

So many of the properties of matter, especially when in the gaseous form, can be deduced from the hypothesis that their minute parts are in rapid motion, the velocity increasing with the temperature, that the precise nature of this motion becomes a subject of rational curiosity.

—JAMES CLERK MAXWELL (1831–1879)

CHAPTER

# 6

## Gases

Top speed 1,322 km/h

**H**uman beings can survive for weeks without food, days without water, but only minutes without air. Fortunately, we live at the bottom of a vast ocean of air, held to Earth by gravity. We inhale a lungful of this air every few seconds, keep some of the molecules for our own uses, add some molecules that our bodies no longer need, and exhale the mixture back into the surrounding air. The air around us is matter in the gaseous state. What are the fundamental properties of these gases? What laws describe their behavior? What theory explains these properties and laws? Recall that the scientific approach (see Section 1.2) proceeds in this way—from observations to laws to theories—exactly the way we will proceed in this chapter. The gaseous state is the simplest and best-understood state of matter. In this chapter, we examine that state.

# MINUTES 27 SECONDS



## JUMP 41,400 METERS

A pressurized suit protected Alan Eustace from the vacuum of space during his record-breaking skydive. Any significant damage to the suit carried the risk of uncontrolled decompression, which would likely result in Eustace's death.

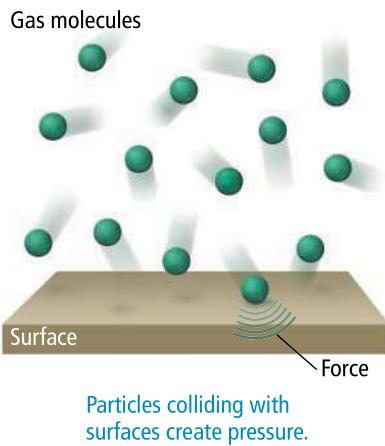
- 6.1** Supersonic Skydiving and the Risk of Decompression 211
- 6.2** Pressure: The Result of Molecular Collisions 212
- 6.3** The Simple Gas Laws: Boyle's Law, Charles's Law, and Avogadro's Law 215
- 6.4** The Ideal Gas Law 222
- 6.5** Applications of the Ideal Gas Law: Molar Volume, Density, and Molar Mass of a Gas 225

- 6.6** Mixtures of Gases and Partial Pressures 228
- 6.7** Gases in Chemical Reactions: Stoichiometry Revisited 235
- 6.8** Kinetic Molecular Theory: A Model for Gases 238
- 6.9** Mean Free Path, Diffusion, and Effusion of Gases 245
- 6.10** Real Gases: The Effects of Size and Intermolecular Forces 246

### LEARNING OUTCOMES 252

#### **6.1** Supersonic Skydiving and the Risk of Decompression

On October 24, 2014, Google executive Alan Eustace released himself into the dark void of space 25.7 miles (41.4 km) above Earth's surface. Eustace's 15-minute journey back to Earth's surface broke the sound barrier and broke the previous skydiving record of 24 miles (38.6 km) set just two years before by Felix Baumgartner.



### ▲ FIGURE 6.1 Gas Pressure

For a gas, pressure is the force per unit area exerted by gas molecules colliding with the surfaces around them.

While Eustace was in space, he was protected from the surrounding vacuum by a pressurized suit. The suit contained air at a pressure similar to that found on the surface of Earth. **Pressure** is the amount of force per unit area, so the pressure exerted by a gas equals the force exerted by its particles divided by the surface area they strike (Figure 6.1◀). Just as a ball exerts a force when it bounces against a wall, a gaseous atom or molecule exerts a force when it collides with a surface. The average of these countless collisions over time results in pressure. The total pressure exerted by a gas depends on several factors, including the concentration of gas particles in the sample; the lower the concentration, the lower the pressure. At 25.7 miles (41.4 km) above Earth's surface, the concentration of gas particles is much lower than it is at sea level; consequently, the pressure at that high altitude is very low. Without the pressurized suit, Eustace could not have survived the space-like conditions.

One of the primary risks that Eustace faced while in space was uncontrolled decompression. Any significant damage to his suit could have caused the air within the suit to escape, resulting in a large pressure drop. While the effects of a large pressure drop are sometimes exaggerated—one urban myth claims that a person can explode—they are nonetheless lethal. A sudden pressure drop would have produced a large pressure difference between the air in Eustace's lungs and the surrounding vacuum. The pressure difference would have caused his lungs to expand too much, resulting in severe lung damage. Fortunately, Eustace's suit worked just as it was designed to do, and he plunged safely back to Earth.

### Pressure and Density



### ▲ FIGURE 6.2 Pressure and Particle Density

A low density of gas particles results in low pressure; a high density of gas particles results in high pressure.

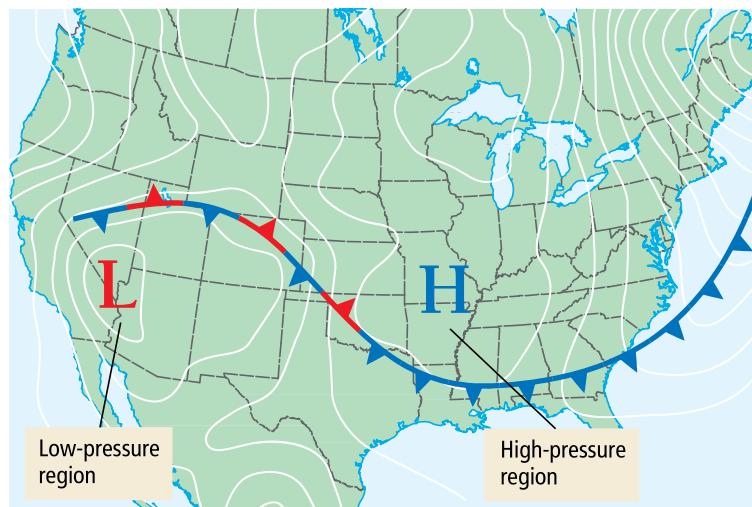
## 6.2 Pressure: The Result of Molecular Collisions

Air can hold up a jumbo jet or knock down a building. How? As we discussed in Section 6.1, air contains gaseous atoms and molecules in constant motion. The particles collide with each other and with the surfaces around them creating pressure. Because of pressure, we can drink from straws, inflate basketballs, and breathe. Variation in pressure in Earth's atmosphere creates wind, and changes in pressure help us to predict weather. Pressure is all around us and inside of us.

The pressure that a gas sample exerts is the *force* that results from the collisions of gas particles divided by the *area* of the surface with which they collide:

$$\text{Pressure} = \frac{\text{force}}{\text{area}} = \frac{F}{A} \quad [6.1]$$

The pressure exerted by a gas sample, therefore, depends on the number of gas particles in a given volume—the fewer the gas particles, the lower the force per unit area and the lower the pressure (Figure 6.2◀). Since the number of gas particles in a given volume



► Pressure variations in Earth's atmosphere create wind and weather. The H on this map indicates a region of high pressure, usually associated with clear weather. The L indicates a region of low pressure, usually associated with unstable weather.

generally decreases with increasing altitude, *pressure decreases with increasing altitude*. Above 30,000 ft, for example, where most commercial airplanes fly, the pressure is so low that a person could pass out due to a lack of oxygen. For this reason, most airplane cabins are artificially pressurized.

You may sometimes feel the effect of a drop in pressure as a brief pain in your ears. This pain arises within the air-containing cavities in your ear (Figure 6.3►). When you ascend to a higher altitude, the external pressure (the pressure that surrounds you) drops, while the pressure within your ear cavities (the internal pressure) remains the same. This creates an imbalance—the greater internal pressure forces your eardrum to bulge outward, causing pain. With time, and with the help of a yawn or two, the excess air within your ear's cavities escapes, equalizing the internal and external pressure and relieving the pain.

## Pressure Units

We measure pressure in several different units. A common unit of pressure, the **millimeter of mercury (mmHg)**, originates from how pressure is measured with a **barometer** (Figure 6.4▼). A barometer is an evacuated glass tube, the tip of which is submerged in a pool of mercury (Hg). Atmospheric pressure on the liquid mercury's surface forces the mercury upward into the evacuated tube. Because mercury is so dense (13.5 times more dense than water), atmospheric pressure can support a column of Hg that is only about 0.760 m or 760 mm (about 30 in) tall. By contrast, atmospheric pressure can support a column of water that is about 10.3 m tall. This makes a column of mercury a convenient way to measure pressure.

In a barometer, when the atmospheric pressure rises, the height of the mercury column rises as well. Similarly, when atmospheric pressure falls, the height of the column falls. The unit *millimeter of mercury* is often called a **torr**, after the Italian physicist Evangelista Torricelli (1608–1647) who invented the barometer:

$$1 \text{ mmHg} = 1 \text{ torr}$$

A second unit of pressure is the **atmosphere (atm)**, the average pressure at sea level. Since one atmosphere of pressure pushes a column of mercury to a height of 760 mm, 1 atm and 760 mmHg are equal:

$$1 \text{ atm} = 760 \text{ mmHg}$$

A fully inflated bike tire has a pressure of about 6 atm, and the pressure at the top of Mount Everest is about 0.31 atm.

The SI unit of pressure is the **pascal (Pa)**, defined as 1 newton (N) per square meter:

$$1 \text{ Pa} = 1 \text{ N/m}^2$$

The pascal is a much smaller unit of pressure than the atmosphere:

$$1 \text{ atm} = 101,325 \text{ Pa}$$

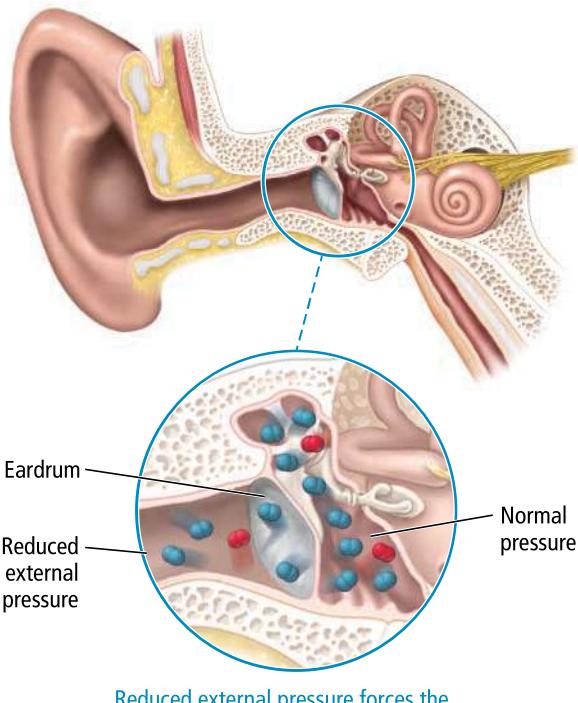
Other common units of pressure include inches of mercury (in Hg) and pounds per square inch (psi):

$$1 \text{ atm} = 29.92 \text{ in Hg}$$

$$1 \text{ atm} = 14.7 \text{ psi}$$

Table 6.1 summarizes common pressure units.

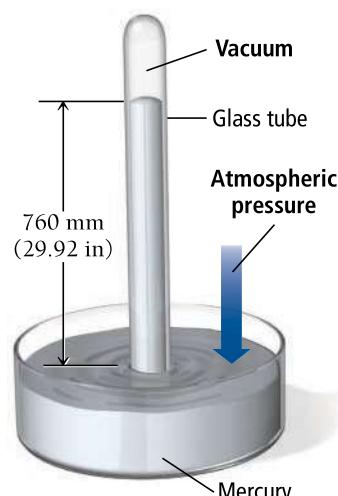
## Pressure Imbalance



Reduced external pressure forces the eardrum to bulge outward, causing pain.

**▲ FIGURE 6.3 Pressure Imbalance** The discomfort you may feel in your ears upon ascending a mountain is caused by a pressure imbalance between the cavities in your ears and the outside air.

## The Mercury Barometer



**▲ FIGURE 6.4 The Mercury Barometer** Average atmospheric pressure at sea level can support a column of mercury 760 mm in height.

**TABLE 6.1** Common Units of Pressure

| Unit                         | Abbreviation | Average Air Pressure at Sea Level |
|------------------------------|--------------|-----------------------------------|
| Pascal (1 N/m <sup>2</sup> ) | Pa           | 101,325 Pa                        |
| Pounds per square inch       | psi          | 14.7 psi                          |
| Torr (1 mmHg)                | torr         | 760 torr (exact)                  |
| Inches of mercury            | in Hg        | 29.92 in Hg                       |
| Atmosphere                   | atm          | 1 atm                             |

**EXAMPLE 6.1** Converting between Pressure Units

A high-performance road bicycle tire is inflated to a total pressure of 132 psi. What is this pressure in mmHg?

**SORT** The problem gives a pressure in psi and asks you to convert the units to mmHg.

**STRATEGIZE** Since Table 6.1 does not have a direct conversion factor between psi and mmHg but does provide relationships between both of these units and atmospheres, convert to atm as an intermediate step.

**SOLVE** Follow the conceptual plan to solve the problem. Begin with 132 psi and use the conversion factors to arrive at the pressure in mmHg.

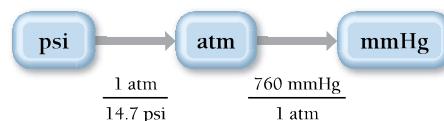
**CHECK** The units of the answer are correct. The magnitude of the answer ( $6.82 \times 10^3$  mmHg) is greater than the given pressure in psi. This is reasonable since mmHg is a much smaller unit than psi.

**FOR PRACTICE 6.1** Your local weather report announces that the barometric pressure is 30.44 in Hg. Convert this pressure to psi.

**FOR MORE PRACTICE 6.1** Convert a pressure of 23.8 in Hg to kPa.

**GIVEN:** 132 psi

**FIND:** mmHg

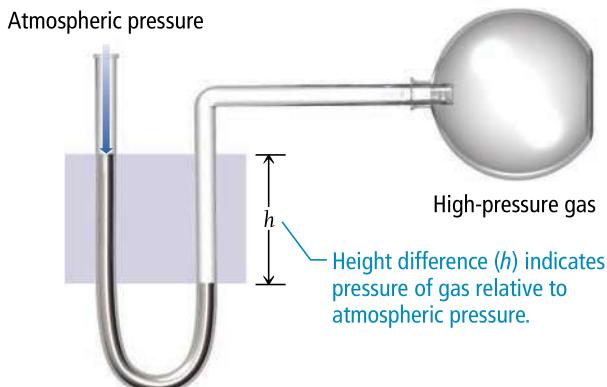
**CONCEPTUAL PLAN****RELATIONSHIPS USED**

$$1 \text{ atm} = 14.7 \text{ psi}$$

$$760 \text{ mmHg} = 1 \text{ atm} \text{ (both from Table 6.1)}$$

**SOLUTION**

$$132 \text{ psi} \times \frac{1 \text{ atm}}{14.7 \text{ psi}} \times \frac{760 \text{ mmHg}}{1 \text{ atm}} = 6.82 \times 10^3 \text{ mmHg}$$

**The Manometer****The Manometer: A Way to Measure Pressure in the Laboratory**

We can measure the pressure of a gas sample in the laboratory with a **manometer**. A manometer is a U-shaped tube containing a dense liquid, usually mercury, as shown in Figure 6.5. In this manometer, one end of the tube is open to atmospheric pressure, and the other is attached to a flask containing the gas sample. If the pressure of the gas sample is exactly equal to atmospheric pressure, then the mercury levels on both sides of the tube are the same. If the pressure of the gas sample is *greater than* atmospheric pressure, the mercury level on the left side of the tube is *higher than* the level on the right. If the pressure of the sample is *less than* atmospheric pressure, the mercury level on the left side is *lower than* the level on the right. This type of manometer always measures the pressure of the gas sample relative to atmospheric pressure. The difference in height between the two levels is equal to the difference between

**▲ FIGURE 6.5** The Manometer A manometer measures the pressure exerted by a sample of gas.

the sample's pressure and atmospheric pressure. To accurately calculate the absolute pressure of the sample, we also need a barometer to measure atmospheric pressure (which can vary from day to day).



## CHEMISTRY AND MEDICINE

### Blood Pressure

**B**lood pressure is the force within our arteries that drives the circulation of blood throughout the body. Blood pressure in the body is analogous to water pressure in a plumbing system. Just as water pressure pushes water through the pipes to faucets and fixtures throughout a house, blood pressure pushes blood to muscles and other tissues throughout the body. However, unlike the water pressure in a plumbing system—which is typically nearly constant—blood pressure varies with each heartbeat. When the heart muscle contracts, blood pressure increases; between contractions it decreases. Systolic blood pressure is the peak pressure during a contraction, and diastolic blood pressure is the lowest pressure between contractions. Just as excessively high water pressure in a plumbing system can damage pipes, blood pressure that is too high in a circulatory system can damage the heart and arteries, resulting in increased risk of stroke and heart attack.

Medical professionals usually measure blood pressure with an instrument called a sphygmomanometer—an inflatable cuff equipped with a pressure gauge—and a stethoscope. The cuff is wrapped around the patient's arm and inflated with air. As air is pumped into the cuff, the pressure in the cuff increases. The cuff tightens around the arm and compresses the artery, momentarily stopping blood flow. The person measuring the blood pressure listens to the artery through the stethoscope while slowly releasing the air pressure in the cuff. When the air pressure in the cuff equals the systolic blood pressure (the peak pressure), a pulse is heard through the stethoscope. The pulse is the sound of blood getting through the compressed artery during a contraction of the heart. The pressure reading at that exact moment is the systolic blood pressure. As the pressure in the cuff continues to decrease, the blood can flow through the compressed artery even between contractions, so the pulsing sound stops. The pressure reading when the pulsing sound stops is the diastolic blood pressure (the lowest pressure).

A blood pressure measurement is usually reported as two pressures, in mmHg, separated by a slash. For example, a blood pressure measurement of 122/84 indicates that the systolic blood pressure is 122 mmHg and the diastolic blood pressure is 84 mmHg. Although the value of blood pressure can vary throughout the day, healthy (or normal) values are usually considered to be below 120 mmHg for systolic and below 80 mmHg for diastolic



▲ A health care provider measures blood pressure with an inflatable cuff that compresses the main artery in the arm. A stethoscope is used to listen for blood flowing through the artery with each heartbeat.

(Table 6.2). High blood pressure, also called hypertension, entails the health risks mentioned previously.

Risk factors for hypertension include obesity, high salt (sodium) intake, high alcohol intake, lack of exercise, stress, a family history of high blood pressure, and age (blood pressure tends to increase as we get older). Mild hypertension can be managed with diet and exercise. Moderate to severe cases require doctor-prescribed medication.

**TABLE 6.2 ■ Blood Pressure Ranges**

| Blood Pressure       | Systolic (mmHg) | Diastolic (mmHg) |
|----------------------|-----------------|------------------|
| Hypotension          | <100            | <60              |
| Normal               | 100–119         | 60–79            |
| Prehypertension      | 120–139         | 80–89            |
| Hypertension Stage 1 | 140–159         | 90–99            |
| Hypertension Stage 2 | >160            | >100             |

### 6.3

## The Simple Gas Laws: Boyle's Law, Charles's Law, and Avogadro's Law

We have learned about pressure and its characteristics in Section 6.2. We now broaden our discussion to include the four basic properties of a gas sample: pressure ( $P$ ), volume ( $V$ ), temperature ( $T$ ), and amount in moles ( $n$ ). These properties are interrelated—when one changes, it affects the others. The *simple gas laws* describe the relationships between pairs of these properties. For example, one simple gas law describes how *volume* varies with *pressure* at constant temperature and amount of gas; another law describes how *volume* varies with *temperature* at constant pressure and

### WATCH NOW!

#### KEY CONCEPT VIDEO 6.3

Simple Gas Laws and Ideal Gas Law

amount of gas. These laws were deduced from observations in which two of the four basic properties were held constant in order to elucidate the relationship between the other two.

## Boyle's Law: Volume and Pressure

In the early 1660s, the pioneering English scientist Robert Boyle (1627–1691) and his assistant Robert Hooke (1635–1703) used a J-tube (Figure 6.6▼) to measure the volume of a sample of gas at different pressures. They trapped a sample of air in the J-tube and added mercury to increase the pressure on the gas. Boyle and Hooke observed an *inverse relationship* between volume and pressure—an increase in one causes a decrease in the other—as shown in Figure 6.7▼. This relationship is now known as **Boyle's law**.

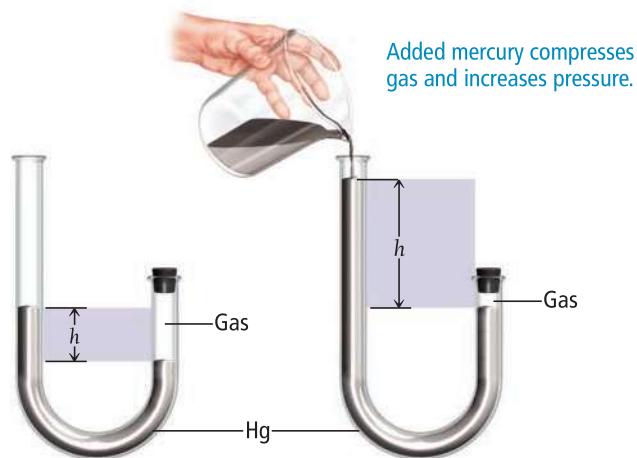
Boyle's law assumes constant temperature and constant amount of gas.

$$\text{Boyle's law: } V \propto \frac{1}{P} \quad (\text{constant } T \text{ and } n)$$

Boyle's law follows from the idea that pressure results from the collisions of the gas particles with the walls of their container. When the volume of a gas sample decreases, the same number of gas particles is crowded into a smaller volume, resulting in more collisions with the walls and therefore an increase in pressure (Figure 6.8►).

Scuba divers learn about Boyle's law during certification because it explains why a diver should not ascend toward the surface without continuous breathing. For every 10 m of depth that a diver descends in water, she experiences an additional 1 atm of pressure due to the weight of the water above her (Figure 6.9►). The pressure regulator used in scuba diving delivers air into the diver's lungs at a pressure that matches the external pressure; otherwise the diver could not inhale the air (see *Chemistry in Your Day: Extra-Long Snorkels* on page 219). For example, when a diver is 20 m below the surface, the regulator delivers air at a pressure of 3 atm to match the 3 atm of pressure around the diver (1 atm due to normal atmospheric pressure and 2 additional atmospheres due to the weight of the water at 20 m). Suppose that a diver inhaled a lungful of air at a pressure of 3 atm and swam quickly to the surface (where the pressure is 1 atm) while holding her breath. What would happen to the volume of air in her lungs?

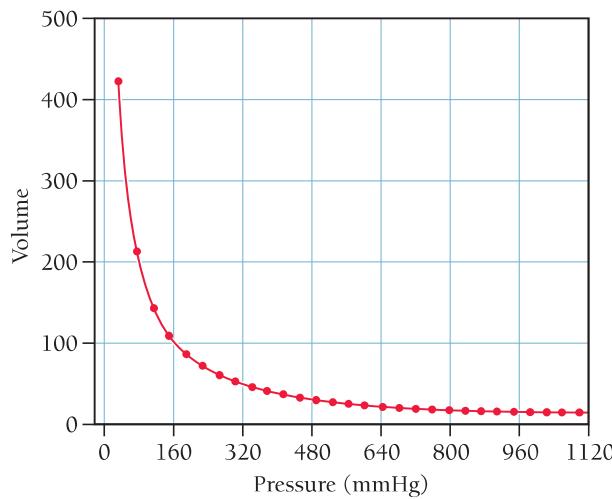
### The J-Tube



▲ **FIGURE 6.6** The J-Tube In a J-tube, a column of mercury traps a sample of gas. The pressure on the gas can be increased by increasing the height ( $h$ ) of mercury in the column.

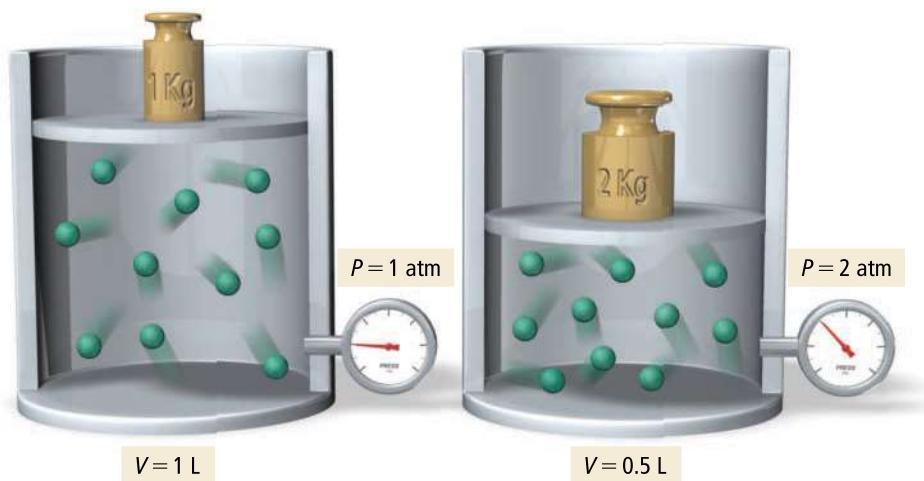
### Boyle's Law

As pressure increases, volume decreases.



▲ **FIGURE 6.7** Volume versus Pressure A plot of the volume of a gas sample versus pressure. The plot shows that volume and pressure are inversely related.

### Volume versus Pressure: A Molecular View



Since the pressure decreases by a factor of 3, the volume of the air in her lungs would increase by a factor of 3—a dangerous situation. The volume increase in the diver's lungs would be so great that she would not be able to hold her breath all the way to the surface—the air would force itself out of her mouth but probably not before the expanded air severely damaged her lungs, possibly killing her. Consequently, the most important rule in diving is *never hold your breath*. To avoid catastrophic results, divers must ascend slowly and breathe continuously, allowing the regulator to bring the air pressure in their lungs back to 1 atm by the time they reach the surface.

