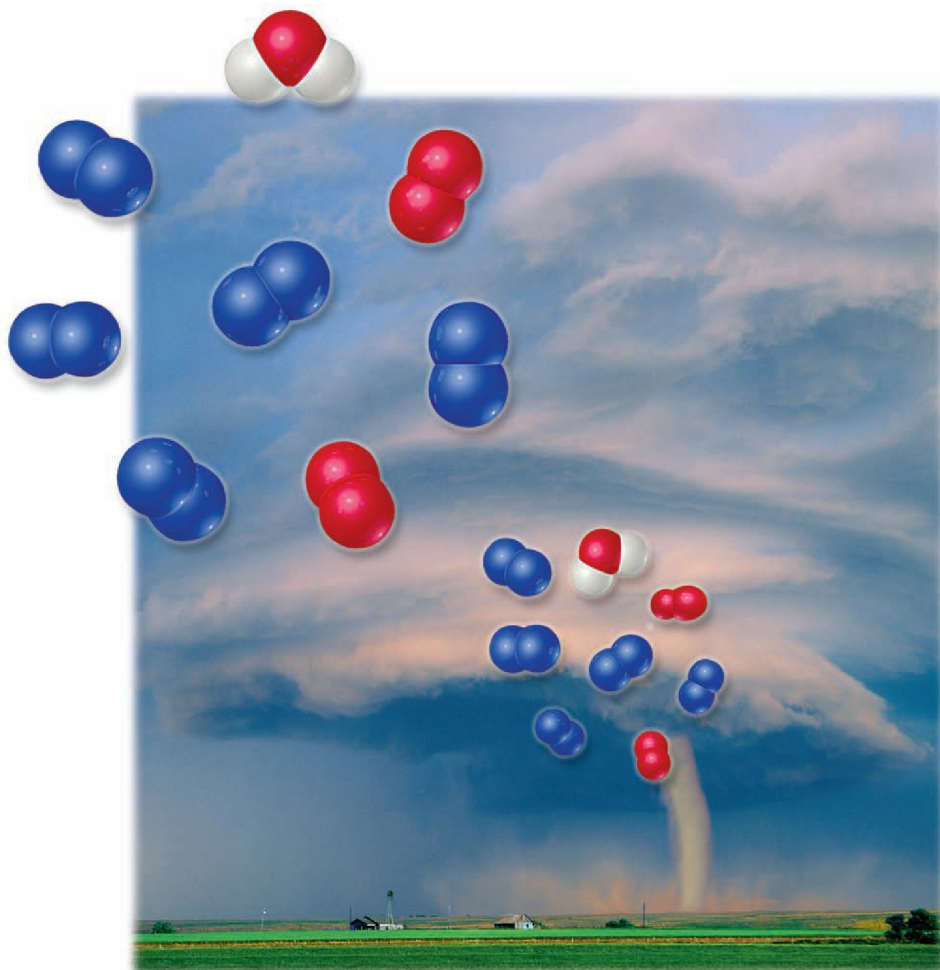


CHAPTER
10
GASES



A TORNADO IS A VIOLENT rotating column of air characterized by a funnel-shaped cloud.
A tornado can be up to 100 m in diameter with wind speeds approaching 500 km/hr.

WHAT'S AHEAD

10.1 Characteristics of Gases

We begin by comparing the distinguishing characteristics of gases with those of liquids and solids.

10.2 Pressure

We will then study gas *pressure*, how it is measured, and the units used to express it, as well as consider Earth's atmosphere and the pressure it exerts.

10.3 The Gas Laws

We consider that the state of a gas can be expressed in terms of its volume (V), pressure (P), temperature (T), and quantity of gas (n). We will examine several empirical relationships among V , P , T , and n that are called *gas laws*.

10.4 The Ideal-Gas Equation

We find that when considered together, the empirical gas laws yield the *ideal-gas equation*, $PV = nRT$. Although the ideal-gas equation is not obeyed exactly by any real gas, most gases come very close to obeying it at temperature and pressure conditions of greatest interest.

10.5 Further Applications of the Ideal-Gas Equation

We can use the ideal-gas equation to make many useful calculations, such as the calculation of the density or molar mass of a gas.

10.6 Gas Mixtures and Partial Pressures

We recognize that in a mixture of gases, each component exerts a pressure that is part of the total pressure. This *partial pressure* is the pressure that the component would exert if it were by itself.

10.7 Kinetic-Molecular Theory

We explore this theory of gases, which helps us understand the behavior of gases, including their pressure and temperature, on a molecular level. According to the theory, the atoms or molecules that make up the gas move about with an average kinetic energy that is proportional to the gas temperature.

10.8 Molecular Effusion and Diffusion


We next observe that the kinetic-molecular theory helps us account for such gas properties as *effusion* through tiny openings and for *diffusion*, the movement through another substance.

10.9 Real Gases: Deviations from Ideal Behavior

We then consider that real gases deviate from ideal behavior because the gas molecules have finite volume and because attractive forces exist between molecules. The *van der Waals equation* gives a more accurate account of real gas behavior at high pressures and low temperatures.

IN THE PAST SEVERAL CHAPTERS we have learned about electronic structures of atoms and about how atoms combine to form molecules and ionic substances. In everyday life, however, we do not have direct experiences with atoms. Instead, we encounter matter as collections of enormous

numbers of atoms or molecules that make up gases, liquids, and solids. Large collections of atoms and molecules in the atmosphere, for example, are responsible for our weather—the gentle breezes and the gales, the humidity and the rain. Tornadoes, such as the one shown in the chapter-opening photo, form when moist, warm air at lower elevations converges with cooler, dry air above. The resultant air flows produce winds that can approach speeds up to 500 km/hr (300 mph).

It was John Dalton's interest in the weather that motivated him to study gases and eventually to propose the atomic theory of matter.  (Section 2.1) We now know that the properties of gases, liquids, and solids are readily

understood in terms of the behavior of their component atoms, ions, and molecules. In this chapter we will examine the physical properties of gases and consider how we can understand these properties in terms of the behavior of gas molecules. In Chapter 11 we will turn our attention to the physical properties of liquids and solids.

10.1 CHARACTERISTICS OF GASES

In many ways gases are the most easily understood form of matter. Even though different gaseous substances may have very different *chemical* properties, they behave quite similarly as far as their *physical* properties are concerned. For example, we live in an atmosphere composed of a mixture of gases that we refer to as air. Air is a complex mixture of several substances, primarily N_2 (78%) and O_2 (21%), with small amounts of several other gases, including Ar (0.9%). We breathe air to absorb oxygen, O_2 , which supports human life. Air also contains nitrogen, N_2 , which has very different chemical properties from oxygen, yet this mixture behaves physically as one gaseous material.

Only a few elements exist as gases under ordinary conditions of temperature and pressure: The noble gases (He, Ne, Ar, Kr, and Xe) are all monatomic gases, whereas H_2 , N_2 , O_2 , F_2 , and Cl_2 are diatomic gases. Many molecular compounds are also gases. Table 10.1 ▼ lists a few of the more common gaseous compounds. Notice that all of these gases are composed entirely of nonmetallic elements. Furthermore, all have simple molecular formulas and, therefore, low molar masses. Substances that are liquids or solids under ordinary conditions can also exist in the gaseous state, where they are often referred to as **vapors**. The substance H_2O , for example, can exist as liquid water, solid ice, or water vapor.

Gases differ significantly from solids and liquids in several respects. For example, a gas expands spontaneously to fill its container. Consequently, the volume of a gas equals the volume of the container in which it is held. Gases also are highly compressible: When pressure is applied to a gas, its volume readily decreases. Solids and liquids, on the other hand, do not expand to fill their containers, and solids and liquids are not readily compressible.

Gases form homogeneous mixtures with each other regardless of the identities or relative proportions of the component gases. The atmosphere serves as an excellent example. As a further example, when water and gasoline are mixed, the two liquids remain as separate layers. In contrast, the water vapor and gasoline vapors above the liquids form a homogeneous gas mixture.

The characteristic properties of gases arise because the individual molecules are relatively far apart. In the air we breathe, for example, the molecules take up only about 0.1% of the total volume, with the rest being empty space. Thus, each molecule behaves largely as though the others were not present.

TABLE 10.1 ■ Some Common Compounds That Are Gases at Room Temperature

Formula	Name	Characteristics
HCN	Hydrogen cyanide	Very toxic, slight odor of bitter almonds
H_2S	Hydrogen sulfide	Very toxic, odor of rotten eggs
CO	Carbon monoxide	Toxic, colorless, odorless
CO_2	Carbon dioxide	Colorless, odorless
CH_4	Methane	Colorless, odorless, flammable
C_2H_4	Ethylene	Colorless, ripens fruit
C_3H_8	Propane	Colorless, odorless, bottled gas
N_2O	Nitrous oxide	Colorless, sweet odor, laughing gas
NO_2	Nitrogen dioxide	Toxic, red-brown, irritating odor
NH_3	Ammonia	Colorless, pungent odor
SO_2	Sulfur dioxide	Colorless, irritating odor

As a result, different gases behave similarly, even though they are made up of different molecules. In contrast, the individual molecules in a liquid are close together and occupy perhaps 70% of the total space. The attractive forces among the molecules keep the liquid together.

GIVE IT SOME THOUGHT

What is the major reason that physical properties do not differ much from one gaseous substance to another?

10.2 PRESSURE

Among the most readily measured properties of a gas are its temperature, volume, and pressure. Many early studies of gases focused on relationships among these properties. We have already discussed volume and temperature. (Section 1.4) Let's now consider the concept of pressure.

In general terms, **pressure** conveys the idea of a force, a push that tends to move something in a given direction. Pressure, P , is, in fact, the force, F , that acts on a given area, A .

$$P = \frac{F}{A} \quad [10.1]$$

Gases exert a pressure on any surface with which they are in contact. The gas in an inflated balloon, for example, exerts a pressure on the inside surface of the balloon.

Atmospheric Pressure and the Barometer

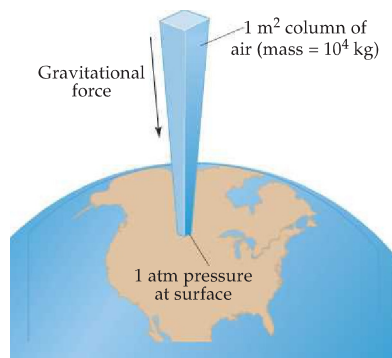
You and I, coconuts and nitrogen molecules, all experience an attractive force that pulls toward the center of Earth. When a coconut comes loose from its place at the top of the tree, for example, the gravitational attractive force causes it to be accelerated toward Earth, increasing in speed as its potential energy is converted into kinetic energy. (Section 5.1) The atoms and molecules of the atmosphere also experience a gravitational acceleration. Because the gas particles have such tiny masses, however, their thermal energies of motion (their kinetic energies) override the gravitational forces, so all the molecules that make up the atmosphere don't just pile up in a thin layer at Earth's surface. Nevertheless, gravity does operate, and it causes the atmosphere as a whole to press down on Earth's surface, creating an atmospheric pressure.

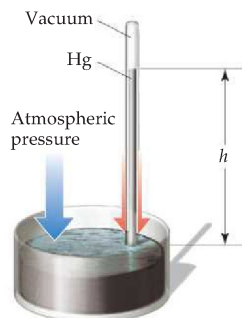
You can demonstrate the existence of atmospheric pressure to yourself with an empty plastic bottle of the sort used to hold water or soft drinks. If you suck on the mouth of the empty bottle, chances are you can cause it to partially cave in. When you break the partial vacuum you have created, the bottle pops out to its original shape. What causes the bottle to cave in when the pressure inside it is reduced, even by the relatively small amount you can manage with your lungs? The gas molecules in the atmosphere are exerting a force on the outside of the bottle that is greater than the force within the bottle when some of the gas has been sucked out.

We can calculate the magnitude of the atmospheric pressure as follows: The force, F , exerted by any object is the product of its mass, m , times its acceleration, a ; that is, $F = ma$. The acceleration given by Earth's gravity to any object located near Earth's surface is 9.8 m/s^2 . Now imagine a column of air 1 m^2 in cross section extending through the entire atmosphere (Figure 10.1). That column has a mass of roughly 10,000 kg. The force exerted by the column on Earth's surface is

$$F = (10,000 \text{ kg})(9.8 \text{ m/s}^2) = 1 \times 10^5 \text{ kg}\cdot\text{m/s}^2 = 1 \times 10^5 \text{ N}$$

▼ **Figure 10.1 Atmospheric pressure.** Illustration of the manner in which Earth's atmosphere exerts pressure at the surface of the planet. The mass of a column of atmosphere exactly 1 m^2 in cross-sectional area and extending to the upper atmosphere exerts a force of $1 \times 10^5 \text{ N}$.





▲ **Figure 10.2 A mercury barometer.** The pressure of the atmosphere on the surface of the mercury (represented by the blue arrow) equals the pressure of the column of mercury (red arrow).

The SI unit for force is $\text{kg}\cdot\text{m}/\text{s}^2$ and is called the *newton* (N): $1 \text{ N} = 1 \text{ kg}\cdot\text{m}/\text{s}^2$. The pressure exerted by the column is the force divided by the cross-sectional area, A , over which the force is applied. Because our air column has a cross-sectional area of 1 m^2 , we have

$$P = \frac{F}{A} = \frac{1 \times 10^5 \text{ N}}{1 \text{ m}^2} = 1 \times 10^5 \text{ N}/\text{m}^2 = 1 \times 10^5 \text{ Pa} = 1 \times 10^2 \text{ kPa}$$

The SI unit of pressure is N/m^2 . It is given the name **pascal** (Pa) after Blaise Pascal (1623–1662), a French mathematician and scientist who studied pressure: $1 \text{ Pa} = 1 \text{ N}/\text{m}^2$. A related unit sometimes used to report pressures is the **bar**, which equals 10^5 Pa . Atmospheric pressure at sea level is about 100 kPa , or 1 bar . The actual atmospheric pressure at any location depends on weather conditions and altitude.

In the early part of the seventeenth century many scientists and philosophers believed that the atmosphere had no weight. Evangelista Torricelli (1608–1647), who was a student of Galileo's, proved this untrue. He invented the *barometer* (Figure 10.2 ◀), which is made from a glass tube more than 760 mm long that is closed at one end, completely filled with mercury, and inverted into a dish that contains additional mercury. (Care must be taken so that no air gets into the tube.) When the mercury-filled tube is inverted into the dish of mercury, some of the mercury flows out of the tube, but a column of mercury remains inside. Torricelli argued that the mercury surface in the dish experiences the full force, or weight, of Earth's atmosphere, which pushes the mercury up the tube until the pressure exerted by the mercury column equals the atmospheric pressure at the base of the tube. So the height, h , of the mercury column is a measure of the atmosphere's pressure, and it changes as the atmospheric pressure changes.

Torricelli's proposed explanation met with fierce opposition. Some argued that there could not possibly be a vacuum at the top of the tube. They said, "Nature does not permit a vacuum!" But Torricelli also had his supporters. Blaise Pascal, for example, had one of the barometers carried to the top of Puy de Dome, a volcanic mountain in central France, and compared its readings with a duplicate barometer kept at the foot of the mountain. As the barometer ascended, the height of the mercury column diminished, as expected, because the amount of atmosphere pressing down on the surface decreases as one moves higher. These and other experiments by scientists eventually prevailed, and the idea that the atmosphere has weight became accepted over many years.

GIVE IT SOME THOUGHT

What happens to the height of the mercury column in a mercury barometer as you move to higher altitude, and why?

Standard atmospheric pressure, which corresponds to the typical pressure at sea level, is the pressure sufficient to support a column of mercury 760 mm high. In SI units this pressure equals $1.01325 \times 10^5 \text{ Pa}$. Standard atmospheric pressure defines some common non-SI units used to express gas pressures, such as the **atmosphere** (atm) and the *millimeter of mercury* (mm Hg). The latter unit is also called the **torr**, after Torricelli.

$$1 \text{ atm} = 760 \text{ mm Hg} = 760 \text{ torr} = 1.01325 \times 10^5 \text{ Pa} = 101.325 \text{ kPa}$$

Note that the units mm Hg and torr are equivalent: $1 \text{ torr} = 1 \text{ mm Hg}$.

We will usually express gas pressure in units of atm, Pa (or kPa), or torr, so you should be comfortable converting gas pressures from one set of units to another.

SAMPLE EXERCISE 10.1 | Converting Units of Pressure

(a) Convert 0.357 atm to torr. (b) Convert 6.6×10^{-2} torr to atm. (c) Convert 147.2 kPa to torr.

SOLUTION

Analyze: In each case we are given the pressure in one unit and asked to convert it to another unit. Our task, therefore, is to choose the appropriate conversion factors.

Plan: We can use dimensional analysis to perform the desired conversions.

Solve:

(a) To convert atmospheres to torr, we use the relationship $760 \text{ torr} = 1 \text{ atm}$:

$$(0.357 \text{ atm}) \left(\frac{760 \text{ torr}}{1 \text{ atm}} \right) = 271 \text{ torr}$$

Note that the units cancel in the required manner.

(b) We use the same relationship as in part (a). To get the appropriate units to cancel, we must use the conversion factor as follows:

$$(6.6 \times 10^{-2} \text{ torr}) \left(\frac{1 \text{ atm}}{760 \text{ torr}} \right) = 8.7 \times 10^{-5} \text{ atm}$$

(c) The relationship $760 \text{ torr} = 101.325 \text{ kPa}$ allows us to write an appropriate conversion factor for this problem:

$$(147.2 \text{ kPa}) \left(\frac{760 \text{ torr}}{101.325 \text{ kPa}} \right) = 1104 \text{ torr}$$

Check: In each case look at the magnitude of the answer and compare it with the starting value. The torr is a much smaller unit than the atmosphere, so we expect the *numerical* answer to be larger than the starting quantity in (a) and smaller in (b). In (c) notice that there are nearly 8 torr per kPa, so the numerical answer in torr should be about 8 times larger than its value in kPa, consistent with our calculation.

PRACTICE EXERCISE

(a) In countries that use the metric system, such as Canada, atmospheric pressure in weather reports is given in units of kPa. Convert a pressure of 745 torr to kPa. (b) An English unit of pressure sometimes used in engineering is pounds per square inch ($\text{lb}/\text{in.}^2$), or psi: $1 \text{ atm} = 14.7 \text{ lb}/\text{in.}^2$. If a pressure is reported as 91.5 psi, express the measurement in atmospheres.

Answer: (a) 99.3 kPa, (b) 6.22 atm

We can use various devices to measure the pressures of enclosed gases. Tire gauges, for example, measure the pressure of air in automobile and bicycle tires. In laboratories we sometimes use a device called a *manometer*. A manometer operates on a principle similar to that of a barometer, as shown in Sample Exercise 10.2.

SAMPLE EXERCISE 10.2 | Using a Manometer to Measure Gas Pressure

On a certain day the barometer in a laboratory indicates that the atmospheric pressure is 764.7 torr. A sample of gas is placed in a flask attached to an open-end mercury manometer, shown in Figure 10.3. A meter stick is used to measure the height of the mercury above the bottom of the manometer. The level of mercury in the open-end arm of the manometer has a height of 136.4 mm, and the mercury in the arm that is in contact with the gas has a height of 103.8 mm. What is the pressure of the gas (a) in atmospheres, (b) in kPa?

SOLUTION

Analyze: We are given the atmospheric pressure (764.7 torr) and the heights of the mercury in the two arms of the manometer and asked to determine the gas pressure in the flask. We know that this pressure must be greater than atmospheric because the manometer level on the flask side (103.8 mm) is lower than that on the side open to the atmosphere (136.4 mm), as indicated in Figure 10.3.

Plan: We will use the difference in height between the two arms (h in Figure 10.3) to obtain the amount by which the pressure of the gas exceeds atmospheric pressure. Because an open-end mercury manometer is used, the height difference directly measures the pressure difference in mm Hg or torr between the gas and the atmosphere.

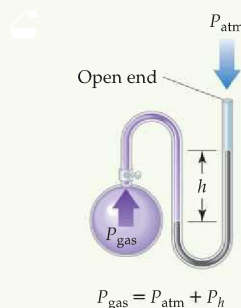


Figure 10.3 A mercury manometer. This device is sometimes employed in the laboratory to measure gas pressures near atmospheric pressure.

