Chapter 2 MATTER AND CHANGE

Hot lava from a volcano changes ocean water to steam.

FEATURES

DISCOVER IT! CLASSIFYING MATTER

You need four small clear containers, sugar, baking powder, flour, baking soda, water, a half-cup measure (about 120 mL), a half-teaspoon measure (about 2.5 mL), a spoon, tape or self-adhesive labels, and a pen.

1. Label the containers with the names of the four white solids (sugar, baking powder, flour, baking soda).
2. Pour a half-cup of water into each of the four containers.
3. Without stirring, add a half-teaspoon of each white solid to its corresponding container. Take note of any reaction between the solid and the water. Wipe the spoon between each addition.
4. Stir the contents of each container for at least 45 seconds, observing what happens to each of the white solids. Rinse the spoon with water between stirrings.

Which solid could you distinguish from the other three after the initial mixing with water? How? Can you distinguish completely among the remaining three solids based on the results of Step 4? Why or why not? Make a list of the distinguishing characteristics you have just explored, and add to it as you read this chapter.

MATTER

section 2.1

As incredibly hot molten lava from a volcano reaches the ocean, the thick liquid lava cools to form solid rock. At the same time, some of the liquid water is heated to form a gas called steam. **What are the characteristics of solids, liquids, and gases, and how do their shapes and volumes differ?**

objectives

- Identify the characteristics of matter and substances
- Differentiate among the three states of matter
- Define physical property and list several common physical properties of substances

key terms

- matter
- mass
- substance
- physical property
- solid
- liquid
- gas
- vapor
- physical change

Properties of Matter

Look around you. All the things you see are examples of matter. But what exactly is matter? What forms does matter take? What can cause matter to change its form? **Matter** is defined as anything that has mass and takes up space. The **mass** of an object is the amount of matter the object contains. A golf ball has a greater mass than a table-tennis ball. The golf ball, therefore, contains more matter.

Everything is made up of matter. However, materials may differ in terms of the kind of matter they contain. For example, table sugar is one particular kind of matter, with the chemical name sucrose. Table sugar is always 100% sucrose. It always has the same makeup, or chemical composition. Matter that has a uniform and definite composition is called a **substance.** Substances, also referred to as pure substances, contain only one kind of matter. Lemonade is not a substance because not all samples of lemonade are identical. Lemonade contains more than one kind of matter and the relative amounts of each kind may differ. For example, different pitchers of lemonade may have varying amounts of sugar, lemon juice, or water, and may taste different. Thus they have different compositions.

All samples of a substance have identical physical properties. For example, all crystals of sucrose taste sweet and dissolve completely in water. A **physical property** is a quality or condition of a substance that can be observed or measured without changing the substance's composition. Physical properties include color, solubility, odor, hardness, density, melting point, and boiling point.

Scan the physical properties of the common substances listed in **Table 2.1.** Such physical properties help chemists identify the substances.

Table 2.1

Physical Properties of Some Common Substances

Substance	Formula	State	Color	Melting point (°C)	Boiling point (°C)	Density (g/cm^3)
Neon	Ne	gas	colorless	−249	−246	0.0009
Oxygen	O_2	gas	colorless	−218	−183	0.0014
Chlorine	Cl_2	gas	greenish-yellow	−101	−34	0.0032
Ethanol	C_2H_5OH	liquid	colorless	−117	78	0.789
Mercury	Hg	liquid	silvery-white	−39	357	13.5
Bromine	Br_2	liquid	red-brown	−7	59	3.12
Water	H_2O	liquid	colorless	0	100	1.00
Sulfur	S	solid	yellow	113	445	2.07
Sucrose	$C_{12}H_{22}O_{11}$	solid	white	185	d*	1.59
Sodium chloride	NaCl	solid	white	801	1413	2.17

d*, decomposes on heating

Table 2.2

Important Properties of the States of Matter			
Property	**Solid**	**Liquid**	**Gas or vapor**
Shape	*definite*	*indefinite*	*indefinite*
Volume	definite	definite	indefinite
Expansion on heating	very slight	moderate	great
Compressibility	almost incompressible	almost incompressible	readily compressible

For example, a colorless liquid that was found to boil at 100 °C and melt at 0 °C would likely be water. A colorless liquid that boiled at 78 °C and melted at −117 °C would most certainly not be water. Which substance in **Table 2.1** has these physical properties?

Chem ASAP!

Animation 1

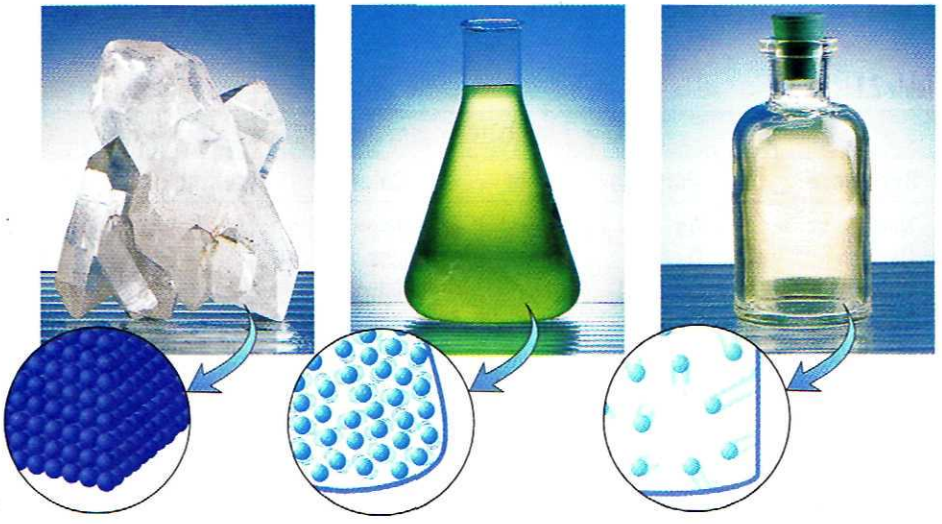
Relate the states of matter to kinetic energy and temperature.

States of Matter

You are very familiar with the substance called water. Under certain circumstances, however, you use the words ice and steam to refer to what is really the same substance. You use these three names because water—like most other substances—can exist in three different physical states: solid, liquid, and gas. The physical state of a substance is a physical property of that substance. Certain characteristics, summarized in **Table 2.2,** distinguish the three states of matter.

Coal, sugar, ice, and iron are examples of materials that are solids. A **solid** is matter that has a definite shape and volume. The shape of a solid does not depend on the shape of its container. The particles in a solid are packed tightly together, as shown in **Figure 2.1a.** As a result, solids are almost incompressible—that is, they cannot be squashed into a smaller volume. In addition, they expand only slightly when heated.

Water, milk, and blood are examples of liquids. The particles in a liquid are in close contact with one another, but unlike solid particles, they are not rigidly packed. A liquid can flow; that is, it can take the shape of the container in which it is placed. The amount of space, or volume, occupied by a sample of a liquid is the same no matter what shape it takes. This



(a) Particles in a solid **(b)** Particles in a liquid **(c)** Particles in a gas

Figure 2.1

Compare the structured arrangement of the particles in a solid, a liquid, and a gas. ***(a)*** *The particles in a solid are packed closely together in a rigid arrangement. A solid thus has a definite shape and volume.* ***(b)*** *The particles in a liquid contact one another yet have more freedom to move than the particles in a solid. A liquid can flow and takes on the shape of its container.* ***(c)*** *The particles in a gas are relatively far apart and are free to move anywhere inside their container. A gas thus has an indefinite shape and indefinite volume.*

unchanging volume is said to be fixed or constant. A **liquid** is thus a form of matter that flows, has a fixed volume, and takes the shape of its container. Liquids are almost incompressible, but they tend to expand when heated.

Like liquids, gases flow to take the shape of the container that holds them. The particles in a gas are spaced far apart. Unlike liquids, gases expand without limit to fill any space. Thus a **gas** is a form of matter that takes both the shape and volume of its container. Gases are also easily compressed.

The words gas and vapor should not be used interchangeably; there is a difference. The term gas is limited to those substances that exist in the gaseous state at ordinary room temperature. For example, air is a mixture of gases including oxygen and nitrogen. The word **vapor** describes the gaseous state of a substance that is generally a liquid or solid at room temperature. Steam, the gaseous form of water, is referred to as a vapor because water is a liquid at room temperature. Moist air contains water vapor.

Figure 2.2
The melting point of gallium metal is 30 °C. What is the approximate Celsius temperature of the hand shown holding the gallium? How can you tell?