We can use Boyle's law to calculate the volume of a gas following a pressure change or the pressure of a gas following a volume change *as long as the temperature and the amount of gas remain constant*. For these types of calculations, we write Boyle's law in a slightly different way.

$$\text{Since } V \propto \frac{1}{P} \quad \text{then} \quad V = \text{constant} \times \frac{1}{P} \quad \text{or} \quad V = \frac{\text{constant}}{P}$$

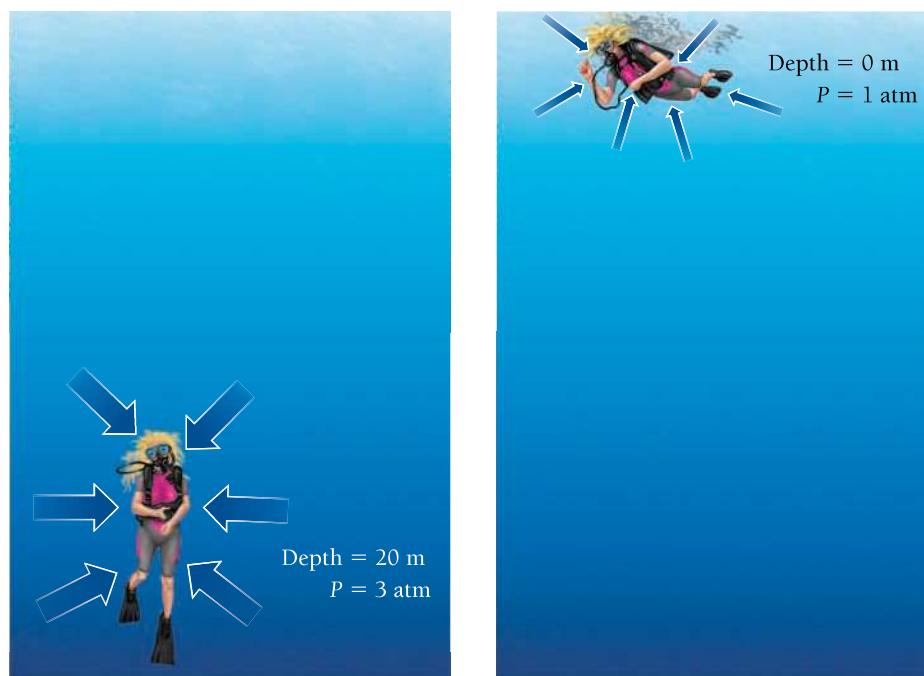
If we multiply both sides by  $P$ , we find that the product of  $P$  and  $V$  is equal to a constant:

$$PV = \text{constant}$$

**◀ FIGURE 6.8 Molecular Interpretation of Boyle's Law**

As the volume of a gas sample decreases, gas molecules collide with surrounding surfaces more frequently, resulting in greater pressure.

If two quantities are proportional, then one is equal to the other multiplied by a constant.



**◀ FIGURE 6.9 Increase in Pressure with Depth**

For every 10 m of depth, a diver experiences approximately 1 additional atmosphere of pressure due to the weight of the surrounding water. At 20 m, for example, the diver experiences approximately 3 atm of pressure (1 atm of normal atmospheric pressure plus an additional 2 atm due to the weight of the water).

This relationship indicates that if the pressure increases, the volume decreases, but the product  $P \times V$  always equals the same constant. For two different sets of conditions, we can say that:

$$P_1V_1 = \text{constant} = P_2V_2$$

or:

$$P_1V_1 = P_2V_2 \quad [6.2]$$

where  $P_1$  and  $V_1$  are the initial pressure and volume of the gas and  $P_2$  and  $V_2$  are the final volume and pressure.

### EXAMPLE 6.2 Boyle's Law

A woman has an initial lung volume of 2.75 L, which is filled with air at an atmospheric pressure of 1.02 atm. If she increases her lung volume to 3.25 L without inhaling any additional air, what is the pressure in her lungs?

To solve the problem, first solve Boyle's law (Equation 6.2) for  $P_2$  and then substitute the given quantities to calculate  $P_2$ .

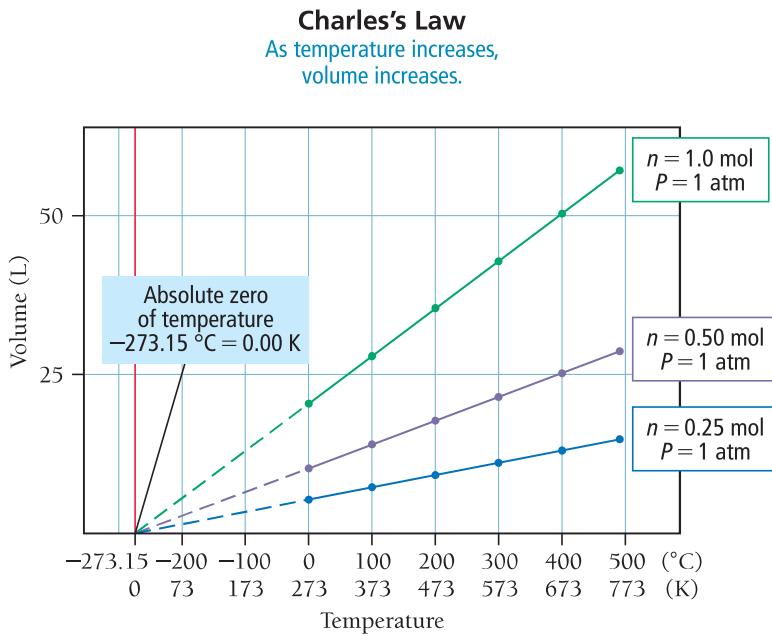
#### SOLUTION

$$\begin{aligned} P_1V_1 &= P_2V_2 \\ P_2 &= \frac{V_1}{V_2}P_1 \\ &= \frac{2.75 \text{ L}}{3.25 \text{ L}} 1.02 \text{ atm} \\ &= 0.863 \text{ atm} \end{aligned}$$

**FOR PRACTICE 6.2** A snorkeler takes a syringe filled with 16 mL of air from the surface, where the pressure is 1.0 atm, to an unknown depth. The volume of the air in the syringe at this depth is 7.5 mL. What is the pressure at the unknown depth? If the pressure increases by 1 atm for every additional 10 m of depth, how deep is the snorkeler?

### Charles's Law: Volume and Temperature

Suppose we keep the pressure of a gas sample constant and measure its volume at a number of different temperatures. Figure 6.10 ▼ shows the results of several such measurements.



► FIGURE 6.10 Volume versus

**Temperature** The volume of a fixed amount of gas at a constant pressure increases linearly with increasing temperature in kelvins. (The extrapolated lines cannot be measured experimentally because all gases condense into liquids before  $-273.15^{\circ}\text{C}$  is reached.)



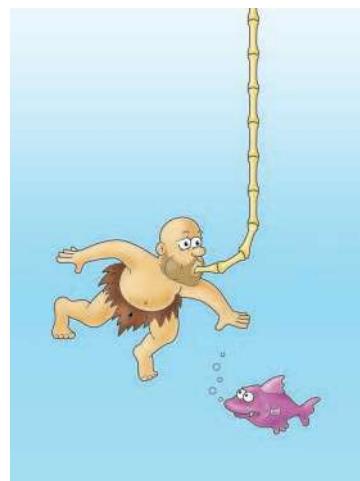
## CHEMISTRY IN YOUR DAY

### Extra-Long Snorkels

Several episodes of *The Flintstones* cartoon featured Fred Flintstone and Barney Rubble snorkeling. Their snorkels, however, were not the modern kind, but long reeds that stretched from the surface of the water down to many meters of depth. Fred and Barney swam around in deep water while breathing air provided to them by these extra-long snorkels. Would this work? Why do people bother with scuba diving equipment if they could instead simply use 10-meter snorkels as Fred and Barney did?

When we breathe, we expand the volume of our chest cavity, reducing the pressure in our lungs to less than 1 atm (Boyle's law). Because of this pressure differential, air from outside our lungs flows into them. Extra-long snorkels do not work because the pressure exerted by water at depth is too great. A diver at 10 m experiences an external pressure of 2 atm. This is more than the muscles of the chest cavity can overcome. Instead, the chest cavity and lungs are compressed, resulting in an air pressure within them of more than 1 atm. If the diver had a snorkel that went to the surface—where the air pressure is 1 atm—air would flow out of the diver's lungs (from greater pressure to less pressure), not into them. It would be impossible for the diver to breathe.

**QUESTION** A diver takes a balloon with a volume of 2.5 L from the surface, where the pressure is 1.0 atm, to a depth of 20 m, where the pressure is 3.0 atm. What happens to the volume of the balloon? What if the end of the submerged balloon is on a long pipe that goes to the surface and is attached to another balloon? Which way would air flow as the diver descended?



▲ In the popular cartoon *The Flintstones*, cavemen used long reeds to breathe surface air while swimming at depth. This would not work because the increased pressure at depth would force air out of their lungs; the pressure would not allow them to inhale.



▲ If two balloons were joined by a long tube and one end was submerged in water, what would happen to the volumes of the two balloons?



From the plot we can see a relationship between volume and temperature: the volume of a gas increases with increasing temperature. Looking at the plot more closely reveals more—volume and temperature are *linearly related*. If two variables are linearly related, plotting one against the other produces a straight line.

Another interesting feature emerges if we extend or *extrapolate* each line in the plot in Figure 6.10 backward from the lowest measured temperature. The dotted extrapolated lines show that the gas should have a zero volume at  $-273.15\text{ }^{\circ}\text{C}$ . Recall from Chapter 1 that  $-273.15\text{ }^{\circ}\text{C}$  corresponds to 0 K (zero on the Kelvin scale), the coldest possible temperature. The extrapolated lines indicate that below  $-273.15\text{ }^{\circ}\text{C}$ , the gas would have a negative volume, which is physically impossible. For this reason, we refer to 0 K as *absolute zero*—colder temperatures do not exist.

The first person to carefully quantify the relationship between the volume of a gas and its temperature was J. A. C. Charles (1746–1823), a French mathematician and physicist. Charles was interested in gases and was among the first people to ascend in a hydrogen-filled balloon. The direct proportionality between volume and temperature is named **Charles's law** after him.

$$\text{Charles's law: } V \propto T \text{ (constant } P \text{ and } n\text{)}$$

When the temperature of a gas sample increases, the gas particles move faster; collisions with the walls are more frequent, and the force exerted with each collision is greater. The only way for the pressure (the force per unit area) to remain constant is for the gas to occupy a larger volume so that collisions become less frequent and occur over a larger area (Figure 6.11►).

Charles's law assumes constant pressure and constant amount of gas.

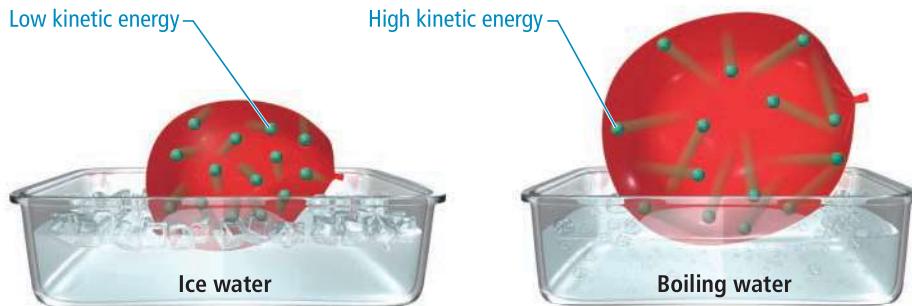


▲ A hot-air balloon floats because the hot air within the balloon is less dense than the surrounding cold air.



▲ If we pour liquid nitrogen (which is at 77 K) onto a balloon, it shrivels up as the air within it cools and occupies less volume at the same external pressure.

### Volume versus Temperature: A Molecular View



▲ **FIGURE 6.11 Molecular Interpretation of Charles's Law** If we move a balloon from an ice water bath to a boiling water bath, its volume expands as the gas particles within the balloon move faster (due to the increased temperature) and collectively occupy more space.

Charles's law explains why the second floor of a house is usually warmer than the ground floor. According to Charles's law, when air is heated, its volume increases, resulting in a lower density. The warm, less dense air tends to rise in a room filled with colder, denser air. Similarly, Charles's law explains why a hot-air balloon can take flight. The gas that fills a hot-air balloon is warmed with a burner, increasing its volume, lowering its density, and causing it to float in the colder, denser surrounding air.

You can experience Charles's law directly by holding a partially inflated balloon over a warm toaster. As the air in the balloon warms, you can feel the balloon expanding. Alternatively, you can put an inflated balloon into liquid nitrogen and watch it become smaller as it cools.

We can use Charles's law to calculate the volume of a gas following a temperature change or the temperature of a gas following a volume change *as long as the pressure and the amount of gas are constant*. For these calculations, we rearrange Charles's law as follows:

$$\text{Since } V \propto T, \text{ then } V = \text{constant} \times T$$

If we divide both sides by  $T$ , we find that  $V/T$  is equal to a constant:

$$V/T = \text{constant}$$

If the temperature increases, the volume increases in direct proportion so that the quotient,  $V/T$ , is always equal to the same constant. So, for two different measurements, we can say that:

$$V_1/T_1 = \text{constant} = V_2/T_2$$

or

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \quad [6.3]$$

where  $V_1$  and  $T_1$  are the initial volume and temperature of the gas and  $V_2$  and  $T_2$  are the final volume and temperature. *The temperatures must always be expressed in kelvins (K)* because, as shown in Figure 6.10, the volume of a gas is directly proportional to its *absolute* temperature, not its temperature in °C. For example, doubling the temperature of a gas sample from 1 °C to 2 °C does not double its volume, but doubling the temperature from 200 K to 400 K does.

**EXAMPLE 6.3** Charles's Law

A sample of gas has a volume of 2.80 L at an unknown temperature. When the sample is submerged in ice water at  $T = 0.00^\circ\text{C}$ , its volume decreases to 2.57 L. What was its initial temperature (in K and in  $^\circ\text{C}$ )?

To solve the problem, first solve Charles's law for  $T_1$ .

**SOLUTION**

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

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

Before you substitute the numerical values to calculate  $T_1$ , convert the temperature to kelvins (K). Remember, gas law problems must always be worked with Kelvin temperatures.

$$T_2(\text{K}) = 0.00 + 273.15 = 273.15 \text{ K}$$

Substitute  $T_2$  and the other given quantities to calculate  $T_1$ .

$$\begin{aligned} T_1 &= \frac{V_1}{V_2} T_2 \\ &= \frac{2.80 \text{ L}}{2.57 \text{ L}} 273.15 \text{ K} \\ &= 297.6 \text{ K} \end{aligned}$$

Calculate  $T_1$  in  $^\circ\text{C}$  by subtracting 273.15 from the value in kelvins.

$$T_1(\text{C}) = 297.6 - 273.15 = 24^\circ\text{C}$$

**FOR PRACTICE 6.3** A gas in a cylinder with a moveable piston has an initial volume of 88.2 mL. If we heat the gas from 35  $^\circ\text{C}$  to 155  $^\circ\text{C}$ , what is its final volume (in mL)?

**BOYLE'S LAW AND CHARLES'S LAW** The pressure exerted on a sample of a fixed amount of gas is doubled at constant temperature, and then the temperature of the gas in kelvins is doubled at constant pressure. What is the final volume of the gas?

- (a) The final volume is twice the initial volume.
- (b) The final volume of the gas is four times the initial volume.
- (c) The final volume of the gas is one-half the initial volume.
- (d) The final volume of the gas is one-fourth the initial volume.
- (e) The final volume of the gas is the same as the initial volume.



ANSWER NOW!

**Avogadro's Law: Volume and Amount (in Moles)**

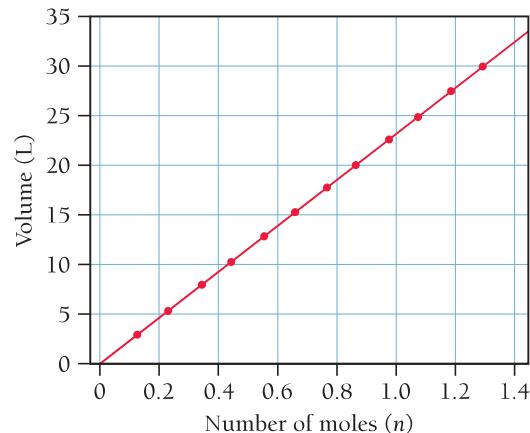
So far, we have discussed the relationships between volume and pressure, and volume and temperature, but we have considered only a constant amount of a gas. What happens when the amount of gas changes? The volume of a gas sample (at constant temperature and pressure) as a function of the amount of gas (in moles) in the sample is shown in Figure 6.12▶. We can see that the relationship between volume and amount is linear. As we might expect, extrapolation to zero moles shows zero volume. This relationship, first stated formally by Amedeo Avogadro (1776–1856), is **Avogadro's law**.

Avogadro's law:  $V \propto n$  (constant  $T$  and  $P$ )

When the amount of gas in a sample increases at constant temperature and pressure, its volume increases in direct proportion because the greater number of gas particles fills more space.

**Avogadro's Law**

As amount of gas increases, volume increases.



► **FIGURE 6.12** Volume versus Number of Moles The volume of a gas sample increases linearly with the number of moles of gas in the sample.

Avogadro's law assumes constant temperature and constant pressure and is independent of the nature of the gas.

## ANSWER NOW!



## 6.2 Cc Conceptual Connection

You experience Avogadro's law when you inflate a balloon. With each exhaled breath, you add more gas particles to the inside of the balloon, increasing its volume. We can use Avogadro's law to calculate the volume of a gas following a change in the amount of the gas *as long as the pressure and temperature of the gas are constant*. For these types of calculations, we express Avogadro's law as:

$$\frac{V_1}{n_1} = \frac{V_2}{n_2} \quad [6.4]$$

where  $V_1$  and  $n_1$  are the initial volume and number of moles of the gas and  $V_2$  and  $n_2$  are the final volume and number of moles. In calculations, we use Avogadro's law in a manner similar to the other gas laws, as demonstrated in Example 6.4.

### SIMPLE GAS LAWS

Which action causes the volume of a gas sample to increase?

- (a) Decreasing the pressure (at constant temperature and number of moles).
- (b) Decreasing the temperature (at constant pressure and number of moles).
- (c) Decreasing the number of moles of gas (at constant temperature and pressure).
- (d) None of the above.

### EXAMPLE 6.4 Avogadro's Law

A male athlete in a kinesiology research study has a lung volume of 6.15 L during a deep inhalation. At this volume, his lungs contain 0.254 mol of air. During exhalation, his lung volume decreases to 2.55 L. How many moles of gas did the athlete exhale? Assume constant temperature and pressure.

To solve the problem, first solve Avogadro's law for the number of moles of gas left in the athlete's lungs after exhalation,  $n_2$ . Then substitute the given quantities to calculate  $n_2$ .

Since the lungs initially contained 0.254 mol of air, calculate the amount of air exhaled by subtracting the result from 0.254 mol. (In Chapter 1, we introduced the practice of underlining the least [rightmost] significant digit of intermediate answers but not rounding the final answer until the very end of the calculation. We continue that practice throughout the book. However, in order to avoid unnecessary notation, we will not carry additional digits in cases, such as this one, where doing so does not affect the final answer.)

**SOLUTION**

$$\begin{aligned}\frac{V_1}{n_1} &= \frac{V_2}{n_2} \\ n_2 &= \frac{V_2}{V_1} n_1 \\ &= \frac{2.55 \text{ L}}{6.15 \text{ L}} 0.254 \text{ mol} \\ &= 0.105 \text{ mol}\end{aligned}$$

$$\begin{aligned}\text{moles exhaled} &= 0.254 \text{ mol} - 0.105 \text{ mol} \\ &= 0.149 \text{ mol}\end{aligned}$$

**FOR PRACTICE 6.4** A chemical reaction occurring in a cylinder equipped with a moveable piston produces 0.621 mol of a gaseous product. If the cylinder contained 0.120 mol of gas before the reaction and had an initial volume of 2.18 L, what is its volume after the reaction? (Assume constant pressure and temperature and that the initial amount of gas completely reacts.)

## 6.4

### The Ideal Gas Law

The relationships that we have discussed so far can be combined into a single law that encompasses all of them. So far, we have shown that:

$$V \propto \frac{1}{P} \quad (\text{Boyle's law})$$

$$V \propto T \quad (\text{Charles's law})$$

$$V \propto n \quad (\text{Avogadro's law})$$

Combining these three expressions, we find that  $V$  is proportional to  $nT/P$ :

$$V \propto \frac{nT}{P}$$

The volume of a gas is directly proportional to the number of moles of gas and to the temperature of the gas but is inversely proportional to the pressure of the gas. We can replace the proportionality sign with an equals sign by incorporating  $R$ , a proportionality constant called the *ideal gas constant*:

$$V = \frac{RnT}{P}$$

Rearranging, we get the equation:

$$PV = nRT \quad [6.5]$$

This equation is the **ideal gas law**, and a hypothetical gas that exactly follows this law is an **ideal gas**. The value of  $R$ , the **ideal gas constant**, is the same for all gases and has the value:

$$R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

The ideal gas law contains within it the simple gas laws that we have discussed. For example, recall that Boyle's law states that  $V \propto 1/P$  when the amount of gas ( $n$ ) and the temperature of the gas ( $T$ ) are kept constant. We can rearrange the ideal gas law as follows:

$$PV = nRT$$

First, divide both sides by  $P$ :

$$V = \frac{nRT}{P}$$

Then put the variables that are constant, along with  $R$ , in parentheses:

$$V = (nRT) \frac{1}{P}$$

Since  $n$  and  $T$  are constant in this case and since  $R$  is always a constant, we can write:

$$V \propto \text{constant} \times \frac{1}{P}$$

which means that  $V \propto 1/P$ .

The ideal gas law also shows how other pairs of variables are related. For example, from Charles's law we know that  $V \propto T$  at constant pressure and constant number of moles. But what if we heat a sample of gas at constant *volume* and constant number of moles? This question applies to the warning labels on aerosol cans such as hair spray or deodorants. These labels warn against excessive heating or incineration of the can, even after the contents are used up. Why? An “empty” aerosol can is not really empty but contains a fixed amount of gas trapped in a fixed volume. What would happen if we were to heat the can? We can rearrange the ideal gas law to clearly see the relationship between pressure and temperature at constant volume and constant number of moles:

$$PV = nRT$$

$$P = \frac{nRT}{V} = \left( \frac{nR}{V} \right) T$$

Since  $n$  and  $V$  are constant and since  $R$  is always a constant:

$$P = \text{constant} \times T$$

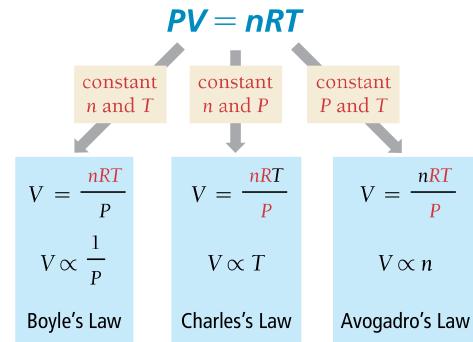
This relationship between pressure and temperature is also known as *Gay-Lussac's law*. As the temperature of a fixed amount of gas in a fixed volume increases, the pressure increases. In an aerosol can, this pressure increase can blow the can apart, which is why aerosol cans should not be heated or incinerated. They might explode.

We can use the ideal gas law to determine the value of any one of the four variables ( $P$ ,  $V$ ,  $n$ , or  $T$ ) given the other three. To do so, we must express each of the quantities in the ideal gas law in the units within  $R$ :

- pressure ( $P$ ) in atm
- volume ( $V$ ) in L
- moles ( $n$ ) in mol
- temperature ( $T$ ) in K

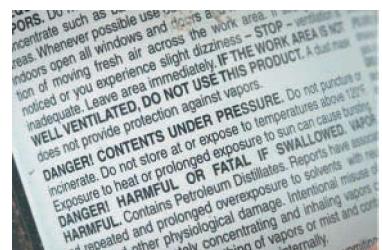
L = liters  
atm = atmospheres  
mol = moles  
K = kelvins

### Ideal Gas Law



▲ The ideal gas law contains the simple gas laws within it.

Divide both sides by  $V$ .



▲ The labels on most aerosol cans warn against incineration. Since the volume of the can is constant, an increase in temperature causes an increase in pressure and possibly an explosion.

## WATCH NOW!

## INTERACTIVE WORKED EXAMPLE 6.5

**EXAMPLE 6.5** Ideal Gas Law I

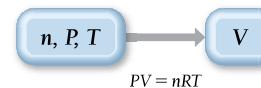
Calculate the volume occupied by 0.845 mol of nitrogen gas at a pressure of 1.37 atm and a temperature of 315 K.

**SORT** The problem gives you the number of moles of nitrogen gas, the pressure, and the temperature. You are asked to find the volume.

**GIVEN:**  $n = 0.845 \text{ mol}$ ,  
 $P = 1.37 \text{ atm}$ ,  $T = 315 \text{ K}$

**FIND:**  $V$

**STRATEGIZE** You are given three of the four variables ( $P$ ,  $T$ , and  $n$ ) in the ideal gas law and asked to find the fourth ( $V$ ). The conceptual plan shows how the ideal gas law provides the relationship between the known quantities and the unknown quantity.

**CONCEPTUAL PLAN****RELATIONSHIPS USED**

$$PV = nRT \text{ (ideal gas law)}$$

**SOLVE** To solve the problem, first solve the ideal gas law for  $V$ .

Then substitute the given quantities to calculate  $V$ .

**SOLUTION**

$$PV = nRT$$

$$V = \frac{nRT}{P}$$

$$V = \frac{0.845 \text{ mol} \times 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 315 \text{ K}}{1.37 \text{ atm}}$$

$$= 15.9 \text{ L}$$

**CHECK** The units of the answer are correct. The magnitude of the answer (15.9 L) makes sense because, as you will see in the next section, one mole of an ideal gas under standard temperature and pressure (273 K and 1 atm) occupies 22.4 L. Although this is not standard temperature and pressure, the conditions are close enough for a ballpark check of the answer. Since this gas sample contains 0.845 mol, a volume of 15.9 L is reasonable.

**FOR PRACTICE 6.5** An 8.50-L tire contains 0.552 mol of gas at a temperature of 305 K. What is the pressure (in atm and psi) of the gas in the tire?

**EXAMPLE 6.6** Ideal Gas Law II

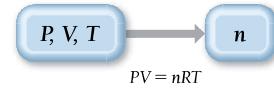
Calculate the number of moles of gas in a 3.24-L basketball inflated to a *total pressure* of 24.3 psi at 25 °C. (Note: The *total pressure* is not the same as the pressure read on a pressure gauge such as the type used for checking a car or bicycle tire. That pressure, called the *gauge pressure*, is the *difference* between the total pressure and atmospheric pressure. In this case, if atmospheric pressure is 14.7 psi, the gauge pressure would be 9.6 psi. However, for calculations involving the ideal gas law, you must use the *total pressure* of 24.3 psi.)

**SORT** The problem gives you the pressure, the volume, and the temperature. You are asked to find the number of moles of gas.

**GIVEN:**  $P = 24.3 \text{ psi}$ ,  $V = 3.24 \text{ L}$ ,  $T(^\circ\text{C}) = 25 \text{ }^\circ\text{C}$

**FIND:**  $n$

**STRATEGIZE** The conceptual plan shows how the ideal gas law provides the relationship between the given quantities and the quantity to be found.

**CONCEPTUAL PLAN****RELATIONSHIPS USED**

$$PV = nRT \text{ (ideal gas law)}$$

**SOLVE** To solve the problem, first solve the ideal gas law for  $n$ .

Before substituting into the equation, convert  $P$  and  $T$  into the correct units.

Finally, substitute into the equation and calculate  $n$ .

