CHAPTER 11

Molecular Composition of Gases

The study of gases led to the formulation of the laws and principles that are the foundations of modern chemistry.

Volume-Mass Relationships of Gases

SECTION 11-1

Objectives

- State the law of combining volumes.
- State Avogadro's law and explain its significance.
- Define *standard molar volume of a gas,* and use it to calculate gas masses and volumes.
- Use standard molar volume to calculate the molar mass of a gas.

In this section, you will study the relationships between the volumes of gases that react with each other. You will also learn how volume, density, and molar mass are related.

Measuring and Comparing the Volumes of Reacting Gases

In the early 1800s, French chemist Joseph Gay-Lussac studied gas volume relationships involving a chemical reaction between hydrogen and oxygen. He observed that 2 L of hydrogen can react with 1 L of oxygen to form 2 L of water vapor at constant temperature and pressure.

hydrogen gas +	oxygen gas ⟶	water vapor
2 L	1 L	2 L
2 volumes	1 volume	2 volumes

In other words, this reaction shows a simple and definite 2:1:2 relationship between the volumes of the reactants and the product. Two volumes of hydrogen react with 1 volume of oxygen to produce 2 volumes of water vapor. The 2:1:2 relationship for this reaction applies to any proportions for volume—for example, 2 mL, 1 mL, and 2 mL; 600 L, 300 L, and 600 L; or 400 cm^3, 200 cm^3, and 400 cm^3.

Gay-Lussac also noticed simple and definite proportions by volume in other reactions of gases, such as in the reaction between hydrogen gas and chlorine gas.

hydrogen gas +	chlorine gas ⟶	hydrogen chloride gas
1 L	1 L	2 L
1 volume	1 volume	2 volumes

In 1808, Gay-Lussac summarized the results of his experiments in a statement known today as **Gay-Lussac's law of combining volumes of gases.** The law states that *at constant temperature and pressure, the volumes of gaseous reactants and products can be expressed as ratios of small whole numbers.* This simple observation, combined with the insight of Avogadro, provided more understanding of how gases react and combine with each other.

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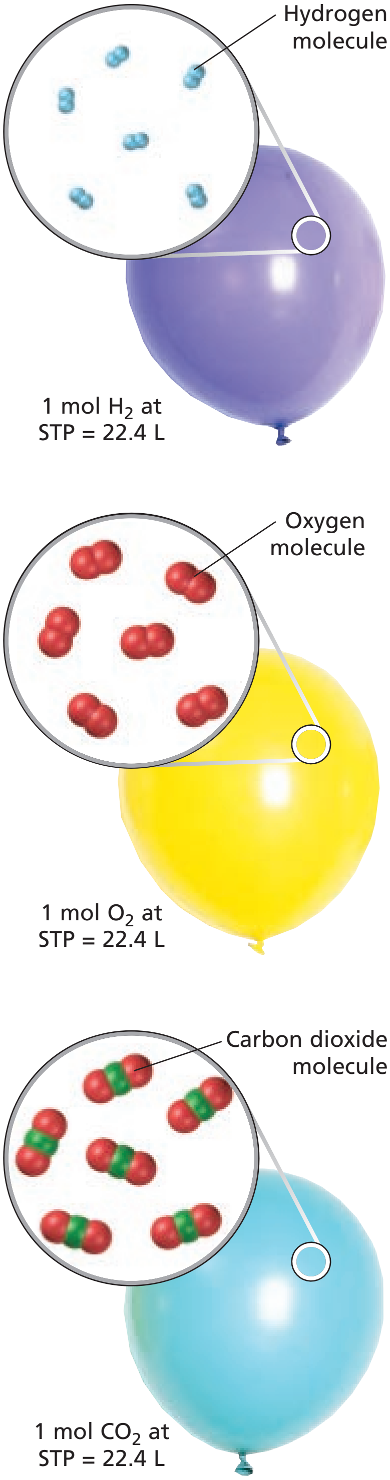

FIGURE 11-1 At the same temperature and pressure, balloons of equal volume have equal numbers of molecules, regardless of which gas they contain.

