chemicals, such as water and carbon dioxide, exist naturally. Others, such as polyethylene, are manufactured. Still others, such as aluminum, are taken from natural materials.

chemical
any substance that has a defined composition

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Figure 1
Chemicals make up everything you see every day.

You Depend on Chemicals Every Day

Many people think of chemicals in negative terms—as the cause of pollution, explosions, and cancer. Some even believe that chemicals and chemical additives should be banned. But just think what such a ban would mean—after all, everything around you is composed of chemicals. Imagine going to buy fruits and vegetables grown without the use of any chemicals at all. Because water is a chemical, the produce section would be completely empty! In fact, the entire supermarket would be empty because all foods are made of chemicals.

The next time you are getting ready for school, look at the list of ingredients in your shampoo or toothpaste. You'll see an impressive list of chemicals. Without chemicals, you would have nothing to wear. The fibers of your clothing are made of chemicals that are either natural, such as cotton or wool, or synthetic, such as polyester. The air you breathe, the food you eat, and the water you drink are made up of chemicals. The paper, inks, and glue used to make the book you are now reading are chemicals, too. You yourself are an incredibly complex mixture of chemicals.

Chemical Reactions Happen All Around You

You will learn in this course that changes in chemicals—or **chemical reactions**—are taking place around you and inside you. Chemical reactions are necessary for living things to grow and for dead things to decay. When you cook food, you are carrying out a chemical reaction. Taking a photograph, striking a match, switching on a flashlight, and starting a gasoline engine require chemical reactions.

chemical reaction

the process by which one or more substances change to produce one or more different substances

Using reactions to manufacture chemicals is a big industry. **Table 1** lists the top eight chemicals made in the United States. Some of these chemicals may be familiar, and some you may have never heard of. By the end of this course, you will know a lot more about them. Chemicals produced on a small scale are important, too. Life-saving antibiotics, cancer-fighting drugs, and many other substances that affect the quality of your life are also products of the chemical industry.

Table 1 **Top Eight Chemicals Made in the United States (by Weight)**

Rank	Name	Formula	Uses
1	sulfuric acid	H_2SO_4	production of fertilizer; metal processing; petroleum refining
2	ethene	C_2H_4	production of plastics; ripening of fruits
3	propylene	C_3H_6	production of plastics
4	ammonia	NH_3	production of fertilizer; refrigeration
5	chlorine	Cl_2	bleaching fabrics; purifying water; disinfectant
6	phosphoric acid (anhydrous)	P_2O_5	production of fertilizer; flavoring agent; rustproofing metals
7	sodium hydroxide	NaOH	petroleum refining; production of plastics
8	1,2-dichloroethene	$C_2H_2Cl_2$	solvent, particularly for rubber

Source: ***Chemical and Engineering News.***

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Physical States of Matter

states of matter

the physical forms of matter, which are solid, liquid, gas, and plasma

All matter is made of particles. The type and arrangement of the particles in a sample of matter determine the properties of the matter. Most of the matter you encounter is in one of three **states of matter:** solid, liquid, or gas. **Figure 2** illustrates water in each of these three states at the macroscopic and microscopic levels. *Macroscopic* refers to what you see with the unaided eye. In this text, *microscopic* refers to what you would see if you could see individual atoms.

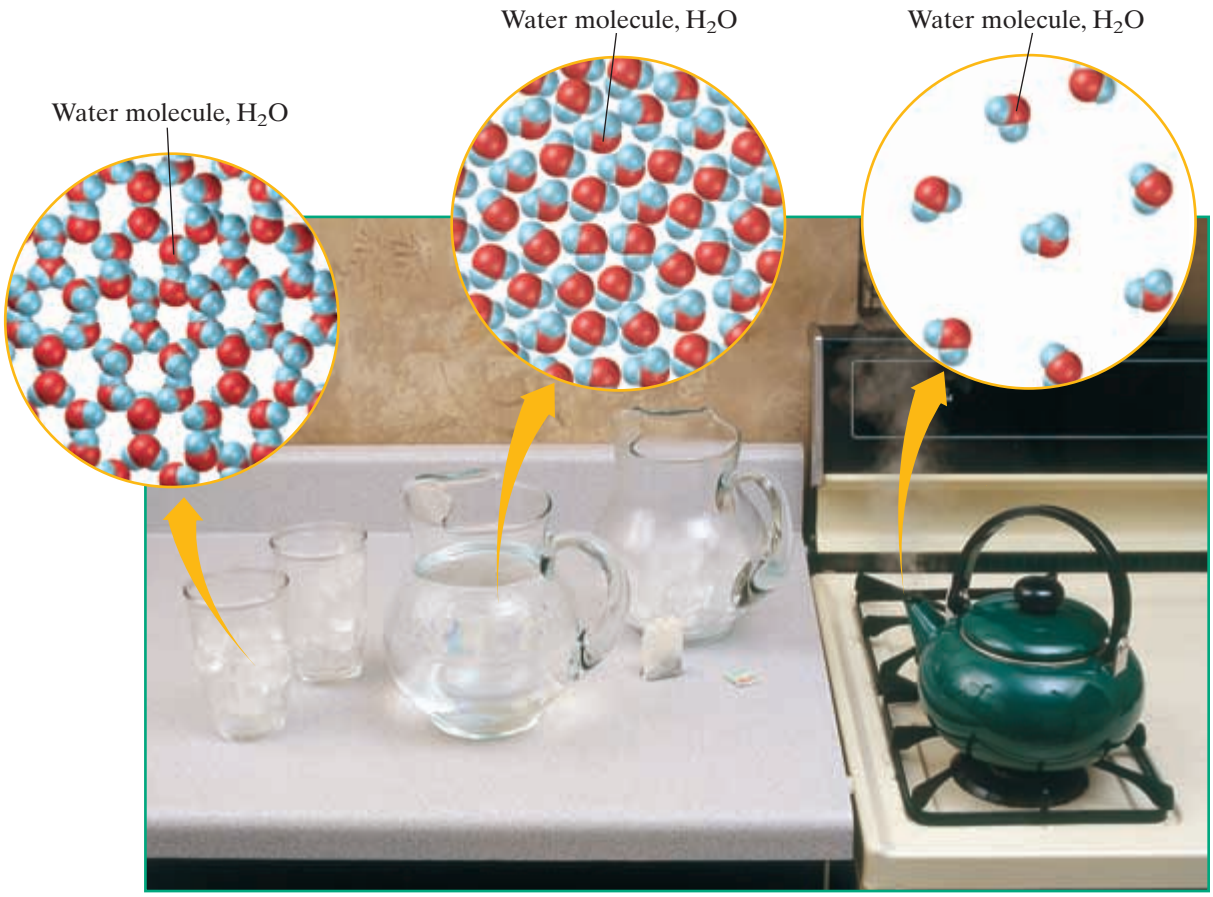
The microscopic views in this book are models that are designed to show you the differences in the arrangement of particles in different states of matter. They also show you the differences in size, shape, and makeup of particles of chemicals. But don't take these models too literally. Think of them as cartoons. Atoms are not really different colors. And groups of connected atoms, or molecules, do not look lumpy. The microscopic views are also limited in that they often show only a single layer of particles whereas the particles are really arranged in three dimensions. Finally, the models cannot show you that particles are in constant motion.

Figure 2

a Below 0°C, water exists as ice. Particles in a solid are in a rigid structure and vibrate in place.

b Between 0°C and 100°C, water exists as a liquid. Particles in a liquid are close together and slide past one another.

c Above 100°C, water is a gas. Particles in a gas move randomly over large distances.



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Properties of the Physical States

Solids have fixed volume and shape that result from the way their particles are arranged. Particles that make up matter in the solid state are held tightly in a rigid structure. They vibrate only slightly.

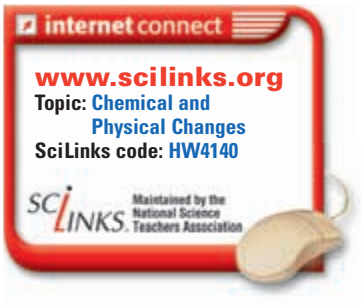
Liquids have fixed volume but not a fixed shape. The particles in a liquid are not held together as strongly as those in a solid. Like grains of sand, the particles of a liquid slip past one another. Thus, a liquid can flow and take the shape of its container.

Gases have neither fixed volume nor fixed shape. Gas particles weakly attract one another and move independently at high speed. Gases will fill any container they occupy as their particles move apart.

There are other states that are beyond the scope of this book. For example, most visible matter in the universe is plasma—a gas whose particles have broken apart and are charged. Bose-Einstein condensates have been described at very low temperatures. A neutron star is also considered by some to be a state of matter.

Changes of Matter

Many changes of matter happen. An ice cube melts. Your bicycle's spokes rust. A red shirt fades. Water fogs a mirror. Milk sours. Scientists who study these and many other events classify them by two broad categories: *physical changes* and *chemical changes*.



Physical Changes

Physical changes are changes in which the identity of a substance doesn't change. However, the arrangement, location, and speed of the particles that make up the substance may change. Changes of state are physical changes. The models in **Figure 2** show that when water changes state, the arrangement of particles changes, but the particles stay water particles. As sugar dissolves in the tea in **Figure 3,** the sugar molecules mix with the tea, but they don't change what they are. The particles are still sugar. Crushing a rock is a physical change because particles separate but do not change identity.

Figure 3
Dissolving sugar in tea is a physical change.