Solve:

(a) The pressure of the gas equals the atmospheric pressure plus h :

$$\begin{aligned} P_{\text{gas}} &= P_{\text{atm}} + h \\ &= 764.7 \text{ torr} + (136.4 \text{ torr} - 103.8 \text{ torr}) \\ &= 797.3 \text{ torr} \end{aligned}$$

We convert the pressure of the gas to atmospheres:

$$P_{\text{gas}} = (797.3 \text{ torr}) \left(\frac{1 \text{ atm}}{760 \text{ torr}} \right) = 1.049 \text{ atm}$$

(b) To calculate the pressure in kPa, we employ the conversion factor between atmospheres and kPa:

$$1.049 \text{ atm} \left(\frac{101.3 \text{ kPa}}{1 \text{ atm}} \right) = 106.3 \text{ kPa}$$

Check: The calculated pressure is a bit more than one atmosphere. This makes sense because we anticipated that the pressure in the flask would be greater than the pressure of the atmosphere acting on the manometer, which is a bit greater than one standard atmosphere.

PRACTICE EXERCISE

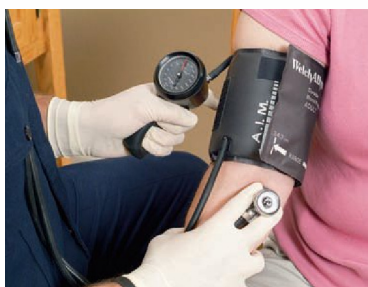
Convert a pressure of 0.975 atm into Pa and kPa.

Answers: 9.88×10^4 Pa and 98.8 kPa

**Chemistry and Life****BLOOD PRESSURE**

The human heart pumps blood to the parts of the body through arteries, and the blood returns to the heart through veins. When your blood pressure is measured, two values are reported, such as 120/80 (120 over 80), which is a normal reading. The first measurement is the *systolic pressure*, the maximum pressure when the heart is pumping. The second is the *diastolic pressure*, the pressure when the heart is in the resting part of its pumping cycle. The units associated with these pressure measurements are torr.

Blood pressure is measured using a pressure gauge attached to a closed, air-filled jacket or cuff that is applied like a tourniquet to the arm (Figure 10.4). The pressure gauge may be a mercury manometer or some other device. The air pressure in the cuff is increased using a small pump until it is above the systolic pressure and prevents the flow of blood. The air pressure inside the cuff is then slowly reduced until blood just begins to pulse through the artery, as detected by the use of a stethoscope. At this point the pressure in the cuff equals the pressure that the blood exerts inside the arteries. Reading the gauge gives the systolic pressure. The pressure in the cuff is then reduced further until the blood flows freely. The pressure at this point is the diastolic pressure.



▲ **Figure 10.4** Measuring blood pressure.

Hypertension is the presence of abnormally high blood pressure. The usual criterion for hypertension is a blood pressure greater than 140/90, although recent studies suggest that health risks increase for systolic readings above 120. Hypertension significantly increases the workload on the heart and places a stress on the walls of the blood vessels throughout the body. These effects increase the risk of aneurysms, heart attacks, and strokes.

10.3 THE GAS LAWS

Experiments with a large number of gases reveal that four variables are needed to define the physical condition, or *state*, of a gas: temperature, T , pressure, P , volume, V , and the amount of gas, which is usually expressed as the number of moles, n . The equations that express the relationships among T , P , V , and n are known as the *gas laws*. Because volume is easily measured, the first gas laws to be studied expressed the effect of one of the variables on volume with the remaining two variables held constant.



◀ **Figure 10.5 An application of the pressure-volume relationship.** The volume of gas in this weather balloon will increase as it ascends into the high atmosphere, where the atmospheric pressure is lower than on Earth's surface.

The Pressure-Volume Relationship: Boyle's Law

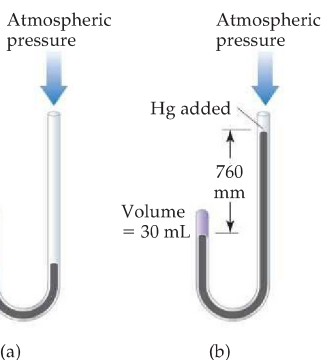
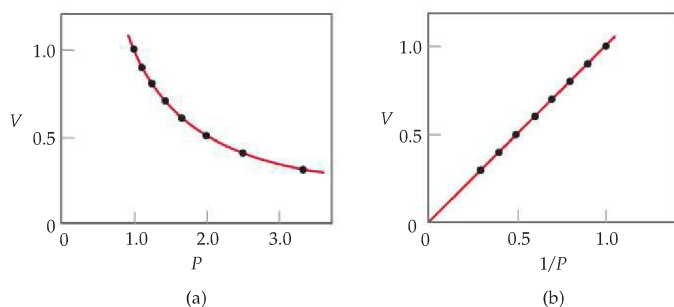
If the pressure on a balloon is decreased, the balloon expands. That is why weather balloons expand as they rise through the atmosphere (Figure 10.5 ▲). Conversely, when a volume of gas is compressed, the pressure of the gas increases. British chemist Robert Boyle (1627–1691) first investigated the relationship between the pressure of a gas and its volume.

To perform his gas experiments, Boyle used a J-shaped tube like that shown in Figure 10.6 ▶. In the tube on the left, a quantity of gas is trapped above a column of mercury. Boyle changed the pressure on the gas by adding mercury to the tube. He found that the volume of the gas decreased as the pressure increased. For example, doubling the pressure caused the gas volume to decrease to half its original value.

Boyle's law, which summarizes these observations, states that *the volume of a fixed quantity of gas maintained at constant temperature is inversely proportional to the pressure*. When two measurements are inversely proportional, one gets smaller as the other gets larger. Boyle's law can be expressed mathematically as

$$V = \text{constant} \times \frac{1}{P} \quad \text{or} \quad PV = \text{constant} \quad [10.2]$$

The value of the constant depends on the temperature and the amount of gas in the sample. The graph of V versus P in Figure 10.7(a) ▼ shows the type of curve obtained for a given quantity of gas at a fixed temperature. A linear relationship is obtained when V is plotted versus $1/P$ [Figure 10.7(b)].

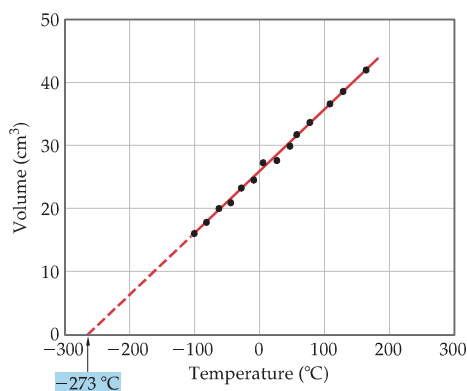


▲ **Figure 10.6 An illustration of Boyle's experiment relating pressure and volume.** In (a) the volume of the gas trapped in the J-tube is 60 mL when the gas pressure is 760 torr. When additional mercury is added, as shown in (b), the trapped gas is compressed. The volume is 30 mL when its total pressure is 1520 torr, corresponding to atmospheric pressure plus the pressure exerted by the 760-mm column of mercury.

◀ **Figure 10.7 Graphs based on Boyle's law.** (a) Volume versus pressure. (b) Volume versus $1/P$.



▲ **Figure 10.8** An illustration of the effect of temperature on volume. As liquid nitrogen (-196°C) is poured over a balloon, the gas in the balloon is cooled and its volume decreases.



▲ **Figure 10.9** Graph based on Charles's law. At constant pressure, the volume of an enclosed gas increases as the temperature increases. The dashed line is an extrapolation to temperatures at which the substance is no longer a gas.

Simple though it is, Boyle's law occupies a special place in the history of science. Boyle was the first to carry out a series of experiments in which one variable was systematically changed to determine the effect on another variable. The data from the experiments were then employed to establish an empirical relationship—a "law."

We apply Boyle's law every time we breathe. The rib cage, which can expand and contract, and the diaphragm, a muscle beneath the lungs, govern the volume of the lungs. Inhalation occurs when the rib cage expands and the diaphragm moves downward. Both of these actions increase the volume of the lungs, thus decreasing the gas pressure inside the lungs. The atmospheric pressure then forces air into the lungs until the pressure in the lungs equals atmospheric pressure. Exhalation reverses the process—the rib cage contracts and the diaphragm moves up, both of which decrease the volume of the lungs. Air is forced out of the lungs by the resulting increase in pressure.

GIVE IT SOME THOUGHT

What happens to the volume of a gas if you double its pressure, say from 1 atm to 2 atm, while its temperature is held constant?

The Temperature-Volume Relationship: Charles's Law

Hot-air balloons rise because air expands as it is heated. The warm air in the balloon is less dense than the surrounding cool air at the same pressure. This difference in density causes the balloon to ascend. Conversely, a balloon will shrink when the gas in it is cooled, as seen in Figure 10.8 ◀.

The relationship between gas volume and temperature was discovered in 1787 by the French scientist Jacques Charles (1746–1823). Charles found that the volume of a fixed quantity of gas at constant pressure increases linearly with temperature. Some typical data are shown in Figure 10.9 ◀. Notice that the extrapolated (extended) line (which is dashed) passes through -273°C . Note also that the gas is predicted to have zero volume at this temperature. This condition is never realized, however, because all gases liquefy or solidify before reaching this temperature.

In 1848 William Thomson (1824–1907), a British physicist whose title was Lord Kelvin, proposed an absolute-temperature scale, now known as the Kelvin scale. On this scale 0 K, which is called *absolute zero*, equals -273.15°C . ∞ (Section 1.4) In terms of the Kelvin scale, **Charles's law** can be stated as follows: *The volume of a fixed amount of gas maintained at constant pressure is directly proportional to its absolute temperature.* Thus, doubling the absolute temperature, say from 200 K to 400 K, causes the gas volume to double. Mathematically, Charles's law takes the following form:

$$V = \text{constant} \times T \quad \text{or} \quad \frac{V}{T} = \text{constant} \quad [10.3]$$

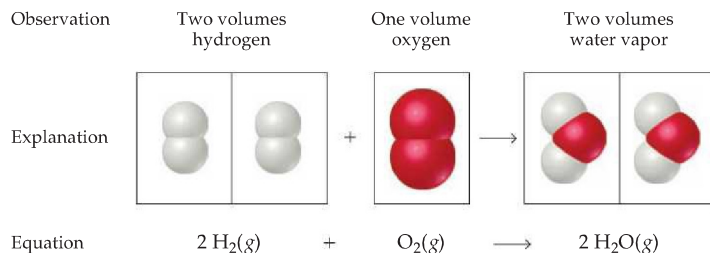
The value of the constant depends on the pressure and amount of gas.

GIVE IT SOME THOUGHT

Does the volume of a fixed quantity of gas decrease to half its original value when the temperature is lowered from 100°C to 50°C ?

The Quantity-Volume Relationship: Avogadro's Law

As we add gas to a balloon, the balloon expands. The volume of a gas is affected by pressure, temperature, and the amount of gas. The relationship between the quantity of a gas and its volume follows from the work of Joseph Louis Gay-Lussac (1778–1823) and Amedeo Avogadro (1776–1856).



◀ **Figure 10.10 The law of combining volumes.** Gay-Lussac's experimental observation of combining volumes is shown together with Avogadro's explanation of this phenomenon.




Gay-Lussac is one of those extraordinary figures in the history of science who could truly be called an adventurer. He was interested in lighter-than-air balloons, and in 1804 he made an ascent to 23,000 ft—an exploit that held the altitude record for several decades. To better control lighter-than-air balloons, Gay-Lussac carried out several experiments on the properties of gases. In 1808 he observed the *law of combining volumes*: At a given pressure and temperature, the volumes of gases that react with one another are in the ratios of small whole numbers. For example, two volumes of hydrogen gas react with one volume of oxygen gas to form two volumes of water vapor, as shown in Figure 10.10 ▲.

Three years later Amedeo Avogadro interpreted Gay-Lussac's observation by proposing what is now known as **Avogadro's hypothesis**: *Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.* For example, experiments show that 22.4 L of any gas at 0 °C and 1 atm contain 6.02×10^{23} gas molecules (that is, 1 mol), as depicted in Figure 10.11 ►.

Avogadro's law follows from Avogadro's hypothesis: *The volume of a gas maintained at constant temperature and pressure is directly proportional to the number of moles of the gas.* That is,

$$V = \text{constant} \times n \quad [10.4]$$

Thus, doubling the number of moles of gas will cause the volume to double if T and P remain constant.

			
Volume	22.4 L	22.4 L	22.4 L
Pressure	1 atm	1 atm	1 atm
Temperature	0 °C	0 °C	0 °C
Mass of gas	4.00 g	28.0 g	16.0 g
Number of gas molecules	6.02×10^{23}	6.02×10^{23}	6.02×10^{23}

▲ **Figure 10.11 A comparison illustrating Avogadro's hypothesis.** Note that helium gas consists of helium atoms. Each gas has the same volume, temperature, and pressure and thus contains the same number of molecules. Because a molecule of one substance differs in mass from a molecule of another, the masses of gas in the three containers differ.

■ SAMPLE EXERCISE 10.3 | Evaluating the Effects of Changes in P , V , n , and T on a Gas

Suppose we have a gas confined to a cylinder as shown in Figure 10.12 ►. Consider the following changes: (a) Heat the gas from 298 K to 360 K, while maintaining the piston in the position shown in the drawing. (b) Move the piston to reduce the volume of gas from 1 L to 0.5 L. (c) Inject additional gas through the gas inlet valve. Indicate how each of these changes will affect the average distance between molecules, the pressure of the gas, and the number of moles of gas present in the cylinder.

SOLUTION

Analyze: We need to think how each of three different changes in the system affects (1) the distance between molecules, (2) the pressure of the gas, and (3) the number of moles of gas in the cylinder.

Plan: We will use our understanding of the gas laws and the general properties of gases to analyze each situation.

Solve:

(a) Heating the gas while maintaining the position of the piston will cause no change in the number of molecules per unit volume. Thus, the distance between molecules and the total moles of gas remain the same. The increase in temperature, however, will cause the pressure to increase (Charles's law).



▲ **Figure 10.12 Cylinder with piston and gas inlet valve.**

(b) Moving the piston compresses the same quantity of gas into a smaller volume. The total number of molecules of gas, and thus the total number of moles, remains the same. The average distance between molecules, however, must decrease because of the smaller volume in which the gas is confined. The reduction in volume causes the pressure to increase (Boyle's law).

(c) Injecting more gas into the cylinder while keeping the volume and temperature the same will result in more molecules and thus a greater number of moles of gas. The average distance between atoms must decrease because their number per unit volume increases. Correspondingly, the pressure increases (Avogadro's law).

■ PRACTICE EXERCISE

What happens to the density of a gas as (a) the gas is heated in a constant-volume container; (b) the gas is compressed at constant temperature; (c) additional gas is added to a constant-volume container?

Answer: (a) no change, (b) increase, (c) increase

10.4 THE IDEAL-GAS EQUATION

In Section 10.3 we examined three historically important gas laws that describe the relationships between the four variables, P , V , T , and n , that define the state of a gas. Each law was obtained by holding two variables constant to see how the remaining two variables affect each other. We can express each law as a proportionality relationship. Using the symbol \propto , which is read "is proportional to," we have

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

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

$$\text{Avogadro's law: } V \propto n \quad (\text{constant } P, T)$$

We can combine these relationships to make a more general gas law.

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

If we call the proportionality constant R , we obtain

$$V = R \left(\frac{nT}{P} \right)$$

Rearranging, we have this relationship in its more familiar form:

$$PV = nRT \quad [10.5]$$

This equation is known as the **ideal-gas equation**. An **ideal gas** is a hypothetical gas whose pressure, volume, and temperature behavior are described completely by the ideal-gas equation.

The term R in the ideal-gas equation is called the **gas constant**. The value and units of R depend on the units of P , V , n , and T . Temperature must *always* be expressed as an absolute temperature when used in the ideal-gas equation. The quantity of gas, n , is normally expressed in moles. The units chosen for pressure and volume are most often atm and liters, respectively. However, other units can be used. In most countries other than the United States, the SI unit of Pa (or kPa) is most commonly employed. Table 10.2 shows the numerical value for R in various units. As we saw in the "A Closer Look" box on P - V work in Section 5.3, the product PV has the units of energy. Therefore, the units of R can include joules or calories. In working problems with the ideal-gas equation, the units of P , V , n , and T must agree with the units in the gas constant. In this chapter we will most often use the value $R = 0.08206$ L-atm/mol-K (four significant figures) or 0.0821 L-atm/mol-K (three significant figures) whenever we use the ideal-gas equation, consistent with the units of atm for pressure. Use of the value $R = 8.314$ J/mol-K, consistent with the units of Pa for pressure, is also very common.

TABLE 10.2 ■ Numerical Values of the Gas Constant, R , in Various Units

Units	Numerical Value
L-atm/mol-K	0.08206
J/mol-K*	8.314
cal/mol-K	1.987
m ³ -Pa/mol-K*	8.314
L-torr/mol-K	62.36

*SI unit

Suppose we have 1.000 mol of an ideal gas at 1.000 atm and 0.00 °C (273.15 K). According to the ideal-gas equation, the volume of the gas is

$$V = \frac{nRT}{P} = \frac{(1.000 \text{ mol})(0.08206 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(273.15 \text{ K})}{1.000 \text{ atm}} = 22.41 \text{ L}$$

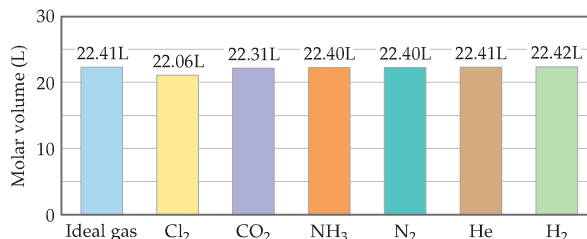
The conditions 0 °C and 1 atm are referred to as the **standard temperature and pressure (STP)**. Many properties of gases are tabulated for these conditions. The volume occupied by one mole of ideal gas at STP, 22.41 L, is known as the *molar volume* of an ideal gas at STP.

GIVE IT SOME THOUGHT

How many molecules are in 22.41 L of an ideal gas at STP?

The ideal-gas equation accounts adequately for the properties of most gases under a wide variety of circumstances. The equation is not exactly correct, however, for any real gas. Thus, the measured volume, V , for given conditions of P , n , and T , might differ from the volume calculated from $PV = nRT$. To illustrate, the measured molar volumes of real gases at STP are compared with the calculated volume of an ideal gas in Figure 10.13. While these real gases do not match the ideal gas behavior exactly, the differences are so small that we can ignore them for all but the most accurate work. We will have more to say about the differences between ideal and real gases in Section 10.9.

▼ **Figure 10.13 Comparison of molar volumes at STP.** One mole of an ideal gas at STP occupies a volume of 22.41 L. One mole of various real gases at STP occupies close to this ideal volume.



SAMPLE EXERCISE 10.4 | Using the Ideal-Gas Equation

Calcium carbonate, $\text{CaCO}_3(\text{s})$, decomposes upon heating to give $\text{CaO}(\text{s})$ and $\text{CO}_2(\text{g})$. A sample of CaCO_3 is decomposed, and the carbon dioxide is collected in a 250-mL flask. After the decomposition is complete, the gas has a pressure of 1.3 atm at a temperature of 31 °C. How many moles of CO_2 gas were generated?

SOLUTION

Analyze: We are given the volume (250 mL), pressure (1.3 atm), and temperature 31 °C of a sample of CO_2 gas and asked to calculate the number of moles of CO_2 in the sample.

Plan: Because we are given V , P , and T , we can solve the ideal-gas equation for the unknown quantity, n .

Solve: In analyzing and solving gas-law problems, it is helpful to tabulate the information given in the problems and then to convert the values to units that are consistent with those for R (0.0821 L·atm/mol·K). In this case the given values are

$$\begin{aligned} V &= 250 \text{ mL} = 0.250 \text{ L} \\ P &= 1.3 \text{ atm} \\ T &= 31 \text{ }^\circ\text{C} = (31 + 273) \text{ K} = 304 \text{ K} \end{aligned}$$

Remember: *Absolute temperature must always be used when the ideal-gas equation is solved.*

We now rearrange the ideal-gas equation (Equation 10.5) to solve for n

$$\begin{aligned} n &= \frac{PV}{RT} \\ n &= \frac{(1.3 \text{ atm})(0.250 \text{ L})}{(0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(304 \text{ K})} = 0.013 \text{ mol CO}_2 \end{aligned}$$

Check: Appropriate units cancel, thus ensuring that we have properly rearranged the ideal-gas equation and have converted to the correct units.

PRACTICE EXERCISE

Tennis balls are usually filled with air or N_2 gas to a pressure above atmospheric pressure to increase their “bounce.” If a particular tennis ball has a volume of 144 cm^3 and contains 0.33 g of N_2 gas, what is the pressure inside the ball at 24 °C?