Physical Changes

Matter can be changed in many ways without altering the chemical composition of the material. Such a change, which alters a given material without changing its composition, is called a **physical change.** Cutting, grinding, and bending a material are examples of physical changes. A change in temperature may also bring about a physical change, such as the melting of the metal gallium, shown in **Figure 2.2.** The melting of ice, the freezing of water, the conversion of water to steam, and the condensation of steam to water are all examples of physical changes—changes that do not alter the chemical composition of the water.

Words such as boil, freeze, dissolve, melt, condense, break, split, crack, grind, cut, crush, and bend usually signify a physical change. Table salt (sodium chloride) is a white solid at room temperature. It can be made to undergo physical changes. For example, it can melt to form a liquid or boil to form a vapor. The temperatures at which these changes of state occur in sodium chloride are very different from the corresponding changes for water. Sodium chloride melts (becomes a liquid) at 801 °C and boils (becomes a vapor) at 1413 °C—high temperatures compared with water's melting and boiling points of 0 °C and 100 °C, respectively.

LINK TO PHYSICS

Changing States of Matter
Changes of state are typically associated with changes in temperature. However, the state of matter can also be affected by other variables. Thixotropic materials are solidlike materials that liquefy when subjected to shearing forces. For example, many paints are thixotropic; they thin out when brushed on a surface and thicken when the brush strokes stop, thus keeping the paint from sliding off the wall! A shearing force has an opposite effect on quicksand. Quick movements "thicken" the quicksand and make it much more difficult for a person or animal trapped in it to move.

section review 2.1

1. Is every sample of matter a substance? Explain.
2. Contrast the characteristics of the three states of matter.
3. Which of the following are physical changes?
   a. making caramel from sugar
   b. carving a wooden figurine
   c. freezing mercury
   d. dissolving salt in water
4. Use **Table 2.1** to answer the following questions.
   a. Which of the liquids listed has the highest boiling point?
   b. What two properties of sucrose distinguish it from sodium chloride?
   c. What single property do neon, oxygen, and ethanol have in common?

Chem ASAP! **Assessment 2.1** Check your understanding of the important ideas and concepts in Section 2.1.

section 2.2

MIXTURES

objectives

- Categorize a sample of matter as a substance or a mixture
- Distinguish between homogeneous and heterogeneous samples of matter

key terms

- mixture
- heterogeneous mixture
- homogeneous mixture
- solutions
- phase
- distillation

In 1848, gold was discovered in the foothills near Placerville, California. This discovery led to a massive gold rush in the following year. Many people in the California foothills still pan for gold as a hobby. Panning separates gold out of a mixture of gold and sand. **What is a mixture, and how can it be separated?**

Classifying Mixtures

You might prepare a salad by tossing lettuce, tomatoes, cucumbers, and celery with some vinegar and oil. The result is not only nutritious; it is also a mixture. In this section, you will learn how to identify and classify mixtures.

Most samples of matter are obviously mixtures. For example, you can easily recognize chicken noodle soup as a mixture of chicken pieces, noodles, and broth. Recognizing other materials as mixtures may be much harder. Air is a mixture of gases, but its components cannot be distinguished by eye, even through a microscope.

A **mixture** is a physical blend of two or more substances. One important characteristic of mixtures is that their compositions may vary. A dinner salad can have varying amounts of tomatoes or celery in it. The composition of air in a forest may differ from that in an industrial city, particularly in the amounts of pollutants. Blood, a mixture of water, various chemicals, and cells, varies somewhat in composition from one individual to another and, from time to time, in a given individual.

Mixtures can be of two basic kinds: heterogeneous or homogeneous. **Figure 2.3** gives examples of each kind. A **heterogeneous mixture** is one that is not uniform in composition. If you were to sample one portion of such a mixture, its composition would be different from that of another portion. Why is the salad described above heterogeneous? A **homogeneous mixture** in contrast, is one that has a completely uniform composition. Its components are evenly distributed throughout the sample. A sample of salt water is the same throughout. Thus salt water is an example of a homogeneous mixture.

Figure 2.3
All of these items are mixtures. The bar of soap and the beverage are homogeneous mixtures; they have uniform compositions. The salad is a heterogeneous mixture; it consists of several phases containing components that are not evenly distributed. What other everyday items can you identify as either homogeneous or heterogeneous mixtures?

section 3.4

DENSITY

objectives

- Calculate the density of an object from experimental data
- List some useful applications of the measurement of specific gravity

key terms

- density
- specific gravity
- hydrometer

Have you ever wondered why some objects float in water, while others sink? If you think that these cranberries float because they are lightweight, you are only partially correct. **What measurements would you need to make to determine whether an object would float in water?**

Determining Density

Has anyone ever tried to trick you with this question: "Which is heavier, a pound of lead or a pound of feathers?" Most people would not give the question much thought and would incorrectly answer "lead." Of course, a pound of lead has the same mass as a pound of feathers. What concept, instead of mass, are people really thinking of when they answer this question?

Most people are incorrectly applying a perfectly correct idea: namely, that if a piece of lead and a feather of the same volume are weighed, the lead would have a greater mass than the feather. It would take a much larger volume of feathers to equal the mass of a given volume of lead.

Table 3.6

Relationship Between Volume and Density for Identical Masses of Common Substances

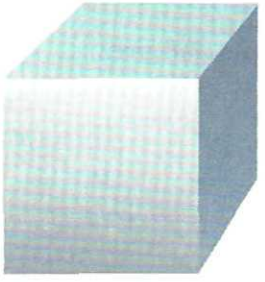
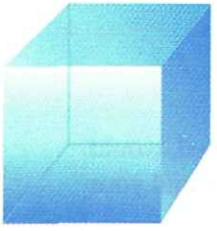
Substance	Cube of substance (face shown actual size)	Mass (g)	Volume (cm^3)	Density (g/cm^3)
Lithium		10	19	0.53
Water		10	10	1.0
Aluminum		10	3.7	2.7
Lead		10	0.88	11.4

Table 3.7

Densities of Some Common Materials			
Solids and Liquids		Gases	
Material	Density at 20 °C (g/cm^3)	Material	Density at 20 °C (g/L)
Gold	19.3	Chlorine	2.95
Mercury	13.6	Carbon dioxide	1.83
Lead	11.4	Oxygen	1.33
Aluminum	2.70	Air	1.20
Table sugar	1.59	Nitrogen	1.17
Water (4 °C)	1.000	Neon	0.84
Corn oil	0.922	Ammonia	0.718
Ice (0 °C)	0.917	Methane	0.665
Ethanol	0.789	Helium	0.166
Gasoline	0.66–0.69	Hydrogen	0.084

The important relationship in this case is between the object's mass and its volume. This relationship is called density. **Density** is the ratio of the mass of an object to its volume.

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

A 10.0-cm^3 piece of lead, for example, has a mass of 114 g. What, then, is the density of lead? If you substitute the mass and volume into the equation above, you can see that the density of lead is as follows.

$$\frac{114\ \text{g}}{10.0\ \text{cm}^3} = 11.4\ \text{g/cm}^3$$

Note that when mass is measured in grams, and volume in cubic centimeters, density has units of grams per cubic centimeter (g/cm^3).

Table 3.6 compares the density of four substances. Why does each 10-g sample have a different volume? The volumes vary because the substances have different densities. Density is a characteristic property that depends only on the composition of a substance, not on the size of the sample. With a mixture, density can vary because the composition of a mixture can vary.

What do you think will happen if glycerine (glycerol) and corn oil are poured into a beaker of water? Using **Table 3.7,** you can see that the density of corn oil is less than the density of water. For that reason, the corn oil floats on top of the water. As you can see in **Figure 3.15,** the glycerol sinks below the water because its density is greater than the density of water. Other liquids with different densities are also shown.

You have probably seen a helium-filled balloon rapidly rise to the ceiling when it is released. Whether a gas-filled balloon will sink or rise when released depends on how the density of the gas compares with the density of air. The densities of various gases are given in **Table 3.7.** Would a carbon dioxide-filled balloon sink or rise in air?