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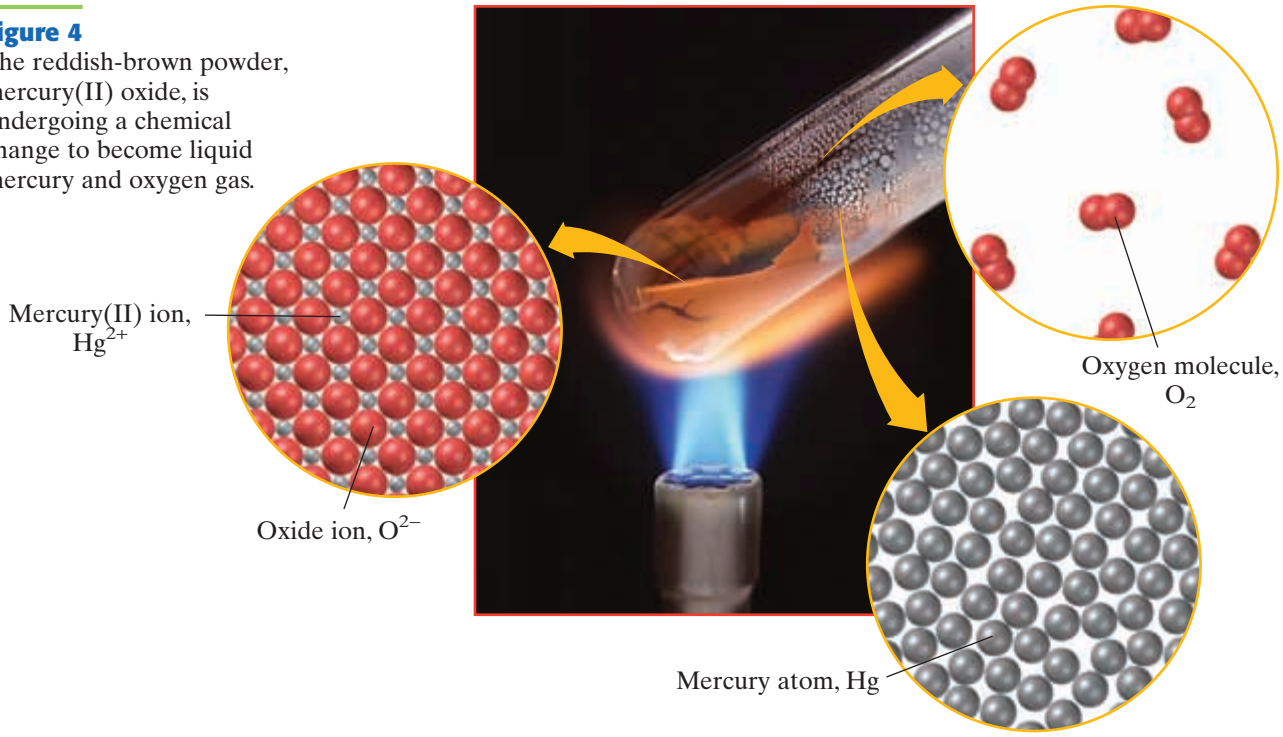

Figure 4
The reddish-brown powder, mercury(II) oxide, is undergoing a chemical change to become liquid mercury and oxygen gas.

Chemical Changes

In a chemical change, the identities of substances change and new substances form. In **Figure 4,** mercury(II) oxide changes into mercury and oxygen as represented by the following word equation:

$$\text{mercury(II) oxide} \longrightarrow \text{mercury} + \text{oxygen}$$

In an equation, the substances on the left-hand side of the arrow are the **reactants.** They are used up in the reaction. Substances on the right-hand side of the arrow are the **products.** They are made by the reaction.

A chemical reaction is a rearrangement of the atoms that make up the reactant or reactants. After rearrangement, those same atoms are present in the product or products. Atoms are not destroyed or created, so mass does not change during a chemical reaction.

reactant
a substance or molecule that participates in a chemical reaction

product
a substance that forms in a chemical reaction

Evidence of Chemical Change

Evidence that a chemical change may be happening generally falls into one of the categories described below and shown in **Figure 5.** The more of these signs you observe, the more likely a chemical change is taking place. But be careful! Some physical changes also have one or more of these signs.

a. The Evolution of a Gas The production of a gas is often observed by bubbling, as shown in **Figure 5a,** or by a change in odor.

b. The Formation of a Precipitate When two clear solutions are mixed and become cloudy, a precipitate has formed, as shown in **Figure 5b.**

c. The Release or Absorption of Energy A change in temperature or the giving off of light energy, as shown in **Figure 5c,** are signs of an energy transfer.

d. A Color Change in the Reaction System Look for a different color when two chemicals react, as shown in **Figure 5d.**

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Figure 5

a When acetic acid, in vinegar, and sodium hydrogen carbonate, or baking soda, are mixed, the solution bubbles as carbon dioxide forms.

b When solutions of sodium sulfide and cadmium nitrate are mixed, cadmium sulfide, a solid precipitate, forms.

c When aluminum reacts with iron(III) oxide in the clay pot, energy is released as heat and light.

d When phenolphthalein is added to ammonia dissolved in water, a color change from colorless to pink occurs.

1 Section Review

UNDERSTANDING KEY IDEAS

1. Name three natural chemicals and three artificial chemicals that are part of your daily life.
2. Describe how chemistry is a part of your morning routine.
3. Classify the following materials as solid, liquid, or gas at room temperature: milk, helium, granite, oxygen, steel, and gasoline.
4. Describe the motions of particles in the three common states of matter.
5. How does a physical change differ from a chemical change?
6. Give three examples of physical changes.
7. Give three examples of chemical changes.
8. Identify each substance in the following word equation as a reactant or a product.

   limestone $\xrightarrow{heat}$ lime + carbon dioxide

9. Sodium salicylate is made from carbon dioxide and sodium phenoxide. Identify each of these substances as a reactant or a product.
10. List four observations that suggest a chemical change is occurring.

CRITICAL THINKING

11. Explain why neither liquids nor gases have permanent shapes.
12. Steam is sometimes used to melt ice. Is this change physical or chemical?
13. Mass does not change during a chemical change. Is the same true for a physical change? Explain your answer, and give an example.
14. In beaker A, water is heated, bubbles of gas form throughout the water, and the water level in the beaker slowly decreases. In beaker B, electrical energy is added to water, bubbles of gas appear on the ends of the wires in the water, and the water level in the beaker slowly decreases.
    a. What signs of a change are visible in each situation?
    b. What type of change is happening in each beaker? Explain your answer.

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SECTION 2

Describing Matter

KEY TERMS

- **matter**
- **volume**
- **mass**
- **weight**
- **quantity**
- **unit**
- **conversion factor**
- **physical property**
- **density**
- **chemical property**

OBJECTIVES

1. **Distinguish** between different characteristics of matter, including mass, volume, and weight.
2. **Identify** and use SI units in measurements and calculations.
3. **Set up** conversion factors, and use them in calculations.
4. **Identify** and describe physical properties, including density.
5. **Identify** chemical properties.

Matter Has Mass and Volume

Matter, the stuff of which everything is made, exists in a dazzling variety of forms. However, matter has a fairly simple definition. **Matter** is anything that has mass and volume. Think about blowing up a balloon. The inflated balloon has more mass and more volume than before. The increase in mass and volume comes from the air that you blew into it. Both the balloon and air are examples of matter.

matter
anything that has mass and takes up space

The Space an Object Occupies Is Its Volume

An object's **volume** is the space the object occupies. For example, this book has volume because it takes up space. Volume can be determined in several different ways. The method used to determine volume depends on the nature of the matter being examined. The book's volume can be found by multiplying the book's length, width, and height. Graduated cylinders are often used in laboratories to measure the volume of liquids, as shown in **Figure 6.** The volume of a gas is the same as that of the container it fills.

volume
a measure of the size of a body or region in three-dimensional space

Figure 6
To read the liquid level in a graduated cylinder correctly, read the level at the bottom part of the *meniscus,* the curved upper surface of the liquid. The volume shown here is 73.0 mL.

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Figure 7
A balance is an instrument that measures mass.

The Quantity of Matter Is the Mass

The **mass** of an object is the quantity of matter contained in that object. Even though a marble is smaller, it has more mass than a ping-pong ball does if the marble contains more matter.

mass
a measure of the amount of matter in an object; a fundamental property of an object that is not affected by the forces that act on the object, such as the gravitational force

Devices used for measuring mass in a laboratory are called *balances*. Balances can be electronic, as shown in **Figure 7,** or mechanical, such as a triple-beam balance.

Balances also differ based on the precision of the mass reading. The balance in **Figure 7** reports readings to the hundredth place. The balance often found in a school chemistry laboratory is the triple-beam balance. If the smallest scale on the triple-beam balance is marked off in 0.1 g increments, you can be certain of the reading to the tenths place, and you can estimate the reading to the hundredths place. The smaller the markings on the balance, the more decimal places you can have in your measurement.

Mass Is Not Weight

Mass is related to weight, but the two are not identical. Mass measures the quantity of matter in an object. As long as the object is not changed, it will have the same mass, no matter where it is in the universe. On the other hand, the weight of that object is affected by its location in the universe. The weight depends on gravity, while mass does not.

Weight is defined as the force produced by gravity acting on mass. Scientists express forces in *newtons,* but they express mass in *kilograms*. Because gravity can vary from one location to another, the weight of an object can vary. For example, an astronaut weighs about six times more on Earth than he weighs on the moon because the effect of gravity is less on the moon. The astronaut's mass, however, hasn't changed because he is still made up of the same amount of matter.

weight
a measure of the gravitational force exerted on an object; its value can change with the location of the object in the universe

The force that gravity exerts on an object is proportional to the object's mass. If you keep the object in one place and double its mass, the weight of the object doubles, too. So, measuring weight can tell you about mass. In fact, when you read the word *weigh* in a laboratory procedure, you probably are determining the mass. Check with your teacher to be sure.

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Figure 8
This graduated cylinder measures a *quantity,* the volume of a liquid, in a *unit,* the milliliter.

quantity
something that has magnitude, size, or amount

unit
a quantity adopted as a standard of measurement

Units of Measurement

Terms such as *heavy, light, rough,* and *smooth* describe matter qualitatively. Some properties of matter, such as color and texture, are usually described in this way. But whenever possible, scientists prefer to describe properties in quantitative terms, that is with numbers.

Mass and volume are properties that can be described in terms of numbers. But numbers alone are not enough because their meanings are unclear. For meaningful descriptions, units are needed with the numbers. For example, describing a quantity of sand as 15 kilograms rather than as 15 bucketfuls or just 15 gives clearer information.

When working with numbers, be careful to distinguish between a quantity and its unit. The graduated cylinder shown in **Figure 8,** for example, is used to measure the volume of a liquid in milliliters. *Volume* is the **quantity** being measured. *Milliliters* (abbreviated mL) is the **unit** in which the measured volume is reported.

Scientists Express Measurements in SI Units

Since 1960, scientists worldwide have used a set of units called the *Système Internationale d'Unités* or *SI.* The system is built on the seven base units listed in **Table 2.** The last two find little use in chemistry, but the first five provide the foundation of all chemical measurements.