Answer: 2.0 atm

In chemistry and throughout your studies of science and math, you may encounter problems that involve several experimentally measured variables as well as several different physical constants. In this chapter we encounter a variety of problems based on the ideal-gas equation, which consists of four experimental quantities— P , V , n , and T —and one constant, R . Depending on the type of problem, we might need to solve for any of the four quantities.

To avoid any difficulty extracting the necessary information from problems when many variables are involved, we suggest you take the following steps as you analyze, plan, and solve such problems:

1. *Tabulate information.* Read the problems carefully to determine which quantity is the unknown and which quantities are given. Every time you encounter a numerical value, jot it down. In many cases, constructing a table of the given information will be useful.
2. *Convert to consistent units.* As you have already seen, we can use several different units to express the same quantity. Make certain that quantities are converted to the proper units by using the correct conversion factors. In using the ideal-gas equation, for example, we usually use the value of R that has units of L-atm/mol-K. If you are given a pressure in torr, you will need to convert it to atmospheres.
3. *If a single equation relates the variables, rearrange the equation to solve for the unknown.* Make certain that you are comfortable using algebra to solve the equation for the desired variable. In the case of the ideal-gas equation the following algebraic rearrangements will all be used at one time or another:

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

4. *Use dimensional analysis.* Carry the units through your calculation. Use of dimensional analysis enables you to check that you have solved the equation correctly. If the units of the quantities in the equation cancel properly to give the units of the desired variable, you have probably used the equation correctly.

Sometimes you will not be given values for the necessary variables directly. Rather, you will be given the values of other quantities that can be used to determine the needed variables. For example, suppose you are trying to use the ideal-gas equation to calculate the pressure of a gas. You are given the temperature of the gas, but you are not given explicit values for n and V . However, the problem states that “the sample of gas contains 0.15 mol of gas per liter.” We can turn this statement into the expression

$$\frac{n}{V} = 0.15 \text{ mol/L}$$

Solving the ideal-gas equation for pressure yields

$$P = \frac{nRT}{V} = \left(\frac{n}{V}\right)RT$$

Thus, we can solve the equation even though we are not given specific values for n and V . We will examine how to use the density and molar mass of a gas in this fashion in Section 10.5.

As we have continuously stressed, the most important thing you can do to become proficient at solving problems is to practice by using practice exercises and assigned exercises at the end of each chapter. By using systematic procedures, such as those described here, you should be able to minimize difficulties in solving problems involving many variables.

Relating the Ideal-Gas Equation and the Gas Laws

The simple gas laws that we discussed in Section 10.3, such as Boyle’s law, are special cases of the ideal-gas equation. For example, when the quantity of gas and the temperature are held constant, n and T have fixed values. Therefore, the product nRT is the product of three constants and must itself be a constant.

$$PV = nRT = \text{constant} \quad \text{or} \quad PV = \text{constant} \quad [10.6]$$

Thus, we have Boyle’s law. We see that if n and T are constant, the individual values of P and V can change, but the product PV must remain constant.

We can use Boyle’s law to determine how the volume of a gas changes when its pressure changes. For example, if a metal cylinder holds 50.0 L of O_2 gas at 18.5 atm and 21 °C, what volume will the gas occupy if the temperature is maintained at 21 °C while the pressure is reduced to 1.00 atm? Because the product PV is a constant when a gas is held at constant n and T , we know that

$$P_1V_1 = P_2V_2 \quad [10.7]$$

where P_1 and V_1 are initial values and P_2 and V_2 are final values. Dividing both sides of this equation by P_2 gives the final volume, V_2 .

$$V_2 = V_1 \times \frac{P_1}{P_2}$$

Substituting the given quantities into this equation gives

$$V_2 = (50.0 \text{ L}) \left(\frac{18.5 \text{ atm}}{1.00 \text{ atm}} \right) = 925 \text{ L}$$

The answer is reasonable because gases expand as their pressures are decreased.

In a similar way we can start with the ideal-gas equation and derive relationships between any other two variables, V and T (Charles's law), n and V (Avogadro's law), or P and T . Sample Exercise 10.5 illustrates how these relationships can be derived and used.

■ SAMPLE EXERCISE 10.5 | Calculating the Effect of Temperature Changes on Pressure

The gas pressure in an aerosol can is 1.5 atm at 25 °C. Assuming that the gas inside obeys the ideal-gas equation, what would the pressure be if the can were heated to 450 °C?

SOLUTION

Analyze: We are given the initial pressure (1.5 atm) and temperature (25 °C) of the gas and asked for the pressure at a higher temperature (450 °C).

Plan: The volume and number of moles of gas do not change, so we must use a relationship connecting pressure and temperature. Converting temperature to the Kelvin scale and tabulating the given information, we have

	P	T
INITIAL	1.5 atm	298 K
FINAL	P_2	723 K

Solve: To determine how P and T are related, we start with the ideal-gas equation and isolate the quantities that do not change (n , V , and R) on one side and the variables (P and T) on the other side.

$$\frac{P}{T} = \frac{nR}{V} = \text{constant}$$

Because the quotient P/T is a constant, we can write

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

(where the subscripts 1 and 2 represent the initial and final states, respectively). Rearranging to solve for P_2 and substituting the given data give

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

$$P_2 = (1.5 \text{ atm}) \left(\frac{723 \text{ K}}{298 \text{ K}} \right) = 3.6 \text{ atm}$$

Check: This answer is intuitively reasonable—increasing the temperature of a gas increases its pressure.

Comment: It is evident from this example why aerosol cans carry a warning not to incinerate.

■ PRACTICE EXERCISE

A large natural-gas storage tank is arranged so that the pressure is maintained at 2.20 atm. On a cold day in December when the temperature is -15 °C (4 °F), the volume of gas in the tank is $3.25 \times 10^3 \text{ m}^3$. What is the volume of the same quantity of gas on a warm July day when the temperature is 31 °C (88 °F)?

Answer: $3.83 \times 10^3 \text{ m}^3$

We are often faced with the situation in which P , V , and T all change for a fixed number of moles of gas. Because n is constant under these circumstances, the ideal-gas equation gives

$$\frac{PV}{T} = nR = \text{constant}$$

If we represent the initial and final conditions of pressure, temperature, and volume by subscripts 1 and 2, respectively, we can write

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad [10.8]$$

This equation is often called the *combined gas law*.

SAMPLE EXERCISE 10.6 | Calculating the Effect of Changing P and T on the Volume of a Gas

An inflated balloon has a volume of 6.0 L at sea level (1.0 atm) and is allowed to ascend in altitude until the pressure is 0.45 atm. During ascent the temperature of the gas falls from 22 °C to -21 °C. Calculate the volume of the balloon at its final altitude.

SOLUTION

Analyze: We need to determine a new volume for a gas sample in a situation where both pressure and temperature change.

Plan: Let's again proceed by converting temperature to the Kelvin scale and tabulating the given information.

	P	V	T
INITIAL	1.0 atm	6.0 L	295 K
FINAL	0.45 atm	V_2	252 K

Because n is constant, we can use Equation 10.8.

Solve: Rearranging Equation 10.8 to solve for V_2 gives

$$V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1} = (6.0 \text{ L}) \left(\frac{1.0 \text{ atm}}{0.45 \text{ atm}} \right) \left(\frac{252 \text{ K}}{295 \text{ K}} \right) = 11 \text{ L}$$

Check: The result appears reasonable. Notice that the calculation involves multiplying the initial volume by a ratio of pressures and a ratio of temperatures. Intuitively, we expect that decreasing pressure will cause the volume to increase. Similarly, decreasing temperature should cause the volume to decrease. Note that the difference in pressures is more dramatic than the difference in temperatures. Thus, we should expect the effect of the pressure change to predominate in determining the final volume, as it does.

PRACTICE EXERCISE

A 0.50-mol sample of oxygen gas is confined at 0 °C in a cylinder with a movable piston, such as that shown in Figure 10.12. The gas has an initial pressure of 1.0 atm. The piston then compresses the gas so that its final volume is half the initial volume. The final pressure of the gas is 2.2 atm. What is the final temperature of the gas in degrees Celsius?

Answer: 27 °C

10.5 FURTHER APPLICATIONS OF THE IDEAL-GAS EQUATION

The ideal-gas equation can be used to determine many relationships involving the physical properties of gases. In this section we use it first to define the relationship between the density of a gas and its molar mass, and then to calculate the volumes of gases formed or consumed in chemical reactions.

Gas Densities and Molar Mass

The ideal-gas equation allows us to calculate gas density from the molar mass, pressure, and temperature of the gas. Recall that density has the units of mass per unit volume ($d = m/V$). $\infty\infty$ (Section 1.4) We can arrange the gas equation to obtain similar units, moles per unit volume, n/V :

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

If we multiply both sides of this equation by the molar mass, \mathcal{M} , which is the number of grams in one mole of a substance, we obtain the following relationship:

$$\frac{n\mathcal{M}}{V} = \frac{P\mathcal{M}}{RT} \quad [10.9]$$

The product of the quantities n/V and M equals the density in g/L, as seen from their units:

$$\frac{\text{moles}}{\text{liter}} \times \frac{\text{grams}}{\text{mole}} = \frac{\text{grams}}{\text{liter}}$$

Thus, the density, d , of the gas is given by the expression on the right in Equation 10.9:

$$d = \frac{P.M}{RT} \quad [10.10]$$

From Equation 10.10 we see that the density of a gas depends on its pressure, molar mass, and temperature. The higher the molar mass and pressure, the more dense the gas. The higher the temperature, the less dense the gas. Although gases form homogeneous mixtures regardless of their identities, a less dense gas will lie above a more dense gas in the absence of mixing. For example, CO_2 has a higher molar mass than N_2 or O_2 and is therefore more dense than air. When CO_2 is released from a CO_2 fire extinguisher, as shown in Figure 10.14, it blankets a fire, preventing O_2 from reaching the combustible material. The fact that a hotter gas is less dense than a cooler one explains why hot air rises. The difference between the densities of hot and cold air is responsible for the lift of hot-air balloons. It is also responsible for many phenomena in weather, such as the formation of large thunderhead clouds during thunderstorms.



▲ **Figure 10.14 A CO_2 fire extinguisher.** The CO_2 gas from a fire extinguisher is denser than air. The CO_2 cools significantly as it emerges from the tank. Water vapor in the air is condensed by the cool CO_2 gas and forms a white fog accompanying the colorless CO_2 .

GIVE IT SOME THOUGHT

Is water vapor more or less dense than N_2 under the same conditions of temperature and pressure?

SAMPLE EXERCISE 10.7 | Calculating Gas Density

What is the density of carbon tetrachloride vapor at 714 torr and 125 °C?

SOLUTION

Analyze: We are asked to calculate the density of a gas given its name, its pressure, and its temperature. From the name we can write the chemical formula of the substance and determine its molar mass.

Plan: We can use Equation 10.10 to calculate the density. Before we can use that equation, however, we need to convert the given quantities to the appropriate units. We must convert temperature to the Kelvin scale and pressure to atmospheres. We must also calculate the molar mass of CCl_4 .

Solve: The temperature on the Kelvin scale is $125 + 273 = 398 \text{ K}$.
The pressure in atmospheres is $(714 \text{ torr})(1 \text{ atm}/760 \text{ torr}) = 0.939 \text{ atm}$.
The molar mass of CCl_4 is $12.0 + (4)(35.5) = 154.0 \text{ g/mol}$.
Using these quantities along with Equation 10.10, we have

$$d = \frac{(0.939 \text{ atm})(154.0 \text{ g/mol})}{(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(398 \text{ K})} = 4.43 \text{ g/L}$$

Check: If we divide the molar mass (g/mol) by the density (g/L), we end up with L/mol. The numerical value is roughly $154/4.4 = 35$. That is in the right ballpark for the molar volume of a gas heated to 125 °C at near atmospheric pressure, so our answer is reasonable.

PRACTICE EXERCISE

The mean molar mass of the atmosphere at the surface of Titan, Saturn's largest moon, is 28.6 g/mol. The surface temperature is 95 K, and the pressure is 1.6 atm. Assuming ideal behavior, calculate the density of Titan's atmosphere.

Answer: 5.9 g/L

Equation 10.10 can be rearranged to solve for the molar mass of a gas:

$$\mathcal{M} = \frac{dRT}{P} \quad [10.11]$$

Thus, we can use the experimentally measured density of a gas to determine the molar mass of the gas molecules, as shown in Sample Exercise 10.8.

SAMPLE EXERCISE 10.8 | Calculating the Molar Mass of a Gas

A series of measurements are made to determine the molar mass of an unknown gas. First, a large flask is evacuated and found to weigh 134.567 g. It is then filled with the gas to a pressure of 735 torr at 31 °C and reweighed. Its mass is now 137.456 g. Finally, the flask is filled with water at 31 °C and found to weigh 1067.9 g. (The density of the water at this temperature is 0.997 g/mL.) Assume that the ideal-gas equation applies, and calculate the molar mass of the unknown gas.

SOLUTION

Analyze: We are given the temperature (31 °C) and pressure (735 torr) for a gas, together with information to determine its volume and mass, and we are asked to calculate its molar mass.

Plan: We need to use the mass information given to calculate the volume of the container and the mass of the gas within it. From this we calculate the gas density and then apply Equation 10.11 to calculate the molar mass of the gas.

Solve: The mass of the gas is the difference between the mass of the flask filled with gas and that of the empty (evacuated) flask:

$$137.456 \text{ g} - 134.567 \text{ g} = 2.889 \text{ g}$$

The volume of the gas equals the volume of water that the flask can hold. The volume of water is calculated from its mass and density. The mass of the water is the difference between the masses of the full and empty flask:

$$1067.9 \text{ g} - 134.567 \text{ g} = 933.3 \text{ g}$$

By rearranging the equation for density ($d = m/V$), we have

$$V = \frac{m}{d} = \frac{(933.3 \text{ g})}{(0.997 \text{ g/mL})} = 936 \text{ mL} = 0.936 \text{ L}$$

Knowing the mass of the gas (2.889 g) and its volume (936 mL), we can calculate the density of the gas:

$$2.889 \text{ g}/0.936 \text{ L} = 3.09 \text{ g/L}$$

After converting pressure to atmospheres and temperature to kelvins, we can use Equation 10.11 to calculate the molar mass:

$$\begin{aligned} \mathcal{M} &= \frac{dRT}{P} \\ &= \frac{(3.09 \text{ g/L})(0.0821 \text{ L}\cdot\text{atm/mol}\cdot\text{K})(304 \text{ K})}{(735/760) \text{ atm}} \\ &= 79.7 \text{ g/mol} \end{aligned}$$

Check: The units work out appropriately, and the value of molar mass obtained is reasonable for a substance that is gaseous near room temperature.

PRACTICE EXERCISE

Calculate the average molar mass of dry air if it has a density of 1.17 g/L at 21 °C and 740.0 torr.

Answer: 29.0 g/mol

Volumes of Gases in Chemical Reactions

We are often concerned with knowing the identity of a gas involved as a reactant or product in a chemical reaction, as well as its quantity. Thus, it is useful to be able to calculate the volumes of gases consumed or produced in reactions. Such calculations are based on the use of the mole concept together with balanced chemical equations. ∞ (Section 3.6) We have seen that the coefficients in balanced chemical equations tell us the relative amounts (in moles) of reactants and products in a reaction. The ideal gas equation relates the number of moles of a gas, to P , V , and T .

Chemistry Put to Work

GAS PIPELINES

Most people are quite unaware of the vast network of underground pipelines that undergirds the developed world. Pipelines are used to move massive quantities of liquids and gases over considerable distances. For example, pipelines move natural gas (methane) from huge natural-gas fields in Siberia to Western Europe. Natural gas from Algeria is moved to Italy through a pipeline 120 cm in diameter and 2500 km in length that stretches across the Mediterranean Sea at depths up to 600 m. In the United States the pipeline systems consist of trunk lines, large-diameter pipes for long-distance transport, with branch lines of lower diameter and lower pressure for local transport to and from the trunk lines.

Essentially all substances that are gases at STP are transported commercially by pipeline, including ammonia, carbon dioxide, carbon monoxide, chlorine, ethane, helium, hydrogen, and methane. The largest volume transport by far, though, is natural gas. (Figure 10.15) The methane-rich gas from oil and gas wells is processed to remove particulates, water, and various gaseous impurities such as hydrogen sulfide and carbon dioxide. The gas is then compressed to pressures ranging from 3.5 MPa (35 atm) to 10 MPa (100 atm), depending on the age and diameter of the pipe. The long-distance pipelines are about 40 cm in diameter and made of steel. Large compressor stations along the pipeline, spaced at 50- to 100-mile intervals, maintain pressure.

Recall from Figure 5.24 that natural gas is a major source of energy for the United States. To meet this demand, methane must be transported from source wells throughout the United



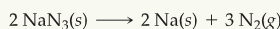
▲ Figure 10.15 A natural gas pipeline junction.

States and Canada to all parts of the nation. The total length of pipeline for natural-gas transport in the United States is about 6×10^5 km, and growing. The United States is divided into seven regions. The total deliverability of natural gas to the seven regions exceeds 2.7×10^{12} L (measured at STP), which is almost 100 billion cubic feet per day! The volume of the pipelines themselves would be entirely inadequate for managing the enormous quantities of natural gas that are placed into and taken out of the system on a continuing basis. For this reason, underground storage facilities, such as salt caverns and other natural formations, are employed to hold large quantities of gas.

Related Exercise: 10.117

SAMPLE EXERCISE 10.9 | Relating the Volume of a Gas to the Amount of Another Substance in a Reaction

The safety air bags in automobiles are inflated by nitrogen gas generated by the rapid decomposition of sodium azide, NaN_3 :



If an air bag has a volume of 36 L and is to be filled with nitrogen gas at a pressure of 1.15 atm at a temperature of 26.0 °C, how many grams of NaN_3 must be decomposed?

SOLUTION

Analyze: This is a multistep problem. We are given the volume, pressure, and temperature of the N_2 gas and the chemical equation for the reaction by which the N_2 is generated. We must use this information to calculate the number of grams of NaN_3 needed to obtain the necessary N_2 .

Plan: We need to use the gas data (P , V , and T) and the ideal-gas equation to calculate the number of moles of N_2 gas that should be formed for the air bag to operate correctly. We can then use the balanced equation to determine the number of moles of NaN_3 . Finally, we can convert the moles of NaN_3 to grams.

Solve: The number of moles of N_2 is determined using the ideal-gas equation:

$$n = \frac{PV}{RT} = \frac{(1.15 \text{ atm})(36 \text{ L})}{(0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(299 \text{ K})} = 1.7 \text{ mol N}_2$$

From here we use the coefficients in the balanced equation to calculate the number of moles of NaN_3 .

$$(1.7 \text{ mol N}_2) \left(\frac{2 \text{ mol NaN}_3}{3 \text{ mol N}_2} \right) = 1.1 \text{ mol NaN}_3$$

Finally, using the molar mass of NaN_3 , we convert moles of NaN_3 to grams:

$$(1.1 \text{ mol NaN}_3) \left(\frac{65.0 \text{ g NaN}_3}{1 \text{ mol NaN}_3} \right) = 72 \text{ g NaN}_3$$



Check: The best way to check our approach is to make sure the units cancel properly at each step in the calculation, leaving us with the correct units in the answer, g NaN₃.

■ PRACTICE EXERCISE

In the first step in the industrial process for making nitric acid, ammonia reacts with oxygen in the presence of a suitable catalyst to form nitric oxide and water vapor:



How many liters of NH₃(g) at 850 °C and 5.00 atm are required to react with 1.00 mol of O₂(g) in this reaction?

Answer: 14.8 L

10.6 GAS MIXTURES AND PARTIAL PRESSURES

Thus far we have considered mainly the behavior of pure gases—those that consist of only one substance in the gaseous state. How do we deal with gases composed of a mixture of two or more different substances? While studying the properties of air, John Dalton (Section 2.1) made an important observation: *The total pressure of a mixture of gases equals the sum of the pressures that each would exert if it were present alone.* The pressure exerted by a particular component of a mixture of gases is called the **partial pressure** of that gas. Dalton's observation is known as **Dalton's law of partial pressures**.



GIVE IT SOME THOUGHT

How is the pressure exerted by N₂ gas affected when some O₂ is introduced into a container if the temperature and volume remain constant?

