Physical Characteristics of Gases

The density of a gas decreases as its temperature increases.

SECTION 10-1

The Kinetic-Molecular Theory of Matter

OBJECTIVES

- State the kinetic-molecular theory of matter, and describe how it explains certain properties of matter.
- List the five assumptions of the kinetic-molecular theory of gases. Define the terms *ideal gas* and *real gas*.
- Describe each of the following characteristic properties of gases: expansion, density, fluidity, compressibility, diffusion, and effusion.
- Describe the conditions under which a real gas deviates from "ideal" behavior.

In Chapter 1, you read that matter exists on Earth in the forms of solids, liquids, and gases. Although it is not usually possible to observe individual particles directly, scientists have studied large groups of these particles as they occur in solids, liquids, and gases.

In the late nineteenth century, scientists developed the kinetic-molecular theory of matter to account for the behavior of the atoms and molecules that make up matter. *The* **kinetic-molecular theory** *is based on the idea that particles of matter are always in motion*. The theory can be used to explain the properties of solids, liquids, and gases in terms of the energy of particles and the forces that act between them. In this section, you will study the theory as it applies to gas molecules. In that form, it is called the kinetic-molecular theory of gases.

The Kinetic-Molecular Theory of Gases

The kinetic-molecular theory can help you understand the behavior of gas molecules and the physical properties of gases. The theory provides a model of what is called an ideal gas. *An* **ideal gas** *is an imaginary gas that perfectly fits all the assumptions of the kinetic-molecular theory.*

The kinetic-molecular theory of gases is based on the following five assumptions:

1. *Gases consist of large numbers of tiny particles that are far apart relative to their size.* These particles, usually molecules or atoms, typically occupy a volume about 1000 times greater than the volume occupied by particles in the liquid or solid state. Thus, molecules of gases are much farther apart than those of liquids or solids. Most of the volume occupied by a gas is empty space. This accounts for the lower density of gases compared with that of liquids and solids. It also explains the fact that gases are easily compressed.
2. *Collisions between gas particles and between particles and container walls are elastic collisions. An* **elastic collision** *is one in which there is no net loss of kinetic energy.* Kinetic energy is transferred between two particles during collisions. However, the total kinetic energy of the two particles remains the same as long as temperature is constant.

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Module 1: States of Matter/Classes of Matter

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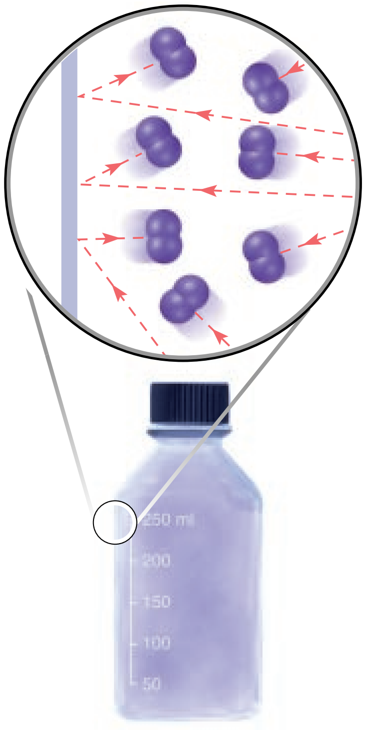

FIGURE 10-1 Gas particles travel in a straight-line motion until they collide with each other or the walls of their container.

3. *Gas particles are in continuous, rapid, random motion. They therefore possess kinetic energy, which is energy of motion.* Gas particles move in all directions, as shown in Figure 10-1. The kinetic energy of the particles overcomes the attractive forces between them, except near the temperature at which the gas condenses and becomes a liquid.
4. *There are no forces of attraction or repulsion between gas particles.* You can think of ideal gas molecules as behaving like small billiard balls. When they collide, they do not stick together but immediately bounce apart.
5. *The average kinetic energy of gas particles depends on the temperature of the gas.* The kinetic energy of any moving object, including a particle, is given by the following equation.

$$KE = \frac{1}{2}mv^2$$

In the equation, m is the mass of the particle and v is its speed. Because all the particles of a specific gas have the same mass, their kinetic energies depend only on their speeds. The average speeds and kinetic energies of gas particles increase with an increase in temperature and decrease with a decrease in temperature.

All gases at the same temperature have the same average kinetic energy. Therefore, at the same temperature, lighter gas particles, such as hydrogen molecules, have higher average speeds than do heavier gas particles, such as oxygen molecules.

The Kinetic-Molecular Theory and the Nature of Gases

The kinetic-molecular theory applies only to ideal gases. Although ideal gases do not actually exist, many gases behave nearly ideally if pressure is not very high or temperature is not very low. In the following sections, you will see how the kinetic-molecular theory accounts for the physical properties of gases.

Expansion

Gases do not have a definite shape or a definite volume. They completely fill any container in which they are enclosed, and they take its shape. A gas transferred from a one-liter vessel to a two-liter vessel will quickly expand to fill the entire two-liter volume. The kinetic-molecular theory explains these facts. According to the theory, gas particles move rapidly in all directions (assumption 3) without significant attraction or repulsion between them (assumption 4).

Fluidity

Because the attractive forces between gas particles are insignificant (assumption 4), gas particles glide easily past one another. This ability to flow causes gases to behave similarly to liquids. *Because liquids and gases flow, they are both referred to as* **fluids.**

Low Density

The density of a substance in the gaseous state is about 1/1000 the density of the same substance in the liquid or solid state. That is because the particles are so much farther apart in the gaseous state (assumption 1).

Compressibility

During compression, the gas particles, which are initially very far apart (assumption 1), are crowded closer together. The volume of a given sample of a gas can be greatly decreased. Steel cylinders containing gases under pressure are widely used in industry. When they are full, such cylinders may contain 100 times as many particles of gas as would be contained in nonpressurized containers of the same size.

Diffusion and Effusion

Gases spread out and mix with one another, even without being stirred. If the stopper is removed from a container of ammonia in a room, ammonia gas will mix uniformly with the air and spread throughout the room. The random and continuous motion of the ammonia molecules (assumption 3) carries them throughout the available space. *Such spontaneous mixing of the particles of two substances caused by their random motion is called* **diffusion.**

The rate of diffusion of one gas through another depends on three properties of the gas particles: their speeds, their diameters, and the attractive forces between them. In Figure 10-2, hydrogen gas diffuses rapidly into other gases at the same temperature because its molecules are lighter and move faster than the molecules of the other gases.

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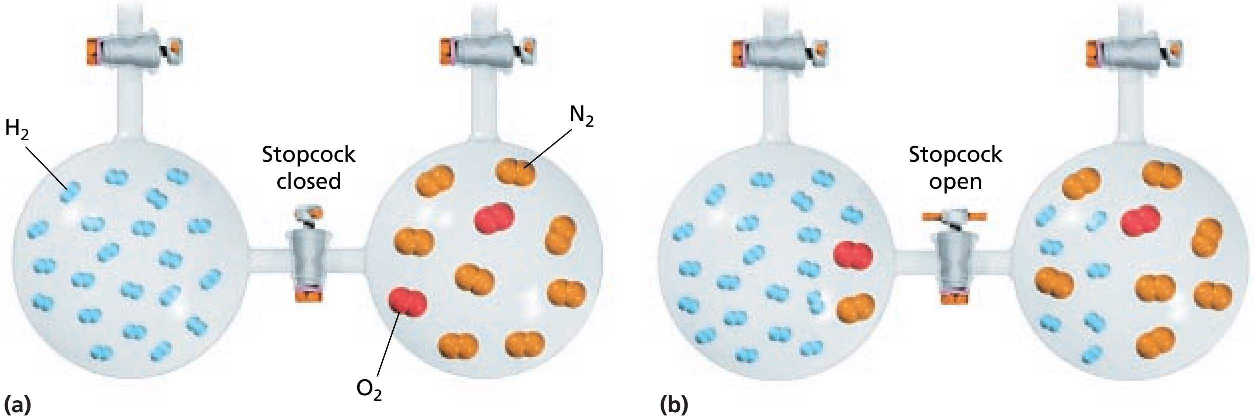

FIGURE 10-2 When hydrogen gas in a flask is allowed to mix with air at the same pressure in another flask, the low-mass molecules of hydrogen diffuse rapidly into the flask with the air. The heavier molecules of nitrogen and oxygen in the air diffuse more slowly into the flask with the hydrogen.

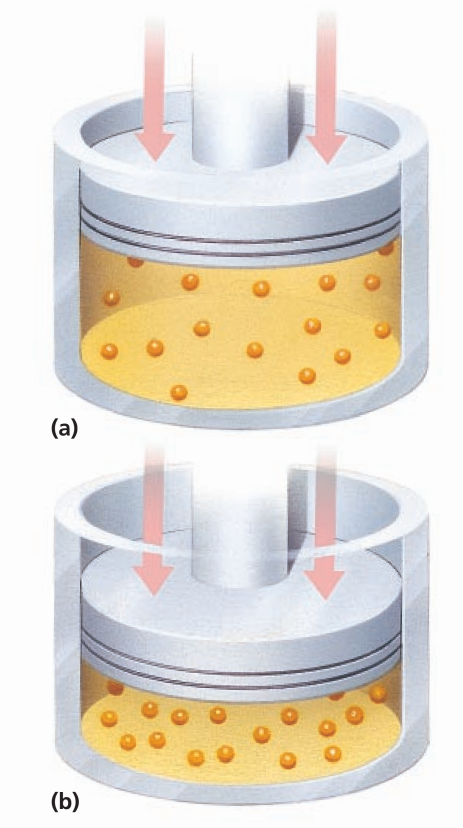

FIGURE 10-3 (a) Gas molecules in a car engine cylinder expand to fill the cylinder. (b) As pressure is exerted on them, the gas molecules move closer together, reducing their volume.

Diffusion is a process by which particles of a gas spread out spontaneously and mix with other gases. In contrast, **effusion** *is a process by which gas particles pass through a tiny opening*. The rates of effusion of different gases are directly proportional to the velocities of their particles. Because of this proportionality, molecules of low mass effuse faster than molecules of high mass.