Base units can be too large or too small for some measurements, so the base units may be modified by attaching prefixes, such as those in **Table 3.** For example, the base unit *meter* is suitable for expressing a person's height. The distance beween cities is more conveniently expressed in kilometers (km), with 1 km being 1000 m. The lengths of many insects are better expressed in millimeters (mm), or one-thousandth of a meter, because of the insects' small size. Additional prefixes can be found in **Appendix A.** Atomic sizes are so small that picometers (pm) are used. A picometer is 0.000 000 000 001 m. The advantage of using prefixes is the ability to use more manageable numbers. So instead of reporting the diameter of a hydrogen atom as 0.000 000 000 120 m, you can report it as 120 pm.

Table 2 SI Base Units

Quantity	Symbol	Unit	Abbreviation
Length	*l*	meter	m
Mass	*m*	kilogram	kg
Time	*t*	second	s
Thermodynamic temperature	*T*	kelvin	K
Amount of substance	*n*	mole	mol
Electric current	*I*	ampere	A
Luminous intensity	I_v	candela	cd

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Topic: SI Units
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Table 3 SI Prefixes

Prefix	Abbreviation	Exponential multiplier	Meaning	Example using length
Kilo-	k	10^3	1000	1 kilometer (km) = 1000 m
Hecto-	h	10^2	100	1 hectometer (hm) = 100 m
Deka-	da	10^1	10	1 dekameter (dam) = 10 m
		10^0	1	1 meter (m)
Deci-	d	10^{-1}	1/10	1 decimeter (dm) = 0.1 m
Centi-	c	10^{-2}	1/100	1 centimeter (cm) = 0.01 m
Milli-	m	10^{-3}	1/1000	1 millimeter (mm) = 0.001 m

Refer to Appendix A for more SI prefixes.

Converting One Unit to Another

In chemistry, you often need to convert a measurement from one unit to another. One way of doing this is to use a **conversion factor.** A conversion factor is a simple ratio that relates two units that express a measurement of the same quantity. Conversion factors are formed by setting up a fraction that has equivalent amounts on top and bottom. For example, you can construct conversion factors between kilograms and grams as follows:

conversion factor

a ratio that is derived from the equality of two different units and that can be used to convert from one unit to the other

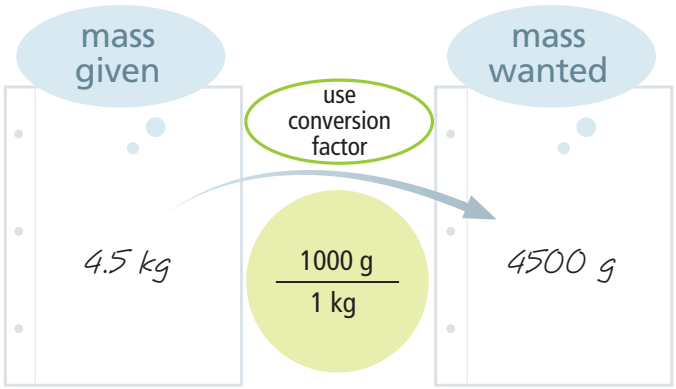
$$1\text{ kg} = 1000\text{ g} \text{ can be written as } \frac{1\text{ kg}}{1000\text{ g}} \text{ or } \frac{1000\text{ g}}{1\text{ kg}}$$

$$0.001\text{ kg} = 1\text{ g} \text{ can be written as } \frac{0.001\text{ kg}}{1\text{ g}} \text{ or } \frac{1\text{ g}}{0.001\text{ kg}}$$

SKILLS Toolkit 1

Using Conversion Factors

1. Identify the quantity and unit given and the unit that you want to convert to.
2. Using the equality that relates the two units, set up the conversion factor that cancels the given unit and leaves the unit that you want to convert to.
3. Multiply the given quantity by the conversion factor. Cancel units to verify that the units left are the ones you want for your answer.



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SAMPLE PROBLEM A

Converting Units

Convert 0.851 L to milliliters.

1 Gather information.

- You are given 0.851 L, which you want to convert to milliliters. This problem can be expressed as this equation:

$$? \text{ mL} = 0.851 \text{ L}$$

- The equality that links the two units is 1000 mL = 1 L. (The prefix *milli-* represents 1/1000 of a base unit.)

2 Plan your work.

The conversion factor needed must cancel liters and leave milliliters. Thus, liters must be on the bottom of the fraction and milliliters must be on the top. The correct conversion factor to use is

$$\frac{1000 \text{ mL}}{1 \text{ L}}$$

3 Calculate.

$$? \text{ mL} = 0.851 \cancel{\text{L}} \times \frac{1000 \text{ mL}}{1 \cancel{\text{L}}} = 851 \text{ mL}$$

4 Verify your results.

The unit of liters cancels out. The answer has the unit of milliliters, which is the unit called for in the problem. Because a milliliter is smaller than a liter, the number of milliliters should be greater than the number of liters for the same volume of material. Thus, the answer makes sense because 851 is greater than 0.851.

PRACTICE HINT

Remember that you can cancel only those units that appear in both the top and the bottom of the fractions you multiply together. Be sure to set up your conversion factors so that the unit you want to cancel is in the correct place.

PRACTICE

PROBLEM SOLVING SKILL

1. Convert each of the following masses to the units requested.
   - **a.** 0.765 g to kilograms
   - **b.** 1.34 g to milligrams
   - **c.** 34.2 mg to grams
   - **d.** 23 745 kg to milligrams (Hint: Use two conversion factors.)
2. Convert each of the following lengths to the units requested.
   - **a.** 17.3 m to centimeters
   - **b.** 2.56 m to kilometers
   - **c.** 567 dm to meters
   - **d.** 5.13 m to millimeters
3. Which of the following lengths is the shortest, and which is the longest: 1583 cm, 0.0128 km, 17 931 mm, and 14 m?

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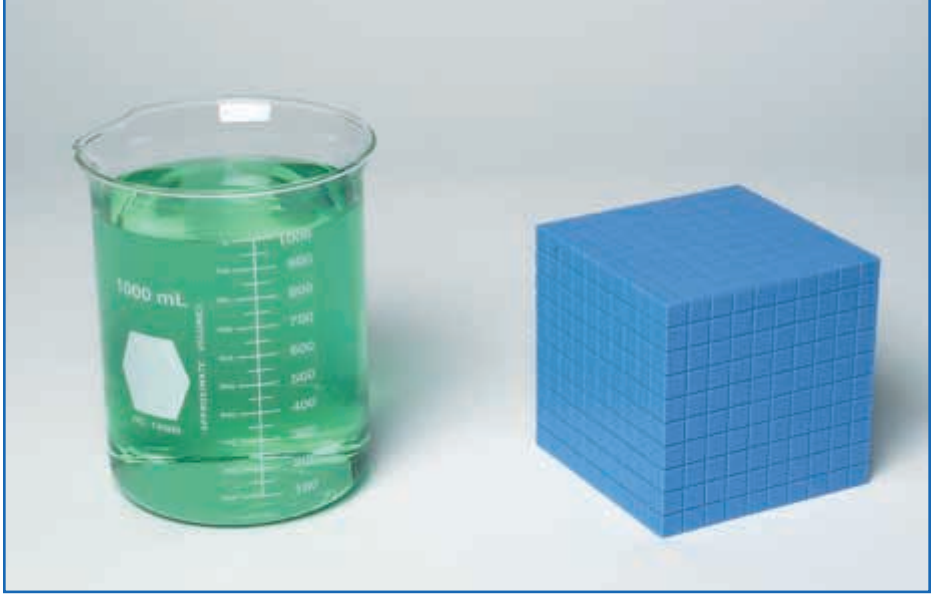

Figure 9
The volume of water in the beaker is 1 L. The model shows the dimensions of a cube that is 10 cm on each side. Its volume is 1000 cm^3, or 1 L.

Derived Units

Many quantities you can measure need units other than the seven basic SI units. These units are derived by multiplying or dividing the base units. For example, *speed* is distance divided by time. The derived unit of speed is meters per second (m/s). A rectangle's *area* is found by multiplying its length (in meters) by its width (also in meters), so its unit is square meters (m^2).

The volume of this book can be found by multiplying its length, width, and height. So the unit of volume is the cubic meter (m^3). But this unit is too large and inconvenient in most labs. Chemists usually use the liter (L), which is one-thousandth of a cubic meter. **Figure 9** shows one liter of liquid and also a cube of one liter volume. Each side of the cube has been divided to show that one liter is exactly 1000 cubic centimeters, which can be expressed in the following equality:

$$1\ L = 1000\ mL = 1000\ cm^3$$

Therefore, a volume of one milliliter (1 mL) is identical to one cubic centimeter (1 cm^3).

Properties of Matter

When examining a sample of matter, scientists describe its properties. In fact, when you describe an object, you are most likely describing it in terms of the properties of matter. Matter has many properties. The properties of a substance may be classified as physical or chemical.

Physical Properties

A **physical property** is a property that can be determined without changing the nature of the substance. Consider table sugar, or sucrose. You can see that it is a white solid at room temperature, so color and state are physical properties. It also has a gritty texture. Because changes of state are physical changes, melting point and boiling point are also physical properties. Even the lack of a physical property, such as air being colorless, can be used to describe a substance.

physical property
a characteristic of a substance that does not involve a chemical change, such as density, color, or hardness

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Density Is the Ratio of Mass to Volume

density
the ratio of the mass of a substance to the volume of the substance; often expressed as grams per cubic centimeter for solids and liquids and as grams per liter for gases

The mass and volume of a sample are physical properties that can be determined without changing the substance. But each of these properties changes depending on how much of the substance you have. The **density** of an object is another physical property: the mass of that object divided by its volume. As a result, densities are expressed in derived units such as g/cm^3 or g/mL. Density is calculated as follows:

$$density = \frac{mass}{volume} \qquad \text{or} \qquad D = \frac{m}{V}$$

The density of a substance is the same no matter what the size of the sample is. For example, the masses and volumes of a set of 10 different aluminum blocks are listed in the table in **Figure 10.** The density of Block 10 is as follows:

$$D = \frac{m}{V} = \frac{36.40\ \text{g}}{13.5\ \text{cm}^3} = 2.70\ \text{g/cm}^3$$

If you divide the mass of any block by the corresponding volume, you will always get an answer close to 2.70 g/cm^3.