What happens to the density of a substance as its temperature increases? Experiments show that the volume of most substances increases as the temperature increases. Meanwhile, the mass remains the same

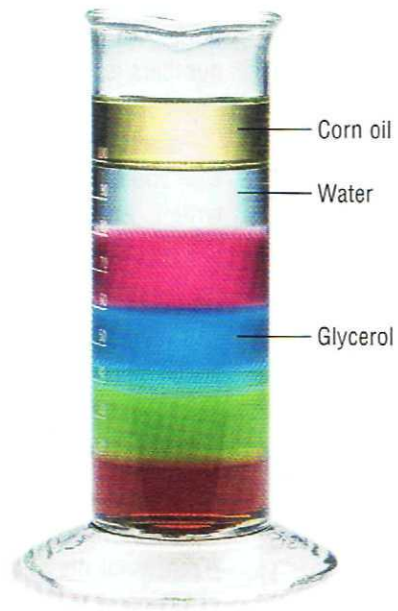


Figure 3.15
Because of differences in density, the blue-colored glycerol sinks below the surface of the water, and the amber-colored corn oil floats on top of the water. Is the red-colored liquid more or less dense than glycerol? How do you know?

Chem ASAP!
Simulation 1
Rank materials according to their densities.

CHEMath

WHAT IS AN EQUATION?

You will encounter numerous equations as you read through this textbook. You should consider every equation as a powerful and easy-to-use tool for simplifying and applying complex relationships. Furthermore, once you understand the basics of what equations are and how they are used, you will find they are usually solved using simple math.

The Equation An equation is a mathematical sentence that uses an equal sign (=) to represent two different expressions that have the same value. In addition to the equal sign, an equation may contain a combination of words, numbers, variables, or mathematical functions. An example of an equation that contains numbers is:

$$5 + 4 = 9$$

Each side of the equation is separated by the equal sign and is equivalent; 5 + 4 is simply another way to write 9. An example of a word equation is:

The perimeter of a square is equal to the sum of the lengths of its four equal sides.

Although the above statement is correct, it can be stated much more concisely by the following mathematical equation:

$$\text{perimeter} = 4 \times \text{length of a side}$$

This can be further simplified by replacing the words with variables that represent the words. Letting p represent the perimeter, and L represent the length of each side yields,

$$p = 4 \times L$$
$$p = 4L$$

Units in Equations 5 + 4 = 9 is an example of an equation that contains no units. Many equations do involve units, however, and you can use the units as a way to check that the equation is correct. Because both sides of an equation are equal, the units on each side of an equation must also be equal. Consider the following equation:

$$\text{mass}_{\text{total}} = \text{mass}_{\text{beaker}} + \text{mass}_{\text{chemical}}$$

Both sides of this equation must contain units of mass, such as grams or kilograms. If the units on both sides of the equation are not the same, the answer must be incorrect.

Solving Equations It is very important to realize that both sides of an equation are always equal to each other. This is important because you can use this fact to help solve for an unknown value.

You only need to remember a single rule when solving an equation for a specific quantity: whatever you do on one side of the equation must be done on the other side of the equation. For example, reconsider the perimeter equation.

$$p = 4L$$

If you need to calculate the length of a single side of the square (L), you must solve the equation for L. In other words the variable L must be isolated from everything else. To isolate L you can divide the right side of the equation by 4. However, remember you must also divide the left side by 4 as shown below.

$$p = 4L \qquad \frac{p}{4} = \frac{4L}{4} \qquad \frac{p}{4} = \frac{\cancel{4}L}{\cancel{4}} \qquad \frac{p}{4} = L$$
$$L = \frac{p}{4}$$

Practice Problems

A. Write a word equation for money earned each week from a part-time job.

B. Solve the equation K = °C + 273 for °C.

C. Using the equation K = °C + 273, calculate K if °C = 200.

D. Solve the equation density = $\frac{\text{mass}}{\text{volume}}$ for mass.

E. Using the equation density = $\frac{\text{mass}}{\text{volume}}$, calculate density if mass = 228 g and volume = 102 cm^3.

F. Repeat problem D, but solve for volume.

G. Solve the following equation for y.

$$\frac{2y}{x^2} + 3 = 4z$$

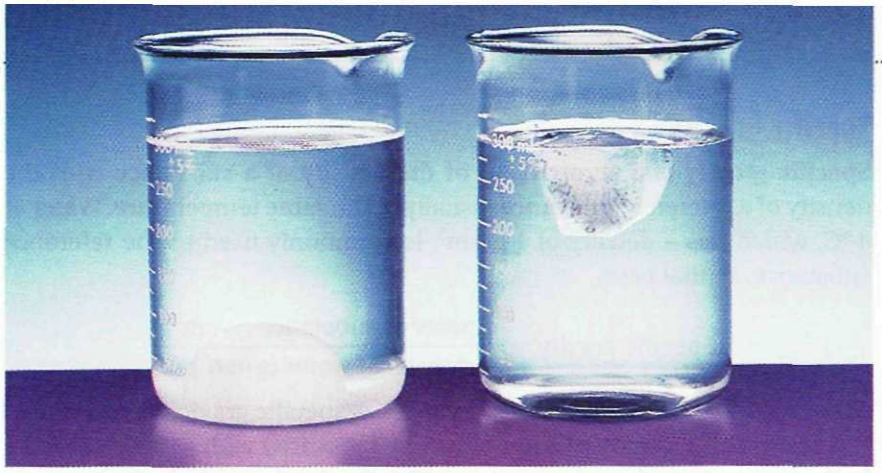


Figure 3.16
Solid paraffin sinks in melted paraffin (left), but ice floats in water (right). What does this tell you about the densities of each of these two substances in the solid and liquid states?

despite the temperature and volume changes. Because density is mass divided by volume, the density of a substance generally decreases as its temperature increases. As you will learn in Chapter 17, water is an important exception. Over a certain range of temperatures, the volume of water increases as its temperature decreases. Ice, or solid water, floats because it is less dense than liquid water. Compare the densities of the substances shown in **Figure 3.16.**

Sample Problem 3-5

A copper penny has a mass of 3.1 g and a volume of 0.35 cm^3. What is the density of copper?

1. ***ANALYZE*** ***List the knowns and the unknown.***

   Knowns:
   - mass = 3.1 g
   - volume = 0.35 cm^3

   Unknown:
   - density = ? g/cm^3

   Use the known values and the definition of density,

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

   to calculate the density of copper.

2. ***CALCULATE*** ***Solve for the unknown.***

   The equation is already set up to solve for the unknown. Substitute the known values for mass and volume, and calculate the density.

   $$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{3.1\ \text{g}}{0.35\ \text{cm}^3} = 8.8571\ \text{g/cm}^3$$

   $$= 8.9\ \text{g/cm}^3 \text{ (rounded to two significant figures)}$$

3. ***EVALUATE*** ***Does the result make sense?***

   A piece of copper with a volume of about 0.3 cm^3 has a mass of about 3 g. Thus, about three times that volume of copper, 1 cm^3, should have a mass three times larger, about 9 g. This estimate agrees with the calculated result.

Practice Problems

23. A student finds a shiny piece of metal that she thinks is aluminum. In the lab, she determines that the metal has a volume of 245 cm^3 and a mass of 612 g. Calculate the density. Is the metal aluminum?
24. The density of silver at 20 °C is 10.5 g/cm^3. What is the volume of a 68-g bar of silver?

Chem ASAP!

Problem-Solving 24

Solve Problem 24 with the help of an interactive guided tutorial.

Specific Gravity

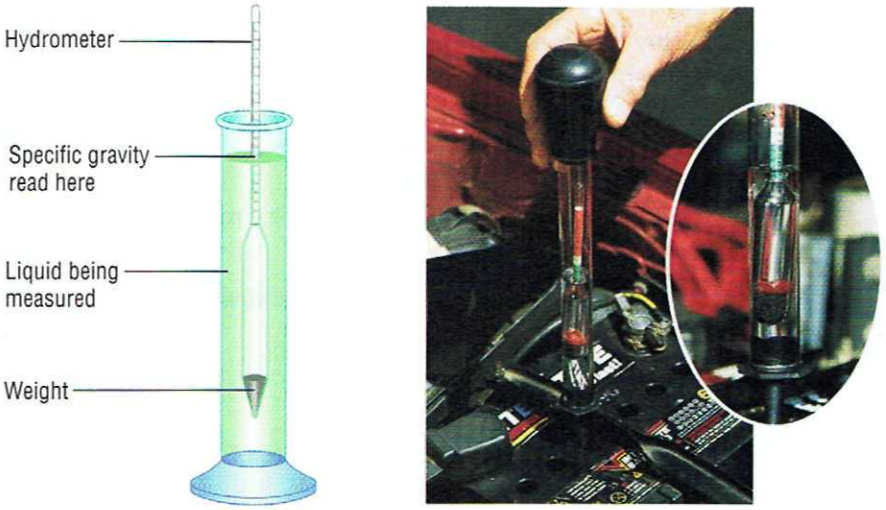
Specific gravity is a comparison of the density of a substance with the density of a reference substance, usually at the same temperature. Water at 4 °C, which has a density of 1 g/cm^3, is commonly used as the reference substance. In that case,

$$\text{Specific gravity} = \frac{\text{density of substance } (\cancel{\text{g/cm}^3})}{\text{density of water } (\cancel{\text{g/cm}^3})}$$

Note that the units in the equation cancel — specific gravity has no units.