If we let P_t be the total pressure of a mixture of gases and P_1, P_2, P_3 , and so forth be the partial pressures of the individual gases, we can write Dalton's law as follows:

$$P_t = P_1 + P_2 + P_3 + \cdots \quad [10.12]$$

This equation implies that each gas in the mixture behaves independently of the others, as we can see by the following analysis. Let n_1, n_2, n_3 , and so forth be the number of moles of each of the gases in the mixture and n_t be the total number of moles of gas ($n_t = n_1 + n_2 + n_3 + \cdots$).

If each of the gases obeys the ideal-gas equation, we can write

$$P_1 = n_1 \left(\frac{RT}{V} \right); \quad P_2 = n_2 \left(\frac{RT}{V} \right); \quad P_3 = n_3 \left(\frac{RT}{V} \right); \quad \text{and so forth}$$

All the gases in the mixture are at the same temperature and occupy the same volume. Therefore, by substituting into Equation 10.12, we obtain

$$P_t = (n_1 + n_2 + n_3 + \cdots) \frac{RT}{V} = n_t \left(\frac{RT}{V} \right) \quad [10.13]$$

That is, at constant temperature and constant volume the total pressure is determined by the total number of moles of gas present, whether that total represents just one substance or a mixture.

SAMPLE EXERCISE 10.10 | Applying Dalton's Law of Partial Pressures

A gaseous mixture made from 6.00 g O₂ and 9.00 g CH₄ is placed in a 15.0-L vessel at 0 °C. What is the partial pressure of each gas, and what is the total pressure in the vessel?

SOLUTION

Analyze: We need to calculate the pressure for two different gases in the same volume and at the same temperature.

Plan: Because each gas behaves independently, we can use the ideal-gas equation to calculate the pressure that each would exert if the other were not present. The total pressure is the sum of these two partial pressures.

Solve: We must first convert the mass of each gas to moles:

$$n_{\text{O}_2} = (6.00 \text{ g O}_2) \left(\frac{1 \text{ mol O}_2}{32.0 \text{ g O}_2} \right) = 0.188 \text{ mol O}_2$$

$$n_{\text{CH}_4} = (9.00 \text{ g CH}_4) \left(\frac{1 \text{ mol CH}_4}{16.0 \text{ g CH}_4} \right) = 0.563 \text{ mol CH}_4$$

We can now use the ideal-gas equation to calculate the partial pressure of each gas:

$$P_{\text{O}_2} = \frac{n_{\text{O}_2} RT}{V} = \frac{(0.188 \text{ mol})(0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(273 \text{ K})}{15.0 \text{ L}} = 0.281 \text{ atm}$$

$$P_{\text{CH}_4} = \frac{n_{\text{CH}_4} RT}{V} = \frac{(0.563 \text{ mol})(0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(273 \text{ K})}{15.0 \text{ L}} = 0.841 \text{ atm}$$

According to Dalton's law (Equation 10.12), the total pressure in the vessel is the sum of the partial pressures:

$$P_t = P_{\text{O}_2} + P_{\text{CH}_4} = 0.281 \text{ atm} + 0.841 \text{ atm} = 1.122 \text{ atm}$$

Check: Performing rough estimates is good practice, even when you may not feel that you need to do it to check an answer. In this case a pressure of roughly 1 atm seems right for a mixture of about 0.2 mol O₂ (that is, 6/32) and a bit more than 0.5 mol CH₄ (that is, 9/16), together in a 15-L volume, because one mole of an ideal gas at 1 atm pressure and 0 °C occupies about 22 L.

PRACTICE EXERCISE

What is the total pressure exerted by a mixture of 2.00 g of H₂ and 8.00 g of N₂ at 273 K in a 10.0-L vessel?

Answer: 2.86 atm

Partial Pressures and Mole Fractions

Because each gas in a mixture behaves independently, we can relate the amount of a given gas in a mixture to its partial pressure. For an ideal gas, $P = nRT/V$ and so we can write

$$\frac{P_1}{P_t} = \frac{n_1 RT/V}{n_t RT/V} = \frac{n_1}{n_t} \quad [10.14]$$

The ratio n_1/n_t is called the mole fraction of gas 1, which we denote X_1 . The **mole fraction**, X , is a dimensionless number that expresses the ratio of the number of moles of one component to the total number of moles in the mixture. We can rearrange Equation 10.14 to give

$$P_1 = \left(\frac{n_1}{n_t} \right) P_t = X_1 P_t \quad [10.15]$$

Thus, the partial pressure of a gas in a mixture is its mole fraction times the total pressure.

The mole fraction of N₂ in air is 0.78 (that is, 78% of the molecules in air are N₂). If the total barometric pressure is 760 torr, then the partial pressure of N₂ is

$$P_{\text{N}_2} = (0.78) (760 \text{ torr}) = 590 \text{ torr}$$

This result makes intuitive sense: Because N₂ makes up 78% of the mixture, it contributes 78% of the total pressure.

SAMPLE EXERCISE 10.11 | Relating Mole Fractions and Partial Pressures

A study of the effects of certain gases on plant growth requires a synthetic atmosphere composed of 1.5 mol percent CO_2 , 18.0 mol percent O_2 , and 80.5 mol percent Ar. (a) Calculate the partial pressure of O_2 in the mixture if the total pressure of the atmosphere is to be 745 torr. (b) If this atmosphere is to be held in a 121-L space at 295 K, how many moles of O_2 are needed?

SOLUTION

Analyze: (a) We first need to calculate the partial pressure of O_2 given its mole percentage and the total pressure of the mixture. (b) We need to calculate the number of moles of O_2 in the mixture given its volume (121 L), temperature (745 torr), and partial pressure (from part (a)).

Plan: (a) We will calculate the partial pressures using Equation 10.15. (b) We will then use P_{O_2} , V , and T together with the ideal-gas equation to calculate the number of moles of O_2 , n_{O_2} .

Solve: (a) The mole percent is just the mole fraction times 100. Therefore, the mole fraction of O_2 is 0.180. Using Equation 10.15, we have

$$P_{\text{O}_2} = (0.180)(745 \text{ torr}) = 134 \text{ torr}$$

(b) Tabulating the given variables and changing them to appropriate units, we have

$$P_{\text{O}_2} = (134 \text{ torr}) \left(\frac{1 \text{ atm}}{760 \text{ torr}} \right) = 0.176 \text{ atm}$$

$$V = 121 \text{ L}$$

$$n_{\text{O}_2} = ?$$

$$R = 0.0821 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$$

$$T = 295 \text{ K}$$

Solving the ideal-gas equation for n_{O_2} , we have

$$\begin{aligned} n_{\text{O}_2} &= P_{\text{O}_2} \left(\frac{V}{RT} \right) \\ &= (0.176 \text{ atm}) \frac{121 \text{ L}}{(0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(295 \text{ K})} = 0.879 \text{ mol} \end{aligned}$$

Check: The units check out satisfactorily, and the answer seems to be the right order of magnitude.

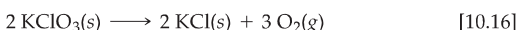
PRACTICE EXERCISE

From data gathered by *Voyager 1*, scientists have estimated the composition of the atmosphere of Titan, Saturn's largest moon. The total pressure on the surface of Titan is 1220 torr. The atmosphere consists of 82 mol percent N_2 , 12 mol percent Ar, and 6.0 mol percent CH_4 . Calculate the partial pressure of each of these gases in Titan's atmosphere.

Answer: 1.0×10^3 torr N_2 , 1.5×10^2 torr Ar, and 73 torr CH_4

Collecting Gases over Water

An experiment that is often encountered in general chemistry laboratories involves determining the number of moles of gas collected from a chemical reaction. Sometimes this gas is collected over water. For example, solid potassium chlorate, KClO_3 , can be decomposed by heating it in a test tube in an arrangement such as that shown in Figure 10.16. The balanced equation for the reaction is

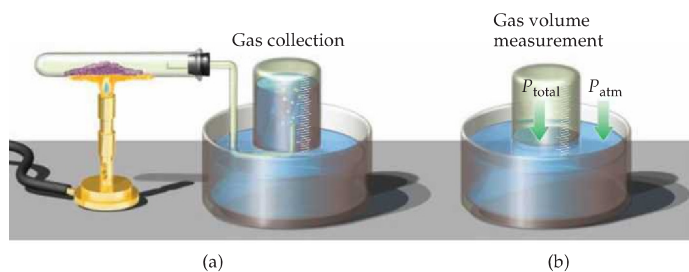


The oxygen gas is collected in a bottle that is initially filled with water and then inverted in a water pan.

The volume of gas collected is measured by raising or lowering the bottle as necessary until the water levels inside and outside the bottle are the same. When this condition is met, the pressure inside the bottle is equal to the atmospheric pressure outside. The total pressure inside is the sum of the pressure of gas collected and the pressure of water vapor in equilibrium with liquid water.

$$P_{\text{total}} = P_{\text{gas}} + P_{\text{H}_2\text{O}} \quad [10.17]$$

The pressure exerted by water vapor, $P_{\text{H}_2\text{O}}$, at various temperatures is listed in Appendix B.



◀ **Figure 10.16 Collecting a water-insoluble gas over water.** (a) A solid is heated, releasing a gas, which is bubbled through water into a collection bottle. (b) When the gas has been collected, the bottle is raised or lowered so that the water levels inside and outside the bottle are equal. The total pressure of the gases inside the bottle is then equal to the atmospheric pressure.

SAMPLE EXERCISE 10.12 | Calculating the Amount of Gas Collected over Water

A sample of KClO_3 is partially decomposed (Equation 10.16), producing O_2 gas that is collected over water as in Figure 10.16. The volume of gas collected is 0.250 L at 26°C and 765 torr total pressure. (a) How many moles of O_2 are collected? (b) How many grams of KClO_3 were decomposed?

SOLUTION

(a) Analyze: We need to calculate the number of moles of O_2 gas in a container that also contains water vapor.

Plan: If we tabulate the information presented, we will see that values are given for V and T . To use the ideal-gas equation to calculate the unknown, n_{O_2} , we also must know the partial pressure of O_2 in the system. We can calculate the partial pressure of O_2 from the total pressure (765 torr) and the vapor pressure of water.

Solve: The partial pressure of the O_2 gas is the difference between the total pressure, 765 torr, and the pressure of the water vapor at 26°C , 25 torr (Appendix B):

$$P_{\text{O}_2} = 765 \text{ torr} - 25 \text{ torr} = 740 \text{ torr}$$

We can use the ideal-gas equation to calculate the number of moles of O_2 :

$$n_{\text{O}_2} = \frac{P_{\text{O}_2} V}{RT} = \frac{(740 \text{ torr})(1 \text{ atm}/760 \text{ torr})(0.250 \text{ L})}{(0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K})(299 \text{ K})} = 9.92 \times 10^{-3} \text{ mol O}_2$$

(b) Analyze: We now need to calculate the number of moles of reactant KClO_3 decomposed.

Plan: We can use the number of moles of O_2 formed and the balanced chemical equation to determine the number of moles of KClO_3 decomposed, which we can then convert to grams of KClO_3 .

Solve: From Equation 10.16, we have $2 \text{ mol KClO}_3 \approx 3 \text{ mol O}_2$. The molar mass of KClO_3 is 122.6 g/mol. Thus, we can convert the moles of O_2 that we found in part (a) to moles of KClO_3 and then to grams of KClO_3 :

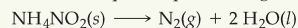
$$(9.92 \times 10^{-3} \text{ mol O}_2) \left(\frac{2 \text{ mol KClO}_3}{3 \text{ mol O}_2} \right) \left(\frac{122.6 \text{ g KClO}_3}{1 \text{ mol KClO}_3} \right) = 0.811 \text{ g KClO}_3$$

Check: As always, we make sure that the units cancel appropriately in the calculations. In addition, the numbers of moles of O_2 and KClO_3 seem reasonable, given the small volume of gas collected.

Comment: Many chemical compounds that react with water and water vapor would be degraded by exposure to wet gas. Thus, in research laboratories gases are often dried by passing wet gas over a substance that absorbs water (a *desiccant*), such as calcium sulfate, CaSO_4 . Calcium sulfate crystals are sold as a desiccant under the trade name Drierite™.

PRACTICE EXERCISE

Ammonium nitrite, NH_4NO_2 , decomposes upon heating to form N_2 gas:



When a sample of NH_4NO_2 is decomposed in a test tube, as in Figure 10.16, 511 mL of N_2 gas is collected over water at 26°C and 745 torr total pressure. How many grams of NH_4NO_2 were decomposed?

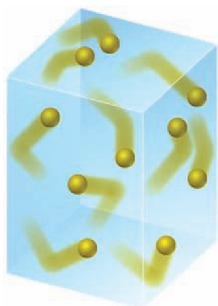
Answer: 1.26 g

10.7 KINETIC-MOLECULAR THEORY

The ideal-gas equation describes *how* gases behave, but it does not explain *why* they behave as they do. Why does a gas expand when heated at constant pressure? Or, why does its pressure increase when the gas is compressed at constant temperature? To understand the physical properties of gases, we need a model that helps us picture what happens to gas particles as experimental conditions such as pressure or temperature change. Such a model, known as the **kinetic-molecular theory**, was developed over a period of about 100 years, culminating in 1857 when Rudolf Clausius (1822–1888) published a complete and satisfactory form of the theory.

The kinetic-molecular theory (the theory of moving molecules) is summarized by the following statements:

1. Gases consist of large numbers of molecules that are in continuous, random motion. (The word *molecule* is used here to designate the smallest particle of any gas; some gases, such as the noble gases, consist of individual atoms.)
2. The combined volume of all the molecules of the gas is negligible relative to the total volume in which the gas is contained.
3. Attractive and repulsive forces between gas molecules are negligible.
4. Energy can be transferred between molecules during collisions, but the *average* kinetic energy of the molecules does not change with time, as long as the temperature of the gas remains constant. In other words, the collisions are perfectly elastic.
5. The average kinetic energy of the molecules is proportional to the absolute temperature. At any given temperature the molecules of all gases have the same average kinetic energy.



▲ **Figure 10.17** The molecular origin of gas pressure. The pressure exerted by a gas is caused by collisions of the gas molecules with the walls of their container.

The kinetic-molecular theory explains both pressure and temperature at the molecular level. The pressure of a gas is caused by collisions of the molecules with the walls of the container, as shown in Figure 10.17 ◀. The magnitude of the pressure is determined by both how often and how forcefully the molecules strike the walls.

The absolute temperature of a gas is a measure of the *average* kinetic energy of its molecules. If two different gases are at the same temperature, their molecules have the same average kinetic energy (statement 5 of the kinetic-molecular theory). If the absolute temperature of a gas is doubled, the average kinetic energy of its molecules doubles. Thus, molecular motion increases with increasing temperature.

GIVE IT SOME THOUGHT

Consider three samples of gas: HCl at 298 K, H₂ at 298 K, and O₂ at 350 K. Compare the average kinetic energies of the molecules in the three samples.

Distributions of Molecular Speed

Although the molecules in a sample of gas have an *average* kinetic energy and hence an average speed, the individual molecules move at varying speeds. The moving molecules collide frequently with other molecules. Momentum is conserved in each collision, but one of the colliding molecules might be deflected off at high speed while the other is nearly stopped. The result is that the molecules at any instant have a wide range of speeds. Figure 10.18 ► illustrates the distribution of molecular speeds for nitrogen gas at 0 °C (blue line) and at 100 °C (red line). The curve shows us the fraction of molecules moving at each

speed. At higher temperatures, a larger fraction of molecules moves at greater speeds; the distribution curve has shifted to the right toward higher speeds and hence toward higher average kinetic energy. The peak of each curve represents the most probable speed (the speed of the largest number of molecules). Notice that the blue curve (0 °C) has a peak at about 4×10^2 m/s, whereas the red curve (100 °C) has a peak at a higher speed, about 5×10^2 m/s.

Figure 10.18 also shows the value of the **root-mean-square (rms) speed**, u , of the molecules at each temperature. This quantity is the speed of a molecule possessing average kinetic energy. The rms speed is not quite the same as the average (mean) speed. The difference between the two, however, is small.* Notice that the rms speed is higher at 100 °C than at 0 °C. Notice also that the distribution curve broadens as we go to a higher temperature.

The rms speed is important because the average kinetic energy of the gas molecules in a sample, ϵ , is related directly to u^2 :

$$\epsilon = \frac{1}{2} mu^2 \quad [10.18]$$

where m is the mass of an individual molecule. Mass does not change with temperature. Thus, the increase in the average kinetic energy as the temperature increases implies that the rms speed (and also the average speed) of molecules likewise increases as temperature increases.

GIVE IT SOME THOUGHT

Consider three samples of gas all at 298 K: HCl, H₂, and O₂. List the molecules in order of increasing average speed.

Application to the Gas Laws

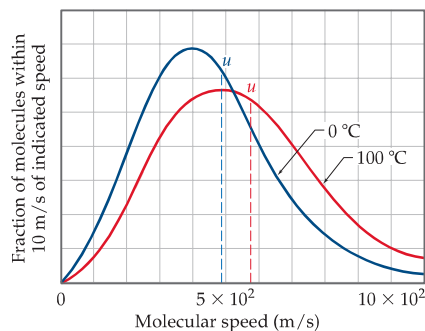
The empirical observations of gas properties as expressed by the various gas laws are readily understood in terms of the kinetic-molecular theory. The following examples illustrate this point:

1. *Effect of a volume increase at constant temperature:* A constant temperature means that the average kinetic energy of the gas molecules remains unchanged. This in turn means that the rms speed of the molecules, u , is unchanged. If the volume is increased, however, the molecules must move a longer distance between collisions. Consequently, there are fewer collisions per unit time with the container walls, and pressure decreases. Thus, the model accounts in a simple way for Boyle's law.
2. *Effect of a temperature increase at constant volume:* An increase in temperature means an increase in the average kinetic energy of the molecules and thus an increase in u . If there is no change in volume, there will be more collisions with the walls per unit time. Furthermore, the change in momentum in each collision increases (the molecules strike the walls more forcefully). Hence, the model explains the observed pressure increase.

*To illustrate the difference between rms speed and average speed, suppose that we have four objects with speeds of 4.0, 6.0, 10.0, and 12.0 m/s. Their average speed is $\frac{1}{4}(4.0 + 6.0 + 10.0 + 12.0) = 8.0$ m/s. The rms speed, u , however, is the square root of the average squared speeds of the molecules:

$$\sqrt{\frac{1}{4}(4.0^2 + 6.0^2 + 10.0^2 + 12.0^2)} = \sqrt{74.0} = 8.6 \text{ m/s}$$

For an ideal gas the average speed equals $0.921 \times u$. Thus, the average speed is directly proportional to the rms speed, and the two are in fact nearly equal.



▲ **Figure 10.18** The effect of temperature on molecular speeds. Distribution of molecular speeds for nitrogen at 0 °C (blue line) and 100 °C (red line). Increasing temperature increases both the most probable speed (curve maximum) and the rms speed, u , which is indicated by the vertical dashed line.

A Closer Look

THE IDEAL-GAS EQUATION

Beginning with the five statements given in the text for the kinetic-molecular theory, it is possible to derive the ideal-gas equation. Rather than proceed through a derivation, let's consider in somewhat qualitative terms how the ideal-gas equation might follow. As we have seen, pressure is force per unit area. [∞∞](#) (Section 10.2) The total force of the molecular collisions on the walls and hence the pressure produced by these collisions depend both on how strongly the molecules strike the walls (impulse imparted per collision) and on the rate at which these collisions occur:

$$P \propto \text{impulse imparted per collision} \times \text{rate of collisions}$$

For a molecule traveling at the rms speed, u , the impulse imparted by a collision with a wall depends on the momentum of the molecule; that is, it depends on the product of its mass and speed, mu . The rate of collisions is proportional to both the number of molecules per unit volume, n/V , and their speed, u . If there are more molecules in a container, there will be more frequent collisions with the container walls. As the molecular speed increases or the volume of the container decreases, the time required for molecules to traverse the distance from one wall to another is reduced, and the molecules collide more frequently with the walls. Thus, we have

$$P \propto mu \times \frac{n}{V} \times u \propto \frac{nm u^2}{V} \quad [10.19]$$

Because the average kinetic energy, $\frac{1}{2} m u^2$, is proportional to temperature, we have $m u^2 \propto T$. Making this substitution into Equation 10.19 gives

$$P \propto \frac{n(mu^2)}{V} \propto \frac{nT}{V} \quad [10.20]$$

Let's now convert the proportionality sign to an equal sign by expressing n as the number of moles of gas. We then insert a proportionality constant, R , the gas constant:

$$P = \frac{nRT}{V} \quad [10.21]$$

This expression is the ideal-gas equation.