### SOLUTION

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$P = 24.3 \text{ psi} \times \frac{1 \text{ atm}}{14.7 \text{ psi}} = 1.6531 \text{ atm}$$

(Since rounding the intermediate answer would result in a slightly different final answer, mark the least significant digit in the intermediate answer but don't round until the end.)

$$T(\text{K}) = 25 + 273 = 298 \text{ K}$$

$$n = \frac{1.6531 \text{ atm} \times 3.24 \text{ L}}{0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 298 \text{ K}} = 0.219 \text{ mol}$$

**CHECK** The units of the answer are correct. The magnitude of the answer (0.219 mol) makes sense because, as you will see in the next section, one mole of an ideal gas under standard temperature and pressure (273 K and 1 atm) occupies 22.4 L. At a pressure that is 65% higher than standard pressure, the volume of 1 mol of gas is proportionally lower. Since this gas sample occupies 3.24 L, the answer of 0.219 mol is reasonable.

**FOR PRACTICE 6.6** What volume does 0.556 mol of gas occupy at a pressure of 715 mmHg and a temperature of 58 °C?

**FOR MORE PRACTICE 6.6** Determine the pressure in mmHg of a 0.133-g sample of helium gas in a 648-mL container at 32 °C.

## 6.5

### Applications of the Ideal Gas Law: Molar Volume, Density, and Molar Mass of a Gas

We just examined how we can use the ideal gas law to calculate one of the variables ( $P$ ,  $V$ ,  $T$ , or  $n$ ) given the other three. We now turn to three other applications of the ideal gas law: molar volume, density, and molar mass.

#### Molar Volume at Standard Temperature and Pressure

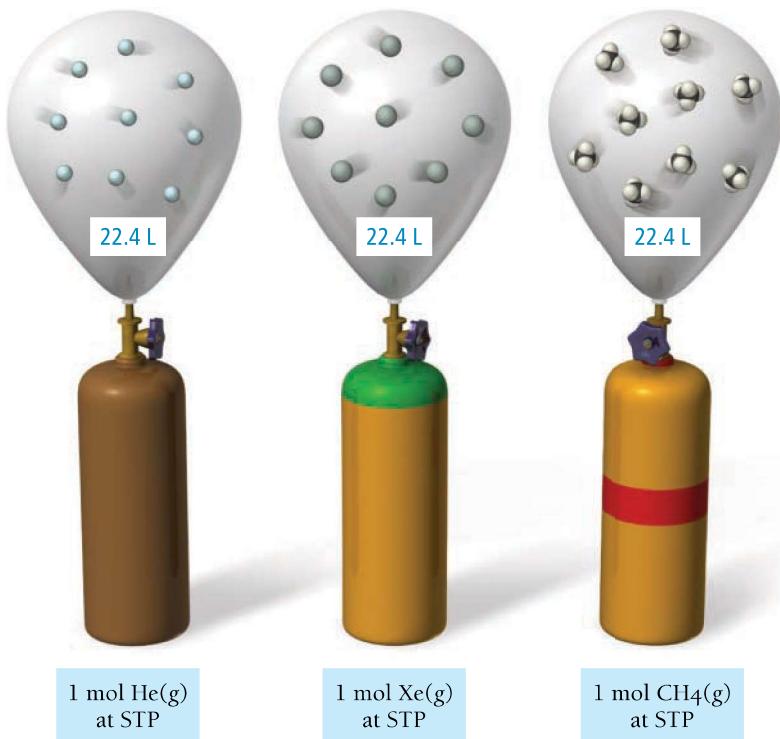
The volume occupied by one mole of a substance is its **molar volume**. For gases, we often specify the molar volume under conditions known as **standard temperature** ( $T = 0^\circ\text{C}$  or 273 K) and **pressure** ( $P = 1.00 \text{ atm}$ ), abbreviated as **STP**. Applying the ideal gas law, we can determine that the molar volume of an ideal gas at STP is:

$$V = \frac{nRT}{P}$$

$$= \frac{1.00 \text{ mol} \times 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \times 273 \text{ K}}{1.00 \text{ atm}}$$

$$= 22.4 \text{ L}$$

The molar volume of 22.4 L only applies at STP.



▲ One mole of any gas occupies approximately 22.4 L at standard temperature (273 K) and pressure (1.0 atm).

ANSWER NOW!



## 6.3 Cc Conceptual Connection

The molar volume of an ideal gas at STP is useful because—as we saw in the *Check* steps of Examples 6.5 and 6.6—it gives us a way to approximate the volume of an ideal gas under conditions that are close to STP.

**MOLAR VOLUME** Assuming ideal behavior, which of these gas samples has the greatest volume at STP?

- (a) 1 g of H<sub>2</sub>
- (b) 1 g of O<sub>2</sub>
- (c) 1 g of Ar

### Density of a Gas

Since one mole of an ideal gas occupies 22.4 L under standard temperature and pressure, we can readily calculate the density of an ideal gas under these conditions. Because density is mass/volume and because the mass of one mole of a gas is simply its molar mass, the *density of a gas* is its molar mass divided by its molar volume:

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

We can calculate the density of a gas at STP by using 22.4 L as the molar volume. For example, the densities of helium and nitrogen gas at STP are:

$$d_{\text{He}} = \frac{4.00 \text{ g/mol}}{22.4 \text{ L/mol}} = 0.179 \text{ g/L} \quad d_{\text{N}_2} = \frac{28.02 \text{ g/mol}}{22.4 \text{ L/mol}} = 1.25 \text{ g/L}$$

The primary components of air are nitrogen (about four-fifths) and oxygen (about one-fifth). We discuss the detailed composition of air in Section 6.6.

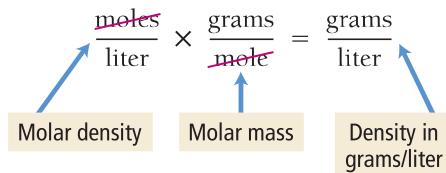
Notice that *the density of a gas is directly proportional to its molar mass*. The greater the molar mass of a gas, the more dense the gas. For this reason, a gas with a molar mass lower than that of air tends to rise in air. For example, both helium and hydrogen gas (molar masses of 4.00 and 2.01 g/mol, respectively) have molar masses that are lower than the average molar mass of air (approximately 28.8 g/mol). Therefore, a balloon filled with either helium or hydrogen gas floats in air.

We can calculate the density of a gas more generally (under any conditions) by using the ideal gas law. To do so, we can arrange the ideal gas law as:

$$PV = nRT$$

$$\frac{n}{V} = \frac{P}{RT}$$

Since the left-hand side of this equation has units of moles/liter, it represents the *molar density*. We obtain the density in grams/liter from the molar density by multiplying by the molar mass ( $\mathcal{M}$ ):



$$\begin{array}{c} \text{Density} \\ \frac{n}{V} \quad \mathcal{M} = \frac{PM}{RT} \\ \text{Molar density} \quad \text{Molar mass} \\ d = \frac{PM}{RT} \end{array}$$

Therefore,

$$d = \frac{PM}{RT} \quad [6.6]$$

Notice that, as expected, density increases with increasing molar mass. Notice also that, as we discussed in Section 6.3, density decreases with increasing temperature.

## WATCH NOW!

## INTERACTIVE WORKED EXAMPLE 6.7

**EXAMPLE 6.7** Density

Calculate the density of nitrogen gas at 125 °C and a pressure of 755 mmHg.

**SORT** The problem gives you the temperature and pressure of nitrogen gas and asks you to find its density.

**STRATEGIZE** Equation 6.6 provides the relationship between the density of a gas and its temperature, pressure, and molar mass. The temperature and pressure are given. You can calculate the molar mass from the formula of the gas, which you know is N<sub>2</sub>.

**GIVEN:** T(°C) = 125 °C, P = 755 mmHg

**FIND:** d

## CONCEPTUAL PLAN

$$\begin{array}{ccc} P, T, M & \longrightarrow & d \\ d = \frac{PM}{RT} \end{array}$$

## RELATIONSHIPS USED

$$d = \frac{PM}{RT} \text{ (density of a gas)}$$

Molar mass N<sub>2</sub> = 28.02 g/mol

**SOLVE** To solve the problem, gather each of the required quantities in the correct units. Convert the temperature to kelvins and the pressure to atmospheres.

Substitute the quantities into the equation to calculate density.

## SOLUTION

$$T(K) = 125 + 273 = 398 K$$

$$P = 755 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.99342 \text{ atm}$$

$$\begin{aligned} d &= \frac{PM}{RT} \\ &= \frac{0.99342 \text{ atm} \left( 28.02 \frac{\text{g}}{\text{mol}} \right)}{0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} (398 \text{ K})} \\ &= 0.852 \text{ g/L} \end{aligned}$$

**CHECK** The units of the answer are correct. The magnitude of the answer (0.852 g/L) makes sense because earlier you calculated the density of nitrogen gas at STP as 1.25 g/L. Since the temperature is higher than standard temperature, it follows that the density is lower.

**FOR PRACTICE 6.7** Calculate the density of xenon gas at a pressure of 742 mmHg and a temperature of 45 °C.

**FOR MORE PRACTICE 6.7** The density of a gas is 1.43 g/L at a temperature of 23 °C and a pressure of 0.789 atm. Calculate the molar mass of the gas.

**DENSITY OF A GAS** Arrange the following gases in order of increasing density at STP: Ne, Cl<sub>2</sub>, F<sub>2</sub>, and O<sub>2</sub>.

- (a) Ne < O<sub>2</sub> < F<sub>2</sub> < Cl<sub>2</sub>
- (b) F<sub>2</sub> < Ne < O<sub>2</sub> < Cl<sub>2</sub>
- (c) Cl<sub>2</sub> < F<sub>2</sub> < O<sub>2</sub> < Ne

## 6.4

## Cc

Conceptual Connection

## ANSWER NOW!

**Molar Mass of a Gas**

We can use the ideal gas law in combination with mass measurements to calculate the molar mass of an unknown gas. First, we measure the mass and volume of an unknown gas under conditions of known pressure and temperature. Then, we determine the amount of the gas in moles from the ideal gas law. Finally, we calculate the molar mass by dividing the mass (in grams) by the amount (in moles) as demonstrated in Example 6.8.

## WATCH NOW!

## INTERACTIVE WORKED EXAMPLE 6.8

**EXAMPLE 6.8** Molar Mass of a Gas

A sample of gas has a mass of 0.311 g. Its volume is 0.225 L at a temperature of 55 °C and a pressure of 886 mmHg. Find its molar mass.

**SORT** The problem gives you the mass of a gas sample, along with its volume, temperature, and pressure. You are asked to find the molar mass.

**STRATEGIZE** The conceptual plan has two parts. In the first part, use the ideal gas law to find the number of moles of gas.

In the second part, use the definition of molar mass to determine the molar mass.

**SOLVE** To find the number of moles, first solve the ideal gas law for  $n$ .

Before substituting into the equation for  $n$ , convert the pressure to atm and the temperature to K.

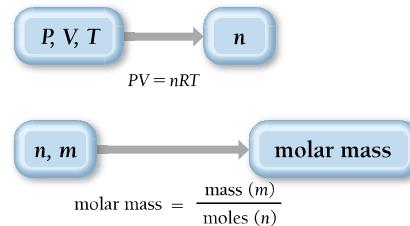
Then substitute into the equation and calculate  $n$ , the number of moles.

Finally, use the number of moles ( $n$ ) and the given mass ( $m$ ) to calculate the molar mass.

**GIVEN:**  $m = 0.311 \text{ g}$ ,  $V = 0.225 \text{ L}$ ,

$T(\text{°C}) = 55 \text{ °C}$ ,  $P = 886 \text{ mmHg}$

**FIND:** molar mass (g/mol)

**CONCEPTUAL PLAN****RELATIONSHIPS USED**

$$PV = nRT$$

$$\text{Molar mass} = \frac{\text{mass (m)}}{\text{moles (n)}}$$

**SOLUTION**

$$PV = nRT$$

$$n = \frac{PV}{RT}$$

$$P = 886 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 1.1658 \text{ atm}$$

$$T(\text{K}) = 55 + 273 = 328 \text{ K}$$

$$n = \frac{1.1658 \text{ atm} \times 0.225 \text{ L}}{0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}} \times 328 \text{ K}$$

$$= 9.7454 \times 10^{-3} \text{ mol}$$

$$\text{molar mass} = \frac{\text{mass (m)}}{\text{moles (n)}}$$

$$= \frac{0.311 \text{ g}}{9.7454 \times 10^{-3} \text{ mol}}$$

$$= 31.9 \text{ g/mol}$$

**CHECK** The units of the answer are correct. The magnitude of the answer (31.9 g/mol) is a reasonable number for a molar mass. If you calculated a very small number (such as any number smaller than 1) or a very large number, you solved the problem incorrectly. Most gases have molar masses between one and several hundred grams per mole.

**FOR PRACTICE 6.8** The mass of a sample of gas is 827 mg. Its volume is 0.270 L at a temperature of 88 °C and a pressure of 975 mmHg. Find its molar mass.

## WATCH NOW!

## KEY CONCEPT VIDEO 6.6

Mixtures of Gases and Partial Pressures

**6.6****Mixtures of Gases and Partial Pressures**

Many gas samples are not pure but are mixtures of gases. Dry air, for example, is a mixture containing nitrogen, oxygen, argon, carbon dioxide, and a few other gases in trace amounts (Table 6.3).

Because the molecules in an ideal gas do not interact (as we will discuss further in Section 6.8), each of the components in an ideal gas mixture acts independently of the others. For example, the nitrogen molecules in air exert a certain pressure—78% of the total pressure—that is independent of the other gases in the mixture. Likewise, the oxygen molecules in air exert a certain pressure—21% of the total pressure—that is also independent of the other gases in the mixture. The pressure due to any individual component in a gas mixture is its **partial pressure** ( $P_n$ ). We can calculate partial pressure from the ideal gas law by assuming that each gas component acts independently.

$$P_n = n_n \frac{RT}{V}$$

For a multicomponent gas mixture, we calculate the partial pressure of each component from the ideal gas law and the number of moles of that component ( $n_n$ ) as follows:

$$P_a = n_a \frac{RT}{V}; P_b = n_b \frac{RT}{V}; P_c = n_c \frac{RT}{V}; \dots \quad [6.7]$$

The sum of the partial pressures of the components in a gas mixture equals the total pressure:

$$P_{\text{total}} = P_a + P_b + P_c + \dots \quad [6.8]$$

where  $P_{\text{total}}$  is the total pressure and  $P_a, P_b, P_c, \dots$  are the partial pressures of the components. This relationship is known as **Dalton's law of partial pressures**.

Combining Equations 6.7 and 6.8, we get:

$$\begin{aligned} P_{\text{total}} &= P_a + P_b + P_c + \dots \\ &= n_a \frac{RT}{V} + n_b \frac{RT}{V} + n_c \frac{RT}{V} + \dots \\ &= (n_a + n_b + n_c + \dots) \frac{RT}{V} \\ &= (n_{\text{total}}) \frac{RT}{V} \end{aligned} \quad [6.9]$$

The total number of moles in the mixture, when substituted into the ideal gas law, indicates the total pressure of the sample.

If we divide Equation 6.7 by Equation 6.9, we get:

$$\frac{P_a}{P_{\text{total}}} = \frac{n_a \cancel{(RT/V)}}{n_{\text{total}} \cancel{(RT/V)}} = \frac{n_a}{n_{\text{total}}} \quad [6.10]$$

The quantity  $n_a/n_{\text{total}}$ , the number of moles of a component in a mixture divided by the total number of moles in the mixture, is the **mole fraction** ( $\chi_a$ ):

$$\chi_a = \frac{n_a}{n_{\text{total}}} \quad [6.11]$$

Rearranging Equation 6.10 and substituting the definition of mole fraction gives:

$$\begin{aligned} \frac{P_a}{P_{\text{total}}} &= \frac{n_a}{n_{\text{total}}} \\ P_a &= \frac{n_a}{n_{\text{total}}} P_{\text{total}} = \chi_a P_{\text{total}} \end{aligned}$$

or simply

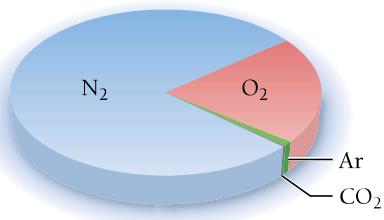
$$P_a = \chi_a P_{\text{total}} \quad [6.12]$$

The partial pressure of a component in a gaseous mixture is its mole fraction multiplied by the total pressure. For gases, the mole fraction of a component is equivalent to its percent by volume divided by 100%. Therefore, based on Table 6.3, we calculate the partial pressure of nitrogen ( $P_{N_2}$ ) in air at 1.00 atm as:

$$\begin{aligned} P_{N_2} &= 0.78 \times 1.00 \text{ atm} \\ &= 0.78 \text{ atm} \end{aligned}$$

**TABLE 6.3 ■ Composition of Dry Air**

| Gas                       | Percent by Volume (%) |
|---------------------------|-----------------------|
| Nitrogen ( $N_2$ )        | 78                    |
| Oxygen ( $O_2$ )          | 21                    |
| Argon (Ar)                | 0.9                   |
| Carbon dioxide ( $CO_2$ ) | 0.04                  |



For these purposes, we can ignore the contribution of the CO<sub>2</sub> and other trace gases because they are so small.

## ANSWER NOW!



## 6.5 Cc Conceptual Connection

Likewise, the partial pressure of oxygen in air at 1.00 atm is 0.21 atm, and the partial pressure of Ar in air is 0.009 atm. Applying Dalton's law of partial pressures to air at 1.00 atm, we can find the total pressure:

$$\begin{aligned} P_{\text{total}} &= P_{\text{N}_2} + P_{\text{O}_2} + P_{\text{Ar}} \\ P_{\text{total}} &= 0.78 \text{ atm} + 0.21 \text{ atm} + 0.009 \text{ atm} \\ &= 1.00 \text{ atm} \end{aligned}$$

**PARTIAL PRESSURES** A gas mixture contains an equal number of moles of He and Ne. The total pressure of the mixture is 3.0 atm. What are the partial pressures of He and Ne?

- (a)  $P_{\text{He}} = 2.0 \text{ atm}; P_{\text{Ne}} = 1.0 \text{ atm}$
- (b)  $P_{\text{He}} = 1.0 \text{ atm}; P_{\text{Ne}} = 2.0 \text{ atm}$
- (c)  $P_{\text{He}} = 1.5 \text{ atm}; P_{\text{Ne}} = 1.5 \text{ atm}$

### EXAMPLE 6.9 Total Pressure and Partial Pressures

A 1.00-L mixture of helium, neon, and argon has a total pressure of 662 mmHg at 298 K. If the partial pressure of helium is 341 mmHg and the partial pressure of neon is 112 mmHg, what mass of argon is present in the mixture?

**SORT** The problem gives you the partial pressures of two of the three components in a gas mixture, along with the total pressure, the volume, and the temperature, and asks you to find the mass of the third component.

**STRATEGIZE** You can find the mass of argon from the number of moles of argon, which you can calculate from the partial pressure of argon and the ideal gas law. Begin by using Dalton's law to determine the partial pressure of argon.

Then use the partial pressure of argon together with the volume of the sample and the temperature to find the number of moles of argon.

Finally, use the molar mass of argon to calculate the mass of argon from the number of moles of argon.

**SOLVE** Follow the conceptual plan. To find the partial pressure of argon, solve the equation for  $P_{\text{Ar}}$  and substitute the values of the other partial pressures to calculate  $P_{\text{Ar}}$ .

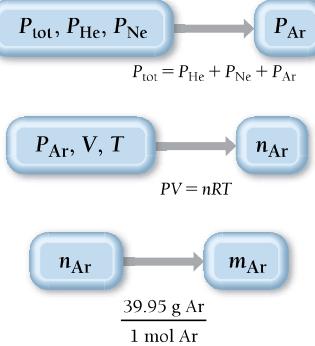
Convert the partial pressure from mmHg to atm and use it in the ideal gas law to calculate the amount of argon in moles.

Use the molar mass of argon to convert from amount of argon in moles to mass of argon.

**GIVEN:**  $P_{\text{He}} = 341 \text{ mmHg}$ ,  
 $P_{\text{Ne}} = 112 \text{ mmHg}$ ,  
 $P_{\text{total}} = 662 \text{ mmHg}$ ,  
 $V = 1.00 \text{ L}, T = 298 \text{ K}$

**FIND:**  $m_{\text{Ar}}$

#### CONCEPTUAL PLAN



#### RELATIONSHIPS USED

$$P_{\text{total}} = P_{\text{He}} + P_{\text{Ne}} + P_{\text{Ar}} \text{ (Dalton's law)}$$

$$PV = nRT \text{ (ideal gas law)}$$

$$\text{molar mass Ar} = 39.95 \text{ g/mol}$$

#### SOLUTION

$$\begin{aligned} P_{\text{total}} &= P_{\text{He}} + P_{\text{Ne}} + P_{\text{Ar}} \\ P_{\text{Ar}} &= P_{\text{total}} - P_{\text{He}} - P_{\text{Ne}} \\ &= 662 \text{ mmHg} - 341 \text{ mmHg} - 112 \text{ mmHg} \\ &= 209 \text{ mmHg} \end{aligned}$$

$$209 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.275 \text{ atm}$$

$$n = \frac{PV}{RT} = \frac{0.275 \text{ atm} (1.00 \text{ L})}{0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} (298 \text{ K})} = 1.125 \times 10^{-2} \text{ mol Ar}$$

$$1.125 \times 10^{-2} \text{ mol Ar} \times \frac{39.95 \text{ g Ar}}{1 \text{ mol Ar}} = 0.449 \text{ g Ar}$$

**CHECK** The units of the answer are correct. The magnitude of the answer makes sense because the volume is 1.0 L, which at STP would contain about 1/22 mol. Since the partial pressure of argon in the mixture is about 1/3 of the total pressure, you can roughly estimate about 1/66 of one molar mass of argon, which is fairly close to your answer.

**FOR PRACTICE 6.9** A sample of hydrogen gas is mixed with water vapor. The mixture has a total pressure of 755 torr, and the water vapor has a partial pressure of 24 torr. What amount (in moles) of hydrogen gas is contained in 1.55 L of this mixture at 298 K?

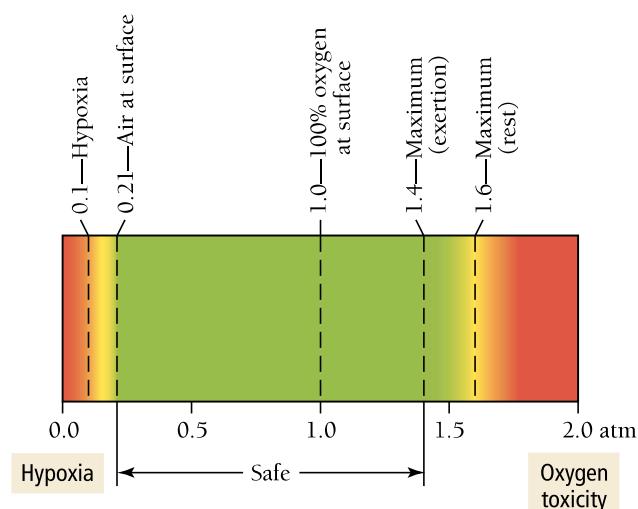
## Deep-Sea Diving and Partial Pressures

Our lungs have evolved to breathe oxygen at a partial pressure of  $P_{O_2} = 0.21 \text{ atm}$ . If the total pressure decreases—when we climb a mountain, for example—the partial pressure of oxygen also decreases. On top of Mount Everest, where the total pressure is 0.311 atm, the partial pressure of oxygen is only 0.065 atm. Low oxygen levels produce a physiological condition called **hypoxia** or oxygen starvation (Figure 6.13►). Mild hypoxia causes dizziness, headache, and shortness of breath. Severe hypoxia, which occurs when  $P_{O_2}$  drops below 0.1 atm, may result in unconsciousness or even death. For this reason, climbers hoping to make the summit of Mount Everest usually carry oxygen to breathe.

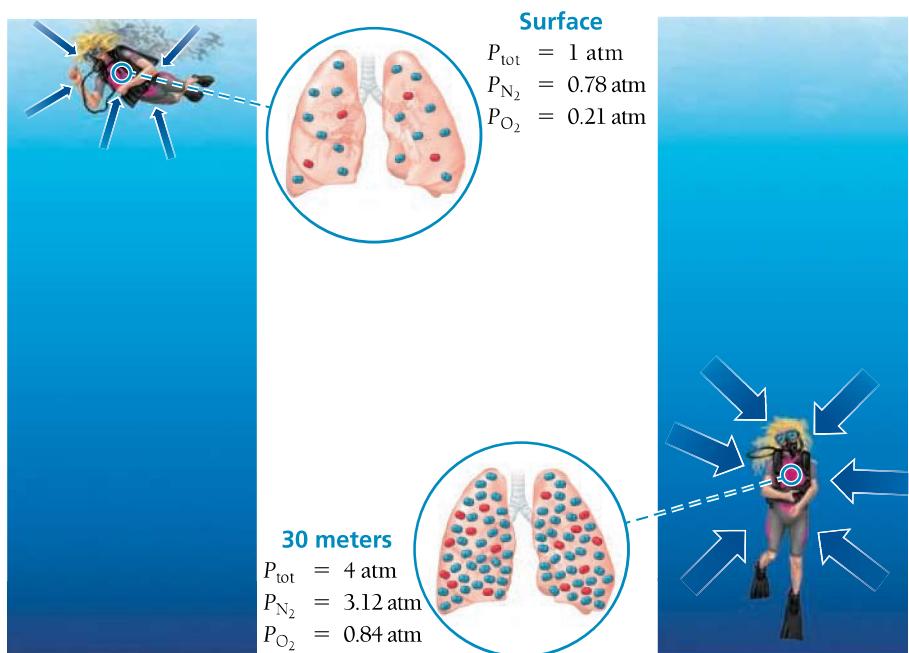
Though not as dangerous as a lack of oxygen, too much oxygen can also cause physiological problems. Recall from Section 6.3 that scuba divers breathe pressurized air. At 30 m, a scuba diver breathes air at a total pressure of 4.0 atm, making  $P_{O_2}$  about 0.84 atm. This elevated partial pressure of oxygen raises the density of oxygen molecules in the lungs, resulting in a higher concentration of oxygen in body tissues. When  $P_{O_2}$  increases beyond 1.4 atm, the increased oxygen concentration in body tissues causes a condition called **oxygen toxicity**, which results in muscle twitching, tunnel vision, and convulsions. Divers who venture too deep without proper precautions have drowned because of oxygen toxicity. A second problem associated with breathing pressurized air is the increase of nitrogen in the lungs. At 30 m, a scuba diver breathes nitrogen at  $P_{N_2} = 3.12 \text{ atm}$ , which increases the nitrogen concentration in body tissues and fluids. When  $P_{N_2}$  increases beyond about 4 atm, a condition called **nitrogen narcosis** or *rapture of the deep* results. Divers describe this condition as feeling inebriated or drunk. A diver breathing compressed air at 60 m feels as if he has consumed too much wine.

To avoid oxygen toxicity and nitrogen narcosis, deep-sea divers—those who descend beyond 50 m—breathe specialized mixtures of gases. One common mixture is heliox, a mixture of helium and oxygen. These mixtures usually contain a smaller percentage of oxygen than would be found in air, thereby lowering the risk of oxygen toxicity. Heliox also contains helium instead of nitrogen, eliminating the risk of nitrogen narcosis.