Deviations of Real Gases from Ideal Behavior

When their particles are far enough apart and have enough kinetic energy, most gases behave ideally. However, all real gases deviate to some degree from ideal-gas behavior. *A* **real gas** *is a gas that does not behave completely according to the assumptions of the kinetic-molecular theory*. In 1873, Johannes van der Waals accounted for this deviation from ideal behavior by pointing out that particles of real gases occupy space and exert attractive forces on each other. At very high pressures and low temperatures, the deviation may be considerable. Under such conditions, the particles will be closer together and their kinetic energy will be insufficient to completely overcome the attractive forces. These conditions are illustrated in Figure 10-3.

The kinetic-molecular theory is more likely to hold true for gases whose particles have little attraction for each other. The noble gases, such as helium, He, and neon, Ne, show essentially ideal gas behavior over a wide range of temperatures and pressures. The particles of these gases are monatomic and thus nonpolar. The particles of gases, such as nitrogen, N_2, and hydrogen, H_2, are nonpolar diatomic molecules. The behavior of these gases most closely approximates that of the ideal gas under certain conditions. The more polar a gas's molecules are, the greater the attractive forces between them and the more the gas will deviate from ideal gas behavior. For example, highly polar gases, such as ammonia, NH_3, and water vapor, deviate from ideal behavior to a larger degree than nonpolar gases.

SECTION REVIEW

1. Use the kinetic-molecular theory to explain each of the following properties of gases: expansion, fluidity, low density, compressibility, and diffusion.
2. Describe the conditions under which a real gas is most likely to behave ideally.
3. State the two factors that van der Waals proposed to explain why real gases deviate from ideal behavior.
4. Which of the following gases would you expect to deviate significantly from ideal behavior: He, O_2, H_2, H_2O, N_2, HCl, or NH_3?

Carbon Monoxide Catalyst—Stopping the Silent Killer

Colorless, odorless, and deadly—carbon monoxide, "the silent killer," causes the deaths of hundreds of Americans every year. When fuel does not burn completely in a combustion process, carbon monoxide is produced. Often this occurs in a malfunctioning heater, furnace, or fireplace. When the carbon monoxide is inhaled, it bonds to the hemoglobin in the blood, leaving the body oxygen starved. Before people realize a combustion device is malfunctioning, it's often too late.

$$O_2Hb + CO \longrightarrow COHb + O_2$$

Carbon monoxide, CO, has almost 200 times the affinity to bind with the hemoglobin, Hb, in the blood as oxygen. This means if the body has a choice, it will bind to carbon monoxide over oxygen. If enough carbon monoxide is present in the blood, it can be fatal.

Carbon monoxide poisoning can be prevented by installing filters that absorb the gas. After a time, however, filters become saturated, and then carbon monoxide can pass freely into the air. The best way to prevent carbon monoxide poisoning is not just to filter out the gas, but to eliminate it completely.

The solution came to research chemists at NASA who were working on a problem with a space-based laser. In order to operate properly, NASA's space-based carbon dioxide laser needed to be fed a continuous supply of CO_2. This was necessary because as a byproduct of its operation, the laser degraded some of the CO_2 into carbon monoxide and oxygen. To address this problem, NASA scientists developed a catalyst made of tin oxide and platinum that oxidized the waste carbon monoxide back into carbon dioxide. The NASA scientists then realized that this catalyst had the potential to be used in many applications here on Earth, including removing carbon monoxide from houses and other buildings.

Typically, a malfunctioning heater circulates the carbon monoxide it produces through its air intake system back into a dwelling space. By installing the catalyst in the air intake, any carbon monoxide would be oxidized to non-toxic carbon dioxide before it reentered the room.

"The form of our catalyst is a very thin coating on some sort of a support, or substrate as we call it," says NASA chemist David Schryer. "And that support, or substrate, can be any one of a number of things. The great thing about a catalyst is that the only thing that matters about it is its surface. So a catalyst can be incredibly thin and still be very effective."

The idea of using catalysts to oxidize gases is not a new one. Catalytic converters in cars oxidize carbon monoxide and unburned hydrocarbons to minimize pollution. Many substances are oxidized into new materials for manufacturing purposes. But both of these types of catalytic reactions occur at very high temperatures. NASA's catalyst is special, because it's able to eliminate carbon monoxide at room temperature.

According to David Schryer, low-temperature catalysts constitute a whole new class of catalysts with abundant applications for the future.

How did NASA's research on the space-based carbon dioxide laser result in a benefit for consumers?

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SECTION 10-2

Pressure

OBJECTIVES

- Define *pressure* and relate it to force.
- Describe how pressure is measured.
- Convert units of pressure.
- State the standard conditions of temperature and pressure.

Suppose you have a one-liter bottle of air. How much air do you actually have? The expression *a liter of air* means little unless the conditions at which the volume is measured are known. A liter of air can be compressed to a few milliliters. It can also be allowed to expand to fill an auditorium.

To describe a gas fully, you need to state four measurable quantities: volume, temperature, number of molecules, and pressure. You already know what is meant by volume, temperature, and number of molecules. In this section, you will learn about pressure and its measurement. Then, in Section 10-3, you will examine the mathematical relationships between volume, temperature, number of gas molecules, and pressure.

Pressure and Force

If you blow air into a rubber balloon, the balloon will increase in size. The volume increase is caused by the collisions of molecules of air with the inside walls of the balloon. The collisions cause an outward push, or force, against the inside walls. **Pressure** *(P) is defined as the force per unit area on a surface*. The equation defining pressure follows.

FIGURE 10-4 The pressure the ballet dancer exerts against the floor depends on the area of contact. The smaller the area of contact, the greater the pressure.

(a) Area of contact = 325 cm²

$$\text{Pressure} = \frac{\text{force}}{\text{area}} = \frac{500\ \text{N}}{325\ \text{cm}^2} = 1.5\ \text{N/cm}^2$$

(b) Area of contact = 13 cm²

$$\text{Pressure} = \frac{\text{force}}{\text{area}} = \frac{500\ \text{N}}{13\ \text{cm}^2} = 38.5\ \text{N/cm}^2$$

(c) Area of contact = 6.5 cm²

$$\text{Pressure} = \frac{\text{force}}{\text{area}} = \frac{500\ \text{N}}{6.5\ \text{cm}^2} = 77\ \text{N/cm}^2$$

$$pressure = \frac{force}{area}$$

The SI unit for force is the **newton,** abbreviated N. *It is the force that will increase the speed of a one kilogram mass by one meter per second each second it is applied.* At Earth's surface, each kilogram of mass exerts 9.8 N of force, due to gravity. Consider a ballet dancer with a mass of 51 kg, as shown in Figure 10-4. A mass of 51 kg exerts a force of 500 N (51×9.8) on Earth's surface. No matter how the dancer stands, she exerts that much force against the floor. However, the pressure she exerts against the floor depends on the area of contact. When she rests her weight on the soles of both feet, as shown in Figure 10-4(a), the area of contact with the floor is about 325 cm^2. The pressure, or force per unit area, when she stands in this manner is 500 N/325 cm^2. That equals roughly 1.5 N/cm^2. When she stands on her toes, as in Figure 10-4(b), the total area of contact with the floor is only 13 cm^2. The pressure exerted is then equal to 500 N/13 cm^2—roughly 38.5 N/cm^2. And when she stands on one toe, as in Figure 10-4(c), the pressure she exerts is twice that, or about 77 N/cm^2. Thus, the same force applied to a smaller area results in a greater pressure.

Gas molecules exert pressure on any surface with which they collide. The pressure exerted by a gas depends on volume, temperature, and the number of molecules present.

The atmosphere—the blanket of air surrounding Earth—exerts pressure. Figure 10-5 shows that atmospheric pressure at sea level is about equal to the weight of a 1.03 kg mass per square centimeter of surface, or 10.1 N/cm^2. The pressure of the atmosphere can be thought of as caused by the weight of the gases that compose the atmosphere. The atmosphere contains about 78% nitrogen, 21% oxygen, and 1% other gases, including

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Module 6: Gas Laws

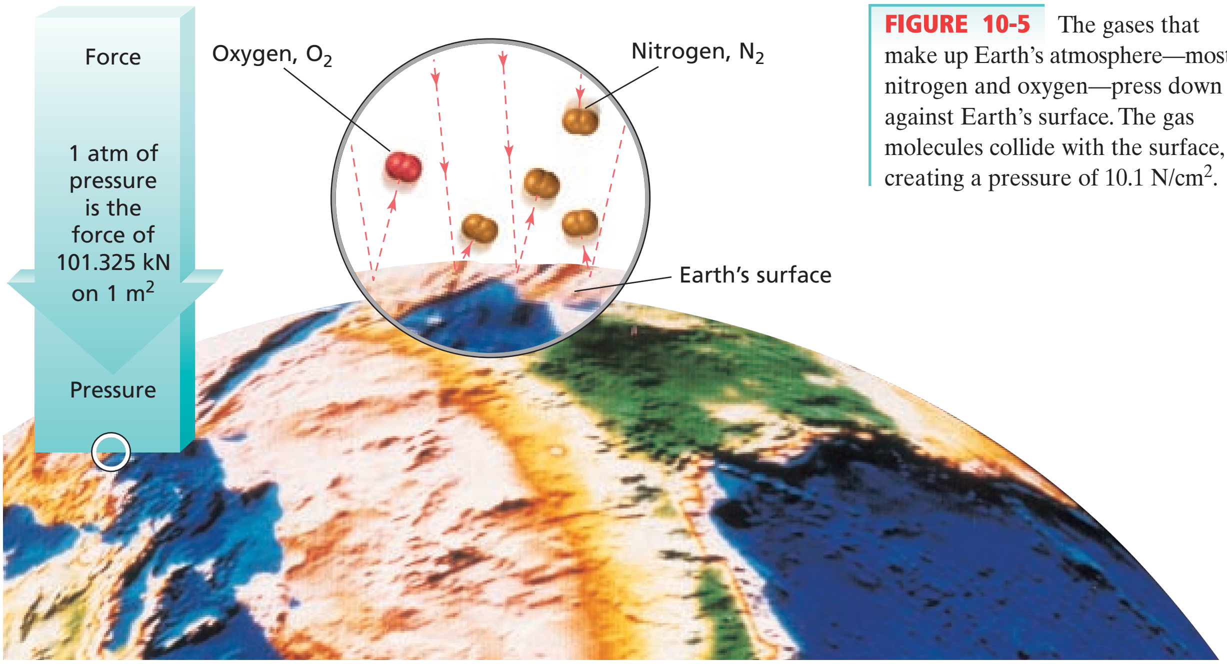

FIGURE 10-5 The gases that make up Earth's atmosphere—mostly nitrogen and oxygen—press down against Earth's surface. The gas molecules collide with the surface, creating a pressure of 10.1 N/cm^2.