The specific gravity of a liquid can be measured with a device called a **hydrometer.** As is shown in **Figure 3.17a,** the depth to which the hydrometer sinks depends on the specific gravity of the liquid tested. The calibration mark on the hydrometer stem at the surface of the liquid indicates the specific gravity of the liquid.

Specific-gravity measurements are commonly used for various practical purposes. A physician uses the measured specific gravity of a patient's urine to help diagnose certain diseases, such as diabetes. You can check the condition of the antifreeze in your car by measuring the specific gravity of the solution in the radiator. The hydrometer in **Figure 3.17** is being used to measure the specific gravity of the acid in an automobile battery.



(a) (b)

Figure 3.17
(a) A hydrometer is a sealed tube with a weight in the bottom. It is used to measure the specific gravity of a liquid. The higher the hydrometer floats, the higher the specific gravity of the liquid tested. (b) In the photograph, the concentration of acid in an automobile battery is checked with a hydrometer.

section review 3.4

25. How is density calculated from measured data?

26. A weather balloon is inflated to a volume of 2.2×10^3 L with 37.4 g of helium. What is the density of helium, in grams per liter?

27. List some applications of the measurement of specific gravity.

28. A plastic ball with a volume of 19.7 cm^3 has a mass of 15.8 g. What is its density? Would this ball sink or float in a container of gasoline?

29. Given samples of gold, gasoline, ice, mercury, lead, and aluminum, which substances have the highest and lowest specific gravities?

Chem ASAP! **Assessment 3.4** Check your understanding of the important ideas and concepts in Section 3.4.

SMALL-SCALE LAB

MEASUREMENT

PURPOSE

To make precise and accurate measurements and to use fundamental data to calculate derived quantities.

MATERIALS

- pencil
- paper
- ruler
- balance
- rectangular block
- calculator

PROCEDURE

1. Use your ruler to measure the length, width, and height of a rectangular block in centimeters. Be sure to estimate between the lines of the ruler's smallest markings and report each measurement so that it includes the estimated number. Draw a diagram of the block and label its dimensions.
2. Measure and record the mass of the block. Estimate between the lines of the smallest markings of the balance and report your measurement in grams. (*Note:* If you are using a digital balance, the last number often blinks, or oscillates, between two or more numbers. Recording this last digit is similar to estimating between the markings on a balance.)

ANALYSIS

Using your experimental data, record the answers to the following questions.

1. How many significant figures do each of your measurements have?
2. Calculate the perimeter of each face of the rectangular block. Report the answer to the correct number of significant figures. (The perimeter of a face is the sum of the four edges of that face.)
3. Calculate the area of each face. Include the proper unit with your calculated value for area. Round your value to the correct number of significant figures.
   (Area = length × width)
4. Calculate the volume of the block in units of cubic centimeters. Round your value to the correct number of significant figures.
   (Volume = length × width × height)
5. Express the volume of your block in liters.
   ($1\ L = 1000\ cm^3$)
6. Express the mass of your block in milligrams.
   (1 g = 1000 mg)

YOU'RE THE CHEMIST

The following small-scale activities allow you to develop your own procedures and analyze the results.

1. **Analyze It!** Density is a measure of how tightly matter is packed together.

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

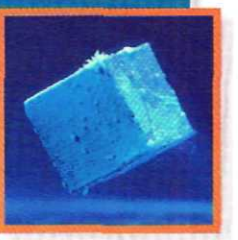
   Calculate the density of the block in units of grams per cubic centimeter (g/cm^3). Convert your calculated density into units of grams per liter (g/L), milligrams per cubic centimeter (mg/cm^3), and kilograms per liter (kg/L).
   (1 kg = 1000 g)
2. **Design It!** Make the necessary measurements to calculate the volume and density of a cylindrical block. The volume is determined from the following equation.

$$V = \pi r^2 h$$

   $\pi \approx 3.1416$
   r = radius
   h = height
3. **Design It!** Make the necessary measurements to calculate the volume and density of a block that has had one or more holes drilled in it. Assume each hole is cylindrical and account for them in your calculations. What approximations must you make?

section 3.5

TEMPERATURE



objective

- Convert between the Celsius and Kelvin temperature scales

key terms

- temperature
- Celsius scale
- Kelvin scale
- absolute zero

In 1911, a Dutch physicist named Heike Kamerlingh Onnes discovered that when mercury is cooled to near absolute zero (−460 °F), it loses all resistance to electrical current. This state is known as superconductivity. If a material could be found that would superconduct at room temperature (around 64 °F), it would revolutionize the computer industry and other electronics industries. **Which temperature scale has its zero point at absolute zero?**

Measuring Temperature

When you hold a glass of hot water, the glass feels hot because heat transfers from the glass to your hand. When you hold an ice cube, it feels cold because heat transfers from your hand to the ice cube. The **temperature** of an object determines the direction of heat transfer. When two objects at different temperatures are in contact, heat moves from the object at the higher temperature to the object at the lower temperature. In Chapter 10, you will learn how the temperature of an object is related to the energy and motion of particles.

Almost all substances expand with an increase in temperature and contract as the temperature decreases (a very important exception is water). These properties are the basis for the common mercury-in-glass thermometer, shown in **Figure 3.19.** The liquid mercury in the thermometer expands and contracts more than the volume of the glass bulb that holds it, producing changes in the column height of the mercury.

Figure 3.18

Temperatures on Earth range from the scorching heat of a desert, to the frigid cold of the Antarctic, to the normal human body temperature of 37 °C. What is the hottest temperature ever recorded? The coldest?

Temperature Scales

Several temperature scales have been devised. The Celsius scale of the metric system is named after the Swedish astronomer Anders Celsius (1701–1744). It uses two readily determined temperatures as reference temperature values: the freezing point and the boiling point of water. The **Celsius scale** sets the freezing point of water at 0 °C and the boiling point of water at 100 °C. The distance between these two fixed points is divided into 100 equal intervals, or degrees Celsius (°C).

Another temperature scale used in the physical sciences is the Kelvin, or absolute, scale. This scale is named for Lord Kelvin (1824–1907), a Scottish physicist and mathematician. On the Kelvin scale, the freezing point of water is 273.15 kelvins (K), and the boiling point is 373.15 K. Notice that with the Kelvin scale, the degree sign is not used. You can use **Figure 3.19** to compare the Celsius and Kelvin scales. The zero point on the Kelvin scale, 0 K, or **absolute zero,** is equal to −273.15 °C. For problems in this text, you can round −273.15 °C to −273 °C. Because one degree on the Celsius scale is equivalent to one kelvin on the Kelvin scale, converting from one temperature scale to the other is easy. You simply add or subtract 273, as shown in the following equations.

$$K = °C + 273$$

$$°C = K - 273$$

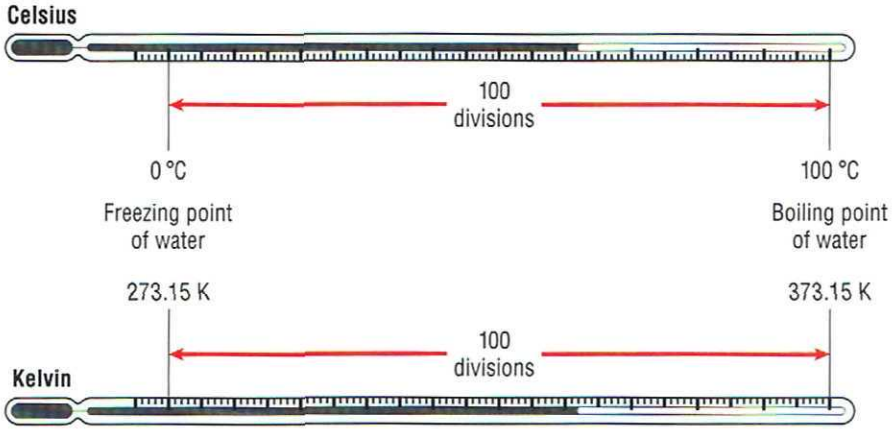


Figure 3.19
These thermometers show a comparison of the Celsius and Kelvin temperature scales. Note that a 1 °C change on the Celsius scale is equal to a 1 K change on the Kelvin scale. What is a change of 10 K equivalent to on the Celsius scale?