► When a diver breathes compressed air, the abnormally high partial pressure of oxygen in the lungs leads to an elevated concentration of oxygen in body tissues.



**▲ FIGURE 6.13 Oxygen Partial Pressure Limits** The partial pressure of oxygen in air at sea level is 0.21 atm. Partial pressures of oxygen below 0.1 atm and above 1.4 atm are dangerous to humans.



## WATCH NOW!

## INTERACTIVE WORKED EXAMPLE 6.10

**EXAMPLE 6.10** Partial Pressures and Mole Fractions

A 12.5-L scuba diving tank contains a helium–oxygen (heliox) mixture made up of 24.2 g of He and 4.32 g of O<sub>2</sub> at 298 K. Calculate the mole fraction and partial pressure of each component in the mixture and the total pressure of the mixture.

**SORT** The problem gives the masses of two gases in a mixture and the volume and temperature of the mixture. You are to find the mole fraction and partial pressure of each component, as well as the total pressure.

**STRATEGIZE** The conceptual plan has several parts. To calculate the mole fraction of each component, you must first find the number of moles of each component. Therefore, in the first part of the conceptual plan, convert the masses to moles using the molar masses.

In the second part, calculate the mole fraction of each component using the mole fraction definition.

To calculate *partial pressures*, calculate the *total pressure* and then use the mole fractions from the previous part to calculate the partial pressures. Calculate the total pressure from the sum of the moles of both components. (Alternatively, you can calculate the partial pressures of the components individually, using the number of moles of each component. Then you can sum them to obtain the total pressure.)

Last, use the mole fractions of each component and the total pressure to calculate the partial pressure of each component.

**SOLVE** Follow the plan to solve the problem. Begin by converting each of the masses to amounts in moles.

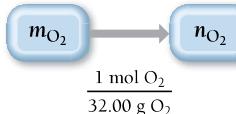
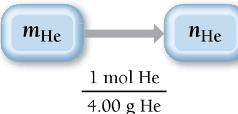
Calculate each of the mole fractions.

Calculate the total pressure.

Finally, calculate the partial pressure of each component.

**GIVEN:**  $m_{\text{He}} = 24.2 \text{ g}$ ,  $m_{\text{O}_2} = 4.32 \text{ g}$ ,  $V = 12.5 \text{ L}$ ,  $T = 298 \text{ K}$

**FIND:**  $\chi_{\text{He}}$ ,  $\chi_{\text{O}_2}$ ,  $P_{\text{He}}$ ,  $P_{\text{O}_2}$ ,  $P_{\text{total}}$

**CONCEPTUAL PLAN**

$$\chi_{\text{He}} = \frac{n_{\text{He}}}{n_{\text{He}} + n_{\text{O}_2}}; \quad \chi_{\text{O}_2} = \frac{n_{\text{O}_2}}{n_{\text{He}} + n_{\text{O}_2}}$$

$$P_{\text{total}} = \frac{(n_{\text{He}} + n_{\text{O}_2})RT}{V}$$

$$P_{\text{He}} = \chi_{\text{He}} P_{\text{total}}; \quad P_{\text{O}_2} = \chi_{\text{O}_2} P_{\text{total}}$$

**RELATIONSHIPS USED**

$$\chi_a = n_a/n_{\text{total}} \text{ (mole fraction definition)}$$

$$P_{\text{total}} V = n_{\text{total}} RT$$

$$P_a = \chi_a P_{\text{total}}$$

**SOLUTION**

$$24.2 \text{ g He} \times \frac{1 \text{ mol He}}{4.00 \text{ g He}} = 6.05 \text{ mol He}$$

$$4.32 \text{ g O}_2 \times \frac{1 \text{ mol O}_2}{32.00 \text{ g O}_2} = 0.135 \text{ mol O}_2$$

$$\chi_{\text{He}} = \frac{n_{\text{He}}}{n_{\text{He}} + n_{\text{O}_2}} = \frac{6.05}{6.05 + 0.135} = 0.97817$$

$$\chi_{\text{O}_2} = \frac{n_{\text{O}_2}}{n_{\text{He}} + n_{\text{O}_2}} = \frac{0.135}{6.05 + 0.135} = 0.021827$$

$$P_{\text{total}} = \frac{(n_{\text{He}} + n_{\text{O}_2})RT}{V}$$

$$= \frac{(6.05 \text{ mol} + 0.135 \text{ mol}) \left( 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (298 \text{ K})}{12.5 \text{ L}}$$

$$= 12.099 \text{ atm}$$

$$P_{\text{He}} = \chi_{\text{He}} P_{\text{total}} = 0.97817 \times 12.099 \text{ atm} = 11.8 \text{ atm}$$

$$P_{\text{O}_2} = \chi_{\text{O}_2} P_{\text{total}} = 0.021827 \times 12.099 \text{ atm} = 0.264 \text{ atm}$$

**CHECK** The units of the answers are correct, and the magnitudes are reasonable.

**FOR PRACTICE 6.10** A diver breathes a heliox mixture with an oxygen mole fraction of 0.050. What must the total pressure be for the partial pressure of oxygen to be 0.21 atm?

## Collecting Gases over Water

When the desired product of a chemical reaction is a gas, the gas is often collected by the displacement of water. For example, suppose we use the reaction of zinc with hydrochloric acid as a source of hydrogen gas:



To collect the gas, we can set up an apparatus like the one shown in Figure 6.14▼. As the hydrogen gas forms, it bubbles through the water and gathers in the collection flask. The hydrogen gas collected in this way is not pure, however. It is mixed with water vapor because some water molecules evaporate and combine with the hydrogen molecules.

The partial pressure of water in the mixture, called its **vapor pressure**, depends on temperature (Table 6.4). Vapor pressure increases with increasing temperature because higher temperatures cause more water molecules to evaporate. We will discuss vapor pressure more thoroughly in Chapter 12.

Suppose we collect hydrogen gas over water at a total pressure of 758.2 mmHg and a temperature of 25 °C. What is the partial pressure of the hydrogen gas? We know that the total pressure is 758.2 mmHg and that the partial pressure of water is 23.78 mmHg (its vapor pressure at 25 °C):

$$P_{\text{total}} = P_{\text{H}_2} + P_{\text{H}_2\text{O}}$$

$$758.2 \text{ mmHg} = P_{\text{H}_2} + 23.78 \text{ mmHg}$$

Therefore,

$$\begin{aligned} P_{\text{H}_2} &= 758.2 \text{ mmHg} - 23.78 \text{ mmHg} \\ &= 734.4 \text{ mmHg} \end{aligned}$$

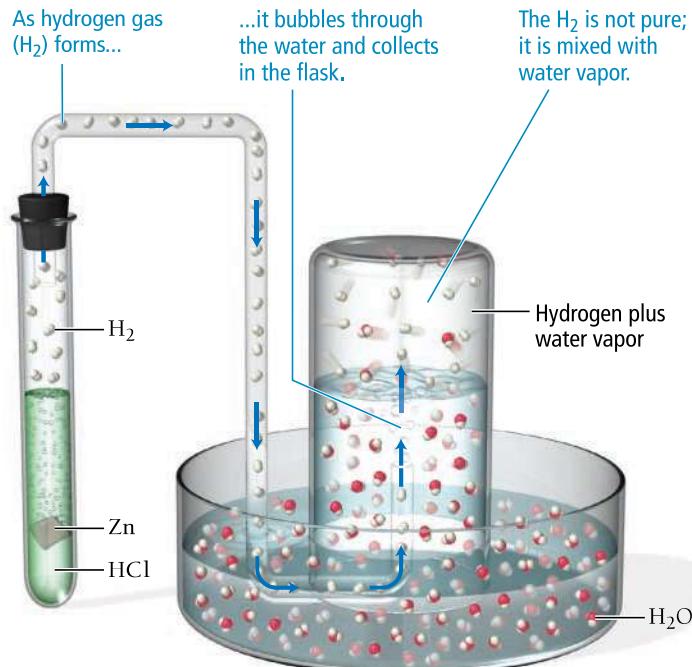
The partial pressure of the hydrogen in the mixture is 734.4 mmHg.

Appendix II E includes a more complete table of the vapor pressure of water versus temperature.

**TABLE 6.4 ■ Vapor Pressure of Water versus Temperature**

| Temperature (°C) | Pressure (mmHg) | Temperature (°C) | Pressure (mmHg) |
|------------------|-----------------|------------------|-----------------|
| 0                | 4.58            | 55               | 118.2           |
| 5                | 6.54            | 60               | 149.6           |
| 10               | 9.21            | 65               | 187.5           |
| 15               | 12.79           | 70               | 233.7           |
| 20               | 17.55           | 75               | 289.1           |
| 25               | 23.78           | 80               | 355.1           |
| 30               | 31.86           | 85               | 433.6           |
| 35               | 42.23           | 90               | 525.8           |
| 40               | 55.40           | 95               | 633.9           |
| 45               | 71.97           | 100              | 760.0           |
| 50               | 92.6            |                  |                 |

### Collecting a Gas over Water



► **FIGURE 6.14 Collecting a Gas over Water** When the gaseous product of a chemical reaction is collected over water, the product molecules (in this case H<sub>2</sub>) mix with water molecules. The pressure of water in the final mixture is equal to the vapor pressure of water at the temperature at which the gas is collected. The partial pressure of the product is the total pressure minus the partial pressure of water.

## WATCH NOW!

## INTERACTIVE WORKED EXAMPLE 6.11

**EXAMPLE 6.11** Collecting Gases over Water

In order to determine the rate of photosynthesis (the conversion by plants of carbon dioxide and water into glucose and oxygen), the oxygen gas emitted by an aquatic plant is collected over water at a temperature of 293 K and a total pressure of 755.2 mmHg. Over a specific time period, a total of 1.02 L of gas is collected. What mass of oxygen gas (in grams) forms?

**SORT** The problem gives the volume of gas collected over water as well as the temperature and the pressure. You are to find the mass in grams of oxygen that forms.

**STRATEGIZE** You can determine the mass of oxygen from the amount of oxygen in moles, which you can calculate from the ideal gas law if you know the partial pressure of oxygen. Since the oxygen is mixed with water vapor, you find the partial pressure of oxygen in the mixture by subtracting the partial pressure of water at 293 K (20 °C) from the total pressure.

Next, use the ideal gas law to determine the number of moles of oxygen from its partial pressure, volume, and temperature.

Finally, use the molar mass of oxygen to convert the number of moles to grams.

**SOLVE** Follow the conceptual plan to solve the problem. Begin by calculating the partial pressure of oxygen in the oxygen/water mixture. You can find the partial pressure of water at 20 °C in Table 6.4.

Next, solve the ideal gas law for number of moles.

Before substituting into the ideal gas law, convert the partial pressure of oxygen from mmHg to atm.

Substitute into the ideal gas law to find the number of moles of oxygen.

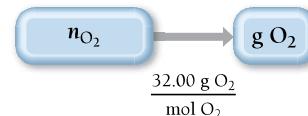
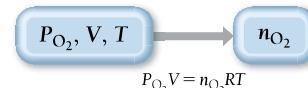
Finally, use the molar mass of oxygen to convert to grams of oxygen.

**GIVEN:**  $V = 1.02 \text{ L}$ ,  $P_{\text{total}} = 755.2 \text{ mmHg}$ ,  $T = 293 \text{ K}$

**FIND:** g O<sub>2</sub>

## CONCEPTUAL PLAN

$$P_{\text{O}_2} = P_{\text{total}} - P_{\text{H}_2\text{O}}(20 \text{ }^{\circ}\text{C})$$



## RELATIONSHIPS USED

$$P_{\text{total}} = P_a + P_b + P_c + \dots \text{ (Dalton's law)}$$

$$PV = nRT \text{ (ideal gas law)}$$

## SOLUTION

$$\begin{aligned} P_{\text{O}_2} &= P_{\text{total}} - P_{\text{H}_2\text{O}}(20 \text{ }^{\circ}\text{C}) \\ &= 755.2 \text{ mmHg} - 17.55 \text{ mmHg} \\ &= 737.65 \text{ mmHg} \end{aligned}$$

$$n_{\text{O}_2} = \frac{P_{\text{O}_2}V}{RT}$$

$$737.65 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.97059 \text{ atm}$$

$$\begin{aligned} n_{\text{O}_2} &= \frac{P_{\text{O}_2}V}{RT} = \frac{0.97059 \text{ atm}(1.02 \text{ L})}{0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}(293 \text{ K})} \\ &= 4.1175 \times 10^{-2} \text{ mol} \end{aligned}$$

$$4.1175 \times 10^{-2} \text{ mol O}_2 \times \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} = 1.32 \text{ g O}_2$$



**CHECK** The answer is in the correct units. You can quickly check the magnitude of the answer by using molar volume. Under STP, one liter is about 1/22 of one mole. Therefore, the answer should be about 1/22 the molar mass of oxygen ( $1/22 \times 32 = 1.45$ ). The magnitude of the answer seems reasonable.

**FOR PRACTICE 6.11** A common way to make hydrogen gas in the laboratory is to place a metal such as zinc in hydrochloric acid. The hydrochloric acid reacts with the metal to produce hydrogen gas, which is then collected over water. Suppose a student carries out this reaction and collects a total of 154.4 mL of gas at a pressure of 742 mmHg and a temperature of 25 °C. What mass of hydrogen gas (in mg) does the student collect?

## 6.7

# Gases in Chemical Reactions: Stoichiometry Revisited

In Chapter 4, we discussed how to use the coefficients in chemical equations as conversion factors between number of moles of reactants and number of moles of products in a chemical reaction. We can use these conversion factors to determine, for example, the mass of product obtained in a chemical reaction based on a given mass of reactant, or the mass of one reactant needed to react completely with a given mass of another reactant. The general conceptual plan for these kinds of calculations is:



where A and B are two different substances involved in the reaction and the conversion factor between amounts (in moles) of each comes from the stoichiometric coefficients in the balanced chemical equation.

In reactions involving *gaseous* reactant or products, we often specify the quantity of a gas in terms of its volume at a given temperature and pressure. As we have seen, stoichiometry involves relationships between amounts in moles. For stoichiometric calculations involving gases, we can use the ideal gas law to determine the amounts in moles from the volumes, or to determine the volumes from the amounts in moles.

$$n = \frac{PV}{RT} \quad V = \frac{nRT}{P}$$

The pressures here could also be partial pressures.

The general conceptual plan for these kinds of calculations is:



Examples 6.12 and 6.13 demonstrate this kind of calculation.

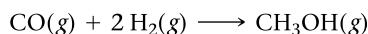
## WATCH NOW!

### INTERACTIVE WORKED EXAMPLE 6.12

#### EXAMPLE 6.12 Gases in Chemical Reactions



Methanol ( $\text{CH}_3\text{OH}$ ) can be synthesized by the reaction:



What volume (in liters) of hydrogen gas, at a temperature of 355 K and a pressure of 738 mmHg, is needed to synthesize 35.7 g of methanol?

**SORT** You are given the mass of methanol, the product of the chemical reaction. You are asked to find the required volume of one of the reactants (hydrogen gas) at a specified temperature and pressure.

**GIVEN:** 35.7 g  $\text{CH}_3\text{OH}$ ,  
 $T = 355 \text{ K}$ ,  $P = 738 \text{ mmHg}$

**FIND:**  $V_{\text{H}_2}$

Continued—

**STRATEGIZE** You can calculate the required volume of hydrogen gas from the number of moles of hydrogen gas, which you can obtain from the number of moles of methanol via the stoichiometry of the reaction.

First, find the number of moles of methanol from its mass by using the molar mass.

Then use the stoichiometric relationship from the balanced chemical equation to find the number of moles of hydrogen you need to form that quantity of methanol.

Finally, substitute the number of moles of hydrogen together with the pressure and temperature into the ideal gas law to find the volume of hydrogen.

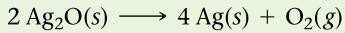
**SOLVE** Follow the conceptual plan to solve the problem. Begin by using the mass of methanol to determine the number of moles of methanol.

Next, convert the number of moles of methanol to moles of hydrogen.

Finally, use the ideal gas law to calculate the volume of hydrogen. Before substituting into the equation, you need to convert the pressure to atmospheres.

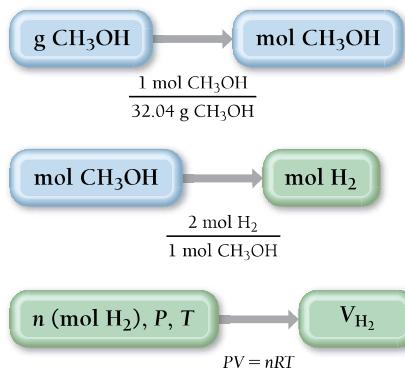
**CHECK** The units of the answer are correct. The magnitude of the answer (66.9 L) seems reasonable. You are given slightly more than one molar mass of methanol, which is therefore slightly more than one mole of methanol. From the equation you can see that you need 2 mol hydrogen to make 1 mol methanol, so the answer must be slightly greater than 2 mol hydrogen. Under standard temperature and pressure, slightly more than 2 mol hydrogen occupies slightly more than  $2 \times 22.4 \text{ L} = 44.8 \text{ L}$ . At a temperature greater than standard temperature, the volume would be even greater; therefore, your answer is reasonable.

**FOR PRACTICE 6.12** In the reaction shown here, 4.58 L of  $\text{O}_2$  forms at  $P = 745 \text{ mmHg}$  and  $T = 308 \text{ K}$ . How many grams of  $\text{Ag}_2\text{O}$  decomposed?



**FOR MORE PRACTICE 6.12** In the previous reaction, what mass of  $\text{Ag}_2\text{O}(s)$  (in grams) is required to form 388 mL of oxygen gas at  $P = 734 \text{ mmHg}$  and  $25.0^\circ\text{C}$ ?

### CONCEPTUAL PLAN



### RELATIONSHIPS USED

$$PV = nRT \text{ (ideal gas law)}$$

2 mol  $\text{H}_2$  : 1 mol  $\text{CH}_3\text{OH}$  (from balanced chemical equation)

molar mass  $\text{CH}_3\text{OH} = 32.04 \text{ g/mol}$

### SOLUTION

$$35.7 \text{ g CH}_3\text{OH} \times \frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g CH}_3\text{OH}} = 1.1142 \text{ mol CH}_3\text{OH}$$

$$1.1142 \text{ mol CH}_3\text{OH} \times \frac{2 \text{ mol H}_2}{1 \text{ mol CH}_3\text{OH}} = 2.2284 \text{ mol H}_2$$

$$V_{\text{H}_2} = \frac{n_{\text{H}_2}RT}{P}$$

$$P = 738 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.97105 \text{ atm}$$

$$V_{\text{H}_2} = \frac{(2.2284 \text{ mol}) \left( 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \right) (355 \text{ K})}{0.97105 \text{ atm}}$$

$$= 66.9 \text{ L}$$

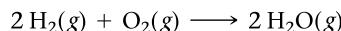
## Molar Volume and Stoichiometry

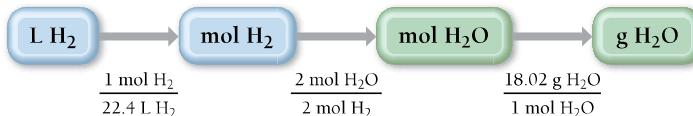
In Section 6.5, we saw that under standard temperature and pressure, 1 mol of an ideal gas occupies 22.4 L. Consequently, if a reaction occurs at or near standard temperature

and pressure, we can use  $1 \text{ mol} = 22.4 \text{ L}$  as a conversion factor in stoichiometric calculations, as demonstrated in Example 6.13.

### EXAMPLE 6.13 Using Molar Volume in Gas Stoichiometric Calculations

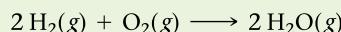
How many grams of water form when 1.24 L of  $\text{H}_2$  gas at STP completely reacts with  $\text{O}_2$ ?



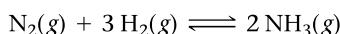
|   |  |
|---|--|
| <b>SORT</b> You are given the volume of hydrogen gas (a reactant) at STP and asked to determine the mass of water that forms upon complete reaction.  | <b>GIVEN:</b> 1.24 L $\text{H}_2$<br><b>FIND:</b> g $\text{H}_2\text{O}$   |
| <b>STRATEGIZE</b> Since the reaction occurs under standard temperature and pressure, you can convert directly from the volume (in liters) of hydrogen gas to the amount in moles. Then use the stoichiometric relationship from the balanced equation to find the number of moles of water formed. Finally, use the molar mass of water to obtain the mass of water formed. | <b>CONCEPTUAL PLAN</b> <br><b>RELATIONSHIPS USED</b><br>$1 \text{ mol} = 22.4 \text{ L}$ (at STP)<br>$2 \text{ mol H}_2 : 2 \text{ mol H}_2\text{O}$ (from balanced equation)<br>molar mass $\text{H}_2\text{O} = 18.02 \text{ g/mol}$ |
| <b>SOLVE</b> Follow the conceptual plan to solve the problem.   | $1.24 \text{ L H}_2 \times \frac{1 \text{ mol H}_2}{22.4 \text{ L H}_2} \times \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} \times \frac{18.02 \text{ g H}_2\text{O}}{1 \text{ mol H}_2\text{O}} = 0.998 \text{ g H}_2\text{O}$   |

**CHECK** The units of the answer are correct. The magnitude of the answer (0.998 g) is about 1/18 of the molar mass of water, which is roughly equivalent to the approximately 1/22 of a mole of hydrogen gas given, as expected for the 1:1 stoichiometric relationship between number of moles of hydrogen and number of moles of water.

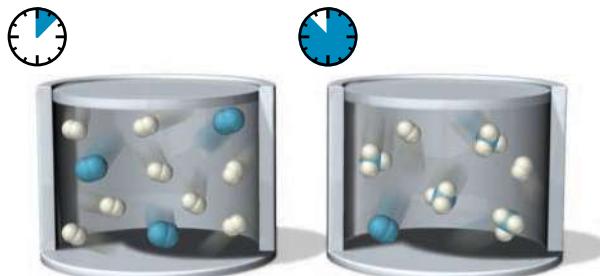
**FOR PRACTICE 6.13** How many liters of oxygen (at STP) are required to form 10.5 g of  $\text{H}_2\text{O}$ ?



**PRESSURE AND NUMBER OF MOLES** Nitrogen and hydrogen react to form ammonia according to the following equation:



Consider the following representations of the initial mixture of reactants and the resulting mixture after the reaction has been allowed to react for some time:



If the volume is kept constant, and nothing is added to the reaction mixture, what happens to the total pressure during the course of the reaction?

- (a) The pressure increases.
- (b) The pressure decreases.
- (c) The pressure does not change.



ANSWER NOW!





## ANALYZING AND INTERPRETING DATA

### Good News about Our Nation's Air Quality

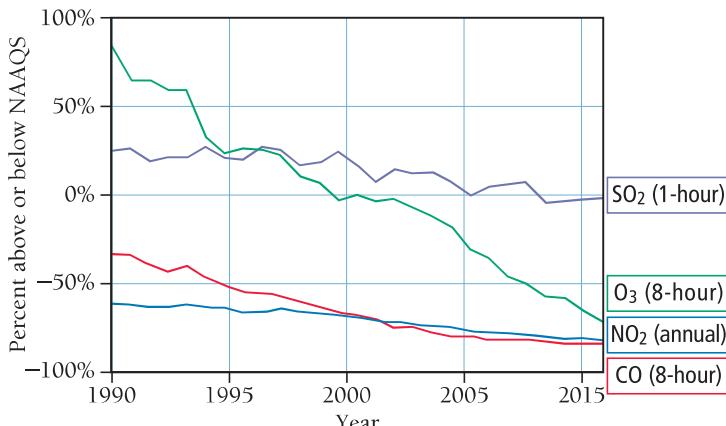
**A**ir pollution has been a major problem in most U.S. cities. These gaseous pollutants are produced (either directly or indirectly) by chemical reactions, especially the combustion of fossil fuels such as petroleum and coal. The major pollutants include carbon monoxide, nitrogen dioxide, sulfur dioxide, and ozone. These pollutants have adverse effects on our health because they affect our respiratory and cardiovascular systems.

The good news is that the concentrations of these pollutants in U.S. cities is decreasing significantly. The U.S. Environmental Protection Agency (EPA) regularly monitors levels of these pollutants relative to standards called the National Ambient Air Quality Standards. The accompanying graph shows progress relative to those standards from 1990 to 2016.

Whenever you encounter data like this, the first thing you should examine are the x- and y-axes. You can't understand a graph if you don't understand its axes. In this case, the x-axis is straightforward: it represents time between 1990 and 2016. The y-axis is a bit more challenging because it represents the percent above or below the National Ambient Air Quality Standards (NAAQS; shown in the following table). By representing the levels this way, the EPA can show the changes in all four of these pollutants on the same graph. If they graphed the actual concentrations of the pollutants, the changes in the pollutants with low concentrations ( $O_3$  and  $SO_2$ ) would not be visible because they would appear as flat lines on the bottom of the graph.

The vast improvements in U.S. air quality have been realized even though the economy, the population, and

### National Air Pollutant Concentration Averages



### National Ambient Air Quality Standards

| Pollutant                | Standard  |
|--------------------------|-----------|
| CO (8 hour)              | 9 ppm     |
| NO <sub>2</sub> (annual) | 53 ppm    |
| O <sub>3</sub> (8 hour)  | 0.070 ppm |
| SO <sub>2</sub> (1 hour) | 75 ppb    |

ppm = parts per million; ppb = parts per billion

overall energy use have grown substantially over the same period. The reasons for these improvements are good legislation (the Clean Air Act and its amendments) and vast improvements in technological innovation. We all breathe cleaner air today thanks to these developments.



### WATCH NOW!

**KEY CONCEPT VIDEO 6.8**  
Kinetic Molecular Theory

6.8

### Kinetic Molecular Theory: A Model for Gases

In Chapter 1, we discussed how the scientific approach proceeds from observations to laws and eventually to theories. Remember that laws summarize behavior—for example, Charles's law summarizes *how* the volume of a gas depends on temperature—whereas theories give the underlying reasons for the behavior. A theory of gas behavior explains, for example, *why* the volume of a gas increases with increasing temperature.

The simplest model for the behavior of gases is the **kinetic molecular theory**. In this theory, a gas is modeled as a collection of particles (either molecules or atoms, depending on the gas) in constant motion (Figure 6.15►). A single particle moves in a straight line until it collides with another particle (or with the wall of the container). Kinetic molecular theory has three basic postulates (or assumptions):

- The size of a particle is negligibly small.** Kinetic molecular theory assumes that the particles themselves occupy negligible volume, even though they have mass. This postulate is justified because, under normal pressures, the space between atoms or molecules in a gas is very large compared to the size of the atoms or molecules themselves. For example, in a sample of argon gas at STP, only about 0.01% of the volume is occupied by atoms, and the average distance from one argon atom to another is 3.3 nm. In comparison, the atomic radius of argon is 97 pm. If an argon atom were the size of a golf ball, its nearest neighbor would be, on average, over 4 ft (1.2 m) away at STP.