Avogadro's Law

Recall an important point of Dalton's atomic theory: atoms are indivisible. Dalton also thought that the particles of gaseous elements exist in the form of isolated single atoms. He believed that one atom of one element always combines with one atom of another element to form a single particle of the product. Accounting for some of the volume relationships observed by Gay-Lussac presented a problem for Dalton's theory. For example, in reactions such as the formation of water vapor, mentioned on the preceding page, it would seem that the oxygen involved would have to divide into two parts.

In 1811, Avogadro found a way to explain Gay-Lussac's simple ratios of combining volumes without violating Dalton's idea of indivisible atoms. He did this by rejecting Dalton's idea that reactant elements are always in monatomic form when they combine to form products. He reasoned that these molecules could contain more than one atom. Avogadro also put forth an idea known today as **Avogadro's law.** The law states that *equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.* Figure 11-1 illustrates Avogadro's law. It follows that at the same temperature and pressure, the volume of any given gas varies directly with the number of molecules.

Consider the reaction of hydrogen and chlorine to produce hydrogen chloride, illustrated in Figure 11-2. According to Avogadro's law, equal volumes of hydrogen and chlorine contain the same number of molecules. Avogadro accepted Dalton's idea that atoms of hydrogen and chlorine are indivisible. However, he rejected Dalton's belief that these elements are monatomic. He concluded that the hydrogen and chlorine components must each consist of two or more atoms joined together. The simplest assumption was that hydrogen and chlorine molecules are composed of two atoms each. That assumption leads to the following balanced equation for the reaction of hydrogen with chlorine.

$$\begin{array}{lclcl} H_2(g) & + & Cl_2(g) & \longrightarrow & 2HCl(g) \\ \text{1 volume} & & \text{1 volume} & & \text{2 volumes} \\ \text{1 molecule} & & \text{1 molecule} & & \text{2 molecules} \end{array}$$

The simplest hypothetical formula for hydrogen chloride, HCl, indicates that the molecule contains one hydrogen atom and one chlorine atom. Given the ratios of the combined volumes, the simplest formulas for hydrogen and chlorine must be H_2 and Cl_2, respectively.

Avogadro's reasoning applied equally well to the combining volumes for the reaction of hydrogen and oxygen to form water vapor. The simplest hypothetical formula for oxygen indicated two oxygen atoms, which turns out to be correct. The simplest possible molecule of water indicated two hydrogen atoms and one oxygen atom per molecule, which is also correct. Experiments eventually showed that all elements that are gases near room temperature, except the noble gases, normally exist as diatomic molecules.

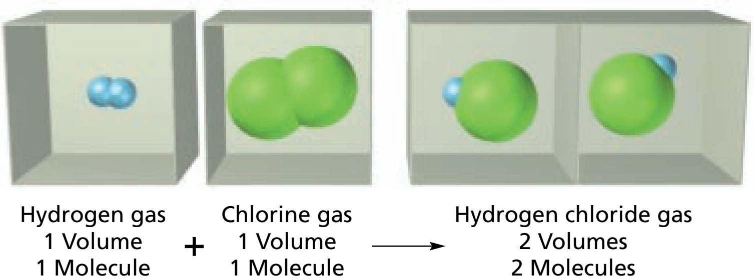

FIGURE 11-2 Hydrogen molecules combine with chlorine molecules in a 1:1 volume ratio to produce 2 volumes of hydrogen chloride.

Avogadro's law also indicates that gas volume is directly proportional to the amount of gas, at a given temperature and pressure. Note the equation for this relationship.

$$V = kn$$

Here, n is the amount of gas, in moles, and k is a constant. As shown below, the coefficients in a chemical reaction involving gases indicate the relative numbers of molecules, the relative numbers of moles, and the relative volumes.

$2H_2(g)$	+ $O_2(g)$	$\longrightarrow 2H_2O(g)$
2 molecules	1 molecule	2 molecules
2 mol	1 mol	2 mol
2 volumes	1 volume	2 volumes

Molar Volume of Gases

Recall that one mole of a molecular substance contains a number of molecules equal to Avogadro's constant (6.022×10^{23}). One mole of oxygen, O_2, contains 6.022×10^{23} diatomic oxygen molecules and has a mass of 31.9988 g. One mole of hydrogen contains the same number of diatomic hydrogen molecules but has a mass of only 2.015 88 g. One mole of helium, a monatomic gas, contains the same number of helium atoms and has a mass of 4.002 602 g.