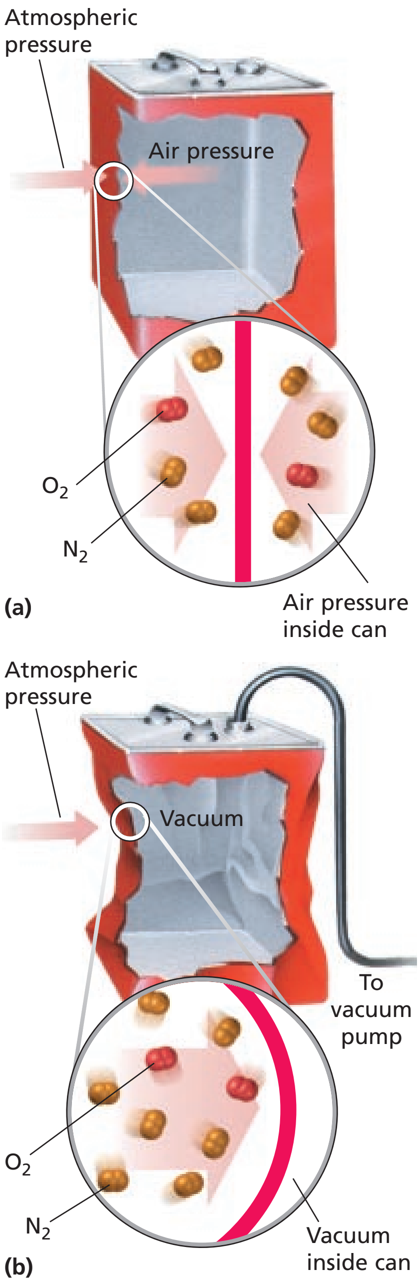

FIGURE 10-6 (a) The "empty" can has the mixture of gases in air inside that push outward and balance the atmospheric pressure that pushes inward. (b) When the air inside the can is removed by a vacuum pump, there is insufficient force to balance the atmospheric pressure. As a result, the can collapses.

argon and carbon dioxide. Atmospheric pressure is the sum of the individual pressures of the various gases in the atmosphere.

To understand the concept of gas pressure and its magnitude, consider the model of an "empty" can shown in Figure 10-6(a). The can does contain a small amount of air. The atmosphere exerts a pressure of 10.1 N/cm^2 against the outside of the can. If the can measures 15 cm × 10 cm × 28 cm, it has a total area of 1700 cm^2. The resulting inward force on the can is greater than 1.0 metric ton of weight. The air inside the can pushes outward and balances the atmosphere's inward-pushing force. If a vacuum pump is used to remove the air from the can, as shown in Figure 10-6(b), the balancing outward force is removed. As a result, the unbalanced force due to atmospheric pressure immediately crushes the can.

Measuring Pressure

A **barometer** *is a device used to measure atmospheric pressure.* The first type of barometer, illustrated in Figure 10-7, was introduced by Evangelista Torricelli during the early 1600s. Torricelli wondered why water pumps could raise water to a maximum height of only about 34 feet. He thought that the height must depend somehow on the weight of water compared with the weight of air. He reasoned that liquid mercury, which is about 14 times as dense as water, could be raised only 1/14 as high as water. To test this idea, Torricelli sealed a long glass tube at one end and filled it with mercury. Holding the open end with his thumb, he inverted the tube into a dish of mercury without allowing any air to enter the tube. When he removed his thumb, the mercury column in the tube dropped to a height of about 30 in. (760 mm) above the surface of the mercury in the dish. He repeated the experiment with tubes of different diameters and lengths longer than 760 mm. In every case, the mercury dropped to a height of about 760 mm.

The space above the mercury in such a tube is nearly a vacuum. The mercury in the tube pushes downward because of gravitational force. The column of mercury in the tube is stopped from falling beyond a certain point because the atmosphere exerts a pressure on the surface of the mercury outside the tube. This pressure is transmitted through the fluid mercury and is exerted upward on the column of mercury. The mercury in the tube falls only until the pressure exerted by its weight is equal to the pressure exerted by the atmosphere.

The exact height of the mercury in the tube depends on the atmospheric pressure, or force per unit area. The pressure is measured directly in terms of the height of the mercury column supported in the barometer tube.

From experiments like Torricelli's, it is known that at sea level at 0°C, the average pressure of the atmosphere can support a 760 mm column of mercury. At any given place on Earth, the specific atmospheric pressure depends on the elevation and the weather conditions at the time. If the atmospheric pressure is greater than the average at sea level, the height of the mercury column in a barometer will be greater than 760 mm. If the atmospheric pressure is less, the height of the mercury column will be less than 760 mm.

All gases, not only those in the atmosphere, exert pressure. A device called a manometer can be used to measure the pressure of an enclosed gas sample, as shown in Figure 10-8. The difference in the height of mercury in the two arms of the U-tube is a measure of the oxygen gas pressure in the container.

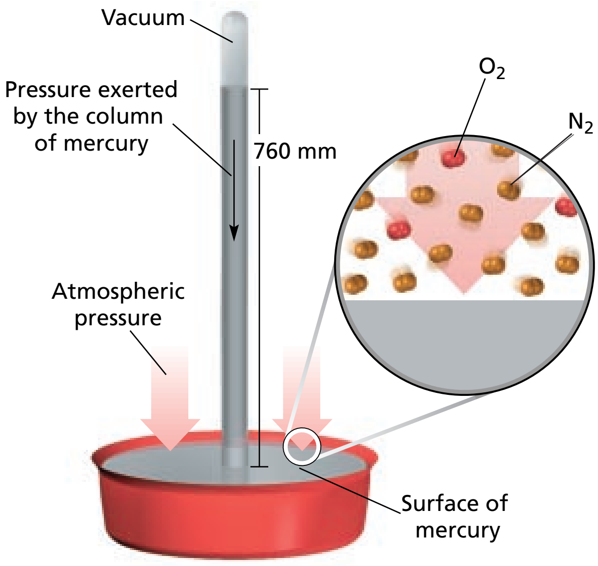

FIGURE 10-7 Torricelli discovered that the pressure of the atmosphere supports a column of mercury about 760 mm above the surface of the mercury in the dish.

Units of Pressure

A number of different units are used to measure pressure. Because atmospheric pressure is often measured by a mercury barometer, pressure can be expressed in terms of the height of a mercury column. *Thus, a common unit of pressure is* **millimeters of mercury,** *symbolized mm Hg. A pressure of 1 mm Hg is now called 1* **torr** in honor of Torricelli for his invention of the barometer. The average atmospheric pressure at sea level at 0°C is 760 mm Hg.

Pressures are often measured in units of atmospheres. *One* **atmosphere of pressure** *(atm) is defined as being exactly equivalent to 760 mm Hg.*

In SI, pressure is expressed in derived units called pascals. The unit is named for Blaise Pascal, a French mathematician and philosopher who studied pressure during the seventeenth century. *One* **pascal** *(Pa) is defined as the pressure exerted by a force of one newton (1 N) acting on an area of one square meter.*

In many cases, it is more convenient to express pressure in kilopascals (kPa). The standard atmosphere (1 atm) is equal to $1.013\,25 \times 10^5$ Pa, or 101.325 kPa. The pressure units used in this book are summarized in Table 10-1.

TABLE 10-1 *Units of Pressure*

Unit	Symbol	Definition/relationship
pascal	Pa	SI pressure unit $1\ \text{Pa} = \frac{1\ \text{N}}{\text{m}^2}$
millimeter of mercury	mm Hg	pressure that supports a 1 mm mercury column in a barometer
torr	torr	1 torr = 1 mm Hg
atmosphere	atm	average atmospheric pressure at sea level and 0°C 1 atm = 760 mm Hg = 760 torr $= 1.013\,25 \times 10^5$ Pa = 101.325 kPa

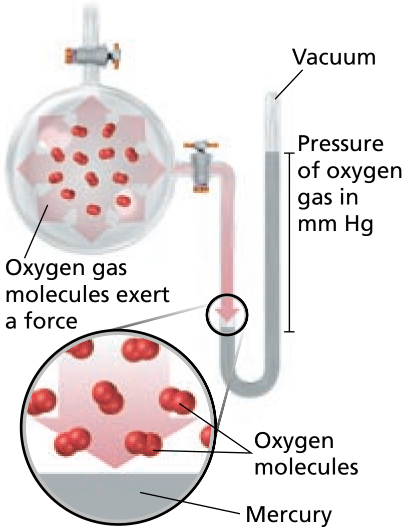

FIGURE 10-8 In the manometer above, the pressure of the oxygen gas in the flask pushes on the mercury column. The difference in the height of the mercury in the two arms of the U-tube indicates the oxygen gas pressure.

Standard Temperature and Pressure

To compare volumes of gases, it is necessary to know the temperature and pressure at which the volumes are measured. *For purposes of comparison, scientists have agreed on standard conditions of exactly 1 atm pressure and 0°C. These conditions are called* **standard temperature and pressure** *and are commonly abbreviated STP.*

SAMPLE PROBLEM 10-1

The average atmospheric pressure in Denver, Colorado, is 0.830 atm. Express this pressure (a) in mm Hg and (b) in kPa.