- 2. The average kinetic energy of a particle is proportional to the temperature in kelvins.** The motion of atoms or molecules in a gas is due to thermal energy, which distributes itself among the particles in the gas. At any given moment, some particles are moving faster than others—there is a distribution of velocities—but the higher the temperature, the faster the overall motion and the greater the average kinetic energy. Notice that *kinetic energy* ( $\frac{1}{2}mv^2$ )—not *velocity*—is proportional to temperature. The atoms in a sample of helium and a sample of argon at the same temperature have the same average *kinetic energy* but not the same average *velocity*. Because the helium atoms are lighter, they must move faster to have the same kinetic energy as argon atoms.

- 3. The collision of one particle with another (or with the walls of its container) is completely elastic.** When two particles collide, they may *exchange energy*, but there is no overall *loss of energy*. Any kinetic energy lost by one particle is completely gained by the other. In other words, the particles have no “stickiness,” and they are not deformed by the collision. An encounter between two particles in kinetic molecular theory is more like the collision between two billiard balls than the collision between two lumps of clay (Figure 6.16▼). Between collisions, the particles do not exert any forces on one another.

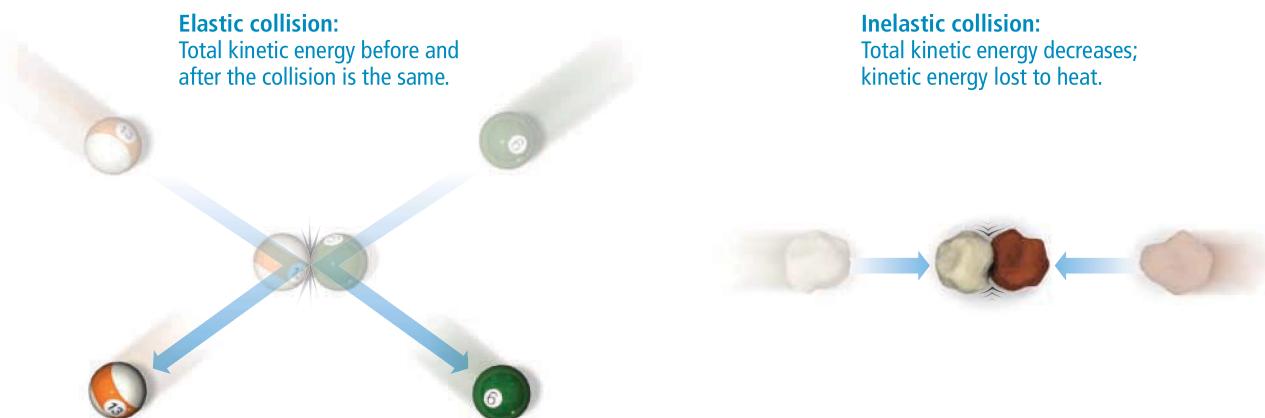
If we start with the postulates of kinetic molecular theory, we can mathematically derive the ideal gas law (as we demonstrate later). Since the ideal gas law follows directly from kinetic molecular theory, we have confidence that the assumptions of the theory are valid, at least under conditions where the ideal gas law applies.

## Kinetic Molecular Theory



▲ FIGURE 6.15 A Model for Gas Behavior

In the kinetic molecular theory of gases, a gas sample is modeled as a collection of particles in constant straight-line motion. The size of each particle is negligibly small, and the particles' collisions are elastic.



## How Kinetic Molecular Theory Explains Pressure and the Simple Gas Laws

Let's examine how the concept of pressure as well as each of the gas laws we have examined in this chapter follow conceptually from kinetic molecular theory.

### The Nature of Pressure

In Section 6.2, we defined pressure as force divided by area:

$$P = \frac{F}{A}$$

From Newton's second law, the force ( $F$ ) associated with an individual collision is given by  $F = ma$ , where  $m$  is the mass of the particle and  $a$  is its acceleration as it changes its direction of travel due to the collision.

According to kinetic molecular theory, a gas is a collection of particles in constant motion. The motion results in collisions between the particles and the surfaces around them. As each particle collides with a surface, it exerts a force upon that surface.

▲ FIGURE 6.16 Elastic versus

Inelastic Collisions When two billiard balls collide, the collision is elastic—the total kinetic energy of the colliding bodies is the same before and after the collision. When two lumps of clay collide, the collision is inelastic—the kinetic energy of the colliding bodies dissipates in the form of heat during the collision.

The result of many particles in a gas sample exerting forces on the surfaces around them is a constant pressure.

### Boyle's Law

Boyle's law states that for a constant number of particles at constant temperature, the volume of a gas is inversely proportional to its pressure. According to kinetic molecular theory, if we decrease the volume of a gas, we force the gas particles to occupy a smaller space. As long as the temperature remains the same, the number of collisions with the surrounding surfaces (per unit surface area) must necessarily increase, resulting in a greater pressure.

### Charles's Law

Charles's law states that for a constant number of particles at constant pressure, the volume of a gas is proportional to its temperature. According to kinetic molecular theory, when we increase the temperature of a gas, the average speed, and thus the average kinetic energy, of the particles increases. Since this greater kinetic energy results in more frequent collisions and more force per collision, the pressure of the gas increases if its volume is held constant (Gay-Lussac's law). The only way for the pressure to remain constant is for the volume to increase. The greater volume spreads the collisions out over a greater surface area so that the pressure (defined as force per unit area) is unchanged.

### Avogadro's Law

Avogadro's law states that at constant temperature and pressure, the volume of a gas is proportional to the number of particles. According to kinetic molecular theory, when we increase the number of particles in a gas sample, the number of collisions with the surrounding surfaces increases. The greater number of collisions results in a greater overall force on surrounding surfaces; the only way for the pressure to remain constant is for the volume to increase so that the number of particles per unit volume (and thus the number of collisions) remains constant.

### Dalton's Law

Dalton's law states that the total pressure of a gas mixture is the sum of the partial pressures of its components. According to Dalton's law, the components in a gas mixture act identically to, and independently of, one another. According to kinetic molecular theory, the particles have negligible size and they do not interact. Consequently, the only property that distinguishes one type of particle from another is its mass. However, even particles of different masses have the same average kinetic energy at a given temperature, so they exert the same force upon collision with a surface. Consequently, adding components to a gas mixture—even different *kinds* of gases—has the same effect as simply adding more particles. The partial pressures of all the components sum to the overall pressure.

## Kinetic Molecular Theory and the Ideal Gas Law

We have just seen how each of the gas laws conceptually follows from kinetic molecular theory. We can also *derive* the ideal gas law from the postulates of kinetic molecular theory. In other words, the kinetic molecular theory is a quantitative model that *implies*  $PV = nRT$ . Let's explore this derivation.

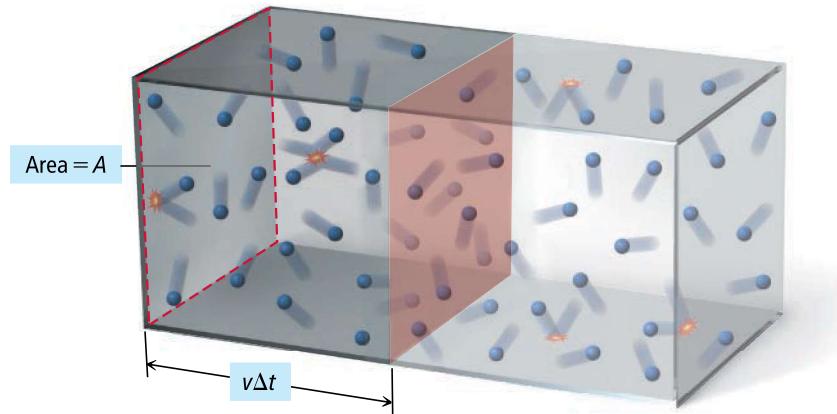
The pressure on a wall of a container (Figure 6.17▶) occupied by particles in constant motion is the total force on the wall (due to the collisions) divided by the area of the wall.

$$P = \frac{F_{\text{total}}}{A} \quad [6.13]$$

Recall that the force ( $F$ ) associated with an individual collision is given by  $F = ma$ , where  $m$  is the mass of the particle and  $a$  is its acceleration as it changes its direction of travel

## Calculating Gas Pressure: A Molecular View

Only particles within the volume defined by  $A \times v\Delta t$  are capable of hitting the end wall within the time period  $\Delta t$ .



**◀ FIGURE 6.17 The Pressure on the Wall of a Container** The pressure on the wall of a container can be calculated by determining the total force due to collisions of the particles with the wall.

due to the collision. The acceleration for each collision is the change in velocity ( $\Delta v$ ) divided by the time interval ( $\Delta t$ ), so the force imparted for each collision is:

$$F_{\text{collision}} = m \frac{\Delta v}{\Delta t} \quad [6.14]$$

If a particle collides elastically with the wall, it bounces off the wall with no loss of energy. For a straight-line collision, the change in velocity is  $2v$  (the particle's velocity was  $v$  before the collision and  $-v$  after the collision; therefore, the change is  $2v$ ). The force per collision is given by:

$$F_{\text{collision}} = m \frac{2v}{\Delta t} \quad [6.15]$$

The total number of collisions in the time interval  $\Delta t$  on a wall of surface area  $A$  is proportional to the number of particles that can reach the wall in this time interval—in other words, all particles within a distance of  $v\Delta t$  of the wall. These particles occupy a volume given by  $v\Delta t \times A$ , and their total number is equal to this volume multiplied by the density of particles in the container ( $n/V$ ):

Number of collisions  $\propto$  number of particles within  $v\Delta t$

$$\propto v\Delta t \times A \times \frac{n}{V}$$

↑  
Volume      Density of particles

[6.16]

The *total force* on the wall is equal to the force per collision multiplied by the number of collisions:

$$\begin{aligned} F_{\text{total}} &= F_{\text{collision}} \times \text{number of collisions} \\ &\propto m \frac{2v}{\Delta t} \times v\Delta t \times A \times \frac{n}{V} \\ &\propto mv^2 \times A \times \frac{n}{V} \end{aligned} \quad [6.17]$$

The pressure on the wall is equal to the total force divided by the surface area of the wall:

$$\begin{aligned} P &= \frac{F_{\text{total}}}{A} \\ &\propto \frac{mv^2 \times A \times \frac{n}{V}}{A} \\ P &\propto mv^2 \times \frac{n}{V} \end{aligned} \quad [6.18]$$

Notice that Equation 6.18 contains within it Boyle's law ( $P \propto 1/V$ ) and Avogadro's law ( $V \propto n$ ). We can get the complete ideal gas law from postulate 2 of the kinetic molecular theory, which states that the average kinetic energy ( $\frac{1}{2}mv^2$ ) is proportional to the temperature in kelvins ( $T$ ):

$$mv^2 \propto T \quad [6.19]$$

By combining Equations 6.18 and 6.19, we get:

$$\begin{aligned} P &\propto \frac{T \times n}{V} \\ PV &\propto nT \end{aligned} \quad [6.20]$$

The proportionality can be replaced by an equals sign if we provide the correct constant,  $R$ :

$$PV = nRT \quad [6.21]$$

In other words, the kinetic molecular theory (a model for how gases behave) predicts behavior that is consistent with our observations and measurements of gases—the theory agrees with the experiment. Recall from Chapter 1 that a scientific theory is the most powerful kind of scientific knowledge. In the kinetic molecular theory, we have a model for what a gas is like. Although the model is not perfect—indeed, it breaks down under certain conditions, as we shall see later in this chapter—it predicts a great deal about the behavior of gases. Therefore, the model is a good approximation of what a gas is actually like. A careful examination of the conditions under which the model breaks down (see Section 6.10) gives us even more insight into the behavior of gases.

## Temperature and Molecular Velocities

According to kinetic molecular theory, particles of different masses have the same average kinetic energy at a given temperature. The kinetic energy of a particle depends on its mass and velocity according to the equation:

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

The only way for particles of different masses to have the same kinetic energy is for them to have different velocities. *In a gas mixture at a given temperature, lighter particles travel faster (on average) than heavier ones.*

In kinetic molecular theory, we define the root mean square velocity ( $u_{\text{rms}}$ ) of a particle as:

$$u_{\text{rms}} = \sqrt{\bar{u}^2} \quad [6.22]$$

where  $\bar{u}^2$  is the average of the squares of the particle velocities. Even though the root mean square velocity of a collection of particles is not identical to the average velocity, the two are close in value and conceptually similar. Root mean square velocity is a special type of average. The average kinetic energy of one mole of gas particles is given by the equation:

$$KE_{\text{avg}} = \frac{1}{2}N_A m \bar{u}^2 \quad [6.23]$$

where  $N_A$  is Avogadro's number and  $m$  is the mass of the particle.

Postulate 2 of the kinetic molecular theory states that the average kinetic energy is proportional to the temperature in kelvins. The constant of proportionality in this relationship is  $(3/2)R$ :

$$KE_{\text{avg}} = (3/2)RT \quad [6.24]$$

where  $R$  is the gas constant but in different units ( $R = 8.314 \text{ J/mol} \cdot \text{K}$ ) than those we use in the ideal gas law. If we combine Equations 6.23 and 6.24, and solve for  $\bar{u}^2$ , we get:

$$\begin{aligned} (1/2)N_A m \bar{u}^2 &= (3/2)RT \\ \bar{u}^2 &= \frac{(3/2)RT}{(1/2)N_A m} = \frac{3RT}{N_A m} \end{aligned}$$

Taking the square root of both sides, we get:

$$\sqrt{\bar{u}^2} = u_{\text{rms}} = \sqrt{\frac{3RT}{N_A m}} \quad [6.25]$$

The  $(3/2)R$  proportionality constant comes from a derivation that is beyond the scope of this book.

The joule (J) is a unit of energy that we will discuss in more detail in Section 7.2.

$$\left(1 \text{ J} = 1 \text{ kg} \frac{\text{m}^2}{\text{s}^2}\right)$$

In Equation 6.25,  $m$  is the mass of a particle in kg and  $N_A$  is Avogadro's number. The product  $N_A m$ , then, is the molar mass in kg/mol. If we call this quantity  $\mathcal{M}$ , then the expression for mean square velocity as a function of temperature becomes the following important result:

$$u_{\text{rms}} = \sqrt{\frac{3RT}{\mathcal{M}}} \quad [6.26]$$

The root mean square velocity of a collection of gas particles is proportional to the square root of the temperature in kelvins and inversely proportional to the square root of the molar mass of the particles (which because of the units of  $R$  must be in kilograms per mole). The root mean square velocity of nitrogen molecules at 25 °C, for example, is 515 m/s (1152 mi/hr). The root mean square velocity of hydrogen molecules at room temperature is 1920 m/s (4295 mi/hr). Notice that the lighter molecules move much faster at a given temperature.

### ROOT MEAN SQUARE VELOCITY

How does the approximate root mean square velocity of neon compare to that of krypton at the same temperature?

- (a) The root mean square velocity of Ne is approximately four times that of Kr.
- (b) The root mean square velocity of Ne is approximately two times that of Kr.
- (c) The root mean square velocity of Ne is approximately one-half that of Kr.
- (d) The root mean square velocity of Ne is approximately one-fourth that of Kr.

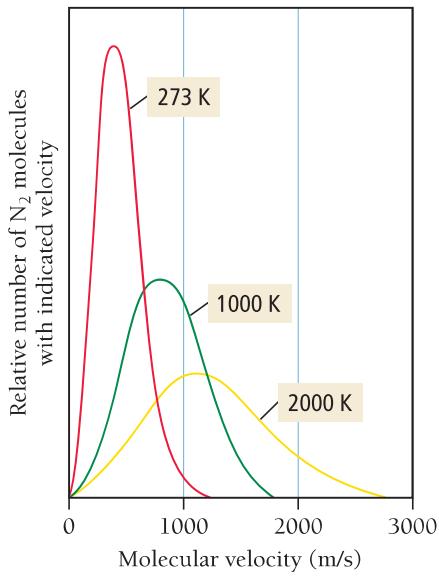


**ANSWER NOW!**

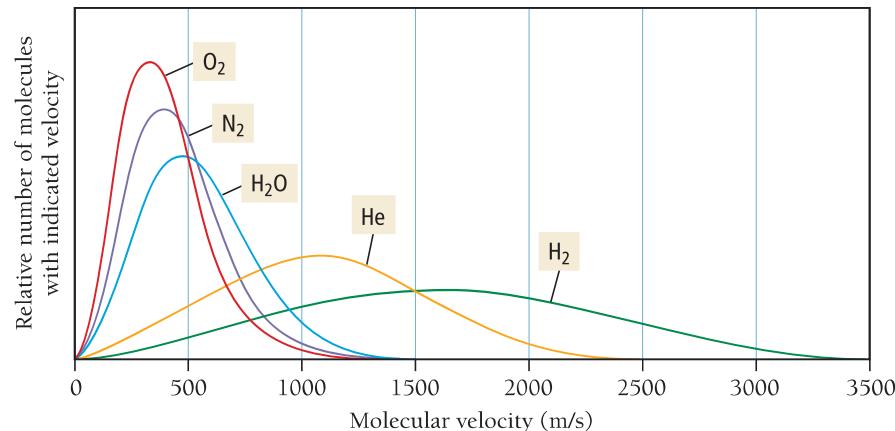


The root mean square velocity, as we have seen, is a kind of average velocity. Some particles move faster, and some move more slowly than this average. The velocities of all the particles in a gas sample form distributions like those shown in Figure 6.18▼. We can see from these distributions that some particles indeed travel at the root mean square velocity. However, many particles travel faster, and many travel more slowly than the root mean square velocity. For lighter particles, such as helium and hydrogen, the velocity distribution is shifted toward higher velocities and the curve becomes broader, indicating a wider range of velocities. Figure 6.19► illustrates the velocity distribution for nitrogen at different temperatures. As the temperature increases, the root mean square velocity increases and the distribution becomes broader.

### Variation of Velocity Distribution with Temperature



### Variation of Velocity Distribution with Molar Mass



**▲ FIGURE 6.18 Velocity Distribution for Several Gases at 25 °C** At a given temperature, there is a distribution of velocities among the particles in a sample of gas. The exact shape and peak of the distribution vary with the molar mass of the gas.

**▲ FIGURE 6.19 Velocity Distribution for Nitrogen at Several Temperatures** As the temperature of a gas sample increases, the velocity distribution of the molecules shifts toward higher velocity and becomes less sharply peaked.

## EXAMPLE 6.14 Root Mean Square Velocity

Calculate the root mean square velocity of oxygen molecules at 25 °C.

**SORT** You are given the kind of molecule and the temperature and asked to find the root mean square velocity.

**STRATEGIZE** The conceptual plan for this problem shows how you can use the molar mass of oxygen and the temperature (in kelvins) with the equation that defines the root mean square velocity to calculate root mean square velocity.

**SOLVE** First, gather the required quantities in the correct units. Note that molar mass must be in kg/mol.

Substitute the quantities into the equation to calculate root mean square velocity. Note that 1 J = 1 kg · m<sup>2</sup>/s<sup>2</sup>.

**GIVEN:** O<sub>2</sub>, t = 25 °C

**FIND:** u<sub>rms</sub>

### CONCEPTUAL PLAN



$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}}$$

### RELATIONSHIP USED

$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}} \quad (\text{Equation 6.26})$$

### SOLUTION

$$T = 25 + 273 = 298 \text{ K}$$

$$\mathcal{M} = \frac{32.00 \text{ g O}_2}{1 \text{ mol O}_2} \times \frac{1 \text{ kg}}{1000 \text{ g}} = \frac{32.00 \times 10^{-3} \text{ kg O}_2}{1 \text{ mol O}_2}$$

$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}}$$

$$= \sqrt{\frac{3 \left( 8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}} \right) (298 \text{ K})}{32.00 \times 10^{-3} \text{ kg O}_2}}$$

$$= \sqrt{2.32 \times 10^5 \frac{\text{J}}{\text{kg}}}$$

$$= \sqrt{2.32 \times 10^5 \frac{\frac{\text{kg} \cdot \text{m}^2}{\text{s}^2}}{\text{kg}}} = 482 \text{ m/s}$$

**CHECK** The units of the answer (m/s) are correct. The magnitude of the answer seems reasonable because oxygen is slightly heavier than nitrogen and should therefore have a slightly lower root mean square velocity at the same temperature. Recall that the root mean square velocity of nitrogen is 515 m/s at 25 °C.

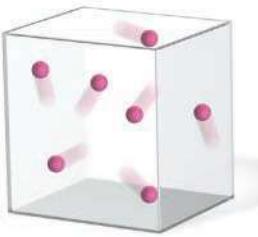
**FOR PRACTICE 6.14** Calculate the root mean square velocity of gaseous xenon atoms at 25 °C.

ANSWER NOW!

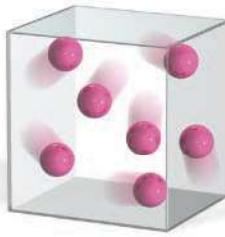


**6.8**  
**Cc**  
Conceptual Connection

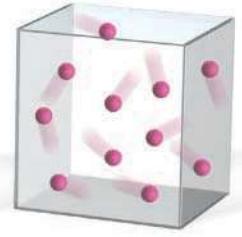
**KINETIC MOLECULAR THEORY** Which sample of an ideal gas has the greatest pressure? Assume that the mass of each particle is proportional to its size and that all the gas samples are at the same temperature.



(a)



(b)



(c)

## 6.9

# Mean Free Path, Diffusion, and Effusion of Gases

We have just seen that the root mean square velocity of gas molecules at room temperature is in the range of hundreds of meters per second. However, suppose that your roommate just put on too much perfume in the bathroom only 6 ft (1.8 m) away. Why does it take a minute or two before you can smell the fragrance? Although most molecules in a perfume bottle have higher molar masses than nitrogen, their velocities are still hundreds of meters per second, so why the delay? The answer is that even though gaseous particles travel at tremendous speeds, they also travel in haphazard paths (Figure 6.20►). To a perfume molecule, the path from the perfume bottle in the bathroom to your nose 6 ft away is much like a bargain hunter's path through a busy shopping mall during a clearance sale. The molecule travels only a short distance before it collides with another molecule and changes direction, only to collide again, and so on. In fact, at room temperature and atmospheric pressure, a molecule in the air experiences several billion collisions per second. The average distance that a molecule travels between collisions is its **mean free path**. At room temperature and atmospheric pressure, the mean free path of a nitrogen molecule with a molecular diameter of 300 pm (four times the covalent radius) is 93 nm, or about 310 molecular diameters. If the nitrogen molecule were the size of a golf ball, it would travel about 40 ft between collisions. Mean free path increases with *decreasing* pressure. Under conditions of ultrahigh vacuum ( $10^{-10}$  torr), the mean free path of a nitrogen molecule is hundreds of kilometers.

The process by which gas molecules spread out in response to a concentration gradient is **diffusion**, and even though the particles undergo many collisions, the root mean square velocity still influences the rate of diffusion. Heavier molecules diffuse more slowly than lighter ones, so the first molecules you smell from a perfume mixture (in a room with no air currents) are the lighter ones.

A process related to diffusion is **effusion**, the process by which a gas escapes from a container into a vacuum through a small hole (Figure 6.21►). The rate of effusion is also related to root mean square velocity—heavier molecules effuse more slowly than lighter ones. The rate of effusion—the amount of gas that effuses in a given time—is inversely proportional to the square root of the molar mass of the gas as follows:

$$\text{rate} \propto \frac{1}{\sqrt{\mathcal{M}}}$$

The ratio of effusion rates of two different gases is given by **Graham's law of effusion**, named after Thomas Graham (1805–1869):

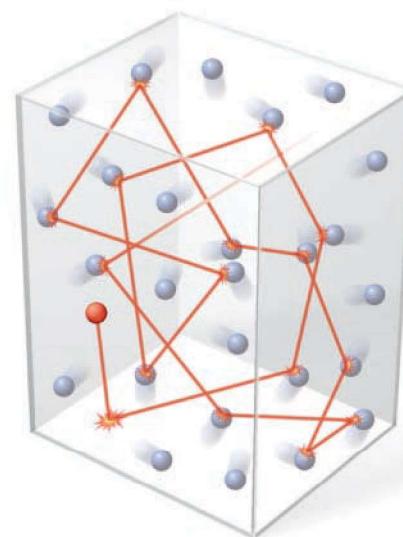
$$\frac{\text{rate}_A}{\text{rate}_B} = \sqrt{\frac{\mathcal{M}_B}{\mathcal{M}_A}} \quad [6.27]$$

In this expression,  $\text{rate}_A$  and  $\text{rate}_B$  are the effusion rates of gases A and B, and  $\mathcal{M}_A$  and  $\mathcal{M}_B$  are their molar masses.

Graham's law explains, in part, why helium balloons only float for a day or so. Because helium has such a low molar mass, it escapes from the balloon quite quickly. A balloon filled with air, by contrast, remains inflated longer.

## Typical Gas Molecule Path

The average distance between collisions is the mean free path.

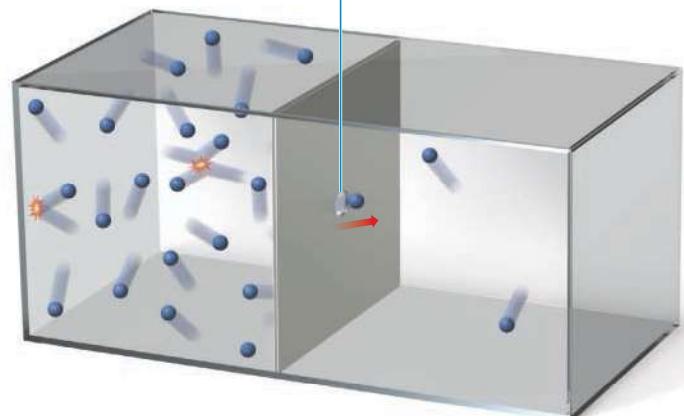


▲ FIGURE 6.20 Mean Free Path A molecule in a volume of gas follows a haphazard path, involving many collisions with other molecules.

In a ventilated room, air currents also enhance the transport of gas molecules.

## Effusion

Gas escapes from container into a vacuum through a small hole.



▲ FIGURE 6.21 Effusion

Effusion is the escape of a gas from a container into a vacuum through a small hole.

## WATCH NOW!

## INTERACTIVE WORKED EXAMPLE 6.15

**EXAMPLE 6.15** Graham's Law of Effusion

An unknown gas effuses at a rate that is 0.462 times that of nitrogen gas (at the same temperature). Calculate the molar mass of the unknown gas in g/mol.

**SORT** You are given the ratio of effusion rates for the unknown gas and nitrogen and asked to find the molar mass of the unknown gas.