According to Avogadro's law, one mole of any gas will occupy the same volume as one mole of any other gas at the same temperature and pressure, despite mass differences. *The volume occupied by one mole of a gas at STP is known as the* **standard molar volume of a gas.** *It has been found to be 22.414 10 L.* For calculations in this book, we use 22.4 L as the standard molar volume.

Knowing the volume of a gas, you can use 1 mol/22.4 L as a conversion factor to find the number of moles, and therefore the mass, of a given volume of a given gas at STP. You can also use the molar volume of a gas to find the volume, at STP, of a known number of moles or a known mass of a gas. These types of problems can also be solved using the ideal gas law, as you will see in Section 11-2.

FIGURE 11-3 One-mole quantities of two different gases each occupy 22.4 L at STP and have equal numbers of molecules. However, their masses are different.

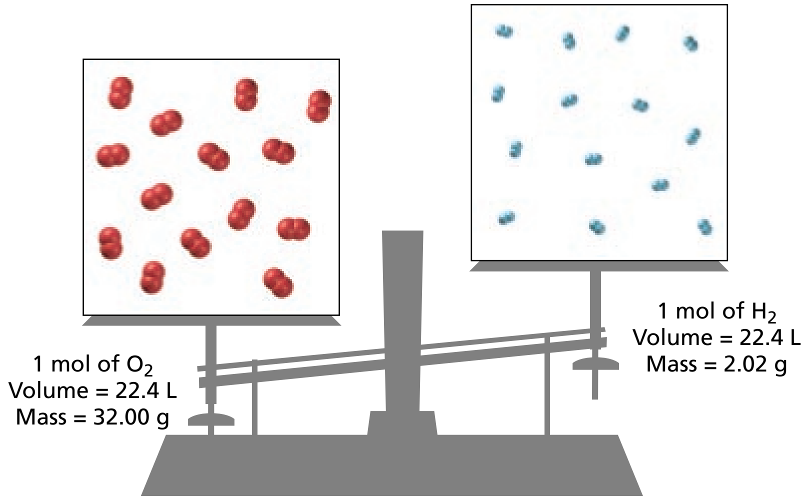

Figure 11-3 shows that 22.4 L of each gas contains the same number of molecules, but the mass of this volume is different for different gases. The mass of each is equal to the molar mass of the gas—the mass of one mole of molecules.

SAMPLE PROBLEM 11-1

A chemical reaction produces 0.0680 mol of oxygen gas. What volume in liters is occupied by this gas sample at STP?

SOLUTION

1 ANALYZE

Given: molar mass of O_2 = 0.0680 mol

Unknown: volume of O_2 in liters at STP

2 PLAN

$$\text{moles of } O_2 \longrightarrow \text{liters of } O_2 \text{ at STP}$$

The standard molar volume can be used to find the volume of a known molar amount of a gas at STP.

$$\text{mol} \times \frac{22.4 \text{ L}}{\text{mol}} = \text{volume of } O_2 \text{ in L}$$

3 COMPUTE

$$0.0680 \cancel{\text{mol}}\ O_2 \times \frac{22.4 \text{ L}}{\cancel{\text{mol}}} = 1.52 \text{ L } O_2$$

4 EVALUATE

The answer is close to an estimated value of 1.4, computed as 0.07 × 20. Units have canceled to yield liters. The calculated result is correctly expressed to three significant figures.

PRACTICE

1. At STP, what is the volume of 7.08 mol of nitrogen gas? *Answer* 159 L N_2

2. A sample of hydrogen gas occupies 14.1 L at STP. How many moles of the gas are present? *Answer* 0.629 mol H_2

3. At STP, a sample of neon gas occupies 550. cm^3. How many moles of neon gas does this represent? *Answer* 0.0246 mol Ne

SAMPLE PROBLEM 11-2

A chemical reaction produced 98.0 mL of sulfur dioxide gas, SO_2, at STP. What was the mass (in grams) of the gas produced?