Sample Problem 3-6

Normal human body temperature is 37 °C. What is that temperature in kelvins?

1. ***ANALYZE*** ***List the knowns and the unknown.***

   Known:
   - Temperature in °C = 37 °C

   Unknown:
   - Temperature in K = ? K

   Use the known value and the equation K = °C + 273 to calculate the temperature in kelvins.

2. ***CALCULATE*** ***Solve for the unknown.***
   Substitute the known value for the Celsius temperature into the equation and solve.

$$K = °C + 273$$
$$= 37 + 273 = 310\ K$$

3. ***EVALUATE*** ***Does the result make sense?***
   You should expect the Kelvin temperature to be in this range because the freezing point of water is 273.15 K and the boiling point of water is 373.15 K; normal body temperature is between these two values.

Practice Problems

30. Liquid nitrogen boils at 77.2 K. What is this temperature in degrees Celsius?

31. The element silver melts at 960.8 °C and boils at 2212 °C. Express these temperatures in kelvins.

Chem ASAP!
Problem-Solving 31
Solve Problem 31 with the help of an interactive guided tutorial.

section review 3.5

32. State the relationship between degrees Celsius and kelvins.

33. Chocolate cookies are baked at 190 °C. Express this temperature in kelvins.

34. Surgical instruments may be sterilized by heating at 170 °C for 1.5 hours. Convert 170 °C to kelvins.

35. The boiling point of the element argon is 87 K. What is the boiling point of argon in degrees Celsius?

Chem ASAP! Assessment 3.5 Check your understanding of the important ideas and concepts in Section 3.5.

Chemistry Serving...the Consumer

THESE STANDARDS ARE WORTH THE WEIGHT

Imagine the confusion if every soft drink manufacturer defined a liter differently, or if the mass of a kilogram depended on where you live. As a consumer, you rely on units of mass, time, and length being unchanging and consistent, whether you buy a kilogram of ice cream, a half-hour of Internet access, or 10 000 square meters of land.

Luckily, you usually do not have to question the consistency of units of measurement because standards for each unit have been defined and agreed upon. But humans have not always had as good a system of standards as we have today. Throughout history, standards of measurement have become more precise and more consistent, and they continue in this direction today.

Using units that were based on body measurements was not a very reliable system of measurement. There could be great variation in the size of people's thumbs or the length of people's feet. In other words, thumbs and feet were not consistent units. People needed standards that did not change.

As a consumer, you rely on units of mass, time, and length being unchanging and consistent.

In 1790, with the establishment of the metric system, the French became the first to adopt measurement standards that were close to being precise. The meter was defined as one ten-millionth of the distance from the equator to the North Pole along the meridian that passes through Paris. The second was defined as $\frac{1}{86\,400}$ of the average day.

As scientific measurement techniques became more precise, the definitions of these and other base units also became more precise. For example, the meter is now defined as the distance traveled by light in a vacuum in $\frac{1}{299\,792\,458}$ of a second. The second itself is defined in terms of the number of cycles of radiation given off by a specific isotope of the element cesium.

The standard for one base unit—the kilogram—has not changed since it was originally established. A piece of metal (a platinum-iridium alloy shown here) that was cast more than 100 years ago is still used as the standard of mass. It resides in a triple bell jar in Sevres, France, and has been removed only three times since 1889. But this kilogram standard has a problem because contaminants can build up on its surface despite the layers of protection. It might seem like an easy task to find a more precise modern standard for the kilogram, but this goal has so far been a challenge to scientists.

How about defining a kilogram as a given number of atoms of a particular element? Scientists thought of that idea a few years ago, using ultra-pure spheres of silicon as the standard. However, the presence of lattice holes (empty spaces in the crystalline structure of silicon) led to different estimates of the number of atoms in a given mass. This problem has almost been solved, and the "atom counters" are getting close to the goal of defining a new kilogram standard. Another group of researchers, the "force measurers," have a competing idea. Their method uses a Watt balance (a type of balance that uses electromagnetic force) to express the mass of a kilogram in terms of the base units of voltage and current. Either way, the old kilogram may soon have a more modern definition.

CHEMISTRY IN CAREERS

ANALYTICAL CHEMIST
Interested in a career with a focus on measurement and problem solving?
See page 869.

CHEMIST NEEDED
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CHEMICAL SPECIALIST
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Chapter 10 STATES OF MATTER

Elk experience three states of water at Yellowstone National Park.

FEATURES

DISCOVER IT! OBSERVING GAS PRESSURE

You need an index card, a graduated cylinder, water, and a small glass with a smooth, even rim.

1. Fill the glass to its rim with water.
2. Place the index card on top of the glass.
3. Working over a sink, use one hand to press the index card firmly to the top of the glass. Then quickly invert the glass, keeping your hand in place.
4. Remove your hand from the index card.

What did you observe when you removed your hand? Measure the volume of water in the glass. How many grams of water are in the glass? (Remember that 1 mL H_2O = 1 g H_2O.) What keeps the water inside the glass? As you read about the properties of gases in this chapter, re-examine this activity and see if you can improve your explanation of what occurred.

THE NATURE OF GASES

section 10.1

You are walking your dog in the woods, enjoying the great outdoors. Suddenly your dog begins to bark at what you believe is a black cat. But before you realize what it actually is, the damage is done. It's too late; the skunk has released its spray! Within seconds you smell that all-too recognizable smell. **How do the gaseous odor molecules travel from one place to another?**

objectives

- Describe the motion of gas particles according to the kinetic theory
- Interpret gas pressure in terms of kinetic theory

key terms

- kinetic energy
- kinetic theory
- gas pressure
- vacuum
- atmospheric pressure
- barometers
- pascal (Pa)
- standard atmosphere (atm)

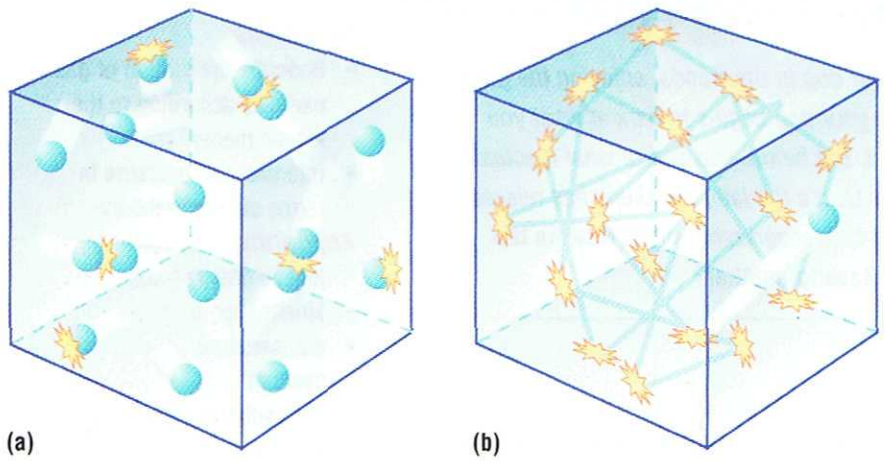
Kinetic Theory

An ice cube is solid water; tap water is liquid water; and steam is gaseous water. Most substances commonly exist in only one of the three states of matter: solid, liquid, or gas. However, a substance in one state may change to another state with a change in temperature. You know that when cooled, water freezes to become ice, and steam condenses to form water. You also know that soon after you put on perfume, anyone nearby will be able to smell it. Obviously, the molecules of the perfume have moved through the air. How do you account for these behaviors? A model called the kinetic theory will help you find the answer.

The word kinetic refers to motion. The energy an object has because of its motion is called **kinetic energy.** The **kinetic theory** states that the tiny particles in all forms of matter are in constant motion. The following are the basic assumptions of the kinetic theory as it applies to gases.

1. A gas is composed of particles, usually molecules or atoms. These particles are considered to be small, hard spheres that have insignificant volume and are relatively far apart from one another. Between the particles there is empty space. No attractive or repulsive forces exist between the particles.
2. The particles in a gas move rapidly in constant random motion. They travel in straight paths and move independently of each other. As a result, gases fill their containers regardless of the shape and volume of the containers; uncontained gases diffuse into space without limit. The gas particles change direction only when they rebound from collisions with one another or with other objects. Measurements indicate that the average speed of oxygen molecules in air at 20 °C is an amazing 1700 km/h! At these high speeds, the odor molecules from a hot cheese pizza in Washington, D.C., should reach Mexico City in about 106 minutes. That does not happen, however, because the odor molecules are constantly striking molecules of air and rebounding in other directions. Their path of uninterrupted travel in a straight line is very short. The aimless path the gas molecules take is called a random walk. **Figure 10.1** on the following page illustrates a typical random walk.
3. All collisions are perfectly elastic. This means that during collisions kinetic energy is transferred without loss from one particle to another, and the total kinetic energy remains constant.