The density of aluminum can also be determined by graphing the data, as shown in **Figure 10.** The straight line rising from left to right indicates that mass increases at a constant rate as volume increases. As the volume of aluminum doubles, its mass doubles; as its volume triples, its mass triples, and so on. In other words, the mass of aluminum is directly proportional to its volume.

The slope of the line equals the ratio of mass (from the vertical *y*-axis) divided by volume (from the horizontal *x*-axis). You may remember this as "rise over run" from math class. The slope between the two points shown is as follows:

$$\text{slope} = \frac{\text{rise}}{\text{run}} = \frac{29.7\ \text{g} - 10.8\ \text{g}}{11\ \text{cm}^3 - 4\ \text{cm}^3} = \frac{18.9\ \text{g}}{7\ \text{cm}^3} = 2.70\ \text{g/cm}^3$$

As you can see, the value of the slope is the density of aluminum.

Figure 10
The graph of mass versus volume shows a relationship of direct proportionality. Notice that the line has been extended to the origin.

Block number	Mass (g)	Volume (cm^3)
1	1.20	0.44
2	3.69	1.39
3	5.72	2.10
4	12.80	4.68
5	15.30	5.71
6	18.80	6.90
7	22.70	8.45
8	26.50	9.64
9	34.00	12.8
10	36.40	13.5

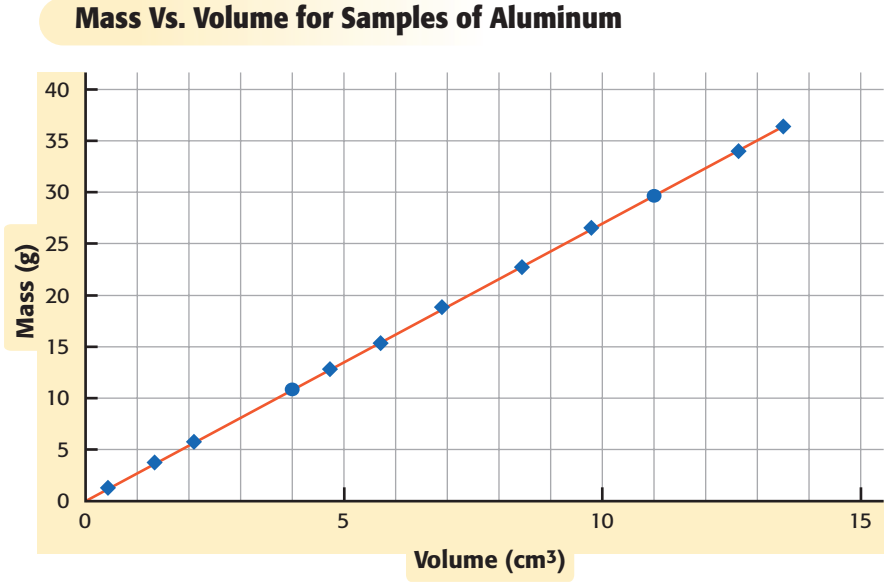

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Table 4 Densities of Various Substances

Substance	Density (g/cm^3) at 25°C
Hydrogen gas, H_2*	0.000 082 4
Carbon dioxide gas, CO_2*	0.001 80
Ethanol (ethyl alcohol), C_2H_5OH	0.789
Water, H_2O	0.997
Sucrose (table sugar), $C_{12}H_{22}O_{11}$	1.587
Sodium chloride, NaCl	2.164
Aluminum, Al	2.699
Iron, Fe	7.86
Copper, Cu	8.94
Silver, Ag	10.5
Gold, Au	19.3
Osmium, Os	22.6

*at 1 atm

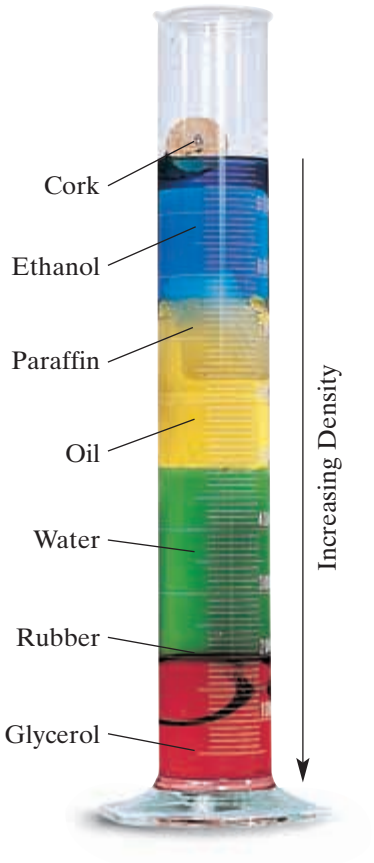

Figure 11
Substances float in layers, and the order of the layers is determined by their densities. Dyes have been added to make the liquid layers more visible.

Density Can Be Used to Identify Substances

Because the density of a substance is the same for all samples, you can use this property to help identify substances. For example, suppose you find a chain that appears to be silver on the ground. To find out if it is pure silver, you can take the chain into the lab and use a balance to measure its mass. One way to find the volume is to use the technique of water displacement. Partially fill a graduated cylinder with water, and note the volume. Place the chain in the water, and watch the water level rise. Note the new volume. The difference in water levels is the volume of the chain. If the mass is 199.0 g, and the volume is 20.5 cm^3, you can calculate the chain's density as follows:

$$D = \frac{m}{V} = \frac{199.0\ \text{g}}{20.5\ \text{cm}^3} = 9.71\ \text{g/cm}^3$$

Comparing this density with the density of silver in **Table 4,** you can see that your find is not pure silver.

Table 4 lists the densities of a variety of substances. Osmium, a bluish white metal, is the densest substance known. A piece of osmium the size of a football would be too heavy to lift. Whether a solid will float or sink in a liquid depends on the relative densities of the solid and the liquid. **Figure 11** shows several things arranged according to densities, with the most dense on the bottom.

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Chemical Properties

You cannot fully describe matter by physical properties alone. You must also describe what happens when matter has the chance to react with other kinds of matter, or the *chemical properties* of matter.

chemical property

a property of matter that describes a substance's ability to participate in chemical reactions

Whereas physical properties can be determined without changing the identity of the substance, **chemical properties** can only be identified by trying to cause a chemical change. Afterward, the substance may have been changed into a new substance.

For example, many substances share the chemical property of reactivity with oxygen. If you have seen a rusty nail or a rusty car, you have seen the result of iron's property of reactivity with oxygen. But gold has a very different chemical property. It does not react with oxygen. This property prevents gold from tarnishing and keeps gold jewelry shiny. If something doesn't react with oxygen, that lack of reaction is also a chemical property.

Not all chemical reactions result from contact between two or more substances. For example, many silver compounds are sensitive to light and undergo a chemical reaction when exposed to light. Photographers rely on silver compounds on film to create photographs. Some sunglasses have silver compounds in their lenses. As a result of this property of the silver compounds, the lenses darken in response to light. Another reaction that involves a single reactant is the reaction you saw earlier in this chapter. The formation of mercury and oxygen when mercury(II) oxide is heated, happens when a single reactant breaks down. Recall that the reaction in this case is described by the following equation:

$$\text{mercury(II) oxide} \longrightarrow \text{mercury} + \text{oxygen}$$

Despite similarities between the names of the products and the reactant, the two products have completely different properties from the starting material, as shown in **Figure 12.**

SAFETY PRECAUTIONS

Thickness of Aluminum Foil

PROCEDURE

1. Using **scissors** and a **metric ruler,** cut a rectangle of **aluminum foil.** Determine the area of the rectangle.
2. Use a **balance** to determine the mass of the foil.
3. Repeat steps **1** and **2** with each brand of aluminum foil available.

ANALYSIS

1. Use the density of aluminum (2.699 g/cm^3) to calculate the volume and the thickness of each piece of foil. Report the thickness in centimeters (cm), meters (m), and micrometers (μm) for each brand of foil. (Hint: 1 μm = 10^{-6}m)
2. Which brand is the thickest?
3. Which unit is the most appropriate unit to use for expressing the thickness of the foil? Explain your reasoning.

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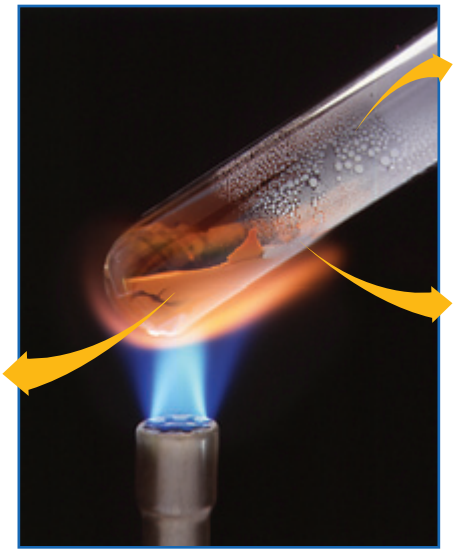
Figure 12
The physical and chemical properties of the components of this reaction system are shown. Decomposition of mercury(II) oxide is a chemical change.

MERCURY(II) OXIDE

Physical properties: Bright red or orange-red, odorless crystalline solid; almost insoluble in water

Chemical properties: Decomposes when exposed to light or at 500°C to form mercury and oxygen gas



OXYGEN

Physical properties: Colorless, odorless gas; soluble in water

Chemical properties: Supports combustion

MERCURY

Physical properties: Silver-white, liquid metal; in the solid state, mercury is ductile and malleable and can be cut with a knife

Chemical properties: Combines readily with sulfur at normal temperatures; reacts with nitric acid and hot sulfuric acid; oxidizes to form mercury(II) oxide upon heating in air

2 Section Review

UNDERSTANDING KEY IDEAS

1. Name two physical properties that characterize matter.
2. How does mass differ from weight?
3. What derived unit is usually used to express the density of liquids?
4. What SI unit would best be used to express the height of your classroom ceiling?
5. Distinguish between a physical property and a chemical property, and give an example of each.
6. Why is density considered a physical property rather than a chemical property of matter?
7. One inch equals 2.54 centimeters. What conversion factor is useful for converting from centimeters to inches?

PRACTICE PROBLEMS

8. What is the mass, in kilograms, of a 22 000 g bag of fertilizer?
9. Convert each of the following measurements to the units indicated. (Hint: Use two conversion factors if needed.)
   - **a.** 17.3 s to milliseconds
   - **b.** 2.56 mm to kilometers
   - **c.** 567 cg to grams
   - **d.** 5.13 m to kilometers
10. Convert 17.3 cm^3 to liters.
11. Five beans have a mass of 2.1 g. How many beans are in 0.454 kg of beans?