SOLUTION

1 ANALYZE

Given: volume of SO_2 at STP = 98.0 mL

Unknown: mass of SO_2 in grams

2 PLAN

$$\text{liters of } SO_2 \text{ at STP} \longrightarrow \text{moles of } SO_2 \longrightarrow \text{grams of } SO_2$$

$$\text{mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{1 \text{ mol } SO_2}{22.4 \text{ L}} \times \frac{\text{g } SO_2}{\text{mol } SO_2} = \text{g } SO_2$$

3 COMPUTE

$$98.0 \cancel{\text{mL}} \times \frac{1 \cancel{\text{L}}}{1000 \cancel{\text{mL}}} \times \frac{1 \cancel{\text{mol } SO_2}}{22.4 \cancel{\text{L}}} \times \frac{64.07 \text{ g } SO_2}{\cancel{\text{mol } SO_2}} = 0.280 \text{ g } SO_2$$

4 EVALUATE

The result is correctly expressed to three significant figures. Units cancel correctly to give the answer in grams. The known volume is roughly 1/200 of the molar volume (22 400 mL/200 = 112 mL). The answer is reasonable: the mass should also be roughly 1/200 of the molar mass (64 g/200 = 0.32 g).

PRACTICE

1. What is the mass of 1.33×10^4 mL of oxygen gas at STP? *Answer* 19.0 g O_2

2. What is the volume of 77.0 g of nitrogen dioxide gas at STP? *Answer* 37.5 L NO_2

3. At STP, 3 L of chlorine is produced during a chemical reaction. What is the mass of this gas? *Answer* 9 g Cl_2

SECTION REVIEW

1. Explain Gay-Lussac's law of combining volumes.
2. State Avogadro's law and explain its significance.
3. Define *molar volume.*
4. How many moles of oxygen gas are there in 135 L of oxygen at STP?
5. What volume (in mL) at STP will be occupied by 0.0035 mol of methane, CH_4?

Chemistry's First Law

HISTORICAL PERSPECTIVE

The notion that nature abhors a vacuum was proposed by the Greek philosopher Aristotle, and his word went unchallenged for nearly 2000 years. Then in the mid-1600s, a new breed of thinkers known as experimental philosophers—later to be called scientists—began testing the long-held assumption that space must contain matter. These investigations represent some of the earliest experiments with gases, and they led to the discovery of the first empirical principle of chemistry, Boyle's law.

Overturning an Ancient Assumption

The first scientist to demonstrate the existence of a vacuum was Evangelista Torricelli. In 1643, he showed that when a glass tube 3 ft long and about an inch in diameter was sealed at one end, filled with mercury, and inverted in a container full of mercury, the mercury in the tube fell to a height of about 30 in. above the level of mercury in the container. Although some thinkers remained skeptical, it was generally accepted that the space between the mercury and the sealed end of the tube was indeed a vacuum.

Torricelli then turned his attention to how the mercury in the glass tube of his apparatus was supported. The known observation that liquids exerted a pressure on objects immersed in them inspired him to hypothesize that a "sea of air" surrounded Earth. He further hypothesized that the air exerted pressure on the mercury in the container and thus supported the mercury in the column.

Evangelista Torricelli invented the mercury barometer.

Support for the New Theory

Although the idea of an atmosphere that has weight and exerts a pressure on the objects within it seems obvious today, it was a radical theory at the time.

To test the effects of the atmosphere, one of the period's great scientists, Robert Boyle, had his talented assistant, Robert Hooke, create a piece of equipment that would revolutionize the study of air. The apparatus was an improved version of a pump designed by the famous German experimenter Otto von Guericke; the pump had a large receptacle in which a partial vacuum could be created.

Boyle placed Torricelli's setup, known today as a barometer, in the receptacle of the pump and observed the mercury column as he reduced the pressure around it. He noted that the height of the mercury decreased as the pressure surrounding the mercury in the container was lowered, strongly supporting Torricelli's atmospheric theory.