As you will learn next, the kinetic theory of gases is very helpful in explaining gas pressure.



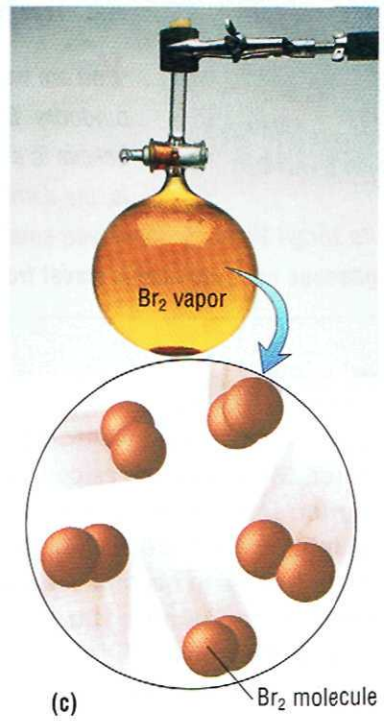


Figure 10.1

(a) *Gas particles move randomly. They are constantly colliding with one another and with the walls of their containers. The collisions cause the particles to change direction frequently.* ***(b)*** *Every gas particle travels in a straight-line path until it collides with another particle, object, or wall.* ***(c)*** *A molecular view of a gas shows that the particles are relatively far apart and fill all available space. In this example, the gas is bromine vapor (Br_2).*

Gas Pressure

Moving bodies exert forces when they collide with other bodies. Although a single gas particle is a moving body, the force it exerts is extremely small. Yet it is not hard to imagine that simultaneous collisions involving many such particles would produce a measurable force on an object. Gas pressure is the result of simultaneous collisions of billions of gas particles with an object. **Gas pressure** is defined as the force exerted by a gas per unit surface area of an object. If there are no gas particles present, there cannot be collisions, and there is consequently no pressure. Such an empty space, with no particles and no pressure, is called a **vacuum.**

Air exerts pressure on Earth because gravity holds air molecules in Earth's atmosphere. **Atmospheric pressure** results from the collisions of air molecules with objects. Atmospheric pressure decreases as you climb a mountain because the air layer around Earth thins out as elevation increases. **Barometers** are devices commonly used to measure atmospheric pressure. This pressure is dependent on weather.

The SI unit of pressure is the **pascal (Pa).** Atmospheric pressure at sea level is about 101.3 kilopascals (kPa). Two older units of pressure are millimeters of mercury (mm Hg) and atmospheres (atm). The origin of these units is the early use of mercury barometers similar to the one shown in **Figure 10.2.** Such a barometer was a straight glass tube filled with mercury and closed at one end. The tube was placed in a dish of mercury so that the open end was below the surface of mercury in the dish. The height of mercury in the tube depended on the pressure created by collisions of air molecules with the surface of the mercury in the dish. At sea level, this pressure is sufficient to support a mercury column about 760 mm high. One **standard atmosphere (atm)** is the pressure required to support 760 mm of mercury in a mercury barometer at 25 °C. That is, 1 atm = 760 mm Hg = 101.3 kPa.

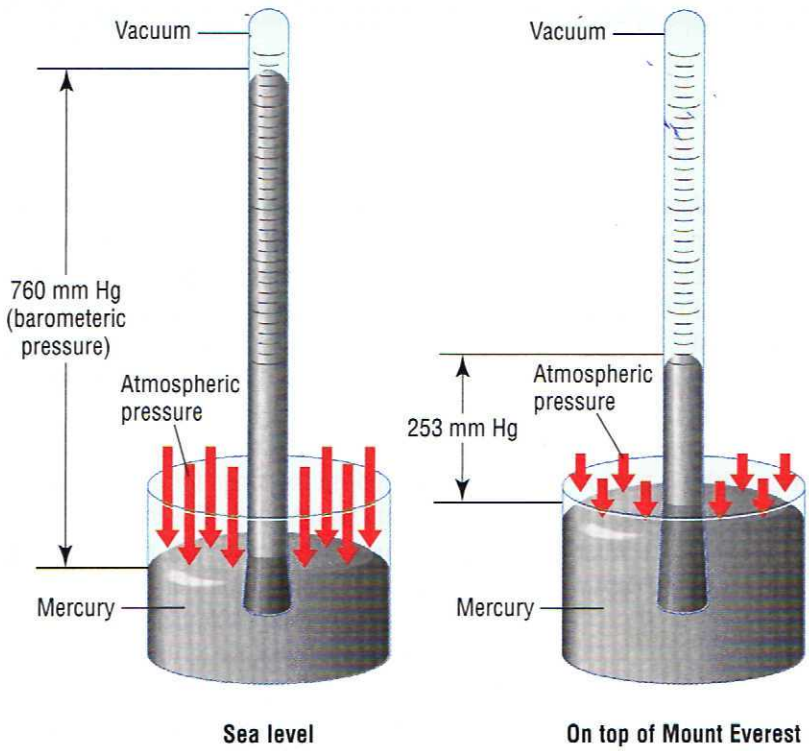


Figure 10.2
Typical atmospheric pressure at sea level in fair weather is 1 atm, which equals 760 mm Hg, or 101.3 kPa. Such atmospheric pressure supports a 760-mm column of mercury (101.3 kPa). On top of Mount Everest (at an altitude of 9000 m), the air exerts only enough pressure to support a 253-mm column of mercury. Express this pressure in kilopascals.

Many modern barometers do not contain mercury and are called aneroid barometers. In these devices, atmospheric pressure is related to the number of collisions of air molecules with a sensitive metal diaphragm. The diaphragm controls the movement of a pointer, which in turn indicates the pressure reading.

In the case of gases, it is important to be able to relate measured values to standards. Standard conditions are defined as a temperature of 0 °C and a pressure of 101.3 kPa, or 1 atm. This set of conditions is called standard temperature and pressure, or STP. What is standard temperature expressed on the Kelvin temperature scale?

Kinetic Energy and Kelvin Temperature

What happens when a substance is heated? The particles of the substance absorb energy, some of which is stored within the particles. This stored portion of the energy, or potential energy, does not raise the temperature of the

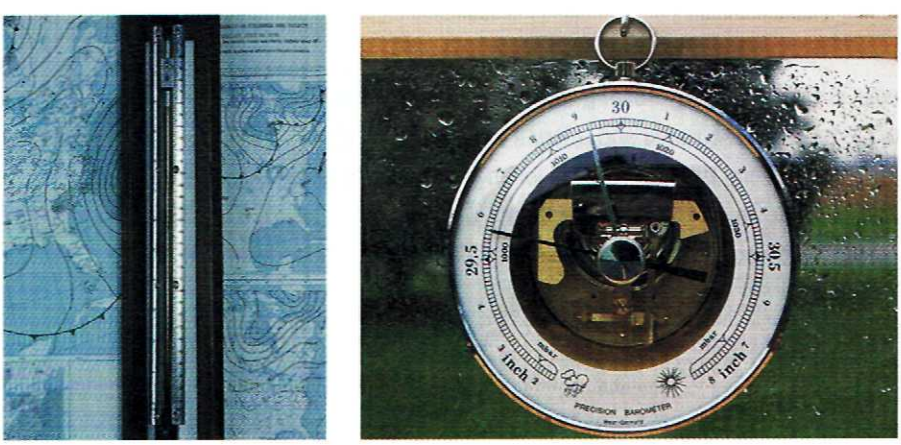


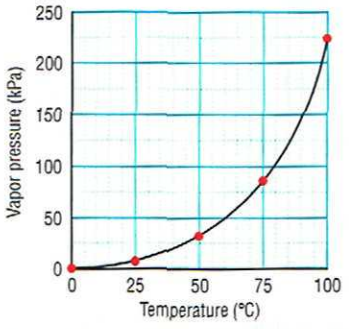
Figure 10.3
The standard lab barometer is still commonly used (left). An aneroid barometer (right) does not contain mercury.

CHEMath

MAKING AND INTERPRETING GRAPHS

A graph is used to display relationships or data visually. You will encounter many graphs in your study of chemistry. Graphs are also used in economics, physics, social studies, and other fields. Your local newspaper probably includes several graphs each week.