**STRATEGIZE** The conceptual plan uses Graham's law of effusion. You are given the ratio of rates and you know the molar mass of the nitrogen. You can use Graham's law to determine the molar mass of the unknown gas.

$$\text{GIVEN: } \frac{\text{rate}_{\text{unk}}}{\text{rate}_{\text{N}_2}} = 0.462$$

$$\text{FIND: } M_{\text{unk}}$$

## CONCEPTUAL PLAN

$$\frac{\text{Rate}_{\text{unk}}}{\text{Rate}_{\text{N}_2}}, M_{\text{N}_2} \rightarrow M_{\text{unk}}$$

$$\frac{\text{Rate}_{\text{unk}}}{\text{Rate}_{\text{N}_2}} = \sqrt{\frac{M_{\text{N}_2}}{M_{\text{unk}}}}$$

## RELATIONSHIPS USED

$$\frac{\text{rate}_A}{\text{rate}_B} = \sqrt{\frac{M_B}{M_A}} \quad (\text{Graham's law})$$

**SOLVE** Solve the equation for  $M_{\text{unk}}$  and substitute the correct values and calculate.

## SOLUTION

$$\frac{\text{rate}_{\text{unk}}}{\text{rate}_{\text{N}_2}} = \sqrt{\frac{M_{\text{N}_2}}{M_{\text{unk}}}}$$

$$M_{\text{unk}} = \frac{M_{\text{N}_2}}{\left(\frac{\text{rate}_{\text{unk}}}{\text{rate}_{\text{N}_2}}\right)^2}$$

$$= \frac{28.02 \text{ g/mol}}{(0.462)^2}$$

$$= 131 \text{ g/mol}$$

**CHECK** The units of the answer are correct. The magnitude of the answer seems reasonable for the molar mass of a gas. From the answer you can even conclude that the gas is probably xenon, which has a molar mass of 131.29 g/mol.

**FOR PRACTICE 6.15** Find the ratio of effusion rates of hydrogen gas and krypton gas.

## ANSWER NOW!


**6.9**  
**CC**  
 Conceptual Connection

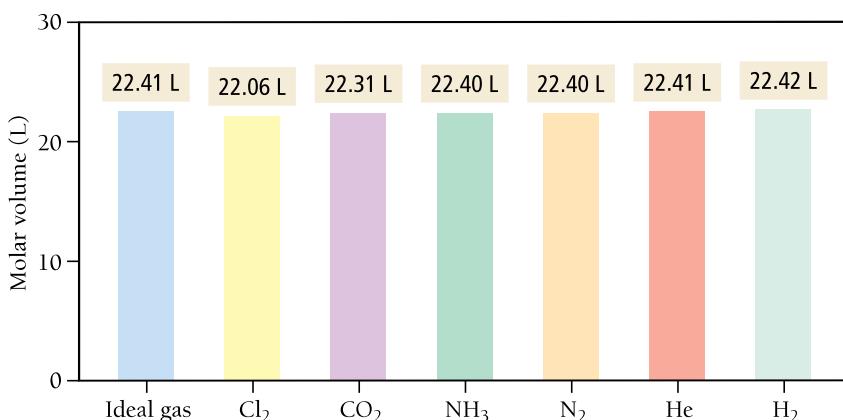
**GRAHAM'S LAW OF EFFUSION** A glass tube contains an equal number of moles of helium and argon. After five minutes, half of the particles escape through a small hole in the glass. What are the *relative* amounts of helium and argon in the tube at five minutes?

- (a) The tube still contains an equal number of moles of helium and argon.
- (b) The tube contains a greater number of moles of helium than argon.
- (c) The tube contains a greater number of moles of argon than helium.

**6.10****Real Gases: The Effects of Size and Intermolecular Forces**

One mole of an ideal gas has a volume of 22.41 L at STP. Figure 6.22 shows the molar volume of several real gases at STP. Note that most of these gases have a volume that is very close to 22.41 L, meaning that they act very nearly as ideal gases. Gases behave ideally when both of the following are true: (a) the volume of the gas particles is small compared to the space between them, and (b) the forces between the gas

## Molar Volume

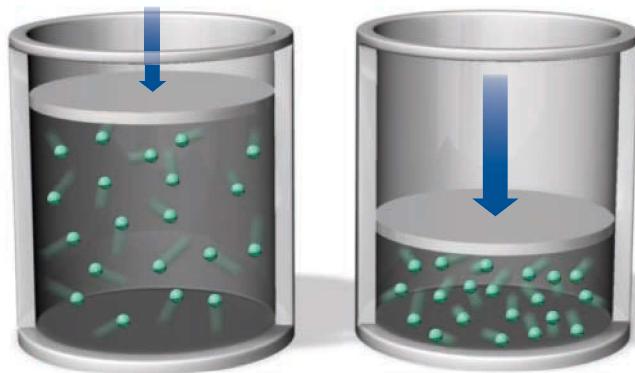


▲ FIGURE 6.22 Molar Volumes of Real Gases The molar volumes of several gases at STP are all close to 22.41 L, indicating that their departures from ideal behavior are small.

particles are not significant. At STP, these assumptions are valid for most common gases. However, these assumptions break down at higher pressures or lower temperatures.

## The Effect of the Finite Volume of Gas Particles

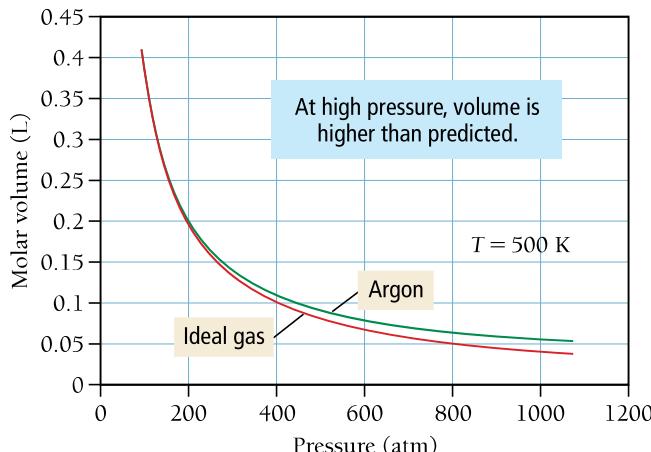
The finite volume of gas particles—that is, their actual *sizes*—becomes important at high pressure because the volume of the particles themselves occupies a significant portion of the total gas volume (Figure 6.23►). We can see the effect of particle volume by comparing the molar volume of argon to the molar volume of an ideal gas as a function of pressure at 500 K as shown in Figure 6.24▼. At low pressures, the molar volume of argon is nearly identical to that of an ideal gas. But as the pressure increases, the molar volume of argon becomes *greater than* that of an ideal gas. At the higher pressures, the argon atoms themselves occupy a significant portion of the gas volume, making the actual volume greater than that predicted by the ideal gas law.



▲ FIGURE 6.23 Particle Volume and Ideal Behavior As a gas is compressed, the gas particles themselves begin to occupy a significant portion of the total gas volume, leading to deviations from ideal behavior.

### Nonideal Behavior: The effect of particle volume

Particle size becomes significant at high pressure.



◀ FIGURE 6.24 The Effect of Particle Volume At high pressures, 1 mol of argon occupies a larger volume than 1 mol of an ideal gas because of the volume of the argon atoms themselves. (This example was chosen to minimize the effects of intermolecular forces, which are very small in argon at 500 K, thereby isolating the effect of particle volume.)

**TABLE 6.5 ■ Van der Waals Constants for Common Gases**

| Gas                  | $a$<br>( $\text{L}^2 \cdot \text{atm/mol}^2$ ) | $b$<br>( $\text{L/mol}$ ) |
|----------------------|--|---------------------------|
| He                   | 0.0342   | 0.02370                   |
| Ne                   | 0.211  | 0.0171                    |
| Ar                   | 1.35   | 0.0322                    |
| Kr                   | 2.32   | 0.0398                    |
| Xe                   | 4.19   | 0.0511                    |
| $\text{H}_2$         | 0.244  | 0.0266                    |
| $\text{N}_2$         | 1.39   | 0.0391                    |
| $\text{O}_2$         | 1.36   | 0.0318                    |
| $\text{Cl}_2$        | 6.49   | 0.0562                    |
| $\text{H}_2\text{O}$ | 5.46   | 0.0305                    |
| $\text{CH}_4$        | 2.25   | 0.0428                    |
| $\text{CO}_2$        | 3.59   | 0.0427                    |
| $\text{CCl}_4$       | 20.4   | 0.1383                    |

In 1873, Johannes van der Waals (1837–1923) modified the ideal gas equation to fit the behavior of real gases. From the graph for argon in Figure 6.24 we can see that the ideal gas law predicts a volume that is too small. Van der Waals suggested a small correction factor that accounts for the volume of the gas particles themselves:

$$\begin{array}{ll} \textbf{Ideal behavior} & V = \frac{nRT}{P} \\ \textbf{Corrected for volume of gas particles} & V = \frac{nRT}{P} + nb \end{array} \quad [6.28]$$

The correction adds the quantity  $nb$  to the volume, where  $n$  is the number of moles and  $b$  is a constant that depends on the gas (see Table 6.5). We can rearrange the corrected equation as follows:

$$(V - nb) = \frac{nRT}{P} \quad [6.29]$$

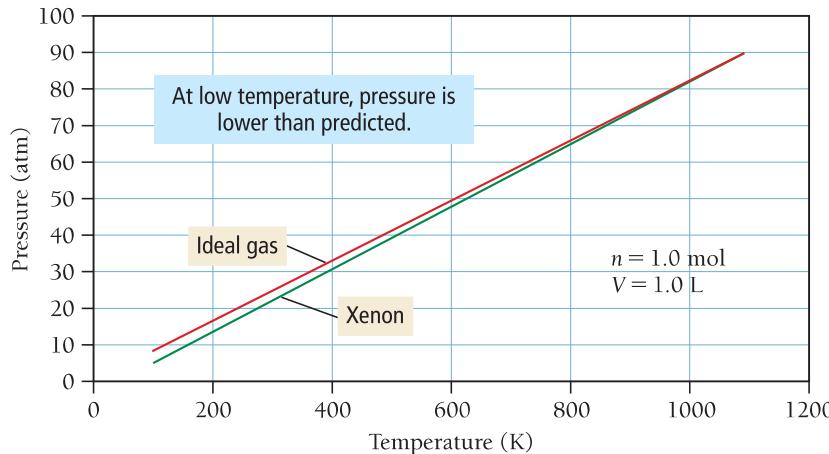
## The Effect of Intermolecular Forces

Intermolecular forces, which we will discuss in more detail in Chapter 12, are attractions between the atoms or molecules that compose any substance. These attractions are typically small in gases and therefore do not matter much at low pressure because the molecules are too far apart to “feel” the attractions. They also do not matter much at high temperatures because the molecules have a lot of kinetic energy. Therefore a weak attraction between them does not affect the collision much. At lower temperatures, however, the collisions occur with less kinetic energy, and weak attractions can affect the collisions. We can understand this difference with an analogy to billiard balls. Imagine two billiard balls coated with a substance that makes them slightly sticky. If they collide when moving at high velocities, the stickiness will not have much of an effect—the balls bounce off one another as if the sticky substance was not even there. However, if the two billiard balls collide when moving very slowly (say barely rolling), the sticky substance would have an effect. The billiard balls might even stick together and not bounce off one another.

The effect of these weak attractions between particles is a decrease in the number of collisions with the surfaces of the container and a corresponding decrease in the pressure compared to that of an ideal gas. We can see the effect of intermolecular forces when we compare the pressure of 1.0 mol of xenon gas to the pressure of 1.0 mol of an ideal gas as a function of temperature and at a fixed volume of 1.0 L, as shown in Figure 6.25▼. At high temperature, the pressure of the xenon gas is nearly identical to that of an ideal gas. But at lower temperatures, the pressure of xenon is *less than* that of an ideal gas. At the lower temperatures, the xenon atoms spend more time interacting with each other and less time colliding with the walls, making the actual pressure less than that predicted by the ideal gas law.

### Nonideal Behavior: The effect of intermolecular forces

Intermolecular forces become significant at low temperature.



► **FIGURE 6.25** The Effect of Intermolecular Forces At low temperatures, the pressure of xenon is less than an ideal gas exerts because interactions among xenon molecules reduce the number of collisions with the walls of the container.

From the graph for xenon shown in Figure 6.25 we can see that the ideal gas law predicts a pressure that is too large at low temperatures. Van der Waals suggested a small correction factor that accounts for the intermolecular forces between gas particles:

**Ideal behavior**

$$P = \frac{nRT}{V}$$

**Corrected for intermolecular forces**

$$P = \frac{nRT}{V} - a\left(\frac{n}{V}\right)^2 \quad [6.30]$$

The correction subtracts the quantity  $a(n/V)^2$  from the pressure, where  $n$  is the number of moles,  $V$  is the volume, and  $a$  is a constant that depends on the gas (see Table 6.5). Notice that the correction factor increases as  $n/V$  (the number of moles of particles per unit volume) increases because a greater concentration of particles makes it more likely that the particles will interact with one another. We can rearrange the corrected equation as:

$$P + a\left(\frac{n}{V}\right)^2 = \frac{nRT}{V} \quad [6.31]$$

## Van der Waals Equation

We can combine the effects of particle volume (Equation 6.29) and particle intermolecular forces (Equation 6.31) into one equation that describes nonideal gas behavior:

$$\left[P + a\left(\frac{n}{V}\right)^2\right] \times \left[V - nb\right] = nRT$$

↑                              ↑

Correction for  
intermolecular forces      Correction for  
particle volume

This equation is the **van der Waals equation**, and we can use it to calculate the properties of a gas under nonideal conditions.

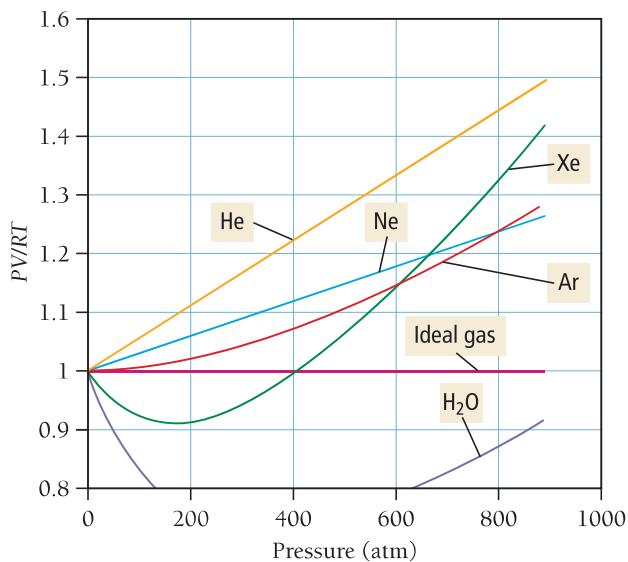
## Real Gases

We can see the combined effects of particle volume and intermolecular forces by examining a plot of  $PV/RT$  versus  $P$  for 1 mol of a number of real gases (Figure 6.26►). For an ideal gas,  $PV/RT = n$ , the number of moles of gas. Therefore, for 1 mol of an ideal gas,  $PV/RT$  is equal to 1, as shown in the plot. For real gases,  $PV/RT$  deviates from 1, but the deviations are not uniform. For example, water displays a large negative deviation from  $PV/RT$  because, for water, the effect of intermolecular forces on lowering the pressure (relative to an ideal gas) is far greater than the effect of particle size on increasing the volume. Notice from Table 6.5 that water has a high value of  $a$ , the constant that corrects for intermolecular forces, but a moderate value of  $b$ , the constant that corrects for particle size. Therefore,  $PV/RT$  for water is lower than predicted from the ideal gas law.

By contrast, consider the behavior of helium, which displays a positive deviation from the ideal behavior. This is because helium has very weak intermolecular forces and the effect of these forces on lowering the pressure (relative to ideal gas) is small compared to the effect of particle size on increasing the volume. Therefore  $PV/RT$  for helium is greater than predicted from the ideal gas law.

## The Behavior of Real Gases

Real gases deviate from ideal gas behavior.



▲ FIGURE 6.26 Real versus Ideal Behavior For 1 mol of an ideal gas,  $PV/RT$  is equal to 1. The combined effects of the volume of gas particles and the interactions among them cause each real gas to deviate from ideal behavior in a slightly different way. These curves were calculated at a temperature of 500 K.

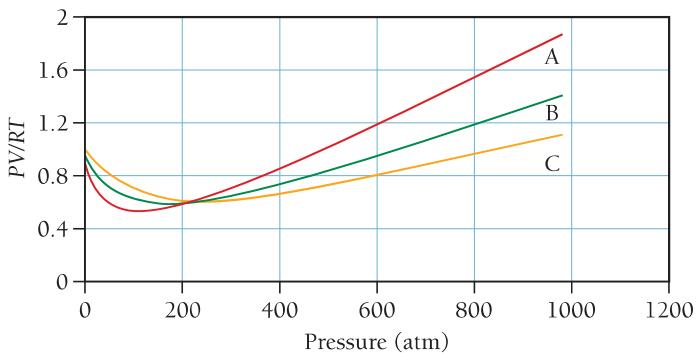
ANSWER NOW!



## 6.10 Cc

Conceptual Connection

**REAL GASES** The graph shows  $PV/RT$  for carbon dioxide at three different temperatures. Rank the curves in order of increasing temperature.



- (a) C < A < B    (b) A < B < C    (c) B < C < A    (d) C < B < A

QUIZ YOURSELF NOW!

## Self-Assessment Quiz

**Q1.** A gas sample has an initial pressure of 547 mmHg and an initial volume of 0.500 L. What is the pressure (in atm) when the volume of the sample is decreased to 225 mL? (Assume constant temperature and constant number of moles of gas.)

**MISSED THIS?** Read Section 6.3; Watch KCV 6.3

- a)  $1.60 \times 10^{-3}$  atm    b) 1.60 atm  
c) 0.324 atm    d) 1.22 atm

**Q2.** A gas sample has a volume of 178 mL at 0.00 °C. The temperature is raised (at constant pressure) until the volume is 211 mL. What is the temperature of the gas sample in °C at this volume? **MISSED THIS?** Read Section 6.3; Watch KCV 6.3

- a) 0.00 °C    b) 324 °C  
c) -43 °C    d) 51 °C

**Q3.** What is the pressure of 1.78 g of nitrogen gas confined to a volume of 0.118 L at 25 °C?

**MISSED THIS?** Read Section 6.4; Watch KCV 6.3, IWE 6.5

- a) 13.2 atm    b) 369 atm  
c) 1.10 atm    d) 26.3 atm

**Q4.** What is the density of a sample of argon gas at 55 °C and 765 mmHg? **MISSED THIS?** Read Section 6.5; Watch IWE 6.7

- a) 2.99 g/L    b)  $1.13 \times 10^3$  g/L  
c) 1.49 g/L    d) 8.91 g/L

**Q5.** Which gas sample has the greatest volume at STP?

**MISSED THIS?** Read Section 6.5

- a) 10.0 g Ar    b) 10.0 g Kr  
c) 10.0 g Xe    d) None of the above (They all have the same volume.)

**Q6.** A 1.25-g gas sample occupies 663 mL at 25 °C and 1.00 atm. What is the molar mass of the gas?

**MISSED THIS?** Read Section 6.5; Watch IWE 6.8

- a) 0.258 g/mol    b) 0.0461 g/mol  
c) 3.87 g/mol    d) 46.1 g/mol

**Q7.** A 255-mL gas sample contains argon and nitrogen at a temperature of 65 °C. The total pressure of the sample is 725 mmHg, and the partial pressure of argon is 231 mmHg. What mass of nitrogen is present in the sample? **MISSED THIS?** Read Section 6.6; Watch KCV 6.6

- a) 0.324 g nitrogen    b) 0.167 g nitrogen  
c) 0.0837 g nitrogen    d) 0.870 g nitrogen

**Q8.** A gas mixture in a 1.55-L container at 298 K contains 10.0 g of Ne and 10.0 g of Ar. Calculate the partial pressure (in atm) of Ne and Ar in the container.

**MISSED THIS?** Read Section 6.6; Watch KCV 6.6, IWE 6.10

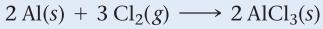
- a)  $P_{\text{Ne}} = 10.5$  atm,  $P_{\text{Ar}} = 5.29$  atm  
b)  $P_{\text{Ne}} = 5.83$  atm,  $P_{\text{Ar}} = 2.95$  atm  
c)  $P_{\text{Ne}} = 5.88$  atm,  $P_{\text{Ar}} = 5.88$  atm  
d)  $P_{\text{Ne}} = 7.82$  atm,  $P_{\text{Ar}} = 3.95$  atm

**Q9.** A gas sample at STP contains 1.15 g oxygen and 1.55 g nitrogen. What is the volume of the gas sample?

**MISSED THIS?** Read Sections 6.5, 6.6

- a) 1.26 L    b) 2.04 L  
c) 4.08 L    d) 61.0 L

**Q10.** Aluminum reacts with chlorine gas to form aluminum chloride.



What minimum volume of chlorine gas (at 298 K and 225 mmHg) is required to completely react with 7.85 g of aluminum?

**MISSED THIS?** Read Section 6.7; Watch IWE 6.12

- a) 36.0 L    b) 24.0 L    c) 0.0474 L    d) 16.0 L

**Q11.** Calculate the root mean square velocity of  $\text{I}_2(g)$  at 373 K.

**MISSED THIS?** Read Section 6.8; Watch KCV 6.8

- a) 19.0 m/s    b) 191 m/s  
c) 6.05 m/s    d) 99.1 m/s

**Q12.** Which gas has the greatest kinetic energy at STP?

**MISSED THIS? Read Section 6.8; Watch KCV 6.8**

- a) He
- b) Ne
- c) Ar
- d) None of the above (All have the same kinetic energy.)

**Q13.** A sample of Xe takes 75 seconds to effuse out of a container. An unknown gas takes 37 seconds to effuse out of the identical container under identical conditions. What is the most likely identity of the unknown gas?

**MISSED THIS? Read Section 6.9; Watch IWE 6.15**

- a) He
- b) O<sub>2</sub>
- c) Br<sub>2</sub>
- d) Kr

**Q14.** Consider the generic reaction: 2 A(g) + B(g) → 2 C(g). If a flask initially contains 1.0 atm of A and 1.0 atm of B, what will be the pressure in the flask if the reaction proceeds to completion? (Assume constant volume and temperature.)

**MISSED THIS? Read Section 6.7**

- a) 1.0 atm
- b) 1.5 atm
- c) 2.0 atm
- d) 3.0 atm

**Q15.** Rank the gases Ar, N<sub>2</sub>, CH<sub>4</sub>, and C<sub>2</sub>H<sub>6</sub> in order of increasing density at STP. **MISSED THIS? Read Section 6.5; Watch IWE 6.7**

- a) CH<sub>4</sub> < C<sub>2</sub>H<sub>6</sub> < N<sub>2</sub> < Ar
- b) CH<sub>4</sub> < N<sub>2</sub> < Ar < C<sub>2</sub>H<sub>6</sub>
- c) Ar < C<sub>2</sub>H<sub>6</sub> < N<sub>2</sub> < CH<sub>4</sub>
- d) CH<sub>4</sub> < N<sub>2</sub> < C<sub>2</sub>H<sub>6</sub> < Ar

**Answers:** 1. (b) 2. (d) 3. (a) 4. (c) 5. (a) 6. (d) 7. (b) 8. (d) 9. (b) 10. (a) 11. (b) 12. (d) 13. (b) 14. (b) 15. (d)

## CHAPTER 6 IN REVIEW

### TERMS

#### Section 6.1

pressure (212)

#### Section 6.2

millimeter of mercury (mmHg) (213)  
barometer (213)  
torr (213)  
atmosphere (atm) (213)  
pascal (Pa) (213)  
manometer (214)

#### Section 6.3

Boyle's law (216)  
Charles's law (219)  
Avogadro's law (221)

#### Section 6.4

ideal gas law (223)  
ideal gas (223)  
ideal gas constant (223)

#### Section 6.5

molar volume (225)

standard temperature and pressure (STP) (225)

#### Section 6.6

partial pressure (229)  
Dalton's law of partial pressures (229)  
mole fraction ( $\chi_a$ ) (229)  
hypoxia (231)  
oxygen toxicity (231)  
nitrogen narcosis (231)  
vapor pressure (233)

#### Section 6.8

kinetic molecular theory (238)

#### Section 6.9

mean free path (245)  
diffusion (245)  
effusion (245)  
Graham's law of effusion (245)

#### Section 6.10

van der Waals equation (249)

### CONCEPTS

#### Pressure (6.1, 6.2)

- Gas pressure is the force per unit area that results from gas particles colliding with the surfaces around them. We use various units to measure pressure, including mmHg, torr, Pa, psi, in Hg, and atm.

#### The Simple Gas Laws (6.3)

- The simple gas laws express relationships between pairs of variables when other variables are constant.
- Boyle's law states that the volume of a gas is inversely proportional to its pressure.
- Charles's law states that the volume of a gas is directly proportional to its temperature.
- Avogadro's law states that the volume of a gas is directly proportional to the amount (in moles).

#### The Ideal Gas Law and Its Applications (6.4, 6.5)

- The ideal gas law,  $PV = nRT$ , describes the relationship among all four gas variables and contains the simple gas laws within it.
- We can use the ideal gas law to find one of the four variables given the other three. We can use it to calculate the molar volume of an ideal gas, which is 22.4 L at STP, and to calculate the density and molar mass of a gas.

#### Mixtures of Gases and Partial Pressures (6.6)

- In a mixture of gases, each gas acts independently of the others so that any overall property of the mixture is the sum of the properties of the individual components.
- The pressure of any individual component is its partial pressure.

#### Gas Stoichiometry (6.7)

- In reactions involving gaseous reactants and products, we often report quantities in volumes at specified pressures and temperatures. We can convert these quantities to amounts (in moles) using the ideal gas law. Then we can use the stoichiometric coefficients from the balanced equation to determine the stoichiometric amounts of other reactants or products.
- The general form for these types of calculations is: volume A → amount A (in moles) → amount B (in moles) → quantity of B (in desired units).
- In cases where the reaction is carried out at STP, we can use the molar volume at STP (22.4 L = 1 mol) to convert between volume in liters and amount in moles.

## Kinetic Molecular Theory and Its Applications (6.8, 6.9)

- Kinetic molecular theory is a quantitative model for gases. The theory has three main assumptions: (1) gas particles are negligibly small, (2) the average kinetic energy of a gas particle is proportional to the temperature in kelvins, and (3) the collision of one gas particle with another is completely elastic (the particles do not stick together). The gas laws all follow from the kinetic molecular theory.
- We can use kinetic molecular theory to derive the expression for the root mean square velocity of gas particles. This velocity is inversely proportional to the molar mass of the gas, and therefore—at a given temperature—smaller gas particles are (on average) moving more quickly than larger ones.

- The kinetic molecular theory also allows us to predict the mean free path of a gas particle (the distance it travels between collisions) and relative rates of diffusion or effusion.

## Real Gases (6.10)

- Real gases differ from ideal gases to the extent that they do not always fit the assumptions of kinetic molecular theory.
- These assumptions tend to break down at high pressures where the volume is higher than predicted for an ideal gas because the particles are no longer negligibly small compared to the space between them.
- The assumptions also break down at low temperatures where the pressure is lower than predicted because the attraction between molecules combined with low kinetic energies causes partially inelastic collisions.
- The van der Waals equation predicts gas properties under nonideal conditions.