Using Hooke's pump, Boyle performed additional studies that verified the idea that air exerted pressure and had weight. His

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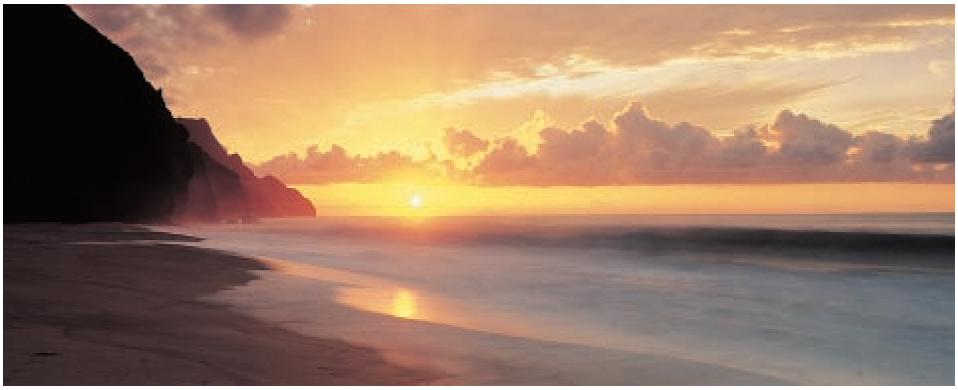
experiments also led him to the important conclusion that air was elastic, that is, it could expand and contract. It was during an investigation into air's elasticity that Boyle discovered the fundamental law that bears his name.



We can measure the difference between the atmospheric pressure on a mountaintop and the atmospheric pressure at sea level thanks to the work of Torricelli and Boyle.

An Ingenious Experiment

In response to a criticism of his findings, Boyle performed an experiment to show that air could be compressed to a pressure greater than that of the atmosphere. First he prepared a glass J tube with the short end sealed off and the long end left open. Then he poured mercury into the tube, making sure that the levels in each end were the same and letting air travel freely between the ends, to ensure that each column was at atmospheric pressure.

Boyle then poured more mercury into the long end of the tube until it was about 30 in. above the level of mercury in the short end, making the trapped air exposed to about twice as much atmospheric pressure. He observed that the volume of the trapped air was halved. He continued to add mercury until the total pressure on the trapped air was about four times that of the atmosphere. Noting that the air had been compressed to about one-quarter of its original value, Boyle discovered the inverse relationship between air's pressure and its volume:

> *It is evident, that as common air, when reduced to half its wonted extent [volume], obtained near about twice as forcible a spring [pressure] as it had before; so this thus comprest air being further thrust into half this narrow room, obtained thereby a spring about . . . four times as strong as that of common air.*

A Long-Standing Contribution

Boyle went on to show that the relationship between air pressure and volume, $P \propto 1/V$ (at constant temperature), held not only when the gas was compressed but also when it was allowed to expand. It would be up to future investigators to show that the law was a principle applying to gases in general. Together with the findings of other researchers, such as Jacques Charles, Joseph Gay-Lussac, and Amadeo Avogadro, Boyle's discovery led chemists to the famous ideal gas law, $PV = nRT$, which serves as a starting point in the study of chemistry today.

SECTION 11-2

The Ideal Gas Law

OBJECTIVES

- State the ideal gas law.
- Derive the ideal gas constant and discuss its units.
- Using the ideal gas law, calculate pressure, volume, temperature, or amount of gas when the other three quantities are known.
- Using the ideal gas law, calculate the molar mass or density of a gas.
- Reduce the ideal gas law to Boyle's law, Charles's law, and Avogadro's law. Describe the conditions under which each applies.

In Section 10-3, you learned about three quantities—pressure, volume, and temperature—needed to describe a gas sample. A gas sample can be further characterized using a fourth quantity—the number of moles. The number of molecules or moles present will always affect at least one of the other three quantities. The collision rate of molecules per unit area of container wall depends on the number of molecules present. If the number of molecules is increased for a sample at constant volume and temperature, the collision rate increases. Therefore, the pressure increases, as shown by the model in Figure 11-4(a). Consider what would happen if the pressure and temperature were kept constant while the number of molecules increased. According to Avogadro's law, the volume would increase. As Figure 11-4(b) shows, an increase in volume keeps the pressure constant at constant temperature. Increasing the volume keeps the collision rate per unit of wall area constant.