SOLUTION

1 *ANALYZE*

Given: P of atmosphere = 0.830 atm

760 mm Hg = 1 atm (definition); 101.325 kPa = 1 atm (definition)

Unknown: **a.** P of atmosphere in mm Hg; **b.** P of atmosphere in kPa

2 *PLAN*

a. $\text{atm} \longrightarrow \text{mm Hg}; \quad \text{atm} \times \frac{\text{mm Hg}}{\text{atm}} = \text{mm Hg}$

b. $\text{atm} \longrightarrow \text{kPa}; \quad \text{atm} \times \frac{\text{kPa}}{\text{atm}} = \text{kPa}$

3 *COMPUTE*

a. $0.830\ \cancel{\text{atm}} \times \frac{760\ \text{mm Hg}}{\cancel{\text{atm}}} = 631\ \text{mm Hg}$

b. $0.830\ \cancel{\text{atm}} \times \frac{101.325\ \text{kPa}}{\cancel{\text{atm}}} = 84.1\ \text{kPa}$

4 *EVALUATE*

Units have canceled to give the desired units, and answers are properly expressed to the correct number of significant figures. The known pressure is roughly 80% of atmospheric pressure. The results are therefore reasonable because each is roughly 80% of the pressure as expressed in the new units.

PRACTICE

1. Convert a pressure of 1.75 atm to kPa and to mm Hg. *Answer* 177 kPa, 1330 mm Hg

2. Convert a pressure of 570. torr to atmospheres and to kPa. *Answer* 0.750 atm, 76.0 kPa

SECTION REVIEW

1. Define *pressure.*
2. What units are used to express pressure measurements?
3. What are standard conditions for gas measurements?
4. Convert the following pressures to pressures in standard atmospheres:
   a. 151.98 kPa
   b. 456 torr
   c. 912 mm Hg

The Gas Laws

SECTION 10-3

Objectives

- Use the kinetic-molecular theory to explain the relationships between gas volume, temperature, and pressure.
- Use Boyle's law to calculate volume-pressure changes at constant temperature.
- Use Charles's law to calculate volume-temperature changes at constant pressure.
- Use Gay-Lussac's law to calculate pressure-temperature changes at constant volume.
- Use the combined gas law to calculate volume-temperature-pressure changes.
- Use Dalton's law of partial pressures to calculate partial pressures and total pressures.

Scientists have been studying physical properties of gases for hundreds of years. In 1662, Robert Boyle discovered that gas pressure and volume are related mathematically. The observations of Boyle and others led to the development of the gas laws. *The* **gas laws** *are simple mathematical relationships between the volume, temperature, pressure, and amount of a gas.*

Boyle's Law: Pressure-Volume Relationship

Robert Boyle discovered that doubling the pressure on a sample of gas at constant temperature reduces its volume by one-half. Tripling the gas pressure reduces its volume to one-third of the original. Reducing the pressure on a gas by one-half allows the volume of the gas to double. As one variable increases, the other decreases. Figure 10-9 shows that as the volume of gas in the syringe decreases, the pressure of the gas increases.

You can use the kinetic-molecular theory to understand why this pressure-volume relationship holds. The pressure of a gas is caused by

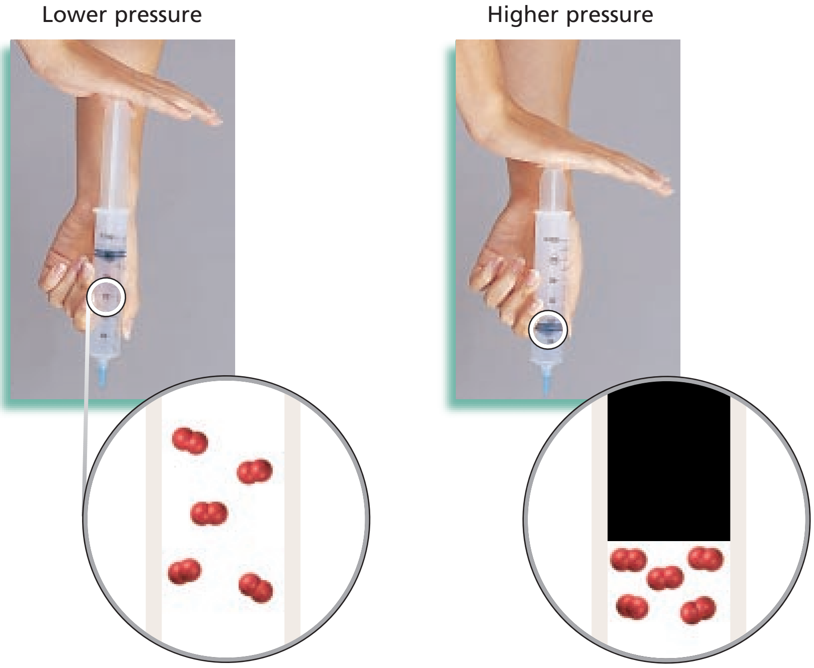

FIGURE 10-9 The volume of gas in the syringe shown in the photo is reduced when the plunger is pushed down. The gas pressure increases as the volume is reduced because the molecules collide more frequently with the walls of the container in a smaller volume.

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Volume (mL)	Pressure (atm)	$P \times V$
1200	0.5	600
600	1.0	600
300	2.0	600
200	3.0	600
150	4.0	600
120	5.0	600
100	6.0	600

TABLE 10-2 *Volume-Pressure Data for a Gas Sample (at Constant Mass and Temperature)*

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moving molecules hitting the container walls. Suppose the volume of a container is decreased but the same number of gas molecules is present at the same temperature. There will be more molecules per unit volume. The number of collisions with a given unit of wall area will increase as a result. Therefore, pressure will also increase.

Table 10-2 shows pressure and volume data for a constant mass of gas at constant temperature. Plotting the values of volume versus pressure gives a curve like that in Figure 10-10. The general volume-pressure relationship that is illustrated is called Boyle's law. **Boyle's law** *states that the volume of a fixed mass of gas varies inversely with the pressure at constant temperature.*

Mathematically, Boyle's law is expressed as follows.

$$V = k\frac{1}{P} \quad \text{or} \quad PV = k$$

The value of k is constant for a given sample of gas and depends only on the mass of gas and the temperature. (Note that for the data in Table 10-2, $k = 600$ mL•atm.) If the pressure of a given gas sample at constant temperature changes, the volume will change. However, the quantity *pressure times volume* will remain equal to the same value of k.

Boyle's law can be used to compare changing conditions for a gas. Using P_1 and V_1 to stand for initial conditions and P_2 and V_2 to stand for new conditions results in the following equations.

$$P_1V_1 = k \qquad P_2V_2 = k$$

Two quantities that are equal to the same thing are equal to each other.

$$P_1V_1 = P_2V_2$$

Given three of the four values P_1, V_1, P_2, and V_2, you can use this equation to calculate the fourth value for a system at constant temperature. For example, suppose that 1.0 L of gas is initially at 1.0 atm pressure (V_1 = 1.0 L, P_1 = 1.0 atm). The gas is allowed to expand fivefold at constant temperature to 5.0 L (V_2 = 5.0 L). You can then calculate the new pressure, P_2, by rearranging the equation as follows.

Volume vs. Pressure for a Gas at Constant Temperature

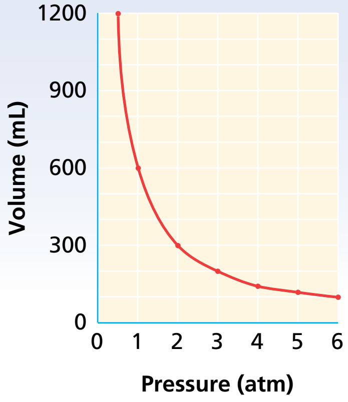


FIGURE 10-10 This graph shows that there is an inverse relationship between volume and pressure. As the pressure drops by half, the volume doubles.

$$P_2 = \frac{P_1V_1}{V_2}$$

$$P_2 = \frac{(1.0 \text{ atm})(1.0 \text{ L})}{5.0 \text{ L}} = 0.20 \text{ atm}$$

The pressure has decreased to one-fifth the original pressure, while the volume increased fivefold.

SAMPLE PROBLEM 10-2

A sample of oxygen gas has a volume of 150. mL when its pressure is 0.947 atm. What will the volume of the gas be at a pressure of 0.987 atm if the temperature remains constant?

SOLUTION

1 ANALYZE

Given: V_1 of O_2 = 150. mL
P_1 of O_2 = 0.947 atm; P_2 of O_2 = 0.987 atm
Unknown: V_2 of O_2 in mL

2 PLAN

$$P_1, V_1, P_2 \longrightarrow V_2$$

Rearrange the equation for Boyle's law ($P_1V_1 = P_2V_2$) to obtain V_2.

$$V_2 = \frac{P_1V_1}{P_2}$$

3 COMPUTE

Substitute values for P_1, V_1, and P_2 to obtain the new volume, V_2.

$$V_2 = \frac{P_1V_1}{P_2} = \frac{(0.947 \cancel{\text{atm}})(150. \text{ mL } O_2)}{0.987 \cancel{\text{atm}}} = 144 \text{ mL } O_2$$

4 EVALUATE

When the pressure is increased slightly at constant temperature, the volume decreases slightly, as expected. Units cancel to give milliliters, a volume unit.

PRACTICE

1. A balloon filled with helium gas has a volume of 500 mL at a pressure of 1 atm. The balloon is released and reaches an altitude of 6.5 km, where the pressure is 0.5 atm. Assuming that the temperature has remained the same, what volume does the gas occupy at this height?
Answer 1000 mL He

2. A gas has a pressure of 1.26 atm and occupies a volume of 7.40 L. If the gas is compressed to a volume of 2.93 L, what will its pressure be, assuming constant temperature?
Answer 3.18 atm

3. Divers know that the pressure exerted by the water increases about 100 kPa with every 10.2 m of depth. This means that at 10.2 m below the surface, the pressure is 201 kPa; at 20.4 m, the pressure is 301 kPa; and so forth. Given that the volume of a balloon is 3.5 L at STP and that the temperature of the water remains the same, what is the volume 51 m below the water's surface?
Answer 0.59 L

CHEMISTRY INTERACTIVE TUTOR

Module 6: Gas Laws

Charles's Law: Volume-Temperature Relationship

Balloonists, such as those in the photo at the beginning of this chapter, are making use of a physical property of gases: if pressure is constant, gases expand when heated. When the temperature increases, the volume of a fixed number of gas molecules must increase if the pressure is to stay constant. At the higher temperature, the gas molecules move faster. They collide with the walls of the container more frequently and with more force. The increased pressure causes the volume of a flexible container to increase; then the molecules must travel farther before reaching the walls. The rate of collisions against each unit of wall area decreases. This lower collision frequency offsets the greater collision force at the higher temperature. The pressure thus stays constant.