CRITICAL THINKING

12. A block of lead, with dimensions 2.0 dm × 8.0 cm × 35 mm, has a mass of 6.356 kg. Calculate the density of lead in g/cm^3.
13. Demonstrate that kg/L and g/cm^3 are equivalent units of density.
14. In the manufacture of steel, pure oxygen is blown through molten iron to remove some of the carbon impurity. If the combustion of carbon is efficient, carbon dioxide (density = 1.80 g/L) is produced. Incomplete combustion produces the poisonous gas carbon monoxide (density = 1.15 g/L) and should be avoided. If you measure a gas density of 1.77 g/L, what do you conclude?

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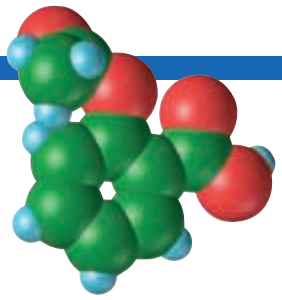
Consumer Focus



Aspirin

For centuries, plant extracts have been used for treating ailments. The bark of the willow tree was found to relieve pain and reduce fever. Writing in 1760, Edward Stone, an English naturalist and clergyman, reported excellent results when he used "twenty grains of powdered bark dissolved in water and administered every four hours" to treat people suffering from an acute, shiver-provoking illness.

The History of Aspirin

Following up on Stone's research, German chemists isolated a tiny amount of the active ingredient of the willow-bark extract, which they called *salicin,* from *Salix,* the botanical name for the willow genus. Researchers in France further purified salicin and converted it to salicylic acid, which proved to be a potent pain-reliever. This product was later marketed as the salt sodium salicylate. Though an effective painkiller, sodium salicylate has the unfortunate side effect of causing nausea and, sometimes, stomach ulcers.

Then back in Germany in the late 1800s, the father of Felix Hoffmann, a skillful organic chemist, developed painful arthritis. Putting aside his research on dyes, the younger Hoffmann looked for a way to prevent the nauseating effects of salicylic acid. He found that a similar compound, acetylsalicylic acid, was effective in treating pain and fever, while having fewer side effects. Under the name *aspirin,* it has been a mainstay in painkillers for over a century.

The FDA and Product Warning Labels

The Federal Drug Administration requires that all over-the-counter drugs carry a warning label. In fact, when you purchase any product, it is your responsibility as a consumer to check the warning label about the hazards of any chemical it may contain. The label on aspirin bottles warns against giving aspirin to children and teenagers who have chickenpox or severe flu. Some reports suggest that aspirin may play a part in Reye's syndrome, a condition in which the brain swells and the liver malfunctions.

Though side effects and allergic responses are rare, the label warns that aspirin may cause nausea and vomiting and should be avoided late in pregnancy. Because aspirin can interfere with blood clotting, it should not be used by hemophiliacs or following surgery of the mouth.

Questions

1. For an adult, the recommended dosage of 325 mg aspirin tablets is "one or two tablets every four hours, up to 12 tablets per day." In grams, what is the maximum dosage of aspirin an adult should take in one day? Why should you not take 12 tablets at once?
2. Research several over-the-counter painkillers, and write a report of your findings. For each product, compare the active ingredient and the price for a day's treatment.
3. Research Reye's syndrome, and write a report of your findings. Include the causes, symptoms, and risk factors.

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SECTION 3

How Is Matter Classified?

Key Terms

- **atom**
- **pure substance**
- **element**
- **molecule**
- **compound**
- **mixture**
- **homogeneous**
- **heterogeneous**

Objectives

1. **Distinguish** between elements and compounds.
2. **Distinguish** between pure substances and mixtures.
3. **Classify** mixtures as homogeneous or heterogeneous.
4. **Explain** the difference between mixtures and compounds.

Classifying Matter

Everything around you—water, air, plants, and your friends—is made of matter. Despite the many examples of matter, all matter is composed of about 110 different kinds of **atoms.** Even the biggest atoms are so small that it would take more than 3 million of them side by side to span just one millimeter. These atoms can be physically mixed or chemically joined together to make up all kinds of matter.

atom
the smallest unit of an element that maintains the properties of that element

Benefits of Classification

Because matter exists in so many different forms, having a way to classify matter is important for studying it. In a store, such as the nursery in **Figure 13,** classification helps you to find what you want. In chemistry, it helps you to predict what characteristics a sample will have based on what you know about others like it.

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Figure 13
Finding the plant you want without the classification scheme adopted by this nursery would be difficult.

Figure 14
Copper, bromine, and dry ice are pure substances. Each is composed of only one type of particle.

Pure Substances

pure substance
a sample of matter, either a single element or a single compound, that has definite chemical and physical properties

Each of the substances shown in **Figure 14** is a **pure substance.** Every pure substance has characteristic properties that can be used to identify it. Characteristic properties can be physical or chemical properties. For example, copper always melts at 1083°C, which is a physical property that is characteristic of copper. There are two types of pure substances: elements and compounds.

Elements Are Pure Substances

element
a substance that cannot be separated or broken down into simpler substances by chemical means; all atoms of an element have the same atomic number

Elements are pure substances that contain only one kind of atom. Copper and bromine are elements. Each element has its own unique set of physical and chemical properties and is represented by a distinct chemical symbol. **Table 5** shows several elements and their symbols and gives examples of how an element got its symbol.

Table 5 Element Names, Symbols, and the Symbols' Origins

Element name	Chemical symbol	Origin of symbol
Hydrogen	H	first letter of element name
Helium	He	first two letters of element name
Magnesium	Mg	first and third letters of element name
Tin	Sn	from *stannum,* the Latin word for "tin"
Gold	Au	from *aurum,* the Latin word meaning "gold"
Tungsten	W	from *Wolfram,* the German word for "tungsten"
Ununpentium	Uup	first letters of root words that describe the digits of the atomic number; used for elements that have not yet been synthesized or whose official names have not yet been chosen

Refer to Appendix A for an alphabetical listing of element names and symbols.

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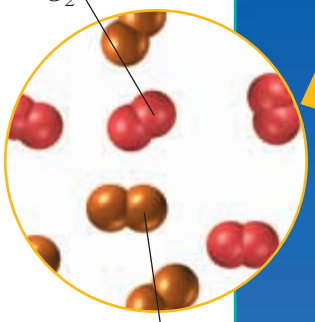


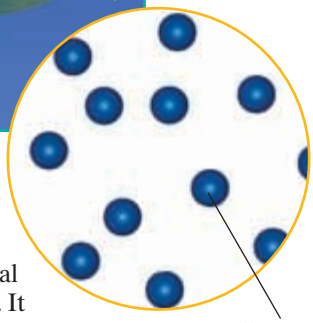


Figure 15
a The element helium, which is used to fill toy balloons, exists as individual atoms in the gaseous state. It is monatomic.

b A hot-air balloon contains a mixture of gases, mostly the elements nitrogen and oxygen. Both are diatomic, which means their molecules are made of two atoms of the element.

Elements as Single Atoms or as Molecules

Some elements exist as single atoms. For example, the helium gas in a balloon consists of individual atoms, as shown by the model in **Figure 15a.** Because it exists as individual atoms, helium gas is known as a *monatomic* gas.

Other elements exist as molecules consisting of as few as two or as many as millions of atoms. A **molecule** usually consists of two or more atoms combined in a definite ratio. If an element consists of molecules, those molecules contain just one type of atom. For example, the element nitrogen, found in air, is an example of a molecular element because it exists as two nitrogen atoms joined together, as shown by the model in **Figure 15b.** Oxygen, another gas found in the air, exists as two oxygen atoms joined together. Nitrogen and oxygen are *diatomic* elements. Other diatomic elements are H_2, F_2, Cl_2, Br_2, and I_2.

molecule

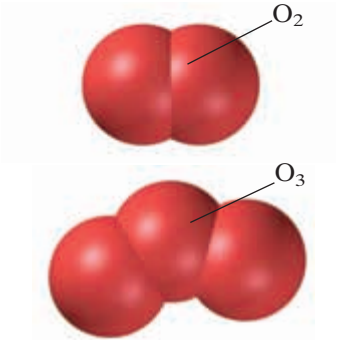
the smallest unit of a substance that keeps all of the physical and chemical properties of that substance; it can consist of one atom or two or more atoms bonded together

Some Elements Have More than One Form

Both oxygen gas and ozone gas are made up of oxygen atoms, and are forms of the element oxygen. However, the models in **Figure 16** show that a molecule of oxygen gas, O_2, is made up of two oxygen atoms, and a molecule of ozone, O_3, is made up of three oxygen atoms.

A few elements, including oxygen, phosphorus, sulfur, and carbon, are unusual because they exist as allotropes. An *allotrope* is one of a number of different molecular forms of an element. The properties of allotropes can vary widely. For example, ozone is a toxic, pale blue gas that has a sharp odor. You often smell ozone after a thunderstorm. But oxygen is a colorless, odorless gas essential to most forms of life.

Figure 16
Two forms of the element oxygen are oxygen gas and ozone gas.



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Compounds Are Pure Substances

compound
a substance made up of atoms of two or more different elements joined by chemical bonds

Pure substances that are not elements are **compounds.** Compounds are composed of more than one kind of atom. For example, the compound carbon dioxide is composed of molecules that consist of one atom of carbon and two atoms of oxygen.

There may be easier ways of preparing them, but compounds can be made from their elements. On the other hand, compounds can be broken down into their elements, though often with great difficulty. The reaction of mercury(II) oxide described earlier in this chapter is an example of the breaking down of a compound into its elements.

Compounds Are Represented by Formulas

Because every molecule of a compound is made up of the same kinds of atoms arranged the same way, a compound has characteristic properties and composition. For example, every molecule of hydrogen peroxide contains two atoms each of hydrogen and oxygen. To emphasize this ratio, the compound can be represented by an abbreviation or *formula:* H_2O_2. *Subscripts* are placed to the lower right of the element's symbol to show the number of atoms of the element in a molecule. If there is just one atom, no subscript is used. For example, the formula for water is H_2O, not H_2O_1.