## EQUATIONS AND RELATIONSHIPS

Relationship between Pressure ( $P$ ), Force ( $F$ ), and Area ( $A$ ) (6.2)

$$P = \frac{F}{A}$$

Boyle's Law: Relationship between Pressure ( $P$ ) and Volume ( $V$ ) (6.3)

$$V \propto \frac{1}{P}$$

$$P_1 V_1 = P_2 V_2$$

Charles's Law: Relationship between Volume ( $V$ ) and Temperature ( $T$ ) (6.3)

$$V \propto T \quad (\text{in K})$$

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

Avogadro's Law: Relationship between Volume ( $V$ ) and Amount in Moles ( $n$ ) (6.3)

$$V \propto n$$

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

Ideal Gas Law: Relationship between Volume ( $V$ ), Pressure ( $P$ ), Temperature ( $T$ ), and Amount ( $n$ ) (6.4)

$$PV = nRT$$

Dalton's Law: Relationship between Partial Pressures ( $P_n$ ) in Mixture of Gases and Total Pressure ( $P_{\text{total}}$ ) (6.6)

$$P_{\text{total}} = P_a + P_b + P_c + \dots$$

$$P_a = \frac{n_a RT}{V} \quad P_b = \frac{n_b RT}{V} \quad P_c = \frac{n_c RT}{V}$$

Mole Fraction ( $\chi_a$ ) (6.6)

$$\chi_a = \frac{n_a}{n_{\text{total}}}$$

$$P_a = \chi_a P_{\text{total}}$$

Average Kinetic Energy ( $KE_{\text{avg}}$ ) (6.8)

$$KE_{\text{avg}} = \frac{3}{2}RT$$

Relationship between Root Mean Square Velocity ( $u_{\text{rms}}$ ) and Temperature ( $T$ ) (6.8)

$$u_{\text{rms}} = \sqrt{\frac{3RT}{M}}$$

Relationship of Effusion Rates of Two Different Gases (6.9)

$$\frac{\text{rate A}}{\text{rate B}} = \sqrt{\frac{M_B}{M_A}}$$

Van der Waals Equation: The Effects of Volume and Intermolecular Forces on Nonideal Gas Behavior (6.10)

$$\left[ P + a\left(\frac{n}{V}\right)^2 \right] \times (V - nb) = nRT$$

## LEARNING OUTCOMES

| Chapter Objectives  | Assessment   |
|---|--|
| Convert between units of pressure (6.2)                       | Example 6.1 For Practice 6.1 For More Practice 6.1 Exercises 25–30                                     |
| Calculate properties of gases using the simple gas laws (6.3) | Examples 6.2, 6.3, 6.4 For Practice 6.2, 6.3, 6.4 Exercises 31–36                                      |
| Calculate properties of gases using the ideal gas law (6.4)   | Examples 6.5, 6.6, 6.7, 6.8 For Practice 6.5, 6.6, 6.7, 6.8 For More Practice 6.6, 6.7 Exercises 37–60 |

|  |  |
|--|--|
| Analyze gas mixtures using Dalton's law of partial pressures (6.5)                               | Examples 6.9, 6.10, 6.11 For Practice 6.9, 6.10, 6.11 Exercises 61–70              |
| Perform stoichiometric calculations involving gas reactions (6.6)                                | Examples 6.12, 6.13 For Practice 6.12, 6.13 For More Practice 6.12 Exercises 71–80 |
| Calculate the root mean square velocity of a gas (6.8)   | Example 6.14 For Practice 6.14 Exercises 81–84                                     |
| Calculate the effusion rate of a gas or the ratio of effusion rates of two different gases (6.9) | Example 6.15 For Practice 6.15 Exercises 85–89                                     |
| Calculate gas properties of real gases (6.10)  | Exercises 91–94  |

## EXERCISES

**Mastering Chemistry** provides end-of-chapter exercises, feedback-enriched tutorial problems, animations, and interactive activities to encourage problem-solving practice and deeper understanding of key concepts and topics.

## REVIEW QUESTIONS

- What is pressure? What causes pressure?
- Explain what happens when a person inhales. What forces air into the lungs?
- Explain what happens when a person exhales. What forces air out of the lungs?
- What are the common units of pressure? List them in order of smallest to largest unit.
- What is a manometer? How does it measure the pressure of a sample of gas?
- Summarize each of the simple gas laws (Boyle's law, Charles's law, and Avogadro's law). For each, explain the relationship between the two variables and also state which variables must be kept constant.
- Explain why people may experience ear pain after a rapid change in altitude.
- Explain why scuba divers should never hold their breath when they ascend to the surface.
- Why is it impossible to breathe air through an extra-long snorkel (longer than a couple of meters) while swimming under water?
- Explain why hot-air balloons float above the ground and why the second story of a two-story home is often warmer than the ground story.
- What is the ideal gas law? Why is it useful?
- Explain how the ideal gas law contains within it the simple gas laws (show an example).
- Define molar volume and list its value for a gas at STP.
- How does the density of a gas depend on temperature? Pressure? How does it depend on the molar mass of the gas?
- What is partial pressure? What is the relationship between the partial pressures of each gas in a sample and the total pressure of gas in the sample?
- Why do deep-sea divers breathe a mixture of helium and oxygen?
- When a gas is collected over water, is the gas pure? Why or why not? How can the partial pressure of the collected gas be determined?
- If a reaction occurs in the gas phase at STP, we can determine the mass of a product from the volumes of reactants. Explain.
- What are the basic postulates of kinetic molecular theory? How does the concept of pressure follow from kinetic molecular theory?
- Explain how Boyle's law, Charles's law, Avogadro's law, and Dalton's law all follow from kinetic molecular theory.
- How is the kinetic energy of a gas related to temperature? How is the root mean square velocity of a gas related to its molar mass?
- Describe how the molecules in a perfume bottle travel from the bottle to your nose. What is mean free path?
- Explain the difference between diffusion and effusion. How is the effusion rate of a gas related to its molar mass?
- Deviations from the ideal gas law are observed at high pressure and low temperature. Explain this in light of kinetic molecular theory.

## PROBLEMS BY TOPIC

### Converting between Pressure Units

25. The pressure in Denver, Colorado (elevation 5280 ft), averages about 24.9 in Hg. Convert this pressure to each indicated unit.

**MISSED THIS? Read Section 6.2**

a. atm      b. mmHg      c. psi      d. Pa

26. The pressure on top of Mount Everest (29,029 ft) averages about 235 mmHg. Convert this pressure to each indicated unit.
- torr
  - psi
  - in Hg
  - atm

27. The North American record for highest recorded barometric pressure is 31.85 in Hg, set in 1989 in Northway, Alaska. Convert this pressure to each indicated unit.

**MISSED THIS?** Read Section 6.2

a. mmHg

b. atm

c. torr

d. kPa (kilopascals)

28. The world record for lowest pressure (at sea level) was 652.5 mmHg recorded inside Typhoon Tip on October 12, 1979, in the western Pacific Ocean. Convert this pressure to each indicated unit.

a. torr

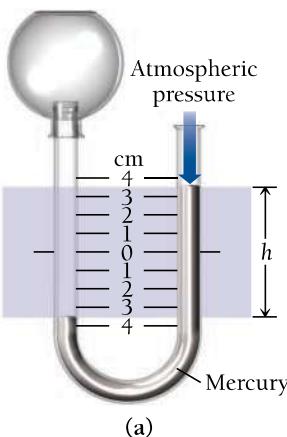
b. atm

c. in Hg

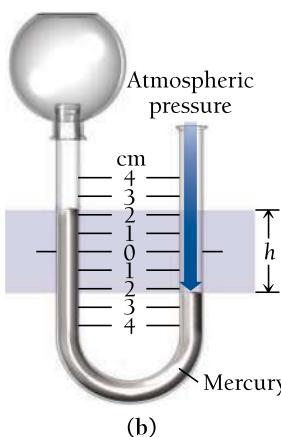
d. psi

29. Given a barometric pressure of 762.4 mmHg, calculate the pressure of each gas sample as indicated by the manometer.

**MISSED THIS?** Read Section 6.2

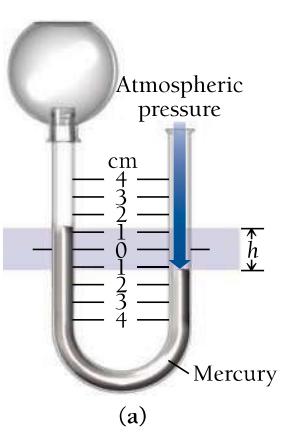


(a)

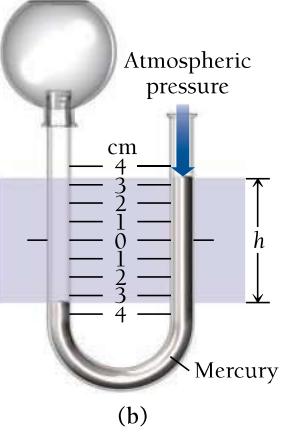


(b)

30. Given a barometric pressure of 751.5 mmHg, calculate the pressure of each gas sample as indicated by the manometer.



(a)



(b)

### Simple Gas Laws

31. A sample of gas has an initial volume of 5.6 L at a pressure of 735 mmHg. If the volume of the gas is increased to 9.4 L, what is its pressure?

**MISSED THIS?** Read Section 6.3; Watch KCV 6.3

32. A sample of gas has an initial volume of 13.9 L at a pressure of 1.22 atm. If the sample is compressed to a volume of 10.3 L, what is its pressure?

33. A 48.3-mL sample of gas in a cylinder is warmed from 22 °C to 87 °C. What is its volume at the final temperature?

**MISSED THIS?** Read Section 6.3; Watch KCV 6.3

34. A syringe containing 1.55 mL of oxygen gas is cooled from 95.3 °C to 0.0 °C. What is the final volume of oxygen gas?

35. A balloon contains 0.158 mol of gas and has a volume of 2.46 L. If an additional 0.113 mol of gas is added to the balloon (at the same temperature and pressure), what is its final volume?

**MISSED THIS?** Read Section 6.3; Watch KCV 6.3

36. A cylinder with a moveable piston contains 0.553 mol of gas and has a volume of 253 mL. What is its volume if an additional 0.365 mol of gas is added to the cylinder? (Assume constant temperature and pressure.)

### Ideal Gas Law

37. What volume is occupied by 0.118 mol of helium gas at a pressure of 0.97 atm and a temperature of 305 K? Would the volume be different if the gas was argon (under the same conditions)?

**MISSED THIS?** Read Section 6.4; Watch KCV 6.3, IWE 6.5

38. What volume is occupied by 12.5 g of argon gas at a pressure of 1.05 atm and a temperature of 322 K? Would the volume be different if the sample were 12.5 g of helium (under identical conditions)?

39. What is the pressure in a 10.0-L cylinder filled with 0.448 mol of nitrogen gas at a temperature of 315 K?

**MISSED THIS?** Read Section 6.4; Watch KCV 6.3, IWE 6.5

40. What is the pressure in a 15.0-L cylinder filled with 32.7 g of oxygen gas at a temperature of 302 K?

41. A cylinder contains 28.5 L of oxygen gas at a pressure of 1.8 atm and a temperature of 298 K. How much gas (in moles) is in the cylinder?

**MISSED THIS?** Read Section 6.4; Watch KCV 6.3, IWE 6.5

42. What is the temperature of 0.52 mol of gas at a pressure of 1.3 atm and a volume of 11.8 L?

43. An automobile tire has a maximum rating of 38.0 psi (gauge pressure). The tire is inflated (while cold) to a volume of 11.8 L and a gauge pressure of 36.0 psi at a temperature of 12.0 °C. On a hot day, the tire warms to 65.0 °C, and its volume expands to 12.2 L. Does the pressure in the tire exceed its maximum rating? (Note: The *gauge pressure* is the *difference* between the total pressure and atmospheric pressure. In this case, assume that atmospheric pressure is 14.7 psi.)

**MISSED THIS?** Read Section 6.4; Watch KCV 6.3, IWE 6.5

44. A weather balloon is inflated to a volume of 28.5 L at a pressure of 748 mmHg and a temperature of 28.0 °C. The balloon rises in the atmosphere to an altitude of approximately 25,000 ft, where the pressure is 385 mmHg and the temperature is -15.0 °C. Assuming the balloon can freely expand, calculate the volume of the balloon at this altitude.

45. A piece of dry ice (solid carbon dioxide) with a mass of 28.8 g sublimes (converts from solid to gas) into a large balloon. Assuming that all of the carbon dioxide ends up in the balloon, what is the volume of the balloon at 22 °C and a pressure of 742 mmHg?

**MISSED THIS?** Read Section 6.4; Watch KCV 6.3, IWE 6.5

46. A 1.0-L container of liquid nitrogen is kept in a closet measuring 1.0 m by 1.0 m by 2.0 m. Assuming that the container is completely full, that the temperature is 25.0 °C, and that the atmospheric pressure is 1.0 atm, calculate the percent (by volume) of air that is displaced if all of the liquid nitrogen evaporates. (Liquid nitrogen has a density of 0.807 g/mL.)

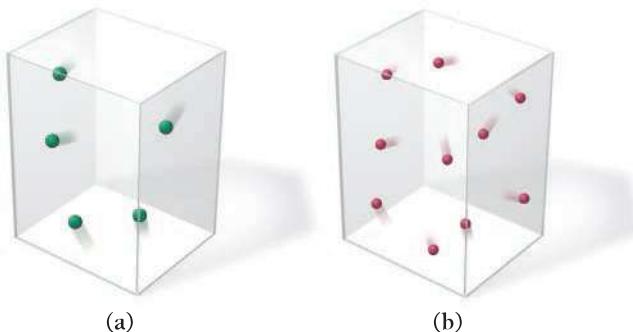
47. A wine-dispensing system uses argon canisters to pressurize and preserve wine in the bottle. An argon canister for the system has a volume of 55.0 mL and contains 26.0 g of argon. Assuming ideal gas behavior, what is the pressure in the canister at 295 K? When the argon is released from the canister, it expands to fill the wine bottle. How many 750.0-mL wine bottles can be purged with the argon in the canister at a pressure of 1.20 atm and a temperature of 295 K?

**MISSED THIS?** Read Section 6.4; Watch KCV 6.3, IWE 6.5

48. Cyclists sometimes use pressurized carbon dioxide inflators to inflate a bicycle tire in the event of a flat. These inflators use metal cartridges that contain 16.0 g of carbon dioxide. At 298 K, to what pressure (in psi) can the carbon dioxide in the cartridge inflate a 3.45-L mountain bike tire? (Note: The *gauge pressure* is the difference between the total pressure and atmospheric pressure. In this case, assume that atmospheric pressure is 14.7 psi.)

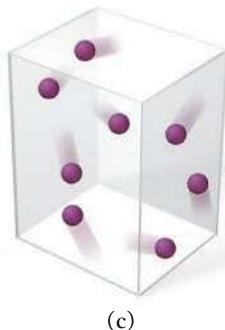
49. Which gas sample has the greatest pressure? Assume that all the samples are at the same temperature. Explain.

**MISSED THIS?** Read Section 6.4; Watch KCV 6.3



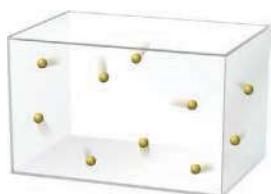
(a)

(b)



(c)

50. This picture represents a sample of gas at a pressure of 1 atm, a volume of 1 L, and a temperature of 25 °C. Draw a similar picture showing what would happen to the sample if the volume were reduced to 0.5 L and the temperature were increased to 250 °C. What would happen to the pressure?



51. Aerosol cans carry clear warnings against incineration because of the high pressures that can develop upon heating. Suppose that a can contains a residual amount of gas at a pressure of 755 mmHg and a temperature of 25 °C. What would the pressure be if the can were heated to 1155 °C? **MISSED THIS?** Read Section 6.4

52. A sample of nitrogen gas in a 1.75-L container exerts a pressure of 1.35 atm at 25 °C. What is the pressure if the volume of the container is maintained constant and the temperature is raised to 355 °C?

### Molar Volume, Density, and Molar Mass of a Gas

53. Use the molar volume of a gas at STP to determine the volume (in L) occupied by 33.6 g of neon at STP.

**MISSED THIS?** Read Section 6.5

54. Use the molar volume of a gas at STP to calculate the density (in g/L) of nitrogen gas at STP.

55. What is the density (in g/L) of hydrogen gas at 20.0 °C and a pressure of 1655 psi?

**MISSED THIS?** Read Section 6.5; Watch IWE 6.7

56. A sample of N<sub>2</sub>O gas has a density of 2.85 g/L at 298 K. What is the pressure of the gas (in mmHg)?

57. A 248-mL gas sample has a mass of 0.433 g at a pressure of 745 mmHg and a temperature of 28 °C. What is the molar mass of the gas? **MISSED THIS?** Read Section 6.5; Watch IWE 6.8

58. A 113-mL gas sample has a mass of 0.171 g at a pressure of 721 mmHg and a temperature of 32 °C. What is the molar mass of the gas?

59. A sample of gas has a mass of 38.8 mg. Its volume is 224 mL at a temperature of 55 °C and a pressure of 886 torr. Find the molar mass of the gas. **MISSED THIS?** Read Section 6.5; Watch IWE 6.8

60. A sample of gas has a mass of 0.555 g. Its volume is 117 mL at a temperature of 85 °C and a pressure of 753 mmHg. Find the molar mass of the gas.

### Partial Pressure

61. A gas mixture contains each of the following gases at the indicated partial pressures: N<sub>2</sub>, 215 torr; O<sub>2</sub>, 102 torr; and He, 117 torr. What is the total pressure of the mixture? What mass of each gas is present in a 1.35-L sample of this mixture at 25.0 °C? **MISSED THIS?** Read Section 6.6; Watch KCV 6.6

62. A gas mixture with a total pressure of 745 mmHg contains each of the following gases at the indicated partial pressures: CO<sub>2</sub>, 125 mmHg; Ar, 214 mmHg; and O<sub>2</sub>, 187 mmHg. The mixture also contains helium gas. What is the partial pressure of the helium gas? What mass of helium gas is present in a 12.0-L sample of this mixture at 273 K?

63. A 1.20-g sample of dry ice is added to a 755 mL flask containing nitrogen gas at a temperature of 25.0 °C and a pressure of 725 mmHg. The dry ice sublimes (converts from solid to gas), and the mixture returns to 25.0 °C. What is the total pressure in the flask? **MISSED THIS?** Read Section 6.6; Watch KCV 6.6

64. A 275-mL flask contains pure helium at a pressure of 752 torr. A second flask with a volume of 475 mL contains pure argon at a pressure of 722 torr. If we connect the two flasks through a stopcock and we open the stopcock, what is the partial pressure of each gas and the total pressure?

65. A gas mixture contains 1.25 g N<sub>2</sub> and 0.85 g O<sub>2</sub> in a 1.55 L container at 18 °C. Calculate the mole fraction and partial pressure of each component in the gas mixture.

**MISSED THIS?** Read Section 6.6; Watch KCV 6.6, IWE 6.10

66. What is the mole fraction of oxygen gas in air (see Table 6.3)? What volume of air contains 10.0 g of oxygen gas at 273 K and 1.00 atm?

67. The hydrogen gas formed in a chemical reaction is collected over water at 30.0 °C at a total pressure of 732 mmHg. What is the partial pressure of the hydrogen gas collected in this way? If the total volume of gas collected is 722 mL, what mass of hydrogen gas is collected?

**MISSED THIS?** Read Section 6.6; Watch KCV 6.6, IWE 6.11

68. The air in a bicycle tire is bubbled through water and collected at 25 °C. If the total volume of gas collected is 5.45 L at a temperature of 25 °C and a pressure of 745 torr, how many moles of gas were in the bicycle tire?

69. The zinc in a copper-plated penny will dissolve in hydrochloric acid if the copper coating is filed down in several spots (so that the hydrochloric acid can get to the zinc). The reaction between the acid and the zinc is  $2\text{H}^+(\text{aq}) + \text{Zn}(\text{s}) \longrightarrow \text{H}_2(\text{g}) + \text{Zn}^{2+}(\text{aq})$ . When the zinc in a certain penny dissolves, the total volume of gas collected over water at 25 °C is 0.951 L at a total pressure of 748 mmHg. What mass of hydrogen gas is collected?

**MISSED THIS?** Read Section 6.6; Watch KCV 6.6, IWE 6.11

70. A heliox deep-sea diving mixture contains 2.0 g of oxygen to every 98.0 g of helium. What is the partial pressure of oxygen when this mixture is delivered at a total pressure of 8.5 atm?

### Reaction Stoichiometry Involving Gases

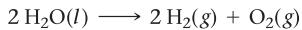
71. Consider the chemical reaction:



How many liters of hydrogen gas are formed from the complete reaction of 15.7 g C? Assume that the hydrogen gas is collected at a pressure of 1.0 atm and a temperature of 355 K.

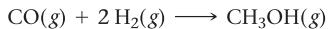
**MISSED THIS?** Read Section 6.7; Watch IWE 6.12

72. Consider the chemical reaction:



What mass of H<sub>2</sub>O is required to form 1.4 L of O<sub>2</sub> at a temperature of 315 K and a pressure of 0.957 atm?

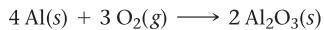
73. CH<sub>3</sub>OH can be synthesized by the reaction:



What volume of H<sub>2</sub> gas (in L), at 748 mmHg and 86 °C, is required to synthesize 25.8 g CH<sub>3</sub>OH? How many liters of CO gas, measured under the same conditions, are required?

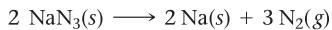
**MISSED THIS?** Read Section 6.7; Watch IWE 6.12

74. Oxygen gas reacts with powdered aluminum according to the reaction:



What volume of O<sub>2</sub> gas (in L), measured at 782 mmHg and 25 °C, completely reacts with 53.2 g Al?

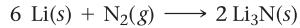
75. Automobile air bags inflate following a serious impact. The impact triggers the chemical reaction:



If an automobile airbag has a volume of 11.8 L, what mass of NaN<sub>3</sub> (in g) is required to fully inflate the airbag upon impact? Assume STP conditions.

**MISSED THIS?** Read Section 6.7; Watch IWE 6.12

76. Lithium reacts with nitrogen gas according to the reaction:



What mass of lithium (in g) reacts completely with 58.5 mL of N<sub>2</sub> gas at STP?

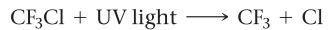
77. Hydrogen gas (a potential future fuel) can be formed by the reaction of methane with water according to the equation:



In a particular reaction, 25.5 L of methane gas (measured at a pressure of 732 torr and a temperature of 25 °C) mixes with 22.8 L of water vapor (measured at a pressure of 702 torr and a temperature of 125 °C). The reaction produces 26.2 L of hydrogen gas at STP. What is the percent yield of the reaction?

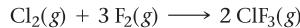
**MISSED THIS?** Read Section 6.7; Watch IWE 6.12

78. Ozone is depleted in the stratosphere by chlorine from CF<sub>3</sub>Cl according to this set of equations:



What total volume of ozone at a pressure of 25.0 mmHg and a temperature of 225 K is destroyed when all of the chlorine from 15.0 g of CF<sub>3</sub>Cl goes through 10 cycles of the given reactions?

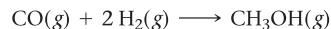
79. Chlorine gas reacts with fluorine gas to form chlorine trifluoride.



A 2.00-L reaction vessel, initially at 298 K, contains chlorine gas at a partial pressure of 337 mmHg and fluorine gas at a partial pressure of 729 mmHg. Identify the limiting reactant and determine the theoretical yield of ClF<sub>3</sub> in grams.

**MISSED THIS?** Read Section 6.7; Watch IWE 6.12

80. Carbon monoxide gas reacts with hydrogen gas to form methanol.



A 1.50-L reaction vessel, initially at 305 K, contains carbon monoxide gas at a partial pressure of 232 mmHg and hydrogen gas at a partial pressure of 397 mmHg. Identify the limiting reactant and determine the theoretical yield of methanol in grams.

### Kinetic Molecular Theory

81. Consider a 1.0-L sample of helium gas and a 1.0-L sample of argon gas, both at room temperature and atmospheric pressure.

- Do the atoms in the helium sample have the same *average kinetic energy* as the atoms in the argon sample?
- Do the atoms in the helium sample have the same *average velocity* as the atoms in the argon sample?
- Do the argon atoms, because they are more massive, exert a greater pressure on the walls of the container? Explain.
- Which gas sample has the faster rate of effusion?

**MISSED THIS?** Read Section 6.8; Watch KCV 6.8

82. A flask at room temperature contains exactly equal amounts (in moles) of nitrogen and xenon.

- Which of the two gases exerts the greater partial pressure?
- The molecules or atoms of which gas have the greater average velocity?
- The molecules or atoms of which gas have the greater average kinetic energy?
- If a small hole were opened in the flask, which gas effuses more quickly?

83. Calculate the root mean square velocity and kinetic energy of F<sub>2</sub>, Cl<sub>2</sub>, and Br<sub>2</sub> at 298 K. Rank these three halogens with respect to their rate of effusion.

**MISSED THIS?** Read Section 6.8; Watch KCV 6.8

84. Calculate the root mean square velocity and kinetic energy of CO, CO<sub>2</sub>, and SO<sub>3</sub> at 298 K. Which gas has the greatest velocity? The greatest kinetic energy? The greatest effusion rate?
85. We separate U-235 from U-238 by fluorinating a sample of uranium to form UF<sub>6</sub> (which is a gas) and then taking advantage of the different rates of effusion and diffusion for compounds containing the two isotopes. Calculate the ratio of effusion rates for <sup>238</sup>UF<sub>6</sub> and <sup>235</sup>UF<sub>6</sub>. The atomic mass of U-235 is 235.054 amu and that of U-238 is 238.051 amu.

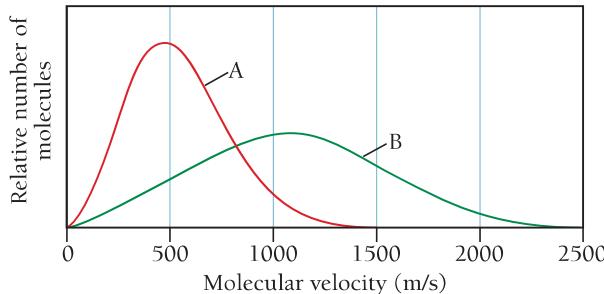
**MISSED THIS?** Read Section 6.9; Watch IWE 6.15

86. Calculate the ratio of effusion rates for Ar and Kr.
87. A sample of neon effuses from a container in 76 seconds. The same amount of an unknown noble gas requires 155 seconds. Identify the second gas.

**MISSED THIS?** Read Section 6.9; Watch IWE 6.15

88. A sample of N<sub>2</sub>O effuses from a container in 42 seconds. How long will it take the same amount of gaseous I<sub>2</sub> to effuse from the same container under identical conditions?
89. The graph shows the distribution of molecular velocities for two different molecules (A and B) at the same temperature. Which molecule has the higher molar mass? Which molecule has the higher rate of effusion?