The quantitative relationship between volume and temperature was discovered by the French scientist Jacques Charles in 1787. Charles's experiments showed that all gases expand to the same extent when heated through the same temperature interval. Charles found that the volume changes by 1/273 of the original volume for each Celsius degree, at constant pressure and an initial temperature of 0°C. For example, raising the temperature to 1°C causes the gas volume to increase by 1/273 of the volume it had at 0°C. A 10°C temperature increase causes the volume to expand by 10/273 of the original volume at 0°C. If the temperature is increased by 273°C, the volume increases by 273/273 of the original, that is, the volume doubles.

The same regularity of volume change occurs if a gas is cooled at constant pressure, as the balloons in Figure 10-11 show. At 0°C, a 1°C decrease in temperature decreases the original volume by 1/273. At this rate of volume decrease, a gas cooled from 0°C to −273°C would be decreased by 273/273. In other words, it would have zero volume, which

FIGURE 10-11 As air-filled balloons are exposed to liquid nitrogen, they shrink greatly in volume. When they are removed from the liquid nitrogen and the air inside them is warmed to room temperature, the balloons expand to their original volume.

TABLE 10-3 Volume-Temperature Data for a Gas Sample (at Constant Mass and Pressure)

Temperature (°C)	Volume (mL)
273	1092
100	746
10	566
1	548
0	546
–1	544
–73	400
–173	200
–223	100

is not actually possible. In fact, real gases cannot be cooled to –273°C. Before they reach that temperature, intermolecular forces exceed the kinetic energy of the molecules, and the gases condense to form liquids or solids.

The data in Table 10-3 illustrate the temperature-volume relationship at constant pressure for a gas sample with a volume of 546 mL at 0°C. When the gas is warmed by 1°C, it expands by 1/273 its original volume. In this case, each 1°C temperature change from 0°C causes a volume change of 2 mL, or 1/273 of 546 mL. Raising the temperature to 100°C from 0°C increases the volume by 200 mL, or 100/273 of 546 mL.

Note that in Table 10-3, the volume does not increase in direct proportion to the Celsius temperature. For example, notice what happens when the temperature is increased tenfold from 10°C to 100°C. The volume does not increase tenfold but increases only from 566 mL to 746 mL.

The Kelvin temperature scale is a scale that starts at a temperature corresponding to –273.15°C. That temperature is the lowest one possible. *The temperature –273.15°C is referred to as* **absolute zero** *and is given a value of zero in the Kelvin scale.* This fact gives the following relationship between the two temperature scales.

$$K = 273.15 + {}^{\circ}C$$

For calculations in this book, 273.15 is rounded off to 273.

The average kinetic energy of gas molecules is more closely related to the Kelvin temperature. Gas volume and Kelvin temperature are directly proportional to each other. For example, quadrupling the Kelvin temperature causes the volume of a gas to quadruple, and reducing the Kelvin temperature by half causes the volume of a gas to decrease by half.

The relationship between Kelvin temperature and gas volume is known as Charles's law. **Charles's law** *states that the volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature.*

TABLE 10-4 Volume-Temperature Data for a Gas Sample (at Constant Mass and Pressure)

Volume (mL)	Kelvin temperature (K)	V/T or k (mL/K)
1092	546	2
746	373	2
566	283	2
548	274	2
546	273	2
544	272	2
400	200	2
100	50	2

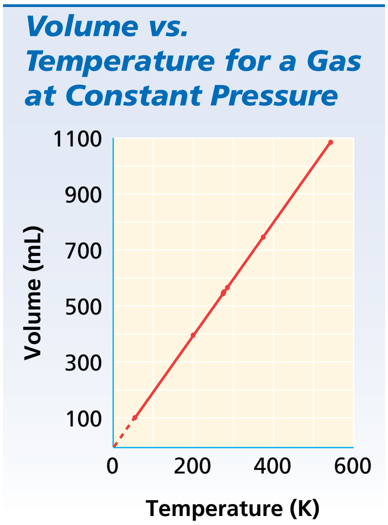


FIGURE 10-12 This graph shows the plot of the volume versus the Kelvin temperature data from Table 10-4. It gives a straight line that, when extended, indicates the volume will become 0 at –273°C. Such a plot is characteristic of directly proportional variables.

Figure 10-12 illustrates the relationship between gas volume and Kelvin temperature by plotting the data from Table 10-4. Charles's law may be expressed as follows.

$$V = kT \quad \text{or} \quad \frac{V}{T} = k$$

The value of T is the Kelvin temperature, and k is a constant. The value of k depends only on the quantity of gas and the pressure. The ratio V/T for any set of volume-temperature values always equals the same k. The form of Charles's law that can be applied directly to most volume-temperature problems involving gases is as follows.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

V_1 and T_1 represent initial conditions. V_2 and T_2 represent a new set of conditions. When three of the four values V_1, T_1, V_2, and T_2 are known, this equation can be used to calculate the fourth value.

SAMPLE PROBLEM 10-3

A sample of neon gas occupies a volume of 752 mL at 25°C. What volume will the gas occupy at 50°C if the pressure remains constant?

SOLUTION

1 *ANALYZE*

Given: V_1 of Ne = 752 mL

T_1 of Ne = 25°C + 273 = 298 K; T_2 of Ne = 50°C + 273 = 323 K

Note that Celsius temperatures have been converted to kelvins. This is a *very important* step for working the problems in this chapter.

Unknown: V_2 of Ne in mL

2 PLAN Because the gas remains at constant pressure, an increase in temperature will cause an increase in volume. To obtain V_2, rearrange the equation for Charles's law.

$$V_2 = \frac{V_1 T_2}{T_1}$$

3 COMPUTE Substitute values for V_1, T_1, and T_2 to obtain the new volume, V_2.

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{(752 \text{ mL Ne})(323 \cancel{\text{K}})}{298 \cancel{\text{K}}} = 815 \text{ mL Ne}$$

4 EVALUATE As expected, the volume of the gas increases as the temperature increases. Units cancel to yield milliliters, as desired. The answer contains the appropriate number of significant figures. It is also reasonably close to an estimated value of 812, calculated as $(750 \times 325)/300$.

PRACTICE

1. A helium-filled balloon has a volume of 2.75 L at 20.°C. The volume of the balloon decreases to 2.46 L after it is placed outside on a cold day. What is the outside temperature in K? in °C?

   Answer 262 K, or −11°C

2. A gas at 65°C occupies 4.22 L. At what Celsius temperature will the volume be 3.87 L, assuming the same pressure?

   Answer 37°C

Gay-Lussac's Law: Pressure-Temperature Relationship

You have just learned about the quantitative relationship between volume and temperature at constant pressure. What would you predict about the relationship between pressure and temperature at constant volume? You have seen that pressure is the result of collisions of molecules with container walls. The energy and frequency of collisions depend on the average kinetic energy of molecules, which depends on temperature. For a fixed quantity of gas at constant volume, the pressure should be directly proportional to the Kelvin temperature, which depends directly on average kinetic energy.

That prediction turns out to be correct. For every kelvin of temperature change, the pressure of a confined gas changes by 1/273 of the pressure at 0°C. Joseph Gay-Lussac is given credit for recognizing this in 1802. The data plotted in Figure 10-13 illustrate **Gay-Lussac's law:** *The pressure of a fixed mass of gas at constant volume varies directly with the Kelvin temperature.* Mathematically, Gay-Lussac's law is expressed as follows.

$$P = kT \quad \text{or} \quad \frac{P}{T} = k$$

Pressure vs. Temperature for a Gas at Constant Volume

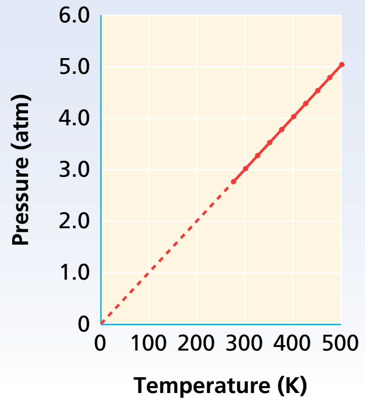


FIGURE 10-13 This graph shows that gas pressure varies directly with Kelvin temperature at constant volume.

The value of T is the temperature in kelvins, and k is a constant that depends on the quantity of gas and the volume. For a given mass of gas at constant volume, the ratio P/T is the same for any set of pressure-temperature values. Unknown values can be found using this form of Gay-Lussac's law.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

When values are known for three of the four quantities, the fourth value can be calculated.

SAMPLE PROBLEM 10-4

The gas in an aerosol can is at a pressure of 3.00 atm at 25°C. Directions on the can warn the user not to keep the can in a place where the temperature exceeds 52°C. What would the gas pressure in the can be at 52°C?

SOLUTION

1 ANALYZE

Given: P_1 of gas = 3.00 atm
T_1 of gas = 25°C + 273 = 298 K; T_2 of gas = 52°C + 273 = 325 K
Unknown: P_2 of gas in atm

2 PLAN

Because the gaseous contents remain at the constant volume of the can, an increase in temperature will cause an increase in pressure. Rearrange Gay-Lussac's law to obtain P_2.

$$P_2 = \frac{P_1 T_2}{T_1}$$

3 COMPUTE

Substitute values for P_1, T_2, and T_1 to obtain the new pressure, P_2.

$$P_2 = \frac{(3.00 \text{ atm})(325 \cancel{\text{K}})}{298 \cancel{\text{K}}} = 3.27 \text{ atm}$$

4 EVALUATE

As expected, a temperature increase at constant volume causes the pressure of the contents in the can to increase. Units cancel correctly. The answer contains the proper number of significant figures. It also is reasonably close to an estimated value of 3.25, calculated as $(3 \times 325)/300$.