Molecular formulas give information only about what makes up a compound. The molecular formula for aspirin is $C_9H_8O_4$. Additional information can be shown by using different models, such as the ones for aspirin shown in **Figure 17.** A *structural formula* shows how the atoms are connected, but the two-dimensional model does not show the molecule's true shape. The distances between atoms and the angles between them are more realistic in a three-dimensional *ball-and-stick model*. However, a *space-filling model* attempts to represent the actual sizes of the atoms and not just their relative positions. A hand-held model can provide even more information than models shown on the flat surface of the page.

Figure 17
These models convey different information about acetylsalicylic acid (aspirin).

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Compounds Are Further Classified

Such a wide variety of compounds exists that scientists classify the compounds to help make sense of them. In later chapters, you will learn that compounds can be classified by their properties, by the type of bond that holds them together, and by whether they are made of certain elements.

Mixtures

A sample of matter that contains two or more pure substances is a **mixture.** Most kinds of food are mixtures, sugar and salt being rare exceptions. Air is a mixture, mostly of nitrogen and oxygen. Water is *not* a mixture of hydrogen and oxygen for two reasons. First, the H and O atoms are chemically bonded together in H_2O molecules, not just physically mixed. Second, the ratio of hydrogen atoms to oxygen atoms is always exactly two to one. In a mixture, such as air, the proportions of the ingredients can vary.

mixture

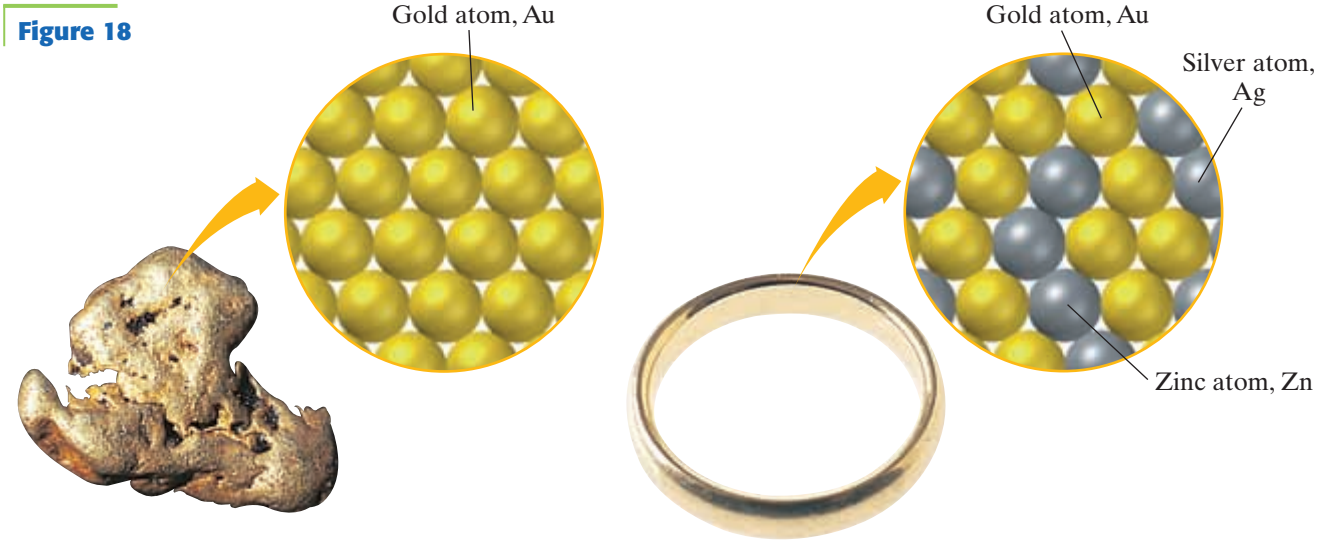
a combination of two or more substances that are not chemically combined

Mixtures Can Vary in Composition and Properties

A glass of sweetened tea is a mixture. If you have ever had a glass of tea that was too sweet or not sweet enough, you have experienced two important characteristics of mixtures. A mixture does not always have the same balance of ingredients. The proportion of the materials in a mixture can change. Because of this, the properties of the mixture may vary.

For example, pure gold, shown in **Figure 18a,** is often mixed with other metals, usually silver, copper, or nickel, in various proportions to change its density, color, and strength. This solid mixture, or *alloy,* is stronger than pure gold. A lot of jewelry is 18-karat gold, meaning that it contains 18 grams of gold per 24 grams of alloy, or 75% gold by mass. A less expensive, and stronger, alloy is 14-karat gold, shown in **Figure 18b.**

Figure 18



a The gold nugget is a pure substance—gold. Pure gold, also called 24-karat gold, is usually considered too soft for jewelry.

b This ring is 14-karat gold, which is 14/24, or 58.3%, gold. This homogeneous mixture is stronger than pure gold and is often used for jewelry.

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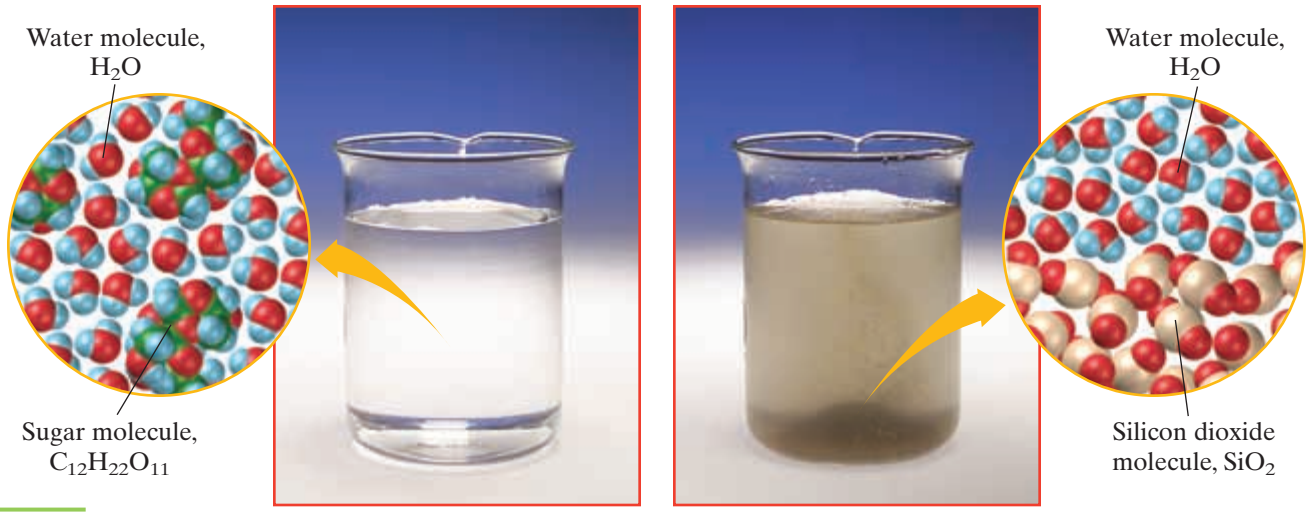

Figure 19
The mixture of sugar and water on the left is a homogeneous mixture, in which there is a uniform distribution of the two components. Sand and water, on the right, do not mix uniformly, so they form a heterogeneous mixture.

Homogeneous Mixtures

Sweetened tea and 14-karat gold are examples of **homogeneous** mixtures. In a homogeneous mixture, the pure substances are distributed uniformly throughout the mixture. Gasoline, syrup, and air are homogeneous mixtures. Their different components cannot be seen—not even using a microscope.

Because of how evenly the ingredients are spread throughout a homogeneous mixture, any two samples taken from the mixture will have the same proprtions of ingredients. As a result, the properties of a homogeneous mixture are the same throughout. Look at the homogeneous mixture in **Figure 19a.** You cannot see the different materials that make up the mixture because the sugar is mixed evenly throughout the water.

homogeneous
describes something that has a uniform structure or composition throughout

Heterogeneous Mixtures

In **Figure 19b** you can clearly see the water and the sand, so the mixture is not homogeneous. It is a **heterogenous** mixture because it contains substances that are not evenly mixed. Different regions of a heterogeneous mixture have different properties. Additional examples of the two types of mixtures are shown in **Table 6.**

heterogeneous
composed of dissimilar components

Table 6 Examples of Mixtures

Homogeneous	Iced tea—uniform distribution of components; components cannot be filtered out and will not settle out upon standing
	Stainless steel—uniform distribution of components
	Maple syrup—uniform distribution of components; components cannot be filtered out and will not settle out upon standing
Heterogeneous	Orange juice or tomato juice—uneven distribution of components; settles out upon standing
	Chocolate chip pecan cookie—uneven distribution of components
	Granite—uneven distribution of components
	Salad—uneven distribution of components; can be easily separated by physical means

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Distinguishing Mixtures from Compounds

A compound is composed of two or more elements chemically joined together. A mixture is composed of two or more substances physically mixed together but not chemically joined. As a result, there are two major differences between mixtures and compounds.

First, the properties of a mixture reflect the properties of the substances it contains, but the properties of a compound often are very different from the properties of the elements that make it up. The oxygen gas that is a component of the mixture air can still support a candle flame. However, the properties of the compound water, including its physical state, do not reflect the properties of hydrogen and oxygen.

Second, a mixture's components can be present in varying proportions, but a compound has a definite composition in terms of the masses of its elements. The composition of milk, for example, will differ from one cow to the next and from day to day. However, the compound sucrose is always exactly 42.107% carbon, 6.478% hydrogen, and 51.415% oxygen no matter what its source is.

Separating Mixtures

One task a chemist often handles is the separation of the components of a mixture based on one or more physical properties. This task is similar to sorting recyclable materials. You can separate glass bottles based on their color and metal cans based on their attraction to a magnet. Techniques used by chemists include filtration, which relies on particle size, and distillation and evaporation, which rely on differences in boiling point.

Quick LAB

Separating a Mixture

SAFETY PRECAUTIONS

PROCEDURE

1. Place the **mixture** of iron, sulfur, and salt on a **watch-glass.** Remove the iron from the mixture with the aid of a **magnet.** Transfer the iron to a 50 mL **beaker.**
2. Transfer the sulfur-salt mixture that remains to a second 50 mL **beaker.** Add 25 mL of **water,** and stir with a **glass stirring rod** to dissolve the salt.
3. Place **filter paper** in a **funnel.** Place the end of the funnel into a third 50 mL **beaker.** Filter the mixture and collect the filtrate—the liquid that passes through the filter.
4. Wash the residue in the filter with 15 mL of water, and collect the rinse water with the filtrate.
5. Set up a **ring stand** and a **Bunsen burner.** Evaporate the water from the filtrate. Stop heating just before the liquid completely disappears.