An eminent Swiss mathematician, Daniel Bernoulli (1700–1782), conceived of a model for gases that was, for all practical purposes, the same as the kinetic theory model. From this model, Bernoulli derived Boyle's law and the ideal-gas equation. His was one of the first examples in science of developing a mathematical model from a set of assumptions, or hypothetical statements. However, Bernoulli's work on this subject was completely ignored, only to be rediscovered a hundred years later by Clausius and others. It was ignored because it conflicted with popular beliefs and was in conflict with Isaac Newton's incorrect model for gases. Those idols of the times had to fall before the way was clear for the kinetic-molecular theory. As this story illustrates, science is not a straight road running from here to the "truth." The road is built by humans, so it zigs and zags.

Related Exercises: 10.71, 10.72, 10.73, and 10.74

SAMPLE EXERCISE 10.13 | Applying the Kinetic-Molecular Theory

A sample of O_2 gas initially at STP is compressed to a smaller volume at constant temperature. What effect does this change have on (a) the average kinetic energy of O_2 molecules, (b) the average speed of O_2 molecules, (c) the total number of collisions of O_2 molecules with the container walls in a unit time, (d) the number of collisions of O_2 molecules with a unit area of container wall per unit time?

SOLUTION

Analyze: We need to apply the concepts of the kinetic-molecular theory to a situation in which a gas is compressed at constant temperature.

Plan: We will determine how each of the quantities in (a)–(d) is affected by the change in volume at constant temperature.

Solve: (a) The average kinetic energy of the O_2 molecules is determined only by temperature. Thus the average kinetic energy is unchanged by the compression of O_2 at constant temperature. (b) If the average kinetic energy of O_2 molecules does not change, the average speed remains constant. (c) The total number of collisions with the container walls per unit time must increase because the molecules are moving within a smaller volume but with the same average speed as before. Under these conditions they must encounter a wall more frequently. (d) The number of collisions with a unit area of wall per unit time increases because the total number of collisions with the walls per unit time increases and the area of the walls decreases.

Check: In a conceptual exercise of this kind, there is no numerical answer to check. All we can check in such cases is our reasoning in the course of solving the problem.

PRACTICE EXERCISE

How is the rms speed of N_2 molecules in a gas sample changed by (a) an increase in temperature, (b) an increase in volume, (c) mixing with a sample of Ar at the same temperature?

Answers: (a) increases, (b) no effect, (c) no effect

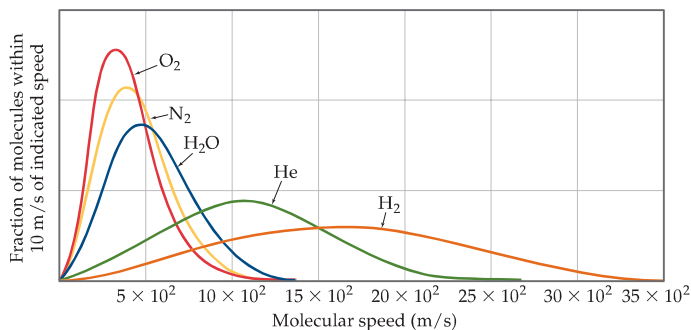
10.8 MOLECULAR EFFUSION AND DIFFUSION

According to the kinetic-molecular theory, the average kinetic energy of any collection of gas molecules, $\frac{1}{2}mu^2$, has a specific value at a given temperature. Thus, a gas composed of lightweight particles, such as He, will have the same average kinetic energy as one composed of much heavier particles, such as Xe, provided the two gases are at the same temperature. The mass, m , of the particles in the lighter gas is smaller than that in the heavier gas. Consequently, the particles of the lighter gas must have a higher rms speed, u , than the particles of the heavier one. The following equation, which expresses this fact quantitatively, can be derived from kinetic-molecular theory:

$$u = \sqrt{\frac{3RT}{\mathcal{M}}} \quad [10.22]$$

Because the molar mass, \mathcal{M} , appears in the denominator, the less massive the gas molecules, the higher the rms speed, u .

Figure 10.19 \blacktriangle shows the distribution of molecular speeds for several gases at 25 °C. Notice how the distributions are shifted toward higher speeds for gases of lower molar masses.



\blacktriangle **Figure 10.19 The effect of molecular mass on molecular speeds.** The distributions of molecular speeds for different gases are compared at 25 °C. The molecules with lower molecular masses have higher rms speeds.

SAMPLE EXERCISE 10.14 | Calculating a Root-Mean-Square Speed

Calculate the rms speed, u , of an N_2 molecule at 25 °C.

SOLUTION

Analyze: We are given the identity of the gas and the temperature, the two quantities we need to calculate the rms speed.

Plan: We will calculate the rms speed using Equation 10.22.

Solve: In using Equation 10.22, we should convert each quantity to SI units so that all the units are compatible. We will also use R in units of J/mol·K (Table 10.2) to make the units cancel correctly.

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

$$\mathcal{M} = 28.0 \text{ g/mol} = 28.0 \times 10^{-3} \text{ kg/mol}$$

$$R = 8.314 \text{ J/mol}\cdot\text{K} = 8.314 \text{ kg}\cdot\text{m}^2/\text{s}^2\cdot\text{mol}\cdot\text{K} \quad (\text{These units follow from the fact that } 1 \text{ J} = 1 \text{ kg}\cdot\text{m}^2/\text{s}^2)$$

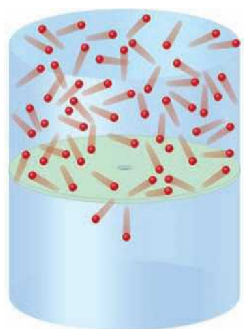
$$u = \sqrt{\frac{3(8.314 \text{ kg}\cdot\text{m}^2/\text{s}^2\cdot\text{mol}\cdot\text{K})(298 \text{ K})}{28.0 \times 10^{-3} \text{ kg/mol}}} = 5.15 \times 10^2 \text{ m/s}$$

Comment: This corresponds to a speed of 1150 mi/hr. Because the average molecular weight of air molecules is slightly greater than that of N_2 , the rms speed of air molecules is a little slower than that for N_2 . The speed at which sound propagates through air is about 350 m/s, a value about two-thirds the average rms speed for air molecules.

PRACTICE EXERCISE

What is the rms speed of an He atom at 25 °C?

Answer: $1.36 \times 10^3 \text{ m/s}$



▲ **Figure 10.20 Effusion.** The top half of this cylinder is filled with a gas, and the bottom half is an evacuated space. Gas molecules effuse through a pinhole in the partitioning wall only when they happen to hit the hole.

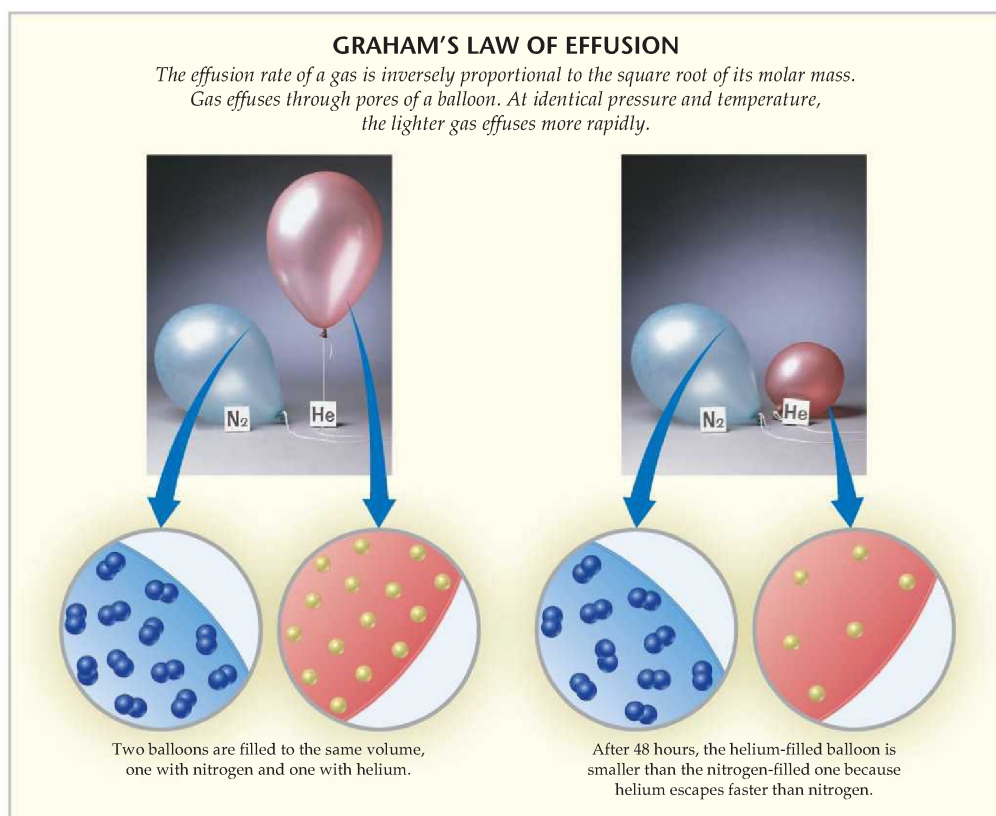
The dependence of molecular speeds on mass has several interesting consequences. The first phenomenon is **effusion**, which is the escape of gas molecules through a tiny hole into an evacuated space, as shown in Figure 10.20. The second is **diffusion**, which is the spread of one substance throughout a space or throughout a second substance. For example, the molecules of a perfume diffuse throughout a room.

Graham's Law of Effusion

In 1846 Thomas Graham (1805–1869) discovered that the effusion rate of a gas is inversely proportional to the square root of its molar mass. Assume that we have two gases at the same temperature and pressure in containers with identical pinholes. If the rates of effusion of the two substances are r_1 and r_2 and their respective molar masses are \mathcal{M}_1 and \mathcal{M}_2 , **Graham's law** states

$$\frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}} \quad [10.23]$$

Equation 10.23 compares the *rates* of effusion of two different gases under identical conditions, and it indicates that the lighter gas effuses more rapidly.



▲ **Figure 10.21** An illustration of Graham's law.

Figure 10.20 illustrates the basis of Graham's law. The only way for a molecule to escape from its container is for it to "hit" the hole in the partitioning wall. The faster the molecules are moving, the greater is the likelihood that a molecule will hit the hole and effuse. This implies that the rate of effusion is directly proportional to the rms speed of the molecules. Because R and T are constant, we have, from Equation 10.22

$$\frac{r_1}{r_2} = \frac{u_1}{u_2} = \sqrt{\frac{3RT/M_1}{3RT/M_2}} = \sqrt{\frac{M_2}{M_1}} \quad [10.24]$$

As expected from Graham's law, helium escapes from containers through tiny pinhole leaks more rapidly than other gases of higher molecular weight (Figure 10.21 ◀).

■ SAMPLE EXERCISE 10.15 | Applying Graham's Law

An unknown gas composed of homonuclear diatomic molecules effuses at a rate that is only 0.355 times that of O_2 at the same temperature. Calculate the molar mass of the unknown, and identify it.

■ SOLUTION

Analyze: We are given the rate of effusion of an unknown gas relative to that of O_2 , and we are asked to find the molar mass and identity of the unknown. Thus, we need to connect relative rates of effusion to relative molar masses.

Plan: We can use Graham's law of effusion, Equation 10.23, to determine the molar mass of the unknown gas. If we let r_x and M_x represent the rate of effusion and molar mass of the unknown gas, Equation 10.23 can be written as follows:

$$\frac{r_x}{r_{O_2}} = \sqrt{\frac{M_{O_2}}{M_x}}$$

Solve: From the information given,

$$r_x = 0.355 \times r_{O_2}$$

Thus,

$$\frac{r_x}{r_{O_2}} = 0.355 = \sqrt{\frac{32.0 \text{ g/mol}}{M_x}}$$

We now solve for the unknown molar mass, M_x

$$\begin{aligned} \frac{32.0 \text{ g/mol}}{M_x} &= (0.355)^2 = 0.126 \\ M_x &= \frac{32.0 \text{ g/mol}}{0.126} = 254 \text{ g/mol} \end{aligned}$$

Because we are told that the unknown gas is composed of homonuclear diatomic molecules, it must be an element. The molar mass must represent twice the atomic weight of the atoms in the unknown gas. We conclude that the unknown gas is I_2 .

■ PRACTICE EXERCISE

Calculate the ratio of the effusion rates of N_2 and O_2 , r_{N_2}/r_{O_2} .

Answer: $r_{N_2}/r_{O_2} = 1.07$

Diffusion and Mean Free Path

Diffusion, like effusion, is faster for lower mass molecules than for higher mass ones. In fact, Graham's law, Equation 10.23, approximates the ratio of rates of diffusion of two gases under identical experimental conditions. Nevertheless, molecular collisions make diffusion more complicated than effusion.

We can see from the horizontal scale in Figure 10.19 that the speeds of molecules are quite high. For example, the average speed of N_2 at room temperature is 515 m/s (1150 mi/hr). In spite of this high speed, if someone opens a vial of perfume at one end of a room, some time elapses—perhaps a few minutes—before the scent is detected at the other end of the room. The diffusion of gases is much slower than molecular speeds because of molecular collisions.* These

*The rate at which the perfume moves across the room also depends on how well stirred the air is from temperature gradients and the movement of people. Nevertheless, even with the aid of these factors, it still takes much longer for the molecules to traverse the room than one would expect from the rms speed alone.

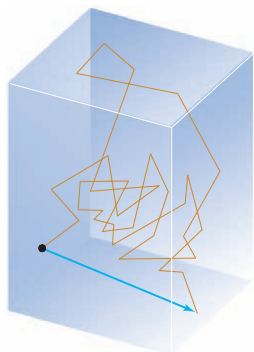
The fact that lighter molecules move at higher average speeds than more massive ones has many interesting consequences and applications. For example, the effort to develop the atomic bomb during World War II required scientists to separate the relatively low-abundance uranium isotope ^{235}U (0.7%) from the much more abundant ^{238}U (99.3%). This separation was accomplished by converting the uranium into a volatile compound, UF_6 , that was then allowed to pass through porous barriers. Because of the diameters of the pores, this process is not a simple effusion. Nevertheless, the dependence on molar mass is essentially the same. The slight difference in molar mass between the two hexafluorides, $^{235}\text{UF}_6$ and $^{238}\text{UF}_6$, caused the molecules to move at slightly different rates:

$$\frac{r_{235}}{r_{238}} = \sqrt{\frac{352.04}{349.03}} = 1.0043$$

Thus, the gas initially appearing on the opposite side of the barrier was very slightly enriched in the lighter molecule. The effusion process was repeated thousands of times, leading to a nearly complete separation of the two isotopes of uranium.

Separation of uranium isotopes by effusion has been largely replaced by a technique that uses centrifuges. In this procedure, cylindrical rotors containing UF_6 vapor spin at high speed inside an evacuated casing. Molecules containing the heavier ^{238}U isotope move closer to the spinning walls, whereas the molecules containing the lighter ^{235}U isotope remain in the middle of the cylinders. A stream of gas moves the UF_6 from the center of one centrifuge into another. Plants that use centrifuges consume less energy and can be constructed in a more compact, modular fashion than those that rely on effusion. Such plants are frequently in the news as countries such as Iran and North Korea enrich uranium in the ^{235}U isotope for both nuclear power and nuclear weaponry.

Related Exercises: 10.79 and 10.80



▲ **Figure 10.22 Diffusion of a gas molecule.** For clarity, no other gas molecules in the container are shown. The path of the molecule of interest begins at the dot. Each short segment of line represents travel between collisions. The blue arrow indicates the net distance traveled by the molecule.

collisions occur quite frequently for a gas at atmospheric pressure—about 10^{10} times per second for each molecule. Collisions occur because real gas molecules have finite volumes.

Because of molecular collisions, the direction of motion of a gas molecule is constantly changing. Therefore, the diffusion of a molecule from one point to another consists of many short, straight-line segments as collisions buffet it around in random directions, as depicted in Figure 10.22. First the molecule moves in one direction, then in another; one instant at high speed, the next at low speed.

The average distance traveled by a molecule between collisions is called the **mean free path** of the molecule. The mean free path varies with pressure as the following analogy illustrates. Imagine walking through a shopping mall. When the mall is very crowded (high pressure), the average distance you can walk before bumping into someone is short (short mean free path). When the mall is empty (low pressure), you can walk a long way (long mean free path) before bumping into someone. The mean free path for air molecules at sea level is about 60 nm (6×10^{-8} m). At about 100 km in altitude, where the air density is much lower, the mean free path is about 10 cm, over 1 million times longer than at Earth's surface.

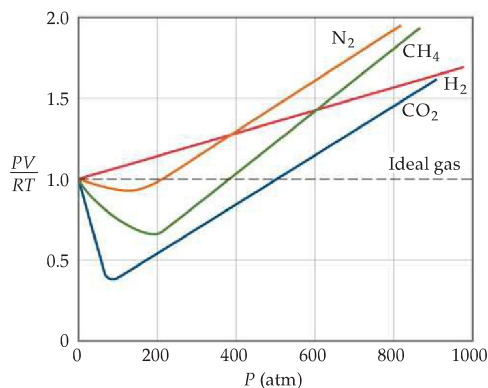
GIVE IT SOME THOUGHT

Will the following changes increase, decrease, or have no effect on the mean free path of the gas molecules in a sample of gas? (a) Increasing pressure, (b) increasing temperature?

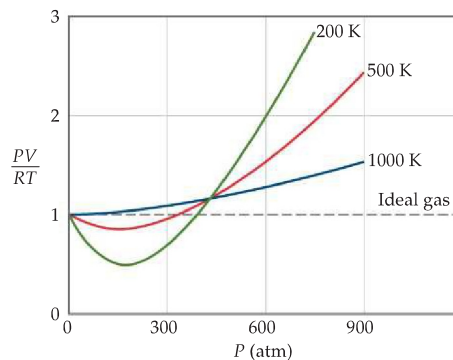
10.9 REAL GASES: DEVIATIONS FROM IDEAL BEHAVIOR

Although the ideal-gas equation is a very useful description of gases, all real gases fail to obey the relationship to some degree. The extent to which a real gas departs from ideal behavior can be seen by rearranging the ideal-gas equation to solve for n :

$$\frac{PV}{RT} = n \quad [10.25]$$



▲ **Figure 10.23** The effect of pressure on the behavior of several gases. The ratios of PV/RT versus pressure are compared for one mole of several gases at 300 K. The data for CO_2 are at 313 K because under high pressure CO_2 liquefies at 300 K. The dashed horizontal line shows the behavior of an ideal gas.



▲ **Figure 10.24** The effect of temperature and pressure on the behavior of nitrogen gas. The ratios of PV/RT versus pressure are shown for 1 mol of nitrogen gas at three temperatures. As temperature increases, the gas more closely approaches ideal behavior, which is represented by the dashed horizontal line.

For one mole of ideal gas ($n = 1$) the quantity PV/RT equals 1 at all pressures. In Figure 10.23 ▲ PV/RT is plotted as a function of P for one mole of several different gases. At high pressures the deviation from ideal behavior ($PV/RT = 1$) is large and is different for each gas. *Real gases, therefore, do not behave ideally at high pressure.* At lower pressures (usually below 10 atm), however, the deviation from ideal behavior is small, and we can use the ideal-gas equation without generating serious error.

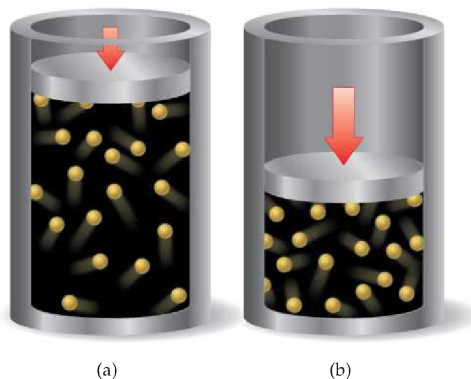
The deviation from ideal behavior also depends on temperature. Figure 10.24 ▲ shows graphs of PV/RT versus P for 1 mol of N_2 at three temperatures. As temperature increases, the behavior of the gas more nearly approaches that of the ideal gas. In general, *the deviations from ideal behavior increase as temperature decreases*, becoming significant near the temperature at which the gas is converted into a liquid.

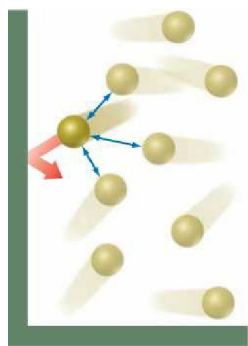
GIVE IT SOME THOUGHT

Would you expect helium gas to deviate from ideal behavior more at (a) 100 K and 1 atm, (b) 100 K and 5 atm, or (c) 300 K and 2 atm?