PRACTICE

1. Before a trip from New York to Boston, the pressure in an automobile tire is 1.8 atm at 20.°C. At the end of the trip, the pressure gauge reads 1.9 atm. What is the new Celsius temperature of the air inside the tire? (Assume tires with constant volume.) *Answer* 36°C
2. At 120.°C, the pressure of a sample of nitrogen is 1.07 atm. What will the pressure be at 205°C, assuming constant volume? *Answer* 1.30 atm
3. A sample of helium gas has a pressure of 1.20 atm at 22°C. At what Celsius temperature will the helium reach a pressure of 2.00 atm? *Answer* 219°C

The Combined Gas Law

A gas sample often undergoes changes in temperature, pressure, and volume all at the same time. When this happens, three variables must be dealt with at once. Boyle's law, Charles's law, and Gay-Lussac's law can be combined into a single expression that is useful in such situations. *The* **combined gas law** *expresses the relationship between pressure, volume, and temperature of a fixed amount of gas.* The combined gas law can be expressed as follows.

$$\frac{PV}{T} = k$$

In the equation, k is constant and depends on the amount of gas. The combined gas law can also be written as follows.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

The subscripts in the equation above indicate two different sets of conditions, and T represents Kelvin temperature. From this expression, any value can be calculated if the other five are known. Note that each of the individual gas laws can be obtained from the combined gas law when the proper variable is constant. Thus, when temperature is constant, T can be canceled out of both sides of the general equation because it represents the same value ($T_1 = T_2$). This leaves Boyle's law.

$$P_1V_1 = P_2V_2$$

If the pressure is held constant, P will cancel out of both sides of the general equation, since $P_1 = P_2$. Charles's law is obtained.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Keeping the volume constant means V can be canceled out of both sides of the general equation because $V_1 = V_2$. This gives Gay-Lussac's law.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

SAMPLE PROBLEM 10-5

A helium-filled balloon has a volume of 50.0 L at 25°C and 1.08 atm. What volume will it have at 0.855 atm and 10.°C?

SOLUTION

1 ANALYZE

Given: V_1 of He = 50.0 L
T_1 of He = 25°C + 273 = 298 K; T_2 of He = 10°C + 273 = 283 K
P_1 of He = 1.08 atm; P_2 of He = 0.855 atm
Unknown: V_2 of He in L

2 PLAN Because the gas changes in both temperature and pressure, the combined gas law is needed. Rearrange the combined gas law to solve for the final volume, V_2.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \longrightarrow V_2 = \frac{P_1V_1T_2}{P_2T_1}$$

3 COMPUTE Substitute the known values into the equation to obtain a value for V_2.

$$V_2 = \frac{(1.08 \cancel{\text{atm}})(50.0 \text{ L He})(283 \cancel{\text{K}})}{(0.855 \cancel{\text{atm}})(298 \cancel{\text{K}})} = 60.0 \text{ L He}$$

4 EVALUATE Here the pressure decreases much more than the temperature decreases. As expected, the net result of the two changes gives an increase in the volume, from 50.0 L to 60.0 L. Units cancel appropriately. The answer is correctly expressed to three significant figures. It is also reasonably close to an estimated value of 50, calculated as $(50 \times 300)/300$.

PRACTICE

1. The volume of a gas is 27.5 mL at 22.0°C and 0.974 atm. What will the volume be at 15.0°C and 0.993 atm?
   Answer 26.3 mL

2. A 700. mL gas sample at STP is compressed to a volume of 200. mL, and the temperature is increased to 30.0°C. What is the new pressure of the gas in Pa?
   Answer 3.94×10^5 Pa, or 394 kPa

Dalton's Law of Partial Pressures

John Dalton, the English chemist who proposed the atomic theory, also studied gas mixtures. He found that *in the absence of a chemical reaction*, the pressure of a gas mixture is the sum of the individual pressures of each gas alone. Figure 10-14 shows a 1.0 L container filled with oxygen gas at a pressure of 0.12 atm at 0°C. In another 1.0 L container, an equal number of molecules of nitrogen gas exert a pressure of 0.12 atm at 0°C. The gas samples are then combined in a 1.0 L container. (At 0°C, oxygen gas and nitrogen gas are unreactive.) The total pressure of the mixture is found to be 0.24 atm at 0°C. The pressure that each gas exerts in the mixture is independent of that exerted by other gases present. *The pressure of each gas in a mixture is called the* **partial pressure** *of that gas.* **Dalton's law of partial pressures** *states that the total pressure of a mixture of gases is equal to the sum of the partial pressures of the component gases.* The law is true regardless of the number of different gases that are present. Dalton's law may be expressed as follows.

$$P_T = P_1 + P_2 + P_3 + \ldots$$

P_T is the total pressure of the mixture. $P_1, P_2, P_3, \ldots$ are the partial pressures of component gases 1, 2, 3, and so on.

You can understand Dalton's law in terms of the kinetic-molecular theory. The rapidly moving particles of each gas in a mixture have an equal chance to collide with the container walls. Therefore, each gas exerts a pressure independent of that exerted by the other gases present. The total pressure is the result of the total number of collisions per unit of wall area in a given time. (Note that because gas particles move independently, the other gas laws, as well as Dalton's law, can be applied to unreacting gas mixtures.)

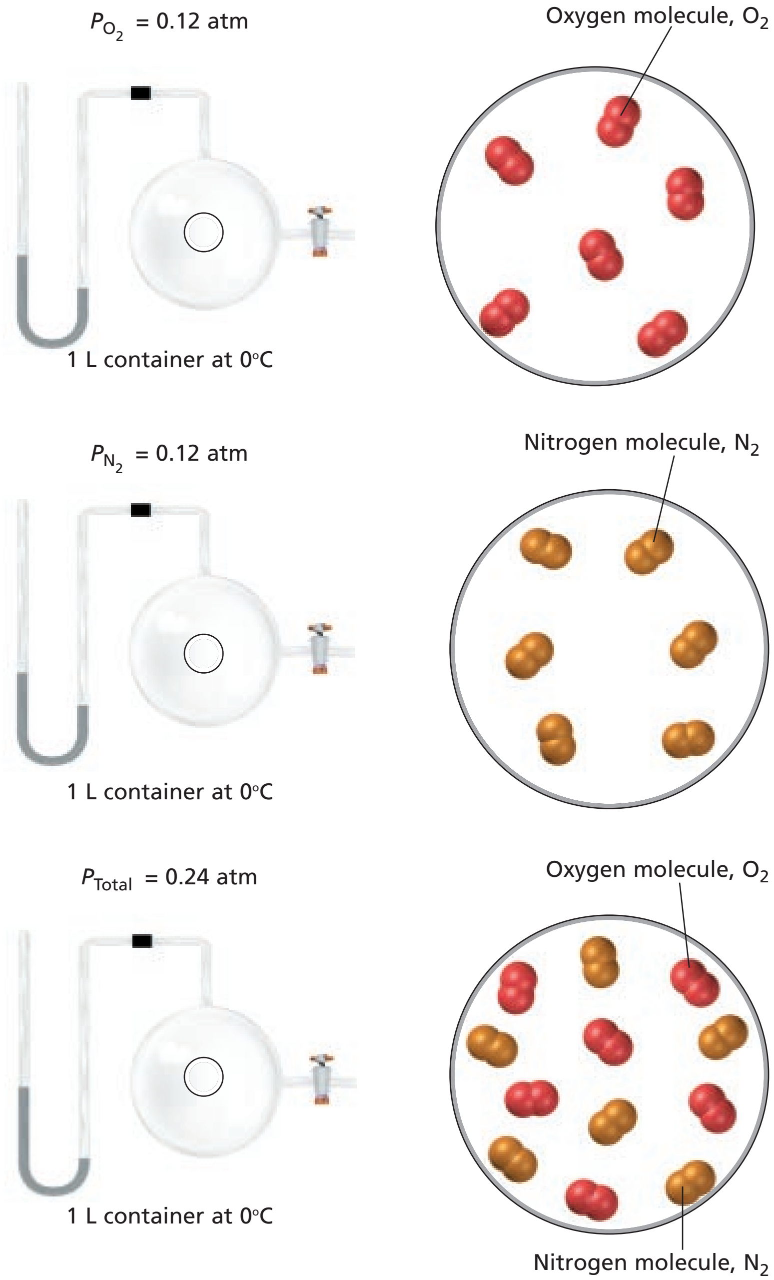


FIGURE 10-14 Samples of oxygen gas and nitrogen gas are mixed. The total pressure of the mixture is the sum of the pressures of the gases.

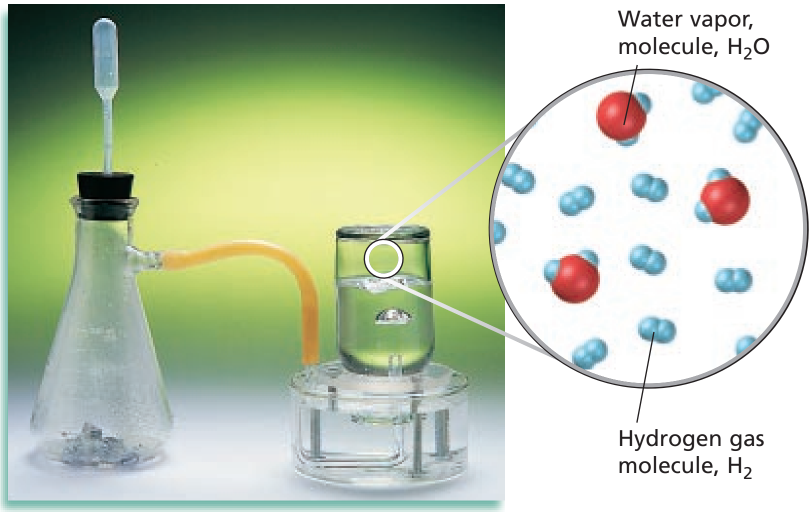


FIGURE 10-15
Hydrogen can be collected by water displacement by reacting zinc with sulfuric acid. The hydrogen gas produced displaces the water in the gas collecting bottle. It now contains some water vapor.

Gases Collected by Water Displacement

Gases produced in the laboratory are often collected over water, as shown in Figure 10-15. The gas produced by the reaction displaces the water, which is more dense, in the collection bottle. You can apply Dalton's law of partial pressures in calculating the pressures of gases collected in this way. A gas collected by water displacement is not pure but is always mixed with water vapor. That is because water molecules at the liquid surface evaporate and mix with the gas molecules. Water vapor, like other gases, exerts a pressure, known as *water-vapor pressure*.