ANALYSIS

1. What properties did you observe in each of the components of the mixture?
2. How did these properties help you to separate the components of the mixture?
3. Did any of the components share similar properties?

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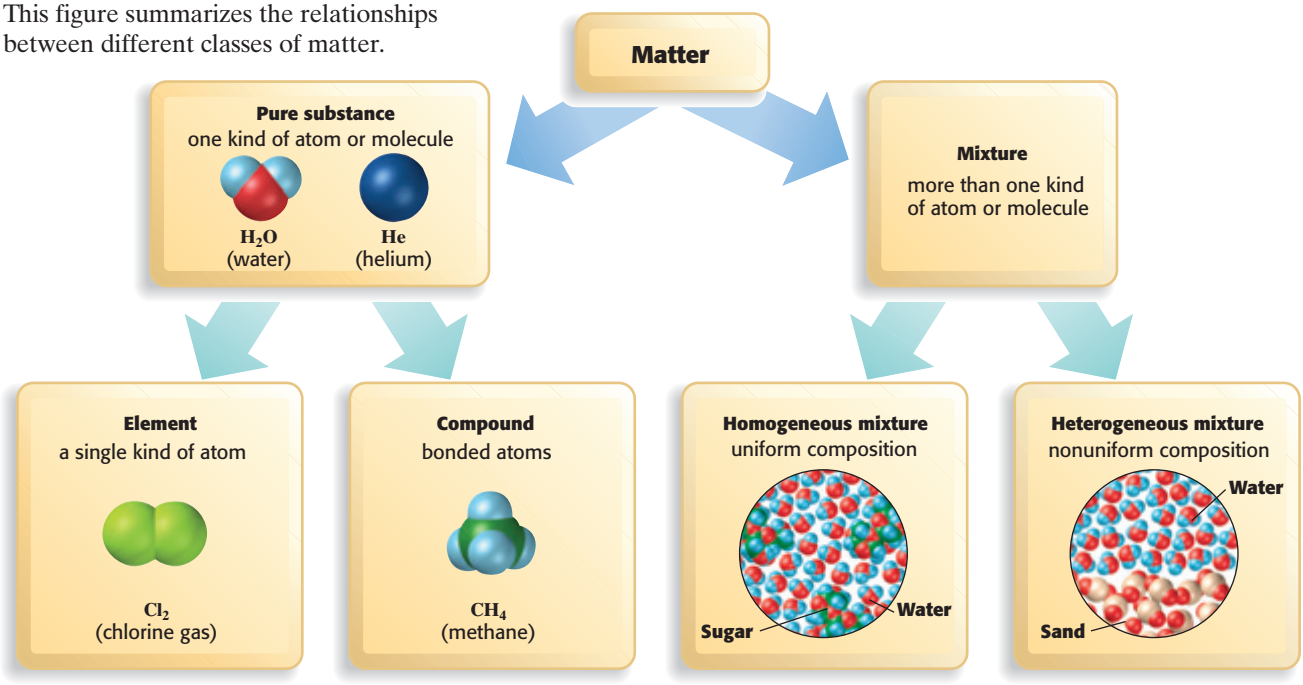
Figure 20
This figure summarizes the relationships between different classes of matter.



3 Section Review

UNDERSTANDING KEY IDEAS

1. What are the two types of pure substances?
2. Define the term *compound*.
3. How does an element differ from a compound?
4. How are atoms and molecules related?
5. What is the smallest number of elements needed to make a compound?
6. What are two differences between compounds and mixtures?
7. Identify each of the following as an element, a compound, a homogeneous mixture, or a heterogeneous mixture.
   - **a.** CH_4
   - **b.** S_8
   - **c.** distilled water
   - **d.** salt water
   - **e.** CH_2O
   - **f.** concrete
8. How is a homogeneous mixture different from a heterogeneous mixture?

CRITICAL THINKING

9. Why is a *monatomic compound* nonsense?
10. Compare the composition of sucrose purified from sugar cane with the composition of sucrose purified from sugar beets. Explain your answer.
11. After a mixture of iron and sulfur are heated and then cooled, a magnet no longer attracts the iron. How would you classify the resulting material? Explain your answer.
12. How could you decide whether a ring was 24-karat gold or 14-karat gold without damaging the ring?
13. Imagine dissolving a spoonful of sugar in a glass of water. Is the sugar-water combination classified as a compound or a mixture? Explain your answer.
14. Four different containers are labeled C + O_2, CO, CO_2, and Co. Based on the labels, classify each as an element, a compound, a homogeneous mixture, or a heterogeneous mixture. Explain your reasoning.

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Where Is Al?

Earth's Crust:
8% by mass

Sea Water:
less than 0.1%

Element Spotlight

13
Al
Aluminum
26.981 538
$[Ne]3s^2 3p^1$

Aluminum's Humble Beginnings

In 1881, Charles Martin Hall was a 22-year-old student at Oberlin College, in Ohio. One day, Hall's chemistry professor mentioned in a lecture that anyone who could discover an inexpensive method for making aluminum metal would become rich. Working in a wooden shed and using a cast-iron frying pan, a blacksmith's forge, and homemade batteries, Hall discovered a practical technique for producing aluminum. Hall's process is the basis for the industrial production of aluminum today.

Aluminum's resistance to corrosion makes it suitable for use outdoors in this statue.

Industrial Uses

- Aluminum is the most abundant metal in Earth's crust. However, it is found in nature only in compounds and never as the pure metal.
- The most important source of aluminum is the mineral bauxite. Bauxite consists mostly of hydrated aluminum oxide.
- Recycling aluminum by melting and reusing it is considerably cheaper than producing new aluminum.
- Aluminum is light, weather-resistant, and easily worked. These properties make aluminum ideal for use in aircraft, cars, cans, window frames, screens, gutters, wire, food packaging, hardware, and tools.

Real-World Connection Recycling just one aluminum can saves enough electricity to run a TV for about four hours.

A Brief History

1824: F. Wöhler, of Germany, isolates aluminum from aluminum chloride.

1827: F. Wöhler describes some of the properties of aluminum.

1854: Henri Saint-Claire Deville, of France, and R. Bunsen, of Germany, independently accomplish the electrolysis of aluminum from sodium aluminum chloride.

1886: Charles Martin Hall, of the United States, and Paul-Louis Héroult, of France, independently discover the process for extracting aluminum from aluminum oxide.

1800 — 1900

Questions

1. Research and identify at least five items that you encounter on a regular basis that are made with aluminum.
2. Research the changes that have occurred in the design and construction of aluminum soft-drink cans and the reasons for the changes. Record a list of items that help illustrate why aluminum is a good choice for this product.

internet connect
www.scilinks.org
Topic: Aluminum
SciLinks code: HW4136
SCILINKS Maintained by the National Science Teachers Association

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1 Chapter Highlights

KEY IDEAS

SECTION ONE What Is Chemistry?

- Chemistry is the study of chemicals, their properties, and the reactions in which they are involved.
- Three of the states of matter are solid, liquid, and gas.
- Matter undergoes both physical changes and chemical changes. Evidence can help to identify the type of change.

SECTION TWO Describing Matter

- Matter has both mass and volume; matter thus has density, which is the ratio of mass to volume.
- Mass and weight are not the same thing. Mass is a measure of the amount of matter in an object. Weight is a measure of the gravitational force exerted on an object.
- SI units are used in science to express quantities. Derived units are combinations of the basic SI units.
- Conversion factors are used to change a given quantity from one unit to another unit.
- Properties of matter may be either physical or chemical.

SECTION THREE How Is Matter Classified?

- All matter is made from atoms.
- All atoms of an element are alike.
- Elements may exist as single atoms or as molecules.
- A molecule usually consists of two or more atoms combined in a definite ratio.
- Matter can be classified as a pure substance or a mixture.
- Elements and compounds are pure substances. Mixtures may be homogeneous or heterogeneous.

KEY TERMS

chemical
chemical reaction
states of matter
reactant
product

matter
volume
mass
weight
quantity
unit
conversion factor
physical property
density
chemical property

atom
pure substance
element
molecule
compound
mixture
homogeneous
heterogeneous

KEY SKILLS

Using Conversion Factors
Skills Toolkit 1 p. 13
Sample Problem A p. 14

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Chapter Review 1

USING KEY TERMS

1. What is chemistry?
2. What are the common physical states of matter, and how do they differ from one another?
3. Explain the difference between a physical change and a chemical change.
4. What units are used to express mass and weight?
5. How does a quantity differ from a unit? Give examples of each in your answer.
6. What is a conversion factor?
7. Explain what derived units are. Give an example of one.
8. Define *density,* and explain why it is considered a physical property rather than a chemical property of matter.
9. Write a brief paragraph that shows that you understand the following terms and the relationships between them: *atom, molecule, compound,* and *element*.

10. What do the terms *homogeneous* and *heterogeneous* mean?

UNDERSTANDING KEY IDEAS

What Is Chemistry?

11. Your friend mentions that she eats only natural foods because she wants her food to be free of chemicals. What is wrong with this reasoning?
12. Determine whether each of the following substances would be a gas, a liquid, or a solid if found in your classroom.
    a. neon
    b. mercury
    c. sodium bicarbonate (baking soda)
    d. carbon dioxide
    e. rubbing alcohol
13. Is toasting bread an example of a chemical change? Why or why not?
14. Classify each of the following as a physical change or a chemical change, and describe the evidence that suggests a change is taking place.
    a. cracking an egg
    b. using bleach to remove a stain from a shirt
    c. burning a candle
    d. melting butter in the sun

Describing Matter

15. Name the five most common SI base units used in chemistry. What quantity is each unit used to express?
16. What derived unit is appropriate for expressing each of the following?
    a. rate of water flow
    b. speed
    c. volume of a room
17. Compare the physical and chemical properties of salt and sugar. What properties do they share? Which properties could you use to distinguish between salt and sugar?
18. What do you need to know to determine the density of a sample of matter?

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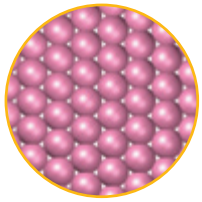
19. Substances A and B are colorless, odorless liquids that are nonconductors and flammable. The density of substance A is 0.97 g/mL; the density of substance B is 0.89 g/mL. Are A and B the same substance? Explain your answer.