The basic assumptions of the kinetic-molecular theory give us insight into why real gases deviate from ideal behavior. The molecules of an ideal gas are assumed to occupy no space and have no attractions for one another. *Real molecules, however, do have finite volumes, and they do attract one another.* As shown in Figure 10.25 ►, the free, unoccupied space in which molecules can move is somewhat less than the container volume. At relatively low pressures the volume of the gas molecules is negligible relative to the container volume. Thus, the free volume available to the molecules is essentially the entire volume of the container. As the pressure increases, however, the free space in which the molecules can move becomes a smaller fraction of the container volume. Under these conditions, therefore, gas volumes tend to be slightly greater than those predicted by the ideal-gas equation.

▼ **Figure 10.25** Comparing the volume of gas molecules to the container volume. In (a), at low pressure, the combined volume of the gas molecules is small relative to the container volume, and we can approximate the empty space between molecules as being equal to the container volume. In (b), at high pressure, the combined volume of the gas molecules is a larger fraction of the total space available. Now we must account for the volume of the molecules themselves in determining the empty space available for the motion of the gas molecules.





▲ **Figure 10.26** The effect of intermolecular forces on gas pressure. The molecule that is about to strike the wall experiences attractive forces from nearby gas molecules, and its impact on the wall is thereby lessened. The lessened impact means the molecule exerts a lower-than-expected pressure on the wall. The attractive forces become significant only under high-pressure conditions, when the average distance between molecules is small.

In addition, the attractive forces between molecules come into play at short distances, as when molecules are crowded together at high pressures. Because of these attractive forces, the impact of a given molecule with the wall of the container is lessened. If we could stop the action in a gas, the positions of the molecules might resemble the illustration in Figure 10.26 ◀. The molecule about to make contact with the wall experiences the attractive forces of nearby molecules. These attractions lessen the force with which the molecule hits the wall. As a result, the pressure is less than that of an ideal gas. This effect serves to decrease PV/RT below its ideal value, as seen in Figure 10.23. When the pressure is sufficiently high, however, the volume effects dominate and PV/RT increases to above the ideal value.

Temperature determines how effective attractive forces between gas molecules are. As a gas is cooled, the average kinetic energy of the molecules decreases, but intermolecular attractions remain constant. In a sense, cooling a gas deprives molecules of the energy they need to overcome their mutual attractive influence. The effects of temperature shown in Figure 10.24 illustrate this point very well. As temperature increases, the negative departure of PV/RT from ideal-gas behavior disappears. The difference that remains at high temperature stems mainly from the effect of the finite volumes of the molecules.

GIVE IT SOME THOUGHT

List two reasons why gases deviate from ideal behavior.

The van der Waals Equation

Engineers and scientists who work with gases at high pressures often cannot use the ideal-gas equation to predict the pressure-volume properties of gases because departures from ideal behavior are too large. One useful equation developed to predict the behavior of real gases was proposed by the Dutch scientist Johannes van der Waals (1837–1923).

Van der Waals recognized that the ideal gas equation could be corrected to account for the effects of attractive forces between gas molecules and for molecular volumes. He introduced two constants, a and b , to make these corrections. The constant a is a measure of how strongly the gas molecules attract each other. The constant b is a measure of the small but finite volume occupied by the gas molecules themselves. His description of gas behavior is known as the **van der Waals equation**:

$$\left(P + \frac{n^2a}{V^2}\right)(V - nb) = nRT \quad [10.26]$$

In this equation the factor n^2a/V^2 accounts for the attractive forces. The van der Waals equation adjusts the pressure, P , upward by adding n^2a/V^2 because attractive forces between molecules tend to reduce the pressure (Figure 10.26). That is, a correction must be added to give the pressure that an ideal gas would have. The form of the correction factor, n^2a/V^2 , results because the attractive forces between pairs of molecules increase as the square of the number of molecules per unit volume, $(n/V)^2$.

The factor nb accounts for the small but finite volume occupied by the gas molecules themselves (Figure 10.25). The van der Waals equation adjusts volume, V , downward (subtracting nb), to give the free volume available to the gas molecules. That is, the particles of an ideal gas have the full volume, V , as free space in which to move, whereas in a real gas only the volume $V - nb$ is available as free space.

The constants a and b , which are called *van der Waals constants*, are experimentally determined positive quantities that differ for each gas. Values of these constants for several gases are listed in Table 10.3 ▶. Notice that the values of

TABLE 10.3 ■ van der Waals Constants for Gas Molecules

Substance	a (L ² -atm/mol ²)	b (L/mol)
He	0.0341	0.02370
Ne	0.211	0.0171
Ar	1.34	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0510
H ₂	0.244	0.0266
N ₂	1.39	0.0391
O ₂	1.36	0.0318
Cl ₂	6.49	0.0562
H ₂ O	5.46	0.0305
CH ₄	2.25	0.0428
CO ₂	3.59	0.0427
CCl ₄	20.4	0.1383

both a and b generally increase with an increase in mass of the molecule and with an increase in the complexity of its structure. Larger, more massive molecules have larger volumes and tend to have greater intermolecular attractive forces. As we will see in the next chapter, substances with large intermolecular attractive forces are relatively easy to liquefy.

■ SAMPLE EXERCISE 10.16 | Using the van der Waals Equation

If 1.000 mol of an ideal gas were confined to 22.41 L at 0.0 °C, it would exert a pressure of 1.000 atm. Use the van der Waals equation and the constants in Table 10.3 to estimate the pressure exerted by 1.000 mol of Cl₂(g) in 22.41 L at 0.0 °C.

SOLUTION

Analyze: The quantity we need to solve for is pressure. Because we will use the van der Waals equation, we must identify the appropriate values for the constants that appear there.

Plan: Solving Equation 10.26 for P , we have

$$P = \frac{nRT}{V - nb} - \frac{n^2a}{V^2}$$

Solve: Substituting $n = 1.000$ mol, $R = 0.08206$ L-atm/mol-K, $T = 273.2$ K, $V = 22.41$ L, $a = 6.49$ L²-atm/mol², and $b = 0.0562$ L/mol:

$$P = \frac{(1.000 \text{ mol})(0.08206 \text{ L-atm/mol-K})(273.2 \text{ K})}{22.41 \text{ L} - (1.000 \text{ mol})(0.0562 \text{ L/mol})} - \frac{(1.000 \text{ mol})^2(6.49 \text{ L}^2\text{-atm/mol}^2)}{(22.41 \text{ L})^2}$$

$$= 1.003 \text{ atm} - 0.013 \text{ atm} = 0.990 \text{ atm}$$

Check: We expect a pressure not far from 1.000 atm, which would be the value for an ideal gas, so our answer seems very reasonable.

Comment: Notice that the first term, 1.003 atm, is the pressure corrected for molecular volume. This value is higher than the ideal value, 1.000 atm, because the volume in which the molecules are free to move is smaller than the container volume, 22.41 L. Thus, the molecules must collide more frequently with the container walls. The second factor, 0.013 atm, corrects for intermolecular forces. The intermolecular attractions between molecules reduce the pressure to 0.990 atm. We can conclude, therefore, that the intermolecular attractions are the main cause of the slight deviation of Cl₂(g) from ideal behavior under the stated experimental conditions.

■ PRACTICE EXERCISE

Consider a sample of 1.000 mol of CO₂(g) confined to a volume of 3.000 L at 0.0 °C. Calculate the pressure of the gas using (a) the ideal-gas equation and (b) the van der Waals equation.

Answers: (a) 7.473 atm, (b) 7.182 atm

SAMPLE INTEGRATIVE EXERCISE | Putting Concepts Together

Cyanogen, a highly toxic gas, is composed of 46.2% C and 53.8% N by mass. At 25 °C and 751 torr, 1.05 g of cyanogen occupies 0.500 L. (a) What is the molecular formula of cyanogen? (b) Predict its molecular structure. (c) Predict the polarity of the compound.

SOLUTION

Analyze: First we need to determine the molecular formula of a compound from elemental analysis data and data on the properties of the gaseous substance. Thus, we have two separate calculations to do.

(a) Plan: We can use the percentage composition of the compound to calculate its empirical formula. [∞ \(Section 3.5\)](#) Then we can determine the molecular formula by comparing the mass of the empirical formula with the molar mass. [∞ \(Section 3.5\)](#)

Solve: To determine the empirical formula, we assume that we have a 100-g sample of the compound and then calculate the number of moles of each element in the sample:

$$\text{Moles C} = (46.2 \text{ g C}) \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 3.85 \text{ mol C}$$

$$\text{Moles N} = (53.8 \text{ g N}) \left(\frac{1 \text{ mol N}}{14.01 \text{ g N}} \right) = 3.84 \text{ mol N}$$

Because the ratio of the moles of the two elements is essentially 1:1, the empirical formula is CN.

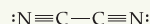
To determine the molar mass of the compound, we use Equation 10.11.

$$M = \frac{dRT}{p} = \frac{(1.05 \text{ g}/0.500 \text{ L}) (0.0821 \text{ L}\cdot\text{atm}/\text{mol}\cdot\text{K}) (298 \text{ K})}{(751/760) \text{ atm}} = 52.0 \text{ g/mol}$$

The molar mass associated with the empirical formula, CN, is 12.0 + 14.0 = 26.0 g/mol. Dividing the molar mass of the compound by that of its empirical formula gives (52.0 g/mol)/(26.0 g/mol) = 2.00. Thus, the molecule has twice as many atoms of each element as the empirical formula, giving the molecular formula C₂N₂.

(b) Plan: To determine the molecular structure of the molecule, we must first determine its Lewis structure. [∞ \(Section 8.5\)](#) We can then use the VSEPR model to predict the structure. [∞ \(Section 9.2\)](#)

Solve: The molecule has 2(4) + 2(5) = 18 valence-shell electrons. By trial and error, we seek a Lewis structure with 18 valence electrons in which each atom has an octet and in which the formal charges are as low as possible. The following structure meets these criteria:



(This structure has zero formal charge on each atom.)

The Lewis structure shows that each atom has two electron domains. (Each nitrogen has a nonbonding pair of electrons and a triple bond, whereas each carbon has a triple bond and a single bond.) Thus the electron-domain geometry around each atom is linear, causing the overall molecule to be linear.

(c) Plan: To determine the polarity of the molecule, we must examine the polarity of the individual bonds and the overall geometry of the molecule.

Solve: Because the molecule is linear, we expect the two dipoles created by the polarity in the carbon–nitrogen bond to cancel each other, leaving the molecule with no dipole moment.

CHAPTER REVIEW

SUMMARY AND KEY TERMS

Section 10.1 Substances that are gases at room temperatures tend to be molecular substances with low molar masses. Air, a mixture composed mainly of N_2 and O_2 , is the most common gas we encounter. Some liquids and solids can also exist in the gaseous state, where they are known as vapors. Gases are compressible; they mix in all proportions because their component molecules are far apart.

Section 10.2 To describe the state, or condition, of a gas, we must specify four variables: pressure (P), volume (V), temperature (T), and quantity (n). Volume is usually measured in liters, temperature in kelvins, and quantity of gas in moles. **Pressure** is the force per unit area. It is expressed in SI units as **pascals**, Pa ($1 \text{ Pa} = 1 \text{ N/m}^2 = 1 \text{ kg/m}\cdot\text{s}^2$). A related unit, the **bar**, equals 10^5 Pa. In chemistry, **standard atmospheric pressure** is used to define the **atmosphere** (atm) and the **torr** (also called the millimeter of mercury). One atmosphere of pressure equals 101.325 kPa, or 760 torr. A barometer is often used to measure the atmospheric pressure. A manometer can be used to measure the pressure of enclosed gases.

Sections 10.3 and 10.4 Studies have revealed several simple gas laws: For a constant quantity of gas at constant temperature, the volume of the gas is inversely proportional to the pressure (**Boyle's law**). For a fixed quantity of gas at constant pressure, the volume is directly proportional to its absolute temperature (**Charles's law**). Equal volumes of gases at the same temperature and pressure contain equal numbers of molecules (**Avogadro's hypothesis**). For a gas at constant temperature and pressure, the volume of the gas is directly proportional to the number of moles of gas (**Avogadro's law**). Each of these gas laws is a special case of the ideal-gas equation.

The **ideal-gas equation**, $PV = nRT$, is the equation of state for an **ideal gas**. The term R in this equation is the **gas constant**. We can use the ideal-gas equation to calculate variations in one variable when one or more of the others are changed. Most gases at pressures of about 1 atm and temperatures near 273 K and above obey the ideal-gas equation reasonably well. The conditions of 273 K (0°C) and 1 atm are known as the **standard temperature and pressure (STP)**. In all applications of the ideal-gas equation we must remember to convert temperatures to the absolute-temperature scale (the Kelvin scale).

Sections 10.5 and 10.6 Using the ideal-gas equation, we can relate the density of a gas to its molar mass: $\mathcal{M} = dRT/P$. We can also use the ideal-gas equation to solve problems involving gases as reactants or products in chemical reactions.

In gas mixtures the total pressure is the sum of the **partial pressures** that each gas would exert if it were present alone under the same conditions (**Dalton's law of partial pressures**). The partial pressure of a component of a mixture is equal to its mole fraction times the total pressure: $P_1 = X_1P_t$. The **mole fraction** is the ratio of the moles of one component of a mixture to the total moles of all components. In calculating the quantity of a gas collected over water, correction must be made for the partial pressure of water vapor in the gas mixture.

Section 10.7 The **kinetic-molecular theory** accounts for the properties of an ideal gas in terms of a set of statements about the nature of gases. Briefly, these statements are as follows: Molecules are in continuous chaotic motion. The volume of gas molecules is negligible compared to the volume of their container. The gas molecules have no attractive forces for one another. Their collisions are elastic. The average kinetic energy of the gas molecules is proportional to the absolute temperature.

The molecules of a gas do not all have the same kinetic energy at a given instant. Their speeds are distributed over a wide range; the distribution varies with the molar mass of the gas and with temperature. The **root-mean-square (rms) speed**, u , varies in proportion to the square root of the absolute temperature and **inversely** with the square root of the molar mass: $u = \sqrt{3RT/\mathcal{M}}$.

Section 10.8 It follows from kinetic-molecular theory that the rate at which a gas undergoes **effusion** (escapes through a tiny hole into a vacuum) is inversely proportional to the square root of its molar mass (**Graham's law**). The **diffusion** of one gas through the space occupied by a second gas is another phenomenon related to the speeds at which molecules move. Because molecules undergo frequent collisions with one another, the **mean free path**—the mean distance traveled between collisions—is short. Collisions between molecules limit the rate at which a gas molecule can diffuse.

Section 10.9 Departures from ideal behavior increase in magnitude as pressure increases and as temperature decreases. The extent of nonideality of a real gas can be seen by examining the quantity $PV = RT$ for one mole of the gas as a function of pressure; for an ideal gas, this quantity is exactly 1 at all pressures. Real gases depart from ideal behavior because the molecules possess finite volume and because the molecules experience attractive forces for one another. The **van der Waals equation** is an equation of state for gases that modifies the ideal-gas equation to account for intrinsic molecular volume and intermolecular forces.

KEY SKILLS

- Convert between pressure units with an emphasis on torr and atmospheres.
- Calculate P , V , n , or T using the ideal gas equation.
- Understand how the gas laws relate to the ideal gas equation and apply the gas laws in calculations.
- Calculate the density or molecular weight of a gas.
- Calculate the volume of gas used or formed in a chemical reaction.
- Calculate the total pressure of a gas mixture given its partial pressures or information for calculating partial pressures.
- Describe the kinetic molecular theory and how it explains the pressure and temperature of a gas, the gas laws, and the rates of effusion and diffusion.
- Explain why intermolecular attractions and molecular volumes cause real gases to deviate from ideal behavior at high pressure or low temperature.

KEY EQUATIONS

- $PV = nRT$ [10.5] Ideal gas equation
- $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ [10.8] The combined gas law, showing how P , V , and T are related for a constant n
- $d = \frac{P\mathcal{M}}{RT}$ [10.10] Calculating the density or molar mass of a gas
- $P_t = P_1 + P_2 + P_3 + \dots$ [10.12] Relating the total pressure of a gas mixture to the partial pressures of its components (Dalton's law of partial pressures)
- $P_i = \left(\frac{n_i}{n_t}\right) P_t = X_i P_t$ [10.15] Relating partial pressure to mole fraction
- $u = \sqrt{\frac{3RT}{\mathcal{M}}}$ [10.22] Relating the root-mean-square (rms) speed of gas molecules to temperature and molar mass
- $\frac{r_1}{r_2} = \sqrt{\frac{\mathcal{M}_2}{\mathcal{M}_1}}$ [10.23] Relating the relative rates of effusion of two gases to their molar masses

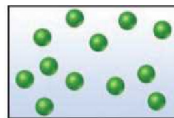
VISUALIZING CONCEPTS

10.1 Mars has an average atmospheric pressure of 0.007 atm. Would it be easier or harder to drink from a straw on Mars than on Earth? Explain. [Section 10.2]

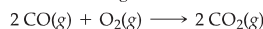
10.2 Assume that you have a sample of gas in a container with a movable piston, such as the one in the drawing. (a) Redraw the container to show what it might look like if the temperature of the gas is increased from 300 K to 500 K while the pressure is kept constant. (b) Redraw the container to show what it might look like if the pressure on the piston is increased from 1.0 atm to 2.0 atm while the temperature is kept constant. [Section 10.3]



10.3 Consider the sample of gas depicted below. What would the drawing look like if the volume and temperature remained constant while you removed enough of the gas to decrease the pressure by a factor of 2? [Section 10.3]

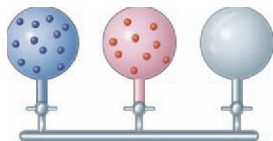


10.4 Consider the following reaction:

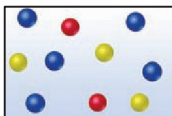


Imagine that this reaction occurs in a container that has a piston that moves to allow a constant pressure to be maintained when the reaction occurs at constant temperature. (a) What happens to the volume of the container as a result of the reaction? Explain. (b) If the piston is not allowed to move, what happens to the pressure as a result of the reaction? [Sections 10.3 and 10.5]

- 10.5** Consider the apparatus below, which shows gases in two containers and one empty container. When the stopcocks are opened and the gases are allowed to mix at constant temperature, what is the distribution of atoms in each container? Assume that the containers are of equal volume, and ignore the volume of the tubing connecting them. Which gas has the greater partial pressure after the stopcocks are opened? [Section 10.6]

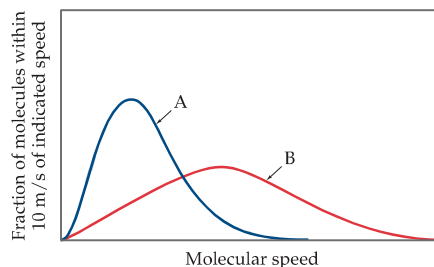


- 10.6** The drawing below represents a mixture of three different gases. (a) Rank the three components in order of increasing partial pressure. (b) If the total pressure of the mixture is 0.90 atm, calculate the partial pressure of each gas. [Section 10.6]

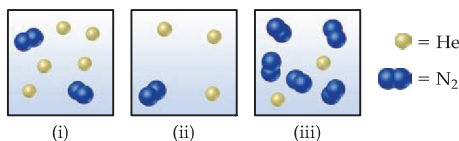


- 10.7** On a single plot, qualitatively sketch the distribution of molecular speeds for (a) Kr(g) at -50°C , (b) Kr(g) at 0°C , (c) Ar(g) at 0°C . [Section 10.7]
- 10.8** Consider the following drawing. (a) If the curves A and B refer to two different gases, He and O_2 at the same

temperature, which is which? Explain. (b) If A and B refer to the same gas at two different temperatures, which represents the higher temperature? [Section 10.7]



- 10.9** Consider the following samples of gases:



If the three samples are all at the same temperature, rank them with respect to (a) total pressure, (b) partial pressure of helium, (c) density, (d) average kinetic energy of particles. [Section 10.6 and 10.7]

- 10.10** Which of the substances listed in Table 10.3 would you expect to deviate the most from ideal-gas behavior at low temperature and high pressure? Which would deviate the least? Explain. [Section 10.9]

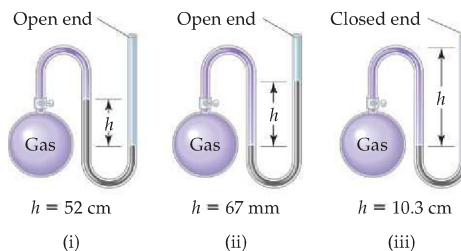
EXERCISES

Gas Characteristics; Pressure

- 10.11** How does a gas differ from a liquid with respect to each of the following properties: (a) density, (b) compressibility, (c) ability to mix with other substances of the same phase to form homogeneous mixtures?
- 10.12** (a) Both a liquid and a gas are moved to larger containers. How does their behavior differ? Explain the difference in molecular terms. (b) Although water and carbon tetrachloride, $\text{CCl}_4(l)$, do not mix, their vapors form homogeneous mixtures. Explain. (c) The densities of gases are generally reported in units of g/L, whereas those for liquids are reported as g/mL. Explain the molecular basis for this difference.
- 10.13** Suppose that a woman weighing 130 lb and wearing high-heeled shoes momentarily places all her weight on the heel of one foot. If the area of the heel is 0.50 in.^2 , calculate the pressure exerted on the underlying surface in kilopascals.
- 10.14** A set of bookshelves rests on a hard floor surface on four legs, each having a cross-sectional dimension of $3.0 \times 4.1\text{ cm}$ in contact with the floor. The total mass of the shelves plus the books stacked on them is 262 kg. Calculate the pressure in pascals exerted by the shelf footings on the surface.
- 10.15** (a) How high in meters must a column of water be to exert a pressure equal to that of a 760-mm column of mercury? The density of water is 1.0 g/mL , whereas that of mercury is 13.6 g/mL . (b) What is the pressure in atmospheres on the body of a diver if he is 39 ft below the surface of the water when atmospheric pressure at the surface is 0.97 atm ?
- 10.16** The compound 1-iodododecane is a nonvolatile liquid with a density of 1.20 g/mL . The density of mercury is 13.6 g/mL . What do you predict for the height of a barometer column based on 1-iodododecane, when the atmospheric pressure is 752 torr?
- 10.17** Each of the following statements concerns a mercury barometer such as that shown in Figure 10.2. Identify any incorrect statements, and correct them. (a) The tube must be 1 cm^2 in cross-sectional area. (b) At equilibrium the force of gravity per unit area acting on the mercury column at the level of the outside mercury equals the force of gravity per unit area acting on the atmosphere. (c) The column of mercury is held up by the vacuum at the top of the column.