Suppose you wished to determine the total pressure of the gas and water vapor inside a collection bottle. You would raise the bottle until the water levels inside and outside the bottle were the same. At that point, the total pressure inside the bottle would be the same as the atmospheric pressure, P_{atm}. According to Dalton's law of partial pressures, the following is true.

$$P_{atm} = P_{gas} + P_{H_2O}$$

Suppose you then needed to calculate the partial pressure of the dry gas collected. You would read the atmospheric pressure, P_{atm}, from a barometer in the laboratory. To make the calculation, subtract the vapor pressure of the water at the given temperature from the total pressure. The vapor pressure of water varies with temperature. You need to look up the value of P_{H_2O} at the temperature of the experiment in a standard reference table like that in Table A-8 of this book.

SAMPLE PROBLEM 10-6

Oxygen gas from the decomposition of potassium chlorate, $KClO_3$, was collected by water displacement. The barometric pressure and the temperature during the experiment were 731.0 torr and 20.0°C, respectively. What was the partial pressure of the oxygen collected?

SOLUTION

1 ANALYZE

Given: $P_T = P_{atm} = 731.0$ torr

$P_{H_2O} = 17.5$ torr (vapor pressure of water at 20.0°C, from Table A-8)

$P_{atm} = P_{O_2} + P_{H_2O}$

Unknown: P_{O_2} in torr

2 PLAN

The partial pressure of the collected oxygen is found by subtracting the partial pressure of water vapor from the atmospheric pressure, according to Dalton's law of partial pressures.

$$P_{O_2} = P_{atm} - P_{H_2O}$$

3 COMPUTE

Substituting values for P_{atm} and P_{H_2O} gives P_{O_2}.

$$P_{O_2} = 731.0 \text{ torr} - 17.5 \text{ torr} = 713.5 \text{ torr}$$

4 EVALUATE

As expected, the oxygen partial pressure is less than atmospheric pressure. It is also much larger than the partial pressure of water vapor at this temperature. The answer has the appropriate number of significant figures. It is reasonably close to an estimated value of 713, calculated as 730 – 17.

PRACTICE

1. Some hydrogen gas is collected over water at 20.0°C. The levels of water inside and outside the gas-collection bottle are the same. The partial pressure of hydrogen is 742.5 torr. What is the barometric pressure at the time the gas is collected?
   Answer 760.0 torr

2. Helium gas is collected over water at 25°C. What is the partial pressure of the helium, given that the barometric pressure is 750.0 mm Hg?
   Answer 726.2 mm Hg

SECTION REVIEW

1. State Boyle's law, Charles's law, and the combined gas law in mathematical terms.
2. A sample of helium gas has a volume of 200.0 mL at 0.960 atm. What pressure, in atm, is needed to reduce the volume at constant temperature to 50.0 mL?
3. A certain quantity of gas has a volume of 0.750 L at 298 K. At what temperature, in degrees Celsius, would this quantity of gas be reduced to 0.500 L, assuming constant pressure?
4. An aerosol can contains gases under a pressure of 4.50 atm at 20.0°C. If the can is left on a hot, sandy beach, the pressure of the gases increases to 4.80 atm. What is the Celsius temperature on the beach?
5. Discuss the significance of the absolute-zero temperature.
6. A certain mass of oxygen was collected over water when potassium chlorate was decomposed by heating. The volume of the oxygen sample collected was 720. mL at 25.0°C and a barometric pressure of 755 torr. What would the volume of the oxygen be at STP? (Hint: First calculate the partial pressure of the oxygen, using Appendix Table A-8. Then use the combined gas law.)

CHAPTER 10 REVIEW

CHAPTER SUMMARY

10-1

- The kinetic-molecular theory of matter can be used to explain the properties of gases, liquids, and solids.
- The kinetic-molecular theory of gases describes a model of an ideal gas. The behavior of most gases is close to ideal except at very high pressures and low temperatures.
- Gases consist of large numbers of tiny, fast-moving particles that are far apart relative to their size. The average kinetic energy of the particles depends on the temperature of the gas.
- Gases exhibit expansion, fluidity, low density, compressibility, diffusion, and effusion.

Vocabulary

diffusion (305)
effusion (306)
elastic collision (303)
fluids (305)
ideal gas (303)
kinetic-molecular theory (303)
real gas (306)

10-2

- Conditions of standard temperature and pressure (STP) allow comparison of volumes of different gases.
- Pressure, volume, temperature, and number of molecules are the four measureable quantities needed to fully describe a gas.
- The gas molecules that make up the atmosphere exert pressure against Earth's surface, varying with weather conditions and elevation.
- A barometer measures the pressure of the atmosphere. The pressure of a gas in a closed container can be measured by a manometer.

Vocabulary

atmosphere of pressure (311)
barometer (310)
millimeters of mercury (311)
newton (309)
pascal (311)
pressure (308)
standard temperature and pressure (312)
torr (311)

10-3

- Boyle's law shows the inverse relationship between the volume and the pressure of a gas.

$$PV = k$$

- Charles's law illustrates the direct relationship between the volume of a gas and its temperature in kelvins.

$$V = kT$$

- Gay-Lussac's law represents the direct relationship between the pressure of a gas and its temperature in kelvins.

$$P = kT$$

- The combined gas law, as its name implies, combines the previous relationships into the following mathematical expression.

$$\frac{PV}{T} = k$$

- A gas exerts pressure on the walls of its container. In a mixture of unreacting gases, the total pressure equals the sum of the partial pressures of each gas.

Vocabulary

absolute zero (317)
Boyle's law (314)
Charles's law (317)
combined gas law (321)
Dalton's law of partial pressures (322)
gas laws (313)
Gay-Lussac's law (319)
partial pressure (322)

REVIEWING CONCEPTS

1. What idea is the kinetic-molecular theory based on? (10-1)

2. What is an ideal gas? (10-1)

3. State the five basic assumptions of the kinetic-molecular theory. (10-1)

4. How do gases compare with liquids and solids in terms of the distance between their molecules? (10-1)

5. What is an elastic collision? (10-1)

6. a. Write and label the equation that relates the average kinetic energy and speed of gas particles.
b. What is the relationship between the temperature, speed, and kinetic energy of gas molecules? (10-1)

7. a. What is diffusion?
b. What factors affect the rate of diffusion of one gas through another?
c. What is the relationship between the mass of a gas particle and the rate at which it diffuses through another gas?
d. What is effusion? (10-1)

8. a. Why does a gas in a closed container exert pressure?
b. What is the relationship between the area a force is applied to and the resulting pressure? (10-2)

9. a. What is atmospheric pressure?
b. Why does the atmosphere exert pressure?
c. What is the value of atmospheric pressure at sea level, in newtons per square centimeter? (10-2)

10. a. Why does a column of mercury in a tube that is inverted in a dish of mercury have a height of about 760 mm at sea level?
b. What height would be maintained by a column of water inverted in a dish of water at sea level?
c. What accounts for the difference in the heights of the mercury and water columns? (10-2)

11. a. Identify three units used to express pressure.
b. Convert one atmosphere to torr.
c. What is a pascal?
d. What is the SI equivalent of one standard atmosphere of pressure? (10-2)

12. a. At constant pressure, how does temperature relate to the volume of a given quantity of gas?
b. How does this explain the danger of throwing an aerosol can into a fire? (10-3)

13. a. What is the Celsius equivalent of absolute zero?
b. What is the significance of this temperature?
c. What is the relationship between Kelvin temperature and the average kinetic energy of gas molecules? (10-3)

14. a. Explain what is meant by the partial pressure of each gas within a mixture of gases.
b. How do the partial pressures of gases in a mixture affect each other? (10-3)

PROBLEMS

Pressure and Temperature Conversions

15. If the atmosphere can support a column of mercury 760 mm high at sea level, what height (in mm) of each of the following could be supported, given the relative density values cited?
a. water, whose density is approximately 1/14 that of mercury
b. a hypothetical liquid with a density 1.40 times that of mercury

16. Convert each of the following into a pressure reading expressed in torr. (Hint: See Sample Problem 10-1.)
a. 1.25 atm
b. 2.48×10^{-3} atm
c. 4.75×10^{4} atm
d. 7.60×10^{6} atm

17. Convert each of the following into the unit specified.
a. 125 mm Hg into atm
b. 3.20 atm into Pa
c. 5.38 kPa into torr

18. Convert each of the following Celsius temperatures to Kelvin temperatures.
a. 0.°C
b. 27°C
c. −50.°C
d. −273°C

19. Convert each of the following Kelvin temperatures to Celsius temperatures.
a. 273 K
b. 350. K
c. 100. K
d. 20. K

Boyle's Law

20. Use Boyle's law to solve for the missing value in each of the following. (Hint: See Sample Problem 10-2.)

a. P_1 = 350. torr, V_1 = 200. mL, P_2 = 700. torr, V_2 = ?

b. P_1 = 0.75 atm, V_2 = 435 mL, P_2 = 0.48 atm, V_1 = ?

c. $V_1 = 2.4 \times 10^5$ L, P_2 = 180 mm Hg, $V_2 = 1.8 \times 10^3$ L, P_1 = ?

21. The pressure exerted on a 240. mL sample of hydrogen gas at constant temperature is increased from 0.428 atm to 0.724 atm. What will the final volume of the sample be?

22. A flask containing 155 cm^3 of hydrogen was collected under a pressure of 22.5 kPa. What pressure would have been required for the volume of the gas to have been 90.0 cm^3, assuming the same temperature?

23. A gas has a volume of 450.0 mL. If the temperature is held constant, what volume would the gas occupy if the pressure were

a. doubled? (Hint: Express P_2 in terms of P_1.)

b. reduced to one-fourth of its original value?

24. A sample of oxygen that occupies 1.00×10^6 mL at 575 mm Hg is subjected to a pressure of 1.25 atm. What will the final volume of the sample be if the temperature is held constant?