How Is Matter Classified?

20. Is a compound a pure substance or a mixture? Explain your answer.

21. Determine if each material represented below is an element, compound, or mixture, and whether the model illustrates a solid, liquid, or gas.

a.



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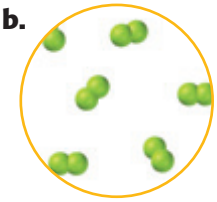

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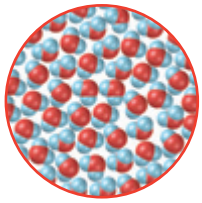
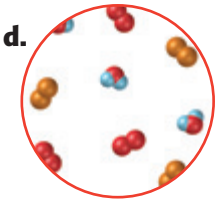


d.



PRACTICE PROBLEMS

PROBLEM SOLVING SKILL

Sample Problem A Converting Units

22. Which quantity of each pair is larger?
a. 2400 cm or 2 m
b. 3 L or 3 mL

23. Using **Appendix A,** convert the following measurements to the units specified.
a. 357 mL = ? L
b. 25 kg = ? mg
c. 35 000 cm^3 = ? L
d. 2.46 L = ? cm^3
e. 250 µg = ? g
f. 250 µg = ? kg

MIXED REVIEW

24. Use particle models to explain why liquids and gases take the shape of their containers.

25. You are given a sample of colorless liquid in a beaker. What type of information could you gather to determine if the liquid is water?

26. Calculate the density of a piece of metal if its mass is 201.0 g and its volume is 18.9 cm^3.

27. The density of CCl_4 (carbon tetrachloride) is 1.58 g/mL. What is the mass of 95.7 mL of CCl_4?

28 What is the volume of 227 g of olive oil if its density is 0.92 g/mL?

CRITICAL THINKING

29. A white, crystalline material that looks like table salt releases gas when heated under certain conditions. There is no change in the appearance of the solid, but the reactivity of the material changes.
a. Did a chemical or physical change occur? How do you know?
b. Was the original material an element or a compound? Explain your answer.

30. A student leaves an uncapped watercolor marker on an open notebook. Later, the student discovers the leaking marker has produced a rainbow of colors on the top page.
a. Is this an example of a physical change or a chemical change? Explain your answer.
b. Should the ink be classified as an element, a compound, or a mixture? Explain your answer.

ALTERNATIVE ASSESSMENT

31. Your teacher will provide you with a sample of a metallic element. Determine its density. Check references that list the density of metals to identify the sample that you analyzed.

32. Make a poster showing the types of product warning labels that are found on products in your home.

CONCEPT MAPPING

concept map

33. Use the following terms to create a concept map: *volume, density, matter, physical property,* and *mass*.

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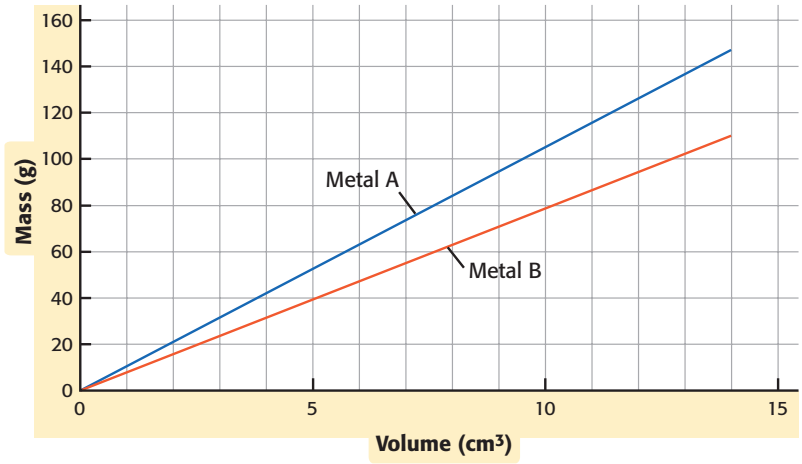
FOCUS ON GRAPHING

Study the graph below, and answer the questions that follow.
For help in interpreting graphs, see Appendix B, "Study Skills for Chemistry."

Mass Versus Volume for Two Metals



34. What does the straight line on the graph indicate about the relationship between volume and mass?

35. What does the slope of each line indicate?

36. What is the density of metal A? of metal B?

37. Based on the density values in **Table 4,** what do you think is the identity of metal A? of metal B? Explain your reasoning.

TECHNOLOGY AND LEARNING

38. Graphing Calculator

Graphing Tabular Data

The graphing calculator can run a program that graphs ordered pairs of data, such as temperature versus time. In this problem, you will answer questions based on a graph of temperature versus time that the calculator will create.

Go to Appendix C. If you are using a TI-83 Plus, you can download the program and data sets and run the application as directed. Press the APPS key on your calculator, and then choose the application CHEMAPPS. Press 1, then highlight ALL on the screen, press 1, then highlight LOAD, and press 2 to load the data into your calculator. Quit the application, and then run the program GRAPH. A set of data points representing degrees Celsius versus time in minutes will be graphed.

If you are using another calculator, your teacher will provide you with keystrokes and data sets to use.

a. Approximately what would the temperature be at the 16-minute interval?

b. Between which two intervals did the temperature increase the most: between 3 and 5 minutes, between 5 and 8 minutes, or between 8 and 10 minutes?

c. If the graph extended to 20 minutes, what would you expect the temperature to be?

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1 Standardized Test Prep

UNDERSTANDING CONCEPTS

Directions (1–3): **For *each* question, write on a separate sheet of paper the letter of the correct answer.**

1 Which of the following is **best** classified as a homogeneous mixture?

A. blood
B. copper wire
C. pizza
D. hot tea

2 Which of the following statements about compounds is true?

F. A compound contains only one element.
G. A compound can be classified as either heterogeneous or homogeneous.
H. A compound has a defined ratio by mass of the elements that it contains.
I. A compound varies in chemical composition depending on the sample size.

3 Which of the following is an element?

A. $BaCl_2$
B. CO
C. He
D. NaOH

Directions (4–6): **For *each* question, write a short response.**

4 Is photosynthesis, in which light energy is captured by plants to make sugar from carbon dioxide and water, a physical change or a chemical change? Explain your answer.

5 A student checks the volume, melting point, and shape of two unlabeled samples of matter and finds that the measurements are identical. He concludes that the samples have the same chemical composition. Is this a valid conclusion? What additional information might be collected to test this conclusion?

6 Describe the physical and chemical changes that occur when a pot of water is boiled over a campfire.

READING SKILLS

Directions (7–8): **Read the passage below. Then answer the questions.**

Willow bark has been a remedy for pain and fever for hundreds of years. In the late eighteenth century, scientists isolated the compound in willow bark that is responsible for its effects. They then converted it to a similar compound, salicylic acid, which is even more effective. In the late nineteenth century, a German chemist, Felix Hoffmann, did research to find a pain reliever that would help his father's arthritis, but not cause the nausea that is a side effect of salicylic acid. Because the technologies used to synthesize chemicals had improved, he had a number of more effective ways to work with chemical compounds than the earlier chemists. The compound that he made, acetylsalicylic acid, is known as aspirin. It is still one of the most common pain relievers more than 100 years later.

7 The main reason willow bark has been used as a painkiller and fever treatment is because

F. chemists can use it to make painkilling compounds
G. it contains elements that have painkilling effects
H. it contains compounds that have painkilling effects
I. no other painkillers were available

8 Why is aspirin normally used as a painkiller instead of salicylic acid?

A. Aspirin tends to cause less nausea.
B. Aspirin is cheaper to make.
C. Only aspirin can be isolated from willow bark.
D. Salicylic acid is less effective as a painkiller.

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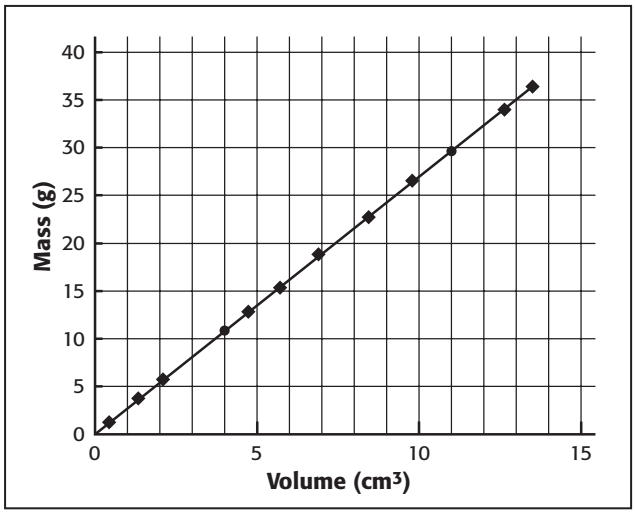
INTERPRETING GRAPHICS

Directions (9–12): **For *each* question below, record the correct answer on a separate sheet of paper.**

The table and graph below show a relationship of direct proportionality between mass (grams) versus volume (cubic centimeters). Use it to answer questions 9 through 12.

Mass Vs. Volume for Samples of Aluminum

Block number	Mass (g)	Volume (cm^3)
1	1.20	0.44
2	3.69	1.39
3	5.72	2.10
4	12.80	4.68
5	15.30	5.71
6	18.80	6.90
7	22.70	8.45
8	26.50	9.64
9	34.00	12.8
10	36.40	13.5



9 Based on information in the table and the graph, what is the relationship between mass and volume of a sample of aluminum?

F. no relationship
G. a linear relationship
H. an inverse relationship
I. an exponential relationship

10 From the data provided, what is the density of aluminum?

A. 0.37 g/cm^3
B. 1.0 g/cm^3
C. 2.0 g/cm^3
D. 2.7 g/cm^3

11 Someone gives you a metal cube that measures 2.0 centimeters on each side and has a mass of 27.5 grams. What can be deduced about the metal from this information and the table?

F. It is not pure aluminum.
G. It has more than one element.
H. It does not contain any aluminum.
I. It is a compound, not an element.

12 The density of nickel is 8.90 g/cm^3. How could this information be applied, along with information from the graph, to determine which of two pieces of metal is aluminum, and which is nickel?

Test TIP

Slow, deep breathing may help you relax. If you suffer from test anxiety, focus on your breathing in order to calm down.

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مكتبة A to Z