- 10.18** Suppose you make a mercury barometer using a glass tube about 50 cm in length, closed at one end. What would you expect to see if the tube is filled with mercury and inverted in a mercury dish, as in Figure 10.2? Explain.
- 10.19** The typical atmospheric pressure on top of Mt. Everest (29,028 ft) is about 265 torr. Convert this pressure to (a) atm, (b) mm Hg, (c) pascals, (d) bars.
- 10.20** Perform the following conversions: (a) 0.850 atm to torr, (b) 785 torr to kilopascals, (c) 655 mm Hg to atmospheres, (d) 1.323×10^5 Pa to atmospheres, (e) 2.50 atm to bars.
- 10.21** In the United States, barometric pressures are generally reported in inches of mercury (in. Hg). On a beautiful summer day in Chicago the barometric pressure is 30.45 in. Hg. (a) Convert this pressure to torr. (b) A meteorologist explains the nice weather by referring to a “high-pressure area.” In light of your answer to part (a), explain why this term makes sense.
- 10.22** (a) On Titan, the largest moon of Saturn, the atmospheric pressure is 1.63105 Pa. What is the atmospheric pressure of Titan in atm? (b) On Venus the surface atmospheric pressure is about 90 Earth atmospheres. What is the Venusian atmospheric pressure in kilopascals?

- 10.23** If the atmospheric pressure is 0.985 atm, what is the pressure of the enclosed gas in each of the three cases depicted in the drawing?

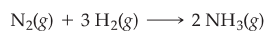


- 10.24** An open-end manometer containing mercury is connected to a container of gas, as depicted in Sample Exercise 10.2. What is the pressure of the enclosed gas in torr in each of the following situations? (a) The mercury in the arm attached to the gas is 15.4 mm higher than in the one open to the atmosphere; atmospheric pressure is 0.966 atm. (b) The mercury in the arm attached to the gas is 8.7 mm lower than in the one open to the atmosphere; atmospheric pressure is 0.99 atm.

The Gas Laws

- 10.25** Assume that you have a cylinder with a movable piston. What would happen to the gas pressure inside the cylinder if you do the following? (a) Decrease the volume to one-fourth the original volume while holding the temperature constant. (b) Reduce the Kelvin temperature to half its original value while holding the volume constant. (c) Reduce the amount of gas to half while keeping the volume and temperature constant.
- 10.26** A fixed quantity of gas at 21 °C exhibits a pressure of 752 torr and occupies a volume of 4.38 L. (a) Use Boyle's law to calculate the volume the gas will occupy if the pressure is increased to 1.88 atm while the temperature is held constant. (b) Use Charles's law to calculate the volume the gas will occupy if the temperature is increased to 175 °C while the pressure is held constant.

- 10.27** (a) How is the law of combining volumes explained by Avogadro's hypothesis? (b) Consider a 1.0-L flask containing neon gas and a 1.5-L flask containing xenon gas. Both gases are at the same pressure and temperature. According to Avogadro's law, what can be said about the ratio of the number of atoms in the two flasks?
- 10.28** Nitrogen and hydrogen gases react to form ammonia gas as follows:



At a certain temperature and pressure, 1.2 L of N_2 reacts with 3.6 L of H_2 . If all the N_2 and H_2 are consumed, what volume of NH_3 , at the same temperature and pressure, will be produced?

The Ideal-Gas Equation

- 10.29** (a) Write the ideal-gas equation, and give the units used for each term in the equation when $R = 0.0821$ L·atm/mol·K. (b) What is an ideal gas?
- 10.30** (a) What conditions are represented by the abbreviation STP? (b) What is the molar volume of an ideal gas at STP? (c) Room temperature is often assumed to be 25 °C. Calculate the molar volume of an ideal gas at 25 °C and 1 atm pressure.
- 10.31** Suppose you are given two 1-L flasks and told that one contains a gas of molar mass 30, the other a gas of molar mass 60, both at the same temperature. The pressure in

flask A is X atm, and the mass of gas in the flask is 1.2 g. The pressure in flask B is 0.5X atm, and the mass of gas in that flask is 1.2 g. Which flask contains gas of molar mass 30, and which contains the gas of molar mass 60?

- 10.32** Suppose you are given two flasks at the same temperature, one of volume 2 L and the other of volume 3 L. The 2-L flask contains 4.8 g of gas, and the gas pressure is X atm. The 3-L flask contains 0.36 g of gas, and the gas pressure is 0.1X. Do the two gases have the same molar mass? If not, which contains the gas of higher molar mass?

10.33 Complete the following table for an ideal gas:

P	V	n	T
2.00 atm	1.00 L	0.500 mol	? K
0.300 atm	0.250 L	? mol	27 °C
650 torr	? L	0.333 mol	350 K
? atm	585 mL	0.250 mol	295 K

- 10.34 Calculate each of the following quantities for an ideal gas: (a) the volume of the gas, in liters, if 1.50 mol has a pressure of 0.985 atm at a temperature of $-6\text{ }^{\circ}\text{C}$; (b) the absolute temperature of the gas at which 3.33×10^{-3} mol occupies 325 mL at 750 torr; (c) the pressure, in atmospheres, if 0.0467 mol occupies 413 mL at $138\text{ }^{\circ}\text{C}$; (d) the quantity of gas, in moles, if 55.7 L at $54\text{ }^{\circ}\text{C}$ has a pressure of 11.25 kPa.
- 10.35 The Goodyear blimps, which frequently fly over sporting events, hold approximately $175,000\text{ ft}^3$ of helium. If the gas is at $23\text{ }^{\circ}\text{C}$ and 1.0 atm, what mass of helium is in the blimp?
- 10.36 A neon sign is made of glass tubing whose inside diameter is 2.5 cm and whose length is 5.5 m. If the sign contains neon at a pressure of 1.78 torr at $35\text{ }^{\circ}\text{C}$, how many grams of neon are in the sign? (The volume of a cylinder is $\pi r^2 h$.)
- 10.37 Calculate the number of molecules in a deep breath of air whose volume is 2.25 L at body temperature, $37\text{ }^{\circ}\text{C}$, and a pressure of 735 torr.
- 10.38 If the pressure exerted by ozone, O_3 , in the stratosphere is 3.0×10^{-3} atm and the temperature is 250 K, how many ozone molecules are in a liter?
- 10.39 A scuba diver's tank contains 0.29 kg of O_2 compressed into a volume of 2.3 L. (a) Calculate the gas pressure inside the tank at $9\text{ }^{\circ}\text{C}$. (b) What volume would this oxygen occupy at $26\text{ }^{\circ}\text{C}$ and 0.95 atm?
- 10.40 An aerosol spray can with a volume of 250 mL contains 2.30 g of propane gas (C_3H_8) as a propellant. (a) If the can is at $23\text{ }^{\circ}\text{C}$, what is the pressure in the can? (b) What volume would the propane occupy at STP? (c) The can says that exposure to temperatures above $130\text{ }^{\circ}\text{F}$ may cause the can to burst. What is the pressure in the can at this temperature?

- 10.41 Chlorine is widely used to purify municipal water supplies and to treat swimming pool waters. Suppose that the volume of a particular sample of Cl_2 gas is 8.70 L at 895 torr and $24\text{ }^{\circ}\text{C}$. (a) How many grams of Cl_2 are in the sample? (b) What volume will the Cl_2 occupy at STP? (c) At what temperature will the volume be 15.00 L if the pressure is 8.76×10^2 torr? (d) At what pressure will the volume equal 6.00 L if the temperature is $58\text{ }^{\circ}\text{C}$?
- 10.42 Many gases are shipped in high-pressure containers. Consider a steel tank whose volume is 65.0 L and which contains O_2 gas at a pressure of 16,500 kPa at $23\text{ }^{\circ}\text{C}$. (a) What mass of O_2 does the tank contain? (b) What volume would the gas occupy at STP? (c) At what temperature would the pressure in the tank equal 150.0 atm? (d) What would be the pressure of the gas, in kPa, if it were transferred to a container at $24\text{ }^{\circ}\text{C}$ whose volume is 55.0 L?
- 10.43 In an experiment reported in the scientific literature, male cockroaches were made to run at different speeds on a miniature treadmill while their oxygen consumption was measured. In one hour the average cockroach running at 0.08 km/hr consumed 0.8 mL of O_2 at 1 atm pressure and $24\text{ }^{\circ}\text{C}$ per gram of insect weight. (a) How many moles of O_2 would be consumed in 1 hr by a 5.2-g cockroach moving at this speed? (b) This same cockroach is caught by a child and placed in a 1-qt fruit jar with a tight lid. Assuming the same level of continuous activity as in the research, will the cockroach consume more than 20% of the available O_2 in a 48-hr period? (Air is 21 mol percent O_2 .)
- 10.44 After the large eruption of Mount St. Helens in 1980, gas samples from the volcano were taken by sampling the downwind gas plume. The unfiltered gas samples were passed over a gold-coated wire coil to absorb mercury (Hg) present in the gas. The mercury was recovered from the coil by heating it, and then analyzed. In one particular set of experiments scientists found a mercury vapor level of 1800 ng of Hg per cubic meter in the plume, at a gas temperature of $10\text{ }^{\circ}\text{C}$. Calculate (a) the partial pressure of Hg vapor in the plume, (b) the number of Hg atoms per cubic meter in the gas, (c) the total mass of Hg emitted per day by the volcano if the daily plume volume was 1600 km^3 .

Further Applications of the Ideal-Gas Equation

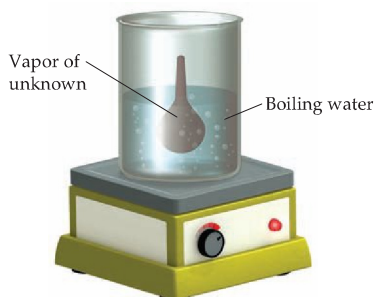
- 10.45 Which gas is most dense at 1.00 atm and 298 K? CO_2 , N_2O , or Cl_2 . Explain.
- 10.46 Rank the following gases from least dense at 1.00 atm and 298 K to most dense under these same conditions: SO_2 , HBr, CO_2 . Explain.
- 10.47 Which of the following statements best explains why a closed balloon filled with helium gas rises in air?
 (a) Helium is a monatomic gas, whereas nearly all the molecules that make up air, such as nitrogen and oxygen, are diatomic.
 (b) The average speed of helium atoms is higher than the average speed of air molecules, and the higher speed of collisions with the balloon walls propels the balloon upward.
 (c) Because the helium atoms are of lower mass than the average air molecule, the helium gas is less dense than air. The balloon thus weighs less than the air displaced by its volume.
 (d) Because helium has a lower molar mass than the average air molecule, the helium atoms are in faster motion. This means that the temperature of the helium is higher than the air temperature. Hot gases tend to rise.

- 10.48** Which of the following statements best explains why nitrogen gas at STP is less dense than Xe gas at STP?
- Because Xe is a noble gas, there is less tendency for the Xe atoms to repel one another, so they pack more densely in the gas state.
 - Xe atoms have a higher mass than N₂ molecules. Because both gases at STP have the same number of molecules per unit volume, the Xe gas must be denser.
 - The Xe atoms are larger than N₂ molecules and thus take up a larger fraction of the space occupied by the gas.
 - Because the Xe atoms are much more massive than the N₂ molecules, they move more slowly and thus exert less upward force on the gas container and make the gas appear denser.

10.49 (a) Calculate the density of NO₂ gas at 0.970 atm and 35 °C. (b) Calculate the molar mass of a gas if 2.50 g occupies 0.875 L at 685 torr and 35 °C.

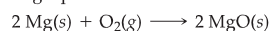
10.50 (a) Calculate the density of sulfur hexafluoride gas at 707 torr and 21 °C. (b) Calculate the molar mass of a vapor that has a density of 7.135 g/L at 12 °C and 743 torr.

10.51 In the Dumas-bulb technique for determining the molar mass of an unknown liquid, you vaporize the sample of a liquid that boils below 100 °C in a boiling-water bath and determine the mass of vapor required to fill the bulb (see drawing). From the following data, calculate the molar mass of the unknown liquid: mass of unknown vapor, 1.012 g; volume of bulb, 354 cm³; pressure, 742 torr; temperature, 99 °C.

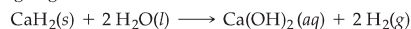


10.52 The molar mass of a volatile substance was determined by the Dumas-bulb method described in Exercise 10.51. The unknown vapor had a mass of 0.846 g; the volume of the bulb was 354 cm³, pressure 752 torr, and temperature 100 °C. Calculate the molar mass of the unknown vapor.

10.53 Magnesium can be used as a “getter” in evacuated enclosures, to react with the last traces of oxygen. (The magnesium is usually heated by passing an electric current through a wire or ribbon of the metal.) If an enclosure of 0.382 L has a partial pressure of O₂ of 3.5×10^{-6} torr at 27 °C, what mass of magnesium will react according to the following equation?

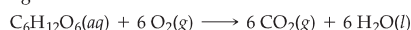


10.54 Calcium hydride, CaH₂, reacts with water to form hydrogen gas:



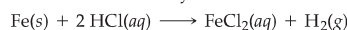
This reaction is sometimes used to inflate life rafts, weather balloons, and the like, where a simple, compact means of generating H₂ is desired. How many grams of CaH₂ are needed to generate 53.5 L of H₂ gas if the pressure of H₂ is 814 torr at 21 °C?

10.55 The metabolic oxidation of glucose, C₆H₁₂O₆, in our bodies produces CO₂, which is expelled from our lungs as a gas:



Calculate the volume of dry CO₂ produced at body temperature (37 °C) and 0.970 atm when 24.5 g of glucose is consumed in this reaction.

10.56 Both Jacques Charles and Joseph Louis Guy-Lussac were avid balloonists. In his original flight in 1783, Jacques Charles used a balloon that contained approximately 31,150 L of H₂. He generated the H₂ using the reaction between iron and hydrochloric acid:



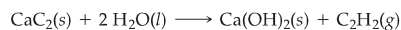
How many kilograms of iron were needed to produce this volume of H₂ if the temperature was 22 °C?

10.57 Hydrogen gas is produced when zinc reacts with sulfuric acid:



If 159 mL of wet H₂ is collected over water at 24 °C and a barometric pressure of 738 torr, how many grams of Zn have been consumed? (The vapor pressure of water is tabulated in Appendix B.)

10.58 Acetylene gas, C₂H₂(g), can be prepared by the reaction of calcium carbide with water:

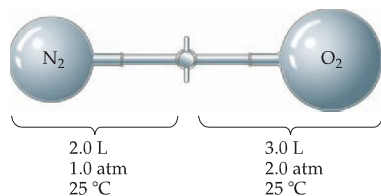


Calculate the volume of C₂H₂ that is collected over water at 23 °C by reaction of 0.752 g of CaC₂ if the total pressure of the gas is 745 torr. (The vapor pressure of water is tabulated in Appendix B.)

Partial Pressures

10.59 Consider the apparatus shown in the drawing on the next page. (a) When the stopcock between the two containers is opened and the gases allowed to mix, how does the volume occupied by the N₂ gas change? What is

the partial pressure of N₂ after mixing? (b) How does the volume of the O₂ gas change when the gases mix? What is the partial pressure of O₂ in the mixture? (c) What is the total pressure in the container after the gases mix?



- 10.60** Consider a mixture of two gases, A and B, confined in a closed vessel. A quantity of a third gas, C, is added to the same vessel at the same temperature. How does the addition of gas C affect the following: (a) the partial pressure of gas A, (b) the total pressure in the vessel, (c) the mole fraction of gas B?
- 10.61** A mixture containing 0.477 mol He(g), 0.280 mol Ne(g), and 0.110 mol Ar(g) is confined in a 7.00-L vessel at 25 °C. (a) Calculate the partial pressure of each of the gases in the mixture. (b) Calculate the total pressure of the mixture.
- 10.62** A deep-sea diver uses a gas cylinder with a volume of 10.0 L and a content of 51.2 g of O_2 and 32.6 g of He. Calculate the partial pressure of each gas and the total pressure if the temperature of the gas is 19 °C.
- 10.63** A piece of solid carbon dioxide with a mass of 5.50 g is placed in a 10.0-L vessel that already contains air at 705 torr and 24 °C. After the carbon dioxide has totally vaporized, what is the partial pressure of carbon dioxide and the total pressure in the container at 24 °C?
- 10.64** A sample of 4.00 mL of diethylether ($C_2H_5OC_2H_5$, density = 0.7134 g/mL) is introduced into a 5.00-L vessel that already contains a mixture of N_2 and O_2 , whose partial pressures are $P_{N_2} = 0.751$ atm and $P_{O_2} = 0.208$ atm. The temperature is held at 35.0 °C, and the diethylether totally evaporates. (a) Calculate the partial pressure of the diethylether. (b) Calculate the total pressure in the container.
- 10.65** A mixture of gases contains 0.75 mol N_2 , 0.30 mol O_2 , and 0.15 mol CO_2 . If the total pressure of the mixture is 1.56 atm, what is the partial pressure of each component?
- 10.66** A mixture of gases contains 10.25 g of N_2 , 1.83 g of H_2 , and 7.95 g of NH_3 . If the total pressure of the mixture is 1.85 atm, what is the partial pressure of each component?
- 10.67** At an underwater depth of 250 ft, the pressure is 8.38 atm. What should the mole percent of oxygen be in the diving gas for the partial pressure of oxygen in the mixture to be 0.21 atm, the same as in air at 1 atm?
- 10.68** (a) What are the mole fractions of each component in a mixture of 5.08 g of O_2 , 7.17 g of N_2 , and 1.32 g of H_2 ? (b) What is the partial pressure in atm of each component of this mixture if it is held in a 12.40-L vessel at 15 °C?
- 10.69** A quantity of N_2 gas originally held at 4.75 atm pressure in a 1.00-L container at 26 °C is transferred to a 10.0-L container at 20 °C. A quantity of O_2 gas originally at 5.25 atm and 26 °C in a 5.00-L container is transferred to this same container. What is the total pressure in the new container?
- 10.70** A sample of 3.00 g of SO_2 (g) originally in a 5.00-L vessel at 21 °C is transferred to a 10.0-L vessel at 26 °C. A sample of 2.35 g N_2 (g) originally in a 2.50-L vessel at 20 °C is transferred to this same 10.0-L vessel. (a) What is the partial pressure of SO_2 (g) in the larger container? (b) What is the partial pressure of N_2 (g) in this vessel? (c) What is the total pressure in the vessel?