In a graph, the horizontal axis (or x-axis) represents the *independent variable* and the vertical axis (or y-axis) represents the *dependent variable.* For example, in the graph shown below, the independent variable is temperature, and the dependent variable is vapor pressure. So, this graph shows vapor pressure vs. temperature. When stating a relationship, you should always give the dependent variable first.



Vapor Pressure of Ethanol vs. Temperature

The *range* and *interval* for each axis determine the appearance of the graph. The range is given by the minimum and maximum values represented in the graph, and the interval is the distance between grid lines or tick marks. In the graph shown, the temperature range is 0 °C to 100 °C, and the temperature interval is 5 °C.

To make a graph, first choose appropriate ranges and scales. Choose the range for each axis so that all important data is included. The intervals should be convenient numbers (such as 1, 5, or 20) that result in an easy-to-read graph. After you draw and label the axes, plot the data. Each data point should be aligned with the correct values on the axes. When appropriate, connect the data with a smooth curve.

In Section 10.4, you will learn about the *phase diagram for a substance (see Figure 10.18).* A phase diagram is a pressure vs. temperature graph that has been divided into regions for each of the solid, liquid, and vapor phases. For a given pressure and temperature, you can determine from the graph if the substance is a solid, a liquid, or a gas.

Example 1

Refer to the graph above. Give the range and interval for vapor pressure. Then find the boiling point *of ethanol when the atmospheric pressure is 60 kPa.*

The vertical axis includes values from 0 kPa to 225 kPa with the grid lines 25 kPa apart. So, the vapor pressure range is 0 kPa to 225 kPa and the vapor pressure interval is 25 kPa.

To find the boiling point of ethanol when the atmospheric pressure is 60 kPa, notice that the curve appears to include the point (65 °C, 60 kPa). Therefore, the boiling point is about 65 °C at 60 kPa.

Practice Problems

Prepare for upcoming problems in this chapter by solving the following problems.

A. For the graph in Figure 10.11, what are the range and interval for temperature?

B. Refer to the graph above. Find the boiling point of ethanol when the atmospheric pressure is 90 kPa.

C. *Refer to the graph above. What is the vapor pressure of ethanol when its temperature is 50 °C?*

D. The table gives the vapor pressure of isopropyl alcohol (rubbing alcohol) at various temperatures. Graph the data. Connect the data points using a smooth curve.

Temperature (°C)	0	25	50	75	100	125
Vapor pressure (kPa)	1.11	6.02	23.9	75.3	198	452

Sample Problem 10-1

A gas is at a pressure of 1.50 atm. Convert this pressure to
a. kilopascals. **b.** millimeters of mercury.

1. ANALYZE ***List the knowns and the unknowns.***

Knowns:
- pressure = 1.50 atm
- 1 atm = 101.3 kPa
- 1 atm = 760 mm Hg

Unknowns:
a. pressure = ? kPa
b. pressure = ? mm Hg

The given pressure is converted into the desired unit by multiplying by the proper conversion factor.

2. CALCULATE ***Solve for the unknowns.***

a. For the conversion atm ⟶ kPa, the conversion factor is $\frac{101.3 \text{ kPa}}{1 \text{ atm}}$.

$$1.50 \text{ atm} \times \frac{101.3 \text{ kPa}}{1 \text{ atm}} = 151.95 \text{ kPa} = 1.52 \times 10^2 \text{ kPa}$$

b. For the conversion atm ⟶ mm Hg, the conversion factor is $\frac{760 \text{ mm Hg}}{1 \text{ atm}}$.

$$1.50 \text{ atm} \times \frac{760 \text{ mm Hg}}{1 \text{ atm}} = 1140 \text{ mm Hg} = 1.14 \times 10^3 \text{ mm Hg}$$

3. EVALUATE ***Do the results make sense?***

Because the conversion factor for calculating mm Hg is larger than the one for kPa, it makes sense that the value expressed in mm Hg is larger than the value in kPa. In each case the given unit cancels, and the answer has the desired unit. Each answer has the three significant figures required.

Practice Problems

1. What pressure, in kilopascals and in atmospheres, does a gas exert at 385 mm Hg?
2. The pressure at the top of Mount Everest is 33.7 kPa. Is that pressure greater or less than 0.25 atm?

Chem ASAP!

Problem-Solving 1

Solve Problem 1 with the help of an interactive guided tutorial.

substance. The remaining absorbed energy speeds up the particles—that is, increases their average kinetic energy—which results in an increase in temperature. The particles in any collection of atoms or molecules at a given temperature have a wide range of kinetic energies, from very low to very high. Most of the particles have kinetic energies somewhere in the middle of this range. Therefore average kinetic energy is used when discussing the kinetic energy of a collection of particles in a substance. **Figure 10.4** on the following page shows the distribution of kinetic energies of gas particles at two different temperatures. Notice that at the higher temperature there is a wider range of kinetic energies.

An increase in the average kinetic energy of particles causes the temperature of a substance to rise. As a substance cools, the particles tend to move more slowly, and their average kinetic energy declines. You could reasonably expect the particles of all substances to stop moving at some very low temperature. The particles would have no kinetic energy at that temperature because they would have no motion. Absolute zero (0 K, or

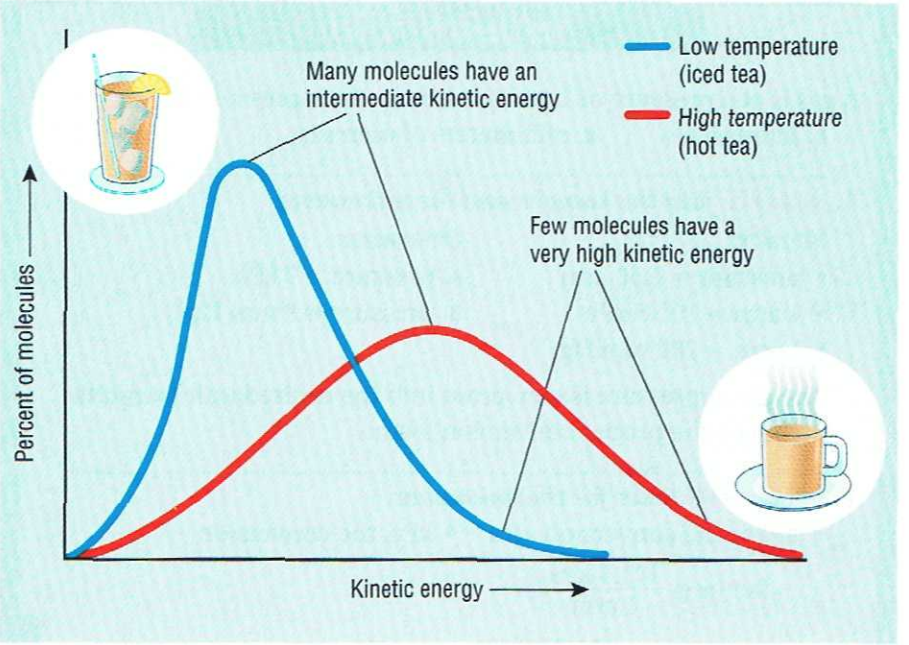


Figure 10.4
The blue curve shows the kinetic energy distribution of a typical collection of molecules at a fairly low temperature, such as the water molecules in iced tea. Notice that most molecules have intermediate energies. The red curve shows the energy distribution of molecules at a higher temperature, such as the water molecules in hot tea. How do the average kinetic energies of the two liquids compare?

Chem ASAP!

Animation 9
Observe particles in motion and discover the connection between temperature and kinetic energy.

−273.15 °C) is the temperature at which the motion of particles theoretically ceases. Absolute zero has never been produced in the laboratory, although temperatures of about 0.000 001 K have been achieved. Would you expect to find negative temperatures on the Kelvin temperature scale?

The Kelvin temperature scale reflects the relationship between temperature and average kinetic energy. The Kelvin temperature of a substance is directly proportional to the average kinetic energy of the particles of the substance. For example, the particles in helium gas at 200 K have twice the average kinetic energy as the particles in helium gas at 100 K. As you will see later in this chapter, the effects of temperature on particle motion in liquids and solids are more complex than in gases. Nevertheless, at any given temperature the particles of all substances, regardless of physical state, have the same average kinetic energy.