You can see that gas pressure, volume, temperature, and the number of moles are all interrelated. There is a mathematical relationship that describes the behavior of a gas sample for any combination of these conditions. *The* **ideal gas law** *is the mathematical relationship among pressure, volume, temperature, and the number of moles of a gas.* It is the equation of state for an ideal gas, because the state of a gas can be defined by its pressure, volume, temperature, and number of moles.

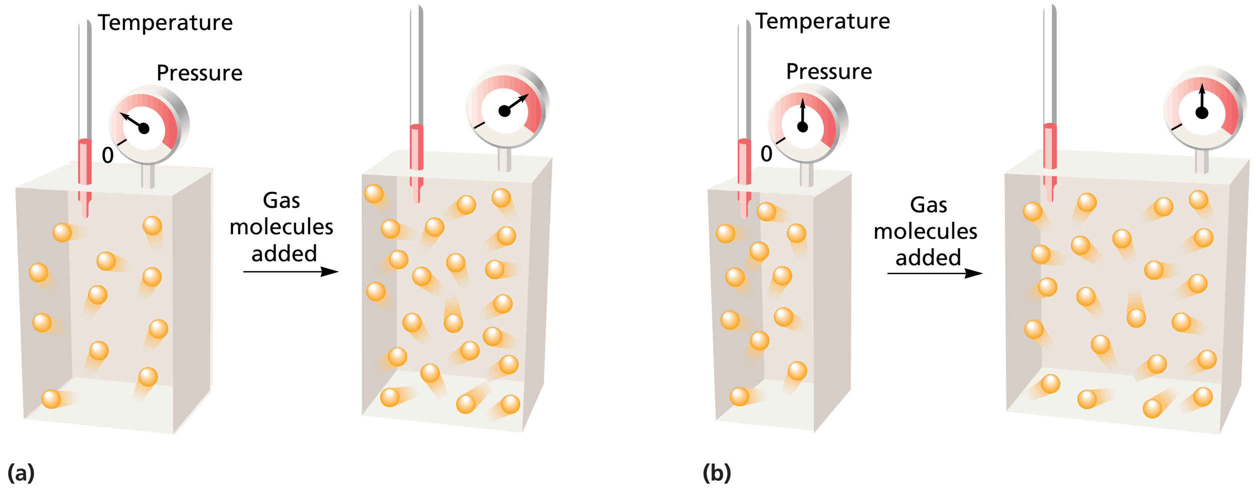


FIGURE 11-4 (a) When volume and temperature are constant, gas pressure increases as the number of molecules increases. (b) When pressure and temperature are constant, gas volume increases as the number of molecules increases.

Derivation of the Ideal Gas Law

The general equation that can be used to calculate unknown information about gas samples can be derived by combining Boyle's law and Charles's law with a statement that follows logically from Avogadro's law. First, consider each of those laws and the principle again.

Boyle's law: At constant temperature, the volume of a given mass of gas is inversely proportional to the pressure.

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

Charles's law: At constant pressure, the volume of a given mass of gas is directly proportional to the Kelvin temperature.

$$V \propto T$$

Avogadro's law implies that at constant pressure and temperature, the volume of a given mass of a gas is directly proportional to the number of moles.

$$V \propto n$$

A quantity—in this case, volume—that is proportional to each of several quantities is also proportional to their product. Therefore, combining the three relationships above gives the following.

$$V \propto \frac{1}{P} \times T \times n$$

Mathematically, you can change a proportionality to an equality by introducing a constant. In this case, the symbol R is used for the constant.

$$V = R \times \frac{1}{P} \times T \times n$$

R represents the value that the quantity PV/nT approaches for any gas whose behavior approaches that of an ideal gas. The equation for the ideal gas law is derived as follows.

$$V = \frac{nRT}{P} \text{ or } PV = nRT$$

This equation states that the volume of a gas varies directly with the number of moles (or molecules) of a gas and its Kelvin temperature. The volume also varies inversely with the pressure. Under ordinary conditions, most gases exhibit behavior that is nearly ideal. The equation can then be applied with reasonable accuracy.