**MISSED THIS?** Read Section 6.8; Watch KCV 6.8



## CUMULATIVE PROBLEMS

95. Modern pennies are composed of zinc coated with copper. A student determines the mass of a penny to be 2.482 g and then makes several scratches in the copper coating (to expose the underlying zinc). The student puts the scratched penny in hydrochloric acid, where the following reaction occurs between the zinc and the HCl (the copper remains undissolved):



The student collects the hydrogen produced over water at 25 °C. The collected gas occupies a volume of 0.899 L at a total pressure of 791 mmHg. Calculate the percent zinc (by mass) in the penny. (Assume that all the Zn in the penny dissolves.)

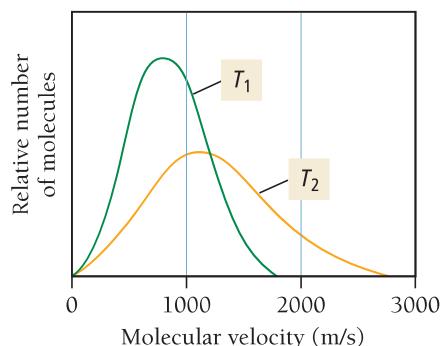
96. A 2.85-g sample of an unknown chlorofluorocarbon decomposes and produces 564 mL of chlorine gas at a pressure of 752 mmHg and a temperature of 298 K. What is the percent chlorine (by mass) in the unknown chlorofluorocarbon?

97. The mass of an evacuated 255 mL flask is 143.187 g. The mass of the flask filled with 267 torr of an unknown gas at 25 °C is 143.289 g. Calculate the molar mass of the unknown gas.

98. A 118-mL flask is evacuated and found to have a mass of 97.129 g. When the flask is filled with 768 torr of helium gas at 35 °C, it has a mass of 97.171 g. Was the helium gas pure?

99. A gaseous hydrogen- and carbon-containing compound is decomposed and found to contain 82.66% carbon and 17.34%

90. The graph shows the distribution of molecular velocities for the same molecule at two different temperatures ( $T_1$  and  $T_2$ ). Which temperature is greater? Explain.



## Real Gases

91. Which postulate of the kinetic molecular theory breaks down under conditions of high pressure? Explain.

**MISSED THIS?** Read Section 6.10

92. Which postulate of the kinetic molecular theory breaks down under conditions of low temperature? Explain.

93. Use the van der Waals equation and the ideal gas equation to calculate the volume of 1.000 mol of neon at a pressure of 500.0 atm and a temperature of 355.0 K. Explain why the two values are different. (Hint: One way to solve the van der Waals equation for  $V$  is to use successive approximations. Use the ideal gas law to get a preliminary estimate for  $V$ .)

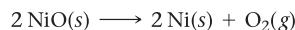
**MISSED THIS?** Read Section 6.10

94. Use the van der Waals equation and the ideal gas equation to calculate the pressure exerted by 1.000 mol of Cl<sub>2</sub> in a volume of 5.000 L at a temperature of 273.0 K. Explain why the two values are different.

hydrogen by mass. The mass of 158 mL of the gas, measured at 556 mmHg and 25 °C, was 0.275 g. What is the molecular formula of the compound?

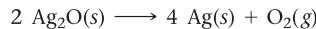
100. A gaseous hydrogen- and carbon-containing compound is decomposed and found to contain 85.63% C and 14.37% H by mass. The mass of 258 mL of the gas, measured at STP, was 0.646 g. What is the molecular formula of the compound?

101. Consider the reaction:



If O<sub>2</sub> is collected over water at 40.0 °C and a total pressure of 745 mmHg, what volume of gas is collected for the complete reaction of 24.78 g of NiO?

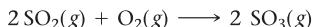
102. Consider the reaction:



If this reaction produces 15.8 g of Ag(s), what total volume of gas can be collected over water at a temperature of 25 °C and a total pressure of 752 mmHg?

103. When hydrochloric acid is poured over potassium sulfide, 42.9 mL of hydrogen sulfide gas is produced at a pressure of 752 torr and 25.8 °C. Write an equation for the gas-evolution reaction and determine how much potassium sulfide (in grams) reacted.

**104.** Consider the reaction:

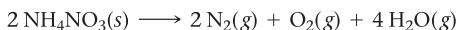


- a. If 285.5 mL of  $\text{SO}_2$  reacts with 158.9 mL of  $\text{O}_2$  (both measured at 315 K and 50.0 mmHg), what is the limiting reactant and the theoretical yield of  $\text{SO}_3$ ?
  - b. If 187.2 mL of  $\text{SO}_3$  is collected (measured at 315 K and 50.0 mmHg), what is the percent yield for the reaction?
- 105.** Ammonium carbonate decomposes upon heating according to the balanced equation:



Calculate the total volume of gas produced at 22 °C and 1.02 atm by the complete decomposition of 11.83 g of ammonium carbonate.

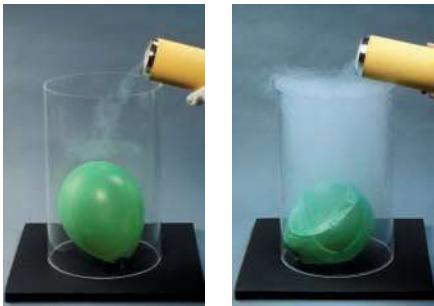
- 106.** Ammonium nitrate decomposes explosively upon heating according to the balanced equation:



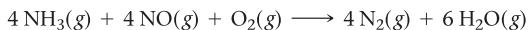
Calculate the total volume of gas (at 125 °C and 748 mmHg) produced by the complete decomposition of 1.55 kg of ammonium nitrate.

- 107.** Olympic cyclists fill their tires with helium to make them lighter. Calculate the mass of air in an air-filled tire and the mass of helium in a helium-filled tire. What is the mass difference between the two? Assume that the volume of the tire is 855 mL, that it is filled to a total pressure of 125 psi, and that the temperature is 25 °C. Also assume an average molar mass for air of 28.8 g/mol.

- 108.** In a common classroom demonstration, a balloon is filled with air and drenched with liquid nitrogen. The balloon contracts as the gases within the balloon cool. Suppose a balloon initially contains 2.95 L of air at a temperature of 25.0 °C and a pressure of 0.998 atm. Calculate the expected volume of the balloon upon cooling to –196 °C (the boiling point of liquid nitrogen). When the demonstration is carried out, the actual volume of the balloon decreases to 0.61 L. How does the observed volume of the balloon compare to your calculated value? Explain the difference.



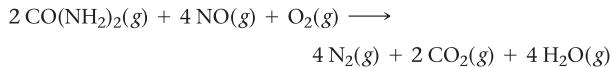
- 109.** Gaseous ammonia is injected into the exhaust stream of a coal-burning power plant to reduce the pollutant NO to  $\text{N}_2$  according to the reaction:



Suppose that the exhaust stream of a power plant has a flow rate of 335 L/s at a temperature of 955 K and that the exhaust contains a partial pressure of NO of 22.4 torr. What should be the flow rate of ammonia delivered at 755 torr and 298 K into

the stream to react completely with the NO if the ammonia is 65.2% pure (by volume)?

- 110.** The emission of  $\text{NO}_2$  by fossil fuel combustion can be prevented by injecting gaseous urea into the combustion mixture. The urea reduces NO (which oxidizes in air to form  $\text{NO}_2$ ) according to the reaction:



Suppose that the exhaust stream of an automobile has a flow rate of 2.55 L/s at 655 K and contains a partial pressure of NO of 12.4 torr. What total mass of urea is necessary to react completely with the NO formed during 8.0 hours of driving?

- 111.** An ordinary gasoline can measuring 30.0 cm by 20.0 cm by 15.0 cm is evacuated with a vacuum pump. Assuming that virtually all of the air can be removed from inside the can and that atmospheric pressure is 14.7 psi, what is the total force (in pounds) on the surface of the can? Do you think that the can could withstand the force?

- 112.** Twenty-five milliliters of liquid nitrogen (density = 0.807 g/mL) is poured into a cylindrical container with a radius of 10.0 cm and a length of 20.0 cm. The container initially contains only air at a pressure of 760.0 mmHg (atmospheric pressure) and a temperature of 298 K. If the liquid nitrogen completely vaporizes, what is the total force (in lb) on the interior of the container at 298 K?

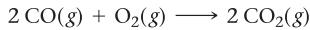
- 113.** A 160.0-L helium tank contains pure helium at a pressure of 1855 psi and a temperature of 298 K. How many 3.5-L helium balloons will the helium in the tank fill? (Assume an atmospheric pressure of 1.0 atm and a temperature of 298 K.)

- 114.** An 11.5-mL sample of liquid butane (density = 0.573 g/mL) is evaporated in an otherwise empty container at a temperature of 28.5 °C. The pressure in the container following evaporation is 892 torr. What is the volume of the container?

- 115.** A scuba diver creates a spherical bubble with a radius of 2.5 cm at a depth of 30.0 m where the total pressure (including atmospheric pressure) is 4.00 atm. What is the radius of the bubble when it reaches the surface of the water? (Assume that the atmospheric pressure is 1.00 atm and the temperature is 298 K.)

- 116.** A particular balloon can be stretched to a maximum surface area of  $1257 \text{ cm}^2$ . The balloon is filled with 3.0 L of helium gas at a pressure of 755 torr and a temperature of 298 K. The balloon is then allowed to rise in the atmosphere. If the atmospheric temperature is 273 K, at what pressure will the balloon burst? (Assume the balloon is the shape of a sphere.)

- 117.** A catalytic converter in an automobile uses a palladium or platinum catalyst (a substance that increases the rate of a reaction without being consumed by the reaction) to convert carbon monoxide gas to carbon dioxide according to the reaction:



A chemist researching the effectiveness of a new catalyst combines a 2.0:1.0 mole ratio mixture of carbon monoxide and oxygen gas, respectively, over the catalyst in a 2.45-L flask at a total pressure of 745 torr and a temperature of 552 °C. When the reaction is complete, the pressure in the flask has dropped to 552 torr. What percentage of the carbon monoxide was converted to carbon dioxide?

- 118.** A quantity of  $\text{N}_2$  occupies a volume of 1.0 L at 300 K and 1.0 atm. The gas expands to a volume of 3.0 L as the result of a change in both temperature and pressure. Find the density of the gas at these new conditions.
- 119.** A mixture of  $\text{CO}(g)$  and  $\text{O}_2(g)$  in a 1.0-L container at  $1.0 \times 10^3$  K has a total pressure of 2.2 atm. After some time, the total pressure falls to 1.9 atm as the result of the formation of  $\text{CO}_2$ . Determine the mass (in grams) of  $\text{CO}_2$  that forms.
- 120.** The radius of a xenon atom is  $1.3 \times 10^{-8}$  cm. A 100-mL flask is filled with Xe at a pressure of 1.0 atm and a temperature of 273 K. Calculate the fraction of the volume that is occupied by Xe atoms. (*Hint:* The atoms are spheres.)
- 121.** A natural gas storage tank is a cylinder with a moveable top. Its volume can change only as its height changes, and its radius remains fixed. The height of the cylinder is 22.6 m on a day when the temperature is 22 °C. The next day the height of the cylinder increases to 23.8 m when the gas expands because of a heat wave. Determine the temperature on the second day, assuming that the pressure and amount of gas in the storage tank have not changed.
- 122.** A mixture of 8.0 g  $\text{CH}_4$  and 8.0 g Xe is placed in a container, and the total pressure is found to be 0.44 atm. Determine the partial pressure of  $\text{CH}_4$ .
- 123.** A steel container of volume 0.35 L can withstand pressures up to 88 atm before exploding. What mass of helium can be stored in this container at 299 K?
- 124.** Binary compounds of alkali metals and hydrogen react with water to liberate  $\text{H}_2(g)$ . The  $\text{H}_2$  from the reaction of a sample of  $\text{NaH}$  with an excess of water fills a volume of 0.490 L above the water. The temperature of the gas is 35 °C, and the total pressure is 758 mmHg. Determine the mass of  $\text{H}_2$  liberated and the mass of  $\text{NaH}$  that reacted.
- 125.** In a given diffusion apparatus, 15.0 mL of  $\text{HBr}$  gas diffuses in 1.0 min. In the same apparatus and under the same conditions, 20.3 mL of an unknown gas diffuses in 1.0 min. The unknown gas is a hydrocarbon. Find its molecular formula.
- 126.** A sample of  $\text{N}_2\text{O}_3(g)$  has a pressure of 0.017 atm. The temperature (in K) is doubled, and the  $\text{N}_2\text{O}_3$  undergoes complete decomposition to  $\text{NO}_2(g)$  and  $\text{NO}(g)$ . Find the total pressure of the mixture of gases assuming constant volume and no additional temperature change.
- 127.** When 0.583 g of neon is added to an 800-cm<sup>3</sup> bulb containing a sample of argon, the total pressure of the gases is 1.17 atm at a temperature of 295 K. Find the mass of the argon in the bulb.
- 128.** A gas mixture composed of helium and argon has a density of 0.670 g/L at 755 mmHg and 298 K. What is the composition of the mixture by volume?
- 129.** A gas mixture contains 75.2% nitrogen and 24.8% krypton by mass. What is the partial pressure of krypton in the mixture if the total pressure is 745 mmHg?

## CHALLENGE PROBLEMS

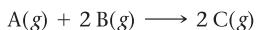
- 130.** A 10-L container is filled with 0.10 mol of  $\text{H}_2(g)$  and heated to 3000 K, causing some of the  $\text{H}_2(g)$  to decompose into  $\text{H}(g)$ . The pressure is found to be 3.0 atm. Find the partial pressure of the  $\text{H}(g)$  that forms from  $\text{H}_2$  at this temperature. (Assume two significant figures for the temperature.)
- 131.** A mixture of  $\text{NH}_3(g)$  and  $\text{N}_2\text{H}_4(g)$  is placed in a sealed container at 300 K. The total pressure is 0.50 atm. The container is heated to 1200 K, at which time both substances decompose completely according to the equations  $2 \text{NH}_3(g) \longrightarrow \text{N}_2(g) + 3 \text{H}_2(g)$ ;  $\text{N}_2\text{H}_4(g) \longrightarrow \text{N}_2(g) + 2 \text{H}_2(g)$ . After decomposition is complete, the total pressure at 1200 K is found to be 4.5 atm. Find the percent of  $\text{N}_2\text{H}_4(g)$  in the original mixture. (Assume two significant figures for the temperature.)
- 132.** A quantity of CO gas occupies a volume of 0.48 L at 1.0 atm and 275 K. The pressure of the gas is lowered, and its temperature is raised until its volume is 1.3 L. Determine the density of the CO under the new conditions.
- 133.** When  $\text{CO}_2(g)$  is put in a sealed container at 701 K and a pressure of 10.0 atm and is heated to 1401 K, the pressure rises to 22.5 atm. Some of the  $\text{CO}_2$  decomposes to CO and  $\text{O}_2$ . Calculate the mole percent of  $\text{CO}_2$  that decomposes.
- 134.** The world burns approximately  $3.7 \times 10^{12}$  kg of fossil fuel per year. Use the combustion of octane as the representative reaction and determine the mass of carbon dioxide (the most significant greenhouse gas) formed per year. The current concentration of carbon dioxide in the atmosphere is approximately 399 ppm (by volume). By what percentage does the concentration increase each year due to fossil fuel combustion? Approximate the average properties of the entire atmosphere by assuming that the atmosphere extends from sea level to 15 km and that it has an average pressure of 381 torr and average temperature of 275 K. Assume Earth is a perfect sphere with a radius of 6371 km.
- 135.** The atmosphere slowly oxidizes hydrocarbons in a number of steps that eventually convert the hydrocarbon into carbon dioxide and water. Part of the process for methane gas is
- $$\text{CH}_4(g) + 5 \text{O}_2(g) + 5 \text{NO}(g) \longrightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g) + 5 \text{NO}(g) + 2 \text{OH}(g)$$
- Suppose that an atmospheric chemist combines 155 mL of methane at STP, 885 mL of oxygen at STP, and 55.5 mL of NO at STP in a 2.0 L flask. The flask is allowed to stand for several weeks at 275 K. If the reaction reaches 90.0% of completion (90.0% of the limiting reactant is consumed), what is the partial pressure of each of the reactants and products in the flask at 275 K? What is the total pressure in the flask?
- 136.** Two identical balloons are filled to the same volume, one with air and one with helium. The next day, the volume of the air-filled balloon has decreased by 5.0%. By what percent has the volume of the helium-filled balloon decreased? (Assume that the air is four-fifths nitrogen and one-fifth oxygen and that the temperature did not change.)
- 137.** A mixture of  $\text{CH}_4(g)$  and  $\text{C}_2\text{H}_6(g)$  has a total pressure of 0.53 atm. Just enough  $\text{O}_2(g)$  is added to the mixture to bring about its complete combustion to  $\text{CO}_2(g)$  and  $\text{H}_2\text{O}(g)$ . The total pressure of the two product gases is 2.2 atm. Assuming constant volume and temperature, find the mole fraction of  $\text{CH}_4$  in the mixture.
- 138.** A sample of  $\text{C}_2\text{H}_2(g)$  has a pressure of 7.8 kPa. After some time a portion of it reacts to form  $\text{C}_6\text{H}_6(g)$ . The total pressure of the mixture of gases is then 3.9 kPa. Assume the volume and the temperature do not change. What fraction of  $\text{C}_2\text{H}_2(g)$  has undergone reaction?

## CONCEPTUAL PROBLEMS

**139.** When the driver of an automobile applies the brakes, the passengers are pushed toward the front of the car, but a helium balloon is pushed toward the back of the car. Upon forward acceleration, the passengers are pushed toward the back of the car, but the helium balloon is pushed toward the front of the car. Why?

**140.** Suppose that a liquid is 10 times denser than water. If you were to sip this liquid at sea level using a straw, what is the maximum length your straw would be?

**141.** This reaction occurs in a closed container:



A reaction mixture initially contains 1.5 L of A and 2.0 L of B. Assuming that the volume and temperature of the reaction mixture remain constant, what is the percent change in pressure if the reaction goes to completion?

**142.** One mole of nitrogen and one mole of neon are combined in a closed container at STP. How big is the container?

**143.** Exactly equal amounts (in moles) of gas A and gas B are combined in a 1-L container at room temperature. Gas B has a molar mass that is twice that of gas A. Which statement is true for the mixture of gases and why?

- a. The molecules of gas B have greater kinetic energy than those of gas A.
- b. Gas B has a greater partial pressure than gas A.
- c. The molecules of gas B have a greater average velocity than those of gas A.
- d. Gas B makes a greater contribution to the average density of the mixture than gas A.

**144.** Which gas would you expect to deviate most from ideal behavior under conditions of low temperature: F<sub>2</sub>, Cl<sub>2</sub>, or Br<sub>2</sub>? Explain.

**145.** The volume of a sample of a fixed amount of gas is decreased from 2.0 L to 1.0 L. The temperature of the gas in kelvins is then doubled. What is the final pressure of the gas in terms of the initial pressure?

**146.** Which gas sample has the greatest volume at STP?  
 a. 10.0 g Kr    b. 10.0 g Xe    c. 10.0 g He

**147.** Draw a depiction of a gas sample, as described by kinetic molecular theory, containing equal molar amounts of helium, neon, and krypton. Use different color dots to represent each element. Give each atom a "tail" to represent its velocity relative to the others in the mixture.

## QUESTIONS FOR GROUP WORK

Discuss these questions with the group and record your consensus answer.

**148.** Assign one of the three simple gas laws to each member of your group. For the assigned gas law, have each member write two equations, draw a graph, and describe it in a complete sentence. Have each group member present his or her law to the group.

**149.** Review the ideal gas law. Without referring back to the text, use algebra to write the ideal gas law and solve for each of the individual variables it contains. Have each group member solve for a different variable and present answers to the group.

**150.** Hydrogen peroxide (H<sub>2</sub>O<sub>2</sub>) decomposes in the presence of a catalyst to form water and oxygen. The catalyst is added to 5.00 mL of a hydrogen peroxide solution at 25.0 °C, and 49.5 mL of gas is collected over water at a total pressure of 763.8 mmHg.

- a. Write and balance the chemical reaction (Note: catalysts do not appear in balanced chemical equations).
- b. Look up the vapor pressure of water under these conditions.
- c. What is the partial pressure of oxygen collected over the water?
- d. How many moles of oxygen are collected?

### Active Classroom Learning

e. How many grams of hydrogen peroxide were in the original sample?

f. What is the concentration (in mol/L) of the hydrogen peroxide solution?

g. Which part of this process is conceptually most difficult for your group?

**151.** A box contains equal amounts of helium, argon, and krypton (all gases) at 25 °C. Using complete sentences, describe the temperatures, masses, average velocities, and average kinetic energy of the three kinds of gas in the mixture. What do they have in common? What are the differences? How are these properties related?

**152.** Calculate the pressure exerted by 1 mol of an ideal gas in a box that is 0.500 L and 298 K. Have each group member calculate the pressure of 1 mol of the following gases in the same box at the same temperature: He, Ne, H<sub>2</sub>, CH<sub>4</sub>, and CO<sub>2</sub>. Compare group members' answers as well as all answers with the pressure of an ideal gas. Assuming that the van der Waals equation predictions are accurate, account for why the pressure of each gas is higher or lower than that predicted for an ideal gas.

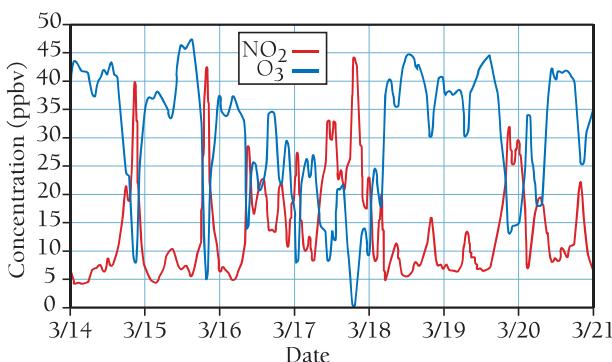


## DATA INTERPRETATION AND ANALYSIS

### Nitrogen Oxide from Automobiles

**153.** When fuels are burned in air, such as in an automobile engine, some of the nitrogen in the air oxidizes to form nitrogen oxide gases such as NO and NO<sub>2</sub> (known collectively as NO<sub>x</sub>). The U.S. Environmental Protection Agency (EPA) sets standards for air quality of several pollutants including NO<sub>2</sub>. According to the EPA, NO<sub>2</sub> levels in U.S. cities are not to exceed a yearly average of 53 ppb or a 1-hour average of 100 ppb. Another pollutant

associated with automobile exhaust is ozone (O<sub>3</sub>). The EPA standard for ozone is an 8-hour average of 70 ppb. Breathing air with elevated levels of NO<sub>2</sub> or O<sub>3</sub> can cause asthma and other respiratory problems. The graph shown here shows the average concentration of nitrogen dioxide (NO<sub>2</sub>) and ozone (O<sub>3</sub>) gases in units of parts per billion by volume (ppbv) over seven days in a city.



### ▲ Concentration of NO<sub>2</sub> and O<sub>3</sub> over Seven Days

Source: <http://www.cas.manchester.ac.uk/resprojects/holmehill/results/fig2/>

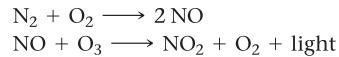
Study the graph and answer the following questions:

- What type of relationship exists between nitrogen dioxide and ozone between March 14 and March 16?
- Calculate the number of moles of NO<sub>2</sub> in 1.00 m<sup>3</sup> produced on March 14. Assume an average temperature of 25.0 °C and a pressure of 1 atm. Note that the number of moles of NO<sub>2</sub> produced is the difference between the existing amount at the start of the day and the peak amount.

c. Calculate the number of moles of O<sub>3</sub> in 1.00 m<sup>3</sup> consumed on March 14. Assume an average temperature of 25.0 °C and a pressure of 1 atm. Note that the number of moles of O<sub>3</sub> consumed is the difference between the existing amount at the start of the day and the minimum amount.

d. What is the mole-to-mole ratio of O<sub>3</sub> consumed to NO<sub>2</sub> produced?

e. The following chemical equations model the interactions of nitrogen dioxide gas and ozone gas. Can this set of equations account for the trends observed in the graph? Explain your answer.



f. Do the concentrations of NO<sub>2</sub> or O<sub>3</sub> exceed the standards set by the EPA?

**Cc**

## ANSWERS TO CONCEPTUAL CONNECTIONS

### Boyle's Law and Charles's Law

**6.1 (e)** The final volume of the gas is the same as the initial volume because doubling the pressure *decreases* the volume by a factor of 2, but doubling the temperature *increases* the volume by a factor of 2. The two changes in volume are equal in magnitude but opposite in sign, resulting in a final volume that is equal to the initial volume.

### Simple Gas Laws

**6.2 (a)** According to Boyle's law, decreasing the pressure increases the volume (at constant temperature and number of moles.)

### Molar Volume

**6.3 (a)** Since one gram of H<sub>2</sub> contains the greatest number of moles (due to H<sub>2</sub> having the lowest molar mass of the listed gases) and since one mole of *any* ideal gas occupies the same volume, the H<sub>2</sub> occupies the greatest volume.

### Density of a Gas

**6.4 (a)** Ne < O<sub>2</sub> < F<sub>2</sub> < Cl<sub>2</sub>

### Partial Pressures

**6.5 (c)** P<sub>He</sub> = 1.5 atm; P<sub>Ne</sub> = 1.5 atm. Since the number of moles of each gas is equal, the mole fraction of each gas is 0.50 and the partial pressure of each gas is 0.50 × P<sub>tot</sub>.

### Pressure and Number of Moles

**6.6 (b)** Since the total number of gas molecules decreases, the total pressure—the sum of all the partial pressures—must also decrease.

### Root Mean Square Velocity

**6.7 (b)** The molar mass of neon (20.18 g/mol) is approximately one-quarter the molar mass of krypton (83.80 g/mol), so the root mean square velocity of neon is  $\sqrt{\frac{1}{1/4}}$  times that of krypton.

### Kinetic Molecular Theory

**6.8 (c)** Since the temperature and the volume are both constant, the ideal gas law tells us that the pressure depends solely on the number of particles. Sample (c) has the greatest number of particles per unit volume and therefore has the greatest pressure. The pressures of samples (a) and (b) at a given temperature are identical. Even though the particles in (b) are more massive than those in (a), they have the same average kinetic energy at a given temperature. The particles in (b) move more slowly than those in (a) and so exert the same pressure as the particles in (a).

### Graham's Law of Effusion

**6.9 (c)** Since the molar mass of helium is less than that of argon, helium effuses faster than argon, resulting in an increase in the relative amount of argon relative to helium left in the tube.

### Real Gases

**6.10 (b)** A < B < C. Curve A is the lowest temperature curve because it deviates the most from ideality. The tendency for the intermolecular forces in carbon dioxide to lower the pressure (relative to that of an ideal gas) is greatest at low temperature (because the molecules are moving more slowly and are therefore less able to overcome the intermolecular forces). As a result, the curve that dips the lowest must correspond to the lowest temperature.