LINK TO MEDICINE

Cryogenics
Cryogenics is the science of producing very low temperatures and studying the behavior of matter at such temperatures. Biological reactions slow down or even stop at very low temperatures. Biologists use this fact to study reactions more easily. In medicine, some tissues can be removed from the body and preserved by rapid freezing. They can later be used when thawed. Blood and cartilage are examples of such materials. Doctors also perform cryosurgery by freezing tissue instead of cutting it. Cryosurgery minimizes bleeding during operations, and healing is rapid with minimal scarring.

section review 10.1

3. According to the assumptions of kinetic theory, how do the particles in a gas move?
4. Use kinetic theory to explain what causes gas pressure.
5. Express the pressure 545 mm Hg in kilopascals.
6. How can you raise the average kinetic energy of the water molecules in a glass of water?
7. A cylinder of oxygen gas is cooled from 300 K (27 °C) to 150 K (−123 °C). By what factor does the average kinetic energy of the oxygen molecules in the cylinder decrease?

Chem ASAP! Assessment 10.1 *Check your understanding of the important ideas and concepts in Section 10.1.*

SMALL-SCALE LAB

KINETIC THEORY IN ACTION

SAFETY

Wear safety glasses and follow the standard safety procedures as outlined on page 18.

PURPOSE

To observe color changes in the chemical reactions of gases and to interpret these changes in terms of kinetic theory.

MATERIALS

- clear plastic cup or Petri dish
- pencil
- paper
- ruler
- cotton swab
- medicine dropper
- reaction surface
- chemicals shown in Figure A and KI, $NaNO_2$, NH_4Cl, and NaOH

PROCEDURE

Use a clear plastic cup or Petri dish as a template to draw a large circle, as shown in Figure A. Place a reaction surface over the grid and add small drops of BTB (bromthymol blue) in the pattern shown by the small circles. Be sure the drops do not touch one another. Mix one drop each of HCl and $NaHSO_3$ in the center of the pattern. Place a clear plastic cup or Petri dish over the grid and observe what happens. Do not clean up until after you complete any assigned YOU'RE THE CHEMIST activities.

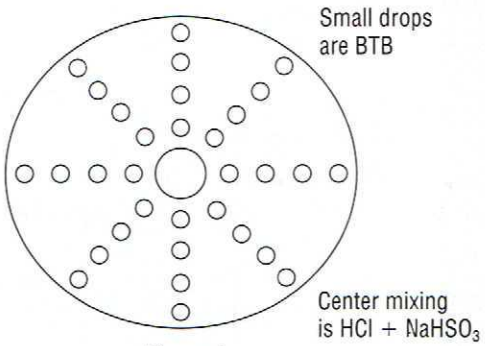


Figure A

ANALYSIS

Using your experimental data, record the answers to the following questions below your data table.

1. Describe in detail the changes you observed in the drops of BTB over time. Draw pictures to illustrate the changes.
2. Draw a series of pictures showing how one of the BTB drops might look over time if you could view the drop from the side.
3. The BTB changed even though you added nothing to it. If the mixture in the center circle produced a gas, would this explain the change in the drops of BTB? Use kinetic theory to explain your answer.
4. Translate the following word equation into a balanced chemical equation: Sodium hydrogen sulfite reacts with hydrochloric acid to produce sulfur dioxide gas, water, and sodium chloride.

YOU'RE THE CHEMIST

The following small-scale activities allow you to develop your own procedures and analyze the results.

1. **Analyze It!** Carefully absorb the center mixture of the original experiment onto a cotton swab and replace it with one drop of NaOH and one drop of NH_4Cl. What happens? Explain in terms of kinetic theory. Ammonium chloride reacts with sodium hydroxide to produce ammonia gas, water, and sodium chloride. Write and balance a chemical equation to describe this reaction.
2. **Design It!** Design an experiment to observe the effect of the size of the BTB drops on the rate at which they change. Explain your results in terms of kinetic theory.
3. **Analyze It!** Repeat the original experiment, except use KI in place of BTB and mix sodium nitrite ($NaNO_2$) with hydrochloric acid (HCl) at the center. Record your results. Write and balance an equation: Sodium nitrite reacts with hydrochloric acid to produce nitrogen monoxide gas, water, sodium nitrate, and sodium chloride.

section 10.2

THE NATURE OF LIQUIDS

objectives

- Describe the nature of a liquid in terms of the attractive forces between the particles
- Differentiate between evaporation and boiling of a liquid, using kinetic theory

key terms

- vaporization
- evaporation
- vapor pressure
- boiling point
- normal boiling point

The Kilauea volcano in Hawaii is the most active volcano in the world—it has been erupting for centuries. The hot lava oozes and flows, scorching everything in its path, occasionally including nearby houses. When the lava cools, it solidifies into rock. **What makes a liquid different from a solid?**

A Model for Liquids

You have just learned from the kinetic theory that gas pressure can be explained by assuming that gas particles have motion and that there is no attraction between the particles. The particles that make up liquids are also in motion, as **Figure 10.5** shows. Liquid particles are free to slide past one another. For that reason, both liquids and gases can flow, as you can see in **Figure 10.6.** However, the particles in a liquid are attracted to each other, while according to kinetic theory, those in a gas are not. The attractive forces between the molecules are called intermolecular forces.

The particles that make up liquids vibrate and spin while they move from place to place. All of these motions contribute to the average kinetic energy of the particles. Even so, most of the particles do not have enough kinetic energy to escape into the gaseous state. To do so, a particle must have sufficient kinetic energy to overcome the intermolecular forces that hold it together with the other particles. The intermolecular forces also reduce the amount of space between the particles in a liquid. Thus liquids are much more dense than gases. Increasing the pressure on a liquid has hardly any effect on its volume. The same is true of solids. For that reason, liquids and solids are known as condensed states of matter.

The interplay between the disruptive motions of particles and the attractive forces between them determines many of the physical properties of liquids. You will now explore two of these properties: vapor pressure and boiling point.

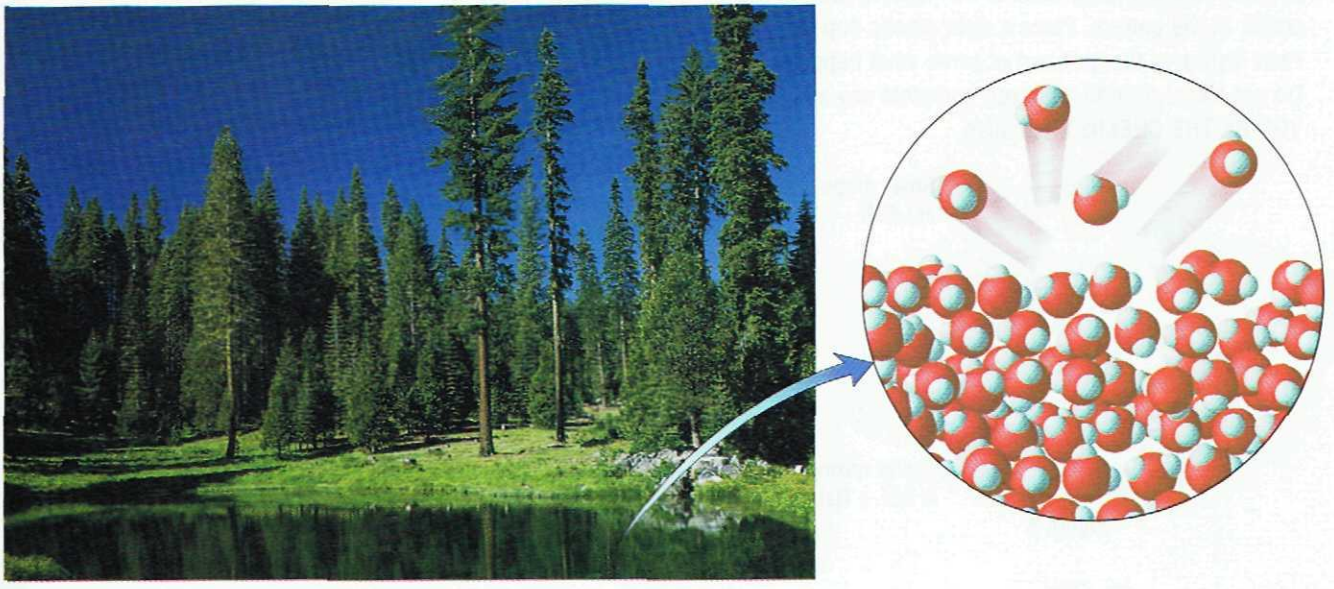


Figure 10.5
The particles in a liquid, such as the water in this lake, are close together but can move and slide around one another. The attractive forces between the particles generally prevent most of the particles from escaping the liquid and entering into the vapor state.