Kinetic-Molecular Theory; Graham's Law

- 10.71** What change or changes in the state of a gas bring about each of the following effects? (a) The number of impacts per unit time on a given container wall increases. (b) The average energy of impact of molecules with the wall of the container decreases. (c) The average distance between gas molecules increases. (d) The average speed of molecules in the gas mixture is increased.
- 10.72** Indicate which of the following statements regarding the kinetic-molecular theory of gases are correct. For those that are false, formulate a correct version of the statement. (a) The average kinetic energy of a collection of gas molecules at a given temperature is proportional to $m^{1/2}$. (b) The gas molecules are assumed to exert no forces on each other. (c) All the molecules of a gas at a given temperature have the same kinetic energy. (d) The volume of the gas molecules is negligible in comparison to the total volume in which the gas is contained.
- 10.73** What property or properties of gases can you point to that support the assumption that most of the volume in a gas is empty space?
- 10.74** Newton had an incorrect theory of gases in which he assumed that all gas molecules repel one another and the walls of their container. Thus, the molecules of a gas are statically and uniformly distributed, trying to get as far apart as possible from one another and the vessel walls. This repulsion gives rise to pressure. Explain why Charles's law argues for the kinetic-molecular theory and against Newton's model.
- 10.75** The temperature of a 5.00-L container of N_2 gas is increased from 20 °C to 250 °C. If the volume is held constant, predict qualitatively how this change affects the following: (a) the average kinetic energy of the molecules; (b) the average speed of the molecules; (c) the strength of the impact of an average molecule with the container walls; (d) the total number of collisions of molecules with walls per second.
- 10.76** Suppose you have two 1-L flasks, one containing N_2 at STP, the other containing CH_4 at STP. How do these systems compare with respect to (a) number of molecules, (b) density, (c) average kinetic energy of the molecules, (d) rate of effusion through a pinhole leak?

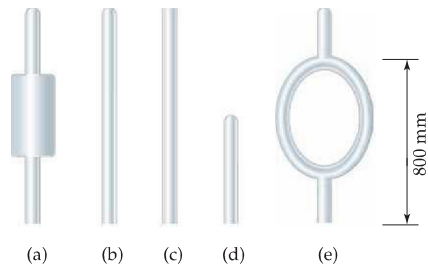
- 10.77** (a) Place the following gases in order of increasing average molecular speed at 25 °C: Ne, HBr, SO₂, NF₃, CO. (b) Calculate the rms speed of NF₃ molecules at 25 °C.
- 10.78** (a) Place the following gases in order of increasing average molecular speed at 300 K: CO, SF₆, H₂S, Cl₂, HBr. (b) Calculate and compare the rms speeds of CO and Cl₂ molecules at 300 K.
- 10.79** Hydrogen has two naturally occurring isotopes, ¹H and ²H. Chlorine also has two naturally occurring isotopes, ³⁵Cl and ³⁷Cl. Thus, hydrogen chloride gas consists of four distinct types of molecules: ¹H³⁵Cl, ¹H³⁷Cl, ²H³⁵Cl, and ²H³⁷Cl. Place these four molecules in order of increasing rate of effusion.
- 10.80** As discussed in the “Chemistry Put to Work” box in Section 10.8, enriched uranium can be produced by gaseous diffusion of UF₆. Suppose a process were developed to allow diffusion of gaseous uranium atoms, U(g). Calculate the ratio of diffusion rates for ²³⁵U and ²³⁸U, and compare it to the ratio for UF₆ given in the essay.
- 10.81** Arsenic(III) sulfide sublimes readily, even below its melting point of 320 °C. The molecules of the vapor phase are found to effuse through a tiny hole at 0.28 times the rate of effusion of Ar atoms under the same conditions of temperature and pressure. What is the molecular formula of arsenic(III) sulfide in the gas phase?
- 10.82** A gas of unknown molecular mass was allowed to effuse through a small opening under constant-pressure conditions. It required 105 s for 1.0 L of the gas to effuse. Under identical experimental conditions it required 31 s for 1.0 L of O₂ gas to effuse. Calculate the molar mass of the unknown gas. (Remember that the faster the rate of effusion, the shorter the time required for effusion of 1.0 L; that is, rate and time are inversely proportional.)

Nonideal-Gas Behavior

- 10.83** (a) List two experimental conditions under which gases deviate from ideal behavior. (b) List two reasons why the gases deviate from ideal behavior. (c) Explain how the function PV/RT can be used to show how gases behave nonideally.
- 10.84** The planet Jupiter has a surface temperature of 140 K and a mass 318 times that of Earth. Mercury has a surface temperature between 600 K and 700 K and a mass 0.05 times that of Earth. On which planet is the atmosphere more likely to obey the ideal-gas law? Explain.
- 10.85** Based on their respective van der Waals constants (Table 10.3), is Ar or CO₂ expected to behave more nearly like an ideal gas at high pressures? Explain.
- 10.86** Briefly explain the significance of the constants a and b in the van der Waals equation.
- 10.87** In Sample Exercise 10.16, we found that one mole of Cl₂ confined to 22.41 L at 0 °C deviated slightly from ideal behavior. Calculate the pressure exerted by 1.00 mol Cl₂ confined to a smaller volume, 5.00 L, at 25 °C. (a) First use the ideal gas equation and (b) then use van der Waals equation for your calculation. (Values for the van der Waals constants are given in Table 10.3.) (c) Why is the difference between the result for an ideal gas and that calculated using van der Waals equation greater when the gas is confined to 5.00 L compared to 22.4 L?
- 10.88** Calculate the pressure that CCl₄ will exert at 40 °C if 1.00 mol occupies 28.0 L, assuming that (a) CCl₄ obeys the ideal-gas equation; (b) CCl₄ obeys the van der Waals equation. (Values for the van der Waals constants are given in Table 10.3.) (c) Which would you expect to deviate more from ideal behavior under these conditions, Cl₂ or CCl₄? Explain.

ADDITIONAL EXERCISES

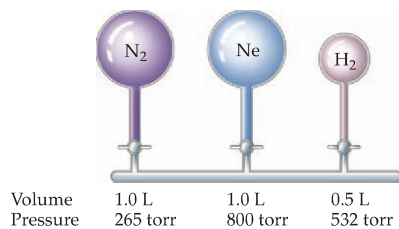
- 10.89** Suppose the mercury used to make a barometer has a few small droplets of water trapped in it that rise to the top of the mercury in the tube. Will the barometer show the correct atmospheric pressure? Explain.
- 10.90** Suppose that when Torricelli had his great idea for constructing a mercury manometer, he had rushed into the laboratory and found the following items of glass:



Which of these would have been satisfactory for his use in forming the first manometer? Explain why the unsatisfactory ones would not have worked.

- 10.91** A gas bubble with a volume of 1.0 mm³ originates at the bottom of a lake where the pressure is 3.0 atm. Calculate its volume when the bubble reaches the surface of the lake where the pressure is 695 torr, assuming that the temperature doesn't change.
- 10.92** A 15.0-L tank is filled with helium gas at a pressure of 1.00×10^2 . How many balloons (each 2.00 L) can be inflated to a pressure of 1.00 atm, assuming that the temperature remains constant and that the tank cannot be emptied below 1.00 atm?
- 10.93** To minimize the rate of evaporation of the tungsten filament, 1.4×10^{-5} mol of argon is placed in a 600-cm³ lightbulb. What is the pressure of argon in the lightbulb at 23 °C?

- 10.94** Carbon dioxide, which is recognized as the major contributor to global warming as a "greenhouse gas," is formed when fossil fuels are combusted, as in electrical power plants fueled by coal, oil, or natural gas. One potential way to reduce the amount of CO_2 added to the atmosphere is to store it as a compressed gas in underground formations. Consider a 1000-megawatt coal-fired power plant that produces about 6×10^6 tons of CO_2 per year. (a) Assuming ideal gas behavior, 1.00 atm, and 27°C , calculate the volume of CO_2 produced by this power plant. (b) If the CO_2 is stored underground as a liquid at 10°C and 120 atm and a density of 1.2 g/cm^3 , what volume does it possess? (c) If it is stored underground as a gas at 36°C and 90 atm, what volume does it occupy?
- 10.95** Propane, C_3H_8 , liquefies under modest pressure, allowing a large amount to be stored in a container. (a) Calculate the number of moles of propane gas in a 110-L container at 3.00 atm and 27°C . (b) Calculate the number of moles of liquid propane that can be stored in the same volume if the density of the liquid is 0.590 g/mL . (c) Calculate the ratio of the number of moles of liquid to moles of gas. Discuss this ratio in light of the kinetic-molecular theory of gases.
- 10.96** Nickel carbonyl, $\text{Ni}(\text{CO})_4$, is one of the most toxic substances known. The present maximum allowable concentration in laboratory air during an 8-hr workday is 1 part in 10^9 parts by volume, which means that there is one mole of $\text{Ni}(\text{CO})_4$ for every 10^9 moles of gas. Assume 24°C and 1.00 atm pressure. What mass of $\text{Ni}(\text{CO})_4$ is allowable in a laboratory that is 54 m^2 in area, with a ceiling height of 3.1 m?
- 10.97** A large flask is evacuated and weighed, filled with argon gas, and then reweighed. When reweighed, the flask is found to have gained 3.224 g. It is again evacuated and then filled with a gas of unknown molar mass. When reweighed, the flask is found to have gained 8.102 g. (a) Based on the molar mass of argon, estimate the molar mass of the unknown gas. (b) What assumptions did you make in arriving at your answer?
- 10.98** Consider the arrangement of bulbs shown in the drawing. Each of the bulbs contains a gas at the pressure shown. What is the pressure of the system when all the stopcocks are opened, assuming that the temperature remains constant? (We can neglect the volume of the capillary tubing connecting the bulbs.)



- 10.99** Assume that a single cylinder of an automobile engine has a volume of 524 cm^3 . (a) If the cylinder is full of air at 74°C and 0.980 atm, how many moles of O_2 are present? (The mole fraction of O_2 in dry air is 0.2095.)

(b) How many grams of C_8H_{18} could be combusted by this quantity of O_2 , assuming complete combustion with formation of CO_2 and H_2O ?

- 10.100** Assume that an exhaled breath of air consists of 74.8% N_2 , 15.3% O_2 , 3.7% CO_2 , and 6.2% water vapor. (a) If the total pressure of the gases is 0.980 atm, calculate the partial pressure of each component of the mixture. (b) If the volume of the exhaled gas is 455 mL and its temperature is 37°C , calculate the number of moles of CO_2 exhaled. (c) How many grams of glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) would need to be metabolized to produce this quantity of CO_2 ? (The chemical reaction is the same as that for combustion of $\text{C}_6\text{H}_{12}\text{O}_6$. See Section 3.2.)
- 10.101** A sample of 1.42 g of helium and an unweighed quantity of O_2 are mixed in a flask at room temperature. The partial pressure of helium in the flask is 42.5 torr, and the partial pressure of oxygen is 158 torr. What is the mass of the oxygen in the container?
- 10.102** A gaseous mixture of O_2 and Kr has a density of 1.104 g/L at 435 torr and 300 K. What is the mole percent O_2 in the mixture?
- 10.103** The density of a gas of unknown molar mass was measured as a function of pressure at 0°C , as in the table below. (a) Determine a precise molar mass for the gas. *Hint:* Graph d/P versus P . (b) Why is d/P not a constant as a function of pressure?

Pressure (atm)	1.00	0.666	0.500	0.333	0.250
Density (g/L)	2.3074	1.5263	1.1401	0.7571	0.5660

- 10.104** A glass vessel fitted with a stopcock has a mass of 337.428 g when evacuated. When filled with Ar, it has a mass of 339.854 g. When evacuated and refilled with a mixture of Ne and Ar, under the same conditions of temperature and pressure, it weighs 339.076 g. What is the mole percent of Ne in the gas mixture?
- 10.105** You have a sample of gas at -33°C . You wish to increase the rms speed by 10.0%. To what temperature should the gas be heated?
- 10.106** Consider the following gases, all at STP: Ne, SF_6 , N_2 , CH_4 . (a) Which gas is most likely to depart from assumption 3 of the kinetic molecular theory (Section 10.7)? (b) Which one is closest to an ideal gas in its behavior? (c) Which one has the highest root-mean-square molecular speed? (d) Which one has the highest total molecular volume relative to the space occupied by the gas? (e) Which has the highest average kinetic molecular energy? (f) Which one would effuse more rapidly than N_2 ?
- 10.107** Does the effect of intermolecular attraction on the properties of a gas become more significant or less significant if (a) the gas is compressed to a smaller volume at constant temperature; (b) the temperature of the gas is increased at constant volume?
- 10.108** For nearly all real gases, the quantity PV/RT decreases below the value of 1, which characterizes an ideal gas, as pressure on the gas increases. At much higher

pressures, however, PV/RT increases and rises above the value of 1. (a) Explain the initial drop in value of PV/RT below 1 and the fact that it rises above 1 for still higher pressures. (b) The effects we have just noted are smaller for gases at higher temperature. Why is this so?

- 10.109** Which of the noble gases other than radon would you expect to depart most readily from ideal behavior? Use the density data in Table 7.8 to show evidence in support of your answer.
- 10.110** It turns out that the van der Waals constant b equals four times the total volume actually occupied by the molecules of a mole of gas. Using this figure, calculate the fraction of the volume in a container actually occupied

by Ar atoms (a) at STP, (b) at 100 atm pressure and 0 °C. (Assume for simplicity that the ideal-gas equation still holds.)

- [10.111]** Large amounts of nitrogen gas are used in the manufacture of ammonia, principally for use in fertilizers. Suppose 120.00 kg of $N_2(g)$ is stored in a 1100.0-L metal cylinder at 280 °C. (a) Calculate the pressure of the gas, assuming ideal-gas behavior. (b) By using data in Table 10.3, calculate the pressure of the gas according to the van der Waals equation. (c) Under the conditions of this problem, which correction dominates, the one for finite volume of gas molecules or the one for attractive interactions?

INTEGRATIVE EXERCISES

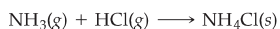
- 10.112** Cyclopropane, a gas used with oxygen as a general anesthetic, is composed of 85.7% C and 14.3% H by mass. (a) If 1.56 g of cyclopropane has a volume of 1.00 L at 0.984 atm and 50.0 °C, what is the molecular formula of cyclopropane? (b) Judging from its molecular formula, would you expect cyclopropane to deviate more or less than Ar from ideal-gas behavior at moderately high pressures and room temperature? Explain.

- 10.113** Consider the combustion reaction between 25.0 mL of liquid methanol (density = 0.850 g/mL) and 12.5 L of oxygen gas measured at STP. The products of the reaction are $CO_2(g)$ and $H_2O(g)$. Calculate the number of moles of H_2O formed if the reaction goes to completion.

- 10.114** An herbicide is found to contain only C, H, N, and Cl. The complete combustion of a 100.0-mg sample of the herbicide in excess oxygen produces 83.16 mL of CO_2 and 73.30 mL of H_2O vapor at STP. A separate analysis shows that the sample also contains 16.44 mg of Cl. (a) Determine the percent composition of the substance. (b) Calculate its empirical formula.

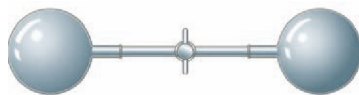
- [10.115]** A 4.00-g sample of a mixture of CaO and BaO is placed in a 1.00-L vessel containing CO_2 gas at a pressure of 730 torr and a temperature of 25 °C. The CO_2 reacts with the CaO and BaO, forming $CaCO_3$ and $BaCO_3$. When the reaction is complete, the pressure of the remaining CO_2 is 150 torr. (a) Calculate the number of moles of CO_2 that have reacted. (b) Calculate the mass percentage of CaO in the mixture.

- [10.116]** Ammonia, $NH_3(g)$, and hydrogen chloride, $HCl(g)$, react to form solid ammonium chloride, $NH_4Cl(s)$:



Two 2.00-L flasks at 25 °C are connected by a stopcock, as shown in the drawing. One flask contains 5.00 g $NH_3(g)$, and the other contains 5.00 g $HCl(g)$. When the stopcock is opened, the gases react until one is completely consumed. (a) Which gas will remain in the sys-

tem after the reaction is complete? (b) What will be the final pressure of the system after the reaction is complete? (Neglect the volume of the ammonium chloride formed.)



- 10.117** The “Chemistry Put to Work” box on pipelines in Section 10.5 mentions that the total deliverability of natural gas (methane, CH_4) to the various regions of the United States is on the order of 2.7×10^{12} L per day, measured at STP. Calculate the total enthalpy change for combustion of this quantity of methane. (Note: Less than this amount of methane is actually combusted daily. Some of the delivered gas is passed through to other regions.)

- [10.118]** A gas forms when elemental sulfur is heated carefully with AgF. The initial product boils at 15 °C. Experiments on several samples yielded a gas density of 0.803 ± 0.010 g/L for the gas at 150 mm pressure and 32 °C. When the gas reacts with water, all the fluorine is converted to aqueous HF. Other products are elemental sulfur, S_8 , and other sulfur-containing compounds. A 480-mL sample of the dry gas at 126 mm pressure and 28 °C, when reacted with 80 mL of water, yielded a 0.081 M solution of HF. The initial gaseous product undergoes a transformation over a period of time to a second compound with the same empirical and molecular formula, which boils at -10 °C. (a) Determine the empirical and molecular formulas of the first compound formed. (b) Draw at least two reasonable Lewis structures that represent the initial compound and the one into which it is transformed over time. (c) Describe the likely geometries of these compounds, and estimate the single bond distances, given that the S—S bond distance in S_8 is 2.04 Å and the F—F distance in F_2 is 1.43 Å.

- 10.119** Chlorine dioxide gas (ClO_2) is used as a commercial bleaching agent. It bleaches materials by oxidizing them. In the course of these reactions, the ClO_2 is itself reduced. (a) What is the Lewis structure for ClO_2 ? (b) Why do you think that ClO_2 is reduced so readily? (c) When a ClO_2 molecule gains an electron, the chlorite ion, ClO_2^- , forms. Draw the Lewis structure for ClO_2^- . (d) Predict the O—Cl—O bond angle in the ClO_2^- ion. (e) One method of preparing ClO_2 is by the reaction of chlorine and sodium chlorite:

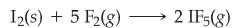


If you allow 10.0 g of NaClO_2 to react with 2.00 L of chlorine gas at a pressure of 1.50 atm at 21 °C, how many grams of ClO_2 can be prepared?

- 10.120** Natural gas is very abundant in many Middle Eastern oil fields. However, the costs of shipping the gas to markets in other parts of the world are high because it is necessary to liquefy the gas, which is mainly methane and thus has a boiling point at atmospheric pressure of -164 °C. One possible strategy is to oxidize the methane to methanol, CH_3OH , which has a boiling point of 65 °C and can therefore be shipped more readily. Suppose that $10.7 \times 10^9 \text{ ft}^3$ of methane at atmospheric pressure and 25 °C are oxidized to methanol. (a) What volume of methanol is formed if the density of CH_3OH is 0.791 g/mL? (b) Write balanced chemical equations for the oxidations of methane and methanol to $\text{CO}_2(\text{g})$ and $\text{H}_2\text{O}(\text{l})$. Calculate the total enthalpy change for complete combustion of the $10.7 \times 10^9 \text{ ft}^3$ of methane described

above and for complete combustion of the equivalent amount of methanol, as calculated in part (a). (c) Methane, when liquefied, has a density of 0.466 g/mL; the density of methanol at 25 °C is 0.791 g/mL. Compare the enthalpy change upon combustion of a unit volume of liquid methane and liquid methanol. From the standpoint of energy production, which substance has the higher enthalpy of combustion per unit volume?

- [10.121]** Gaseous iodine pentafluoride, IF_5 , can be prepared by the reaction of solid iodine and gaseous fluorine:



A 5.00-L flask containing 10.0 g I_2 is charged with 10.0 g F_2 , and the reaction proceeds until one of the reagents is completely consumed. After the reaction is complete, the temperature in the flask is 125 °C. (a) What is the partial pressure of IF_5 in the flask? (b) What is the mole fraction of IF_5 in the flask?

- [10.122]** A 6.53-g sample of a mixture of magnesium carbonate and calcium carbonate is treated with excess hydrochloric acid. The resulting reaction produces 1.72 L of carbon dioxide gas at 28 °C and 743 torr pressure. (a) Write balanced chemical equations for the reactions that occur between hydrochloric acid and each component of the mixture. (b) Calculate the total number of moles of carbon dioxide that forms from these reactions. (c) Assuming that the reactions are complete, calculate the percentage by mass of magnesium carbonate in the mixture.