Charles's Law

25. Use Charles's law to solve for the missing value in each of the following. (Hint: See Sample Problem 10-3.)

a. V_1 = 80.0 mL, T_1 = 27°C, T_2 = 77°C, V_2 = ?

b. V_1 = 125 L, V_2 = 85.0 L, T_2 = 127°C, T_1 = ?

c. T_1 = –33°C, V_2 = 54.0 mL, T_2 = 160.°C, V_1 = ?

26. A sample of air has a volume of 140.0 mL at 67°C. At what temperature will its volume be 50.0 mL at constant pressure?

27. At standard temperature, a gas has a volume of 275 mL. The temperature is then increased to 130.°C, and the pressure is held constant. What is the new volume?

Gay-Lussac's Law

28. A sample of hydrogen at 47°C exerts a pressure of 0.329 atm. The gas is heated to 77°C at constant volume. What will its new pressure be? (Hint: See Sample Problem 10-4.)

29. To what temperature must a sample of nitrogen at 27°C and 0.625 atm be taken so that its pressure becomes 1.125 atm at constant volume?

30. The pressure on a gas at –73°C is doubled, but its volume is held constant. What will the final temperature be in degrees Celsius?

Combined Gas Law

31. A sample of gas at 47°C and 1.03 atm occupies a volume of 2.20 L. What volume would this gas occupy at 107°C and 0.789 atm? (Hint: See Sample Problem 10-5.)

32. A 350. mL air sample collected at 35°C has a pressure of 550. torr. What pressure will the air exert if it is allowed to expand to 425 mL at 57°C?

33. A gas has a volume of 1.75 L at –23°C and 150. kPa. At what temperature would the gas occupy 1.30 L at 210. kPa?

34. A sample of oxygen at 40.°C occupies 820. mL. If this sample later occupies 1250 mL at 60.°C and 1.40 atm, what was its original pressure?

35. A gas at 7.75×10^4 Pa and 17°C occupies a volume of 850. cm^3. At what temperature, in degrees Celsius, would the gas occupy 720. cm^3 at 8.10×10^4 Pa?

36. A meteorological balloon contains 250. L of He at 22°C and 740. mm Hg. If the volume of the balloon can vary according to external conditions, what volume would it occupy at an altitude at which the temperature is –52°C and the pressure is 0.750 atm?

37. The balloon in the previous problem will burst if its volume reaches 400. L. Given the initial conditions specified in that problem, at what temperature, in degrees Celsius, will the balloon burst if its pressure at that bursting point is 0.475 atm?

38. The normal respiratory rate for a human being is 15.0 breaths per minute. The average volume of air for each breath is 505 cm^3 at 20.°C and 9.95×10^4 Pa. What is the volume of air at STP that an individual breathes in one day? Give your answer in cubic meters.

Dalton's Law of Partial Pressures

39. Three of the primary components of air are carbon dioxide, nitrogen, and oxygen. In a sample containing a mixture of only these gases at exactly one atmosphere pressure, the partial pressures of carbon dioxide and nitrogen are given as P_{CO_2} = 0.285 torr and P_{N_2} = 593.525 torr. What is the partial pressure of oxygen? (Hint: See Sample Problem 10-6.)

40. Determine the partial pressure of oxygen collected by water displacement if the water temperature is 20.0°C and the total pressure of the gases in the collection bottle is 730.0 torr.

41. A sample of gas is collected over water at a temperature of 35.0°C when the barometric pressure reading is 742.0 torr. What is the partial pressure of the dry gas?

42. A sample of oxygen is collected in a 175 mL container over water at 15°C, and the barometer reads 752.0 torr. What volume would the dry gas occupy at 770.0 torr and 15°C?

43. Suppose that 120. mL of argon is collected over water at 25°C and 780.0 torr. Compute the volume of the dry argon at STP.

MIXED REVIEW

44. A mixture of three gases, A, B, and C, is at a total pressure of 6.11 atm. The partial pressure of gas A is 1.68 atm; that of gas B is 3.89 atm. What is the partial pressure of gas C?

45. A child receives a balloon filled with 2.30 L of helium from a vendor at an amusement park. The temperature outside is 311 K. What will the volume of the balloon be when the child brings it home to an air-conditioned house at 295 K? Assume that the pressure stays the same.

46. A sample of argon gas occupies a volume of 295 mL at 36°C. What volume will the gas occupy at 55°C, assuming constant pressure?

47. A sample of carbon dioxide gas occupies 638 mL at 0.893 atm and 12°C. What will the pressure be at a volume of 881 mL and a temperature of 18°C?

48. At 84°C, a gas in a container exerts a pressure of 0.503 atm. Assuming the size of the container has not changed, at what Celsius temperature would the pressure be 1.20 atm?

49. A weather balloon at Earth's surface has a volume of 4.00 L at 304 K and 755 mm Hg. If the balloon is released and the volume reaches 4.08 L at 728 mm Hg, what is the temperature?

50. A gas has a pressure of 4.62 atm when its volume is 2.33 L. What will the pressure be when the volume is changed to 1.03 L, assuming constant temperature? Express the final pressure in torr.

51. At a deep-sea station 200. m below the surface of the Pacific Ocean, workers live in a highly pressurized environment. How many liters of gas at STP must be compressed on the surface to fill the underwater environment with 2.00×10^7 L of gas at 20.0 atm? Assume that temperature remains constant.

CRITICAL THINKING

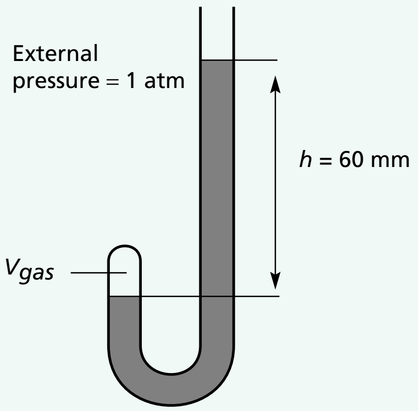
52. Applying Models
a. Why do we say the graph in Figure 10-10 illustrates an inverse relationship?
b. Why does the data plotted in Figure 10-12 show a direct relationship?

53. Relating Ideas Explain how different gases in a mixture can have the same average kinetic energy value, even though the masses of their individual particles differ.

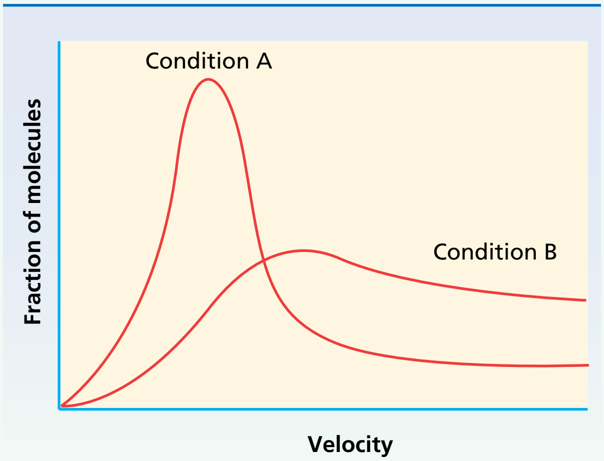
54. Inferring Conclusions If all gases behaved as ideal gases under all conditions of temperature and pressure, there would be no solid or liquid forms of these substances. Explain.

55. Relating Ideas Pressure is defined as force per unit area. Yet Torricelli found that the diameter of the barometer dish and the surface area of contact between the mercury in the tube and in the dish did not affect the height of mercury that was supported. Explain this seemingly inconsistent observation in view of the relationship between pressure and surface areas.

56. Interpreting Graphics Examine Boyle's J-tube apparatus shown here. The tube is open to the atmosphere at the top. The other end is closed and contains a gas with a volume labeled V_{gas}. If h = 60 mm Hg, what is the pressure exerted by the enclosed gas?



57. Velocity distribution curves are shown in the following graph for the same gas under two different conditions, A and B. Compare the behavior of the gas under conditions A and B in relation to each of the following:
a. temperature
b. average kinetic energy
c. average molecular velocity
d. gas volume
e. gas pressure



TECHNOLOGY & LEARNING

58. Graphing Calculator Deriving the Boyle's Law Equation

The graphing calculator can run a program that derives the equation for a curve, given data such as volume vs. pressure. Begin by creating a table of data. Then the program will plot the data.

Go to Appendix C. If you are using a TI 83 Plus, you can download the program and data and run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes and data sets to use. Remember that you will need to name the program and check the display, as explained in Appendix C. You will then be ready to run the program. After you have graphed the data, answer these questions.

Note: Answers are written with two significant figures.

a. What is the pressure if the measured volume is 500 mL?
b. What is the volume if the measured pressure is 22 atm?
c. What is the volume if the measured pressure is 7.5 atm?

HANDBOOK SEARCH

59. Review the melting point data in the properties tables for each group of the *Elements Handbook*. What elements on the periodic table exist as gases at room temperature?

60. Review the listing found in the *Elements Handbook* of the top 10 chemicals produced in the United States. Which of the top 10 chemicals are gases?

61. Though mercury is an ideal liquid to use in barometers, it is highly toxic. Review the transition metals section of the *Elements Handbook*, and answer the following questions:
a. What is the density of mercury? How does the density of Hg compare with that of other transition metals?

b. In what ways is mercury toxic?
c. What mercury compound is becoming more common as a pollutant in lakes and streams? Why is this pollutant so dangerous to freshwater ecosystems?
d. What are the symptoms of mercury poisoning?

RESEARCH & WRITING

62. Prepare a report on the development of the modern submarine. Include a discussion of the technology that enables the submarine to withstand the tremendous pressures at great ocean depths. Also report on the equipment used to ensure a sufficient supply of oxygen for submarine crew members.

63. Design and conduct a meteorological study to examine the interrelationships among barometric pressure, temperature, humidity, and other weather variables. Prepare a report explaining your results.

64. Conduct library research on attempts made to approach absolute zero and on the interesting properties that materials exhibit near that temperature. Write a report on your findings.

ALTERNATIVE ASSESSMENT

65. The air pressure of car tires should be checked regularly for safety reasons and to prevent uneven tire wear. Find out the units of measurement on a typical tire gauge, and determine how gauge pressure relates to atmospheric pressure.