The ideal gas law reduces to Boyle's law, Charles's law, Gay-Lussac's law, or Avogadro's law when the appropriate variables are held constant. For example, if n and T are constant, the product nRT is constant because R is also constant. In this case, the ideal gas law reduces to PV = a constant, which is Boyle's law.

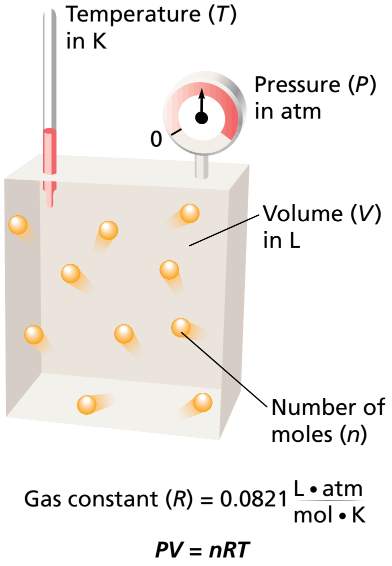


FIGURE 11-5 The ideal gas law relates the pressure, volume, number of moles, and temperature of an ideal gas. Note the value of the gas constant, R, when the other quantities are in the units shown.

The Ideal Gas Constant

In the equation representing the ideal gas law, *the constant R is known as the* **ideal gas constant.** Its value depends on the units chosen for pressure, volume, and temperature. Figure 11-5 shows that measured values of P, V, T, and n for a gas at near-ideal conditions can be used to calculate R. Recall from Section 11-1 that the volume of one mole of an ideal gas at STP (1 atm and 273.15 K) is 22.414 10 L. Substituting these values and solving the ideal gas law equation for R gives the following.

$$R = \frac{PV}{nT} = \frac{(1\ \text{atm})(22.414\ 10\ \text{L})}{(1\ \text{mol})(273.15\ \text{K})} = 0.082\ 057\ 84\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

This calculated value of R is usually rounded to 0.0821 L•atm/(mol•K). Use this value in ideal gas law calculations when the volume is in liters, the pressure is in atmospheres, and the temperature is in kelvins. See Table 11-1 for the value of R when other units for n, P, V, and T are used.

Finding *P, V, T,* or *n* from the Ideal Gas Law

The ideal gas law can be applied to determine the existing conditions of a gas sample when three of the four variables, P, V, T, and n, are known. It can also be used to calculate the molar mass or density of a gas sample.

Be sure to match the units of the known quantities and the units of R. In this book, you will be using R = 0.0821 L•atm/(mol•K). Your first step in solving any ideal gas law problem should be to check the known values to be sure you are working with the correct units. If necessary, you must convert volumes to liters, pressures to atmospheres, temperatures to kelvins, and masses to numbers of moles before using the ideal gas law.

TABLE 11-1 *Numerical Values of the Gas Constant, R*

Unit of *R*	Numerical value of *R*	Unit of *P*	Unit of *V*	Unit of *T*	Unit of *n*
$\frac{\text{L} \cdot \text{mm Hg}}{\text{mol} \cdot \text{K}}$	62.4	mm Hg	L	K	mol
$\frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$	0.0821	atm	L	K	mol
$\frac{\text{J}}{\text{mol} \cdot \text{K}}$	8.314*	Pa	m^3	K	mol
$\frac{\text{L} \cdot \text{kPa}}{\text{mol} \cdot \text{K}}$	8.314	kPa	L	K	mol

Note: 1 L•atm = 101.325 J; 1 J = 1 Pa•m^3
* SI units

SAMPLE PROBLEM 11-3

What is the pressure in atmospheres exerted by a 0.500 mol sample of nitrogen gas in a 10.0 L container at 298 K?

SOLUTION

1 ANALYZE

Given: V of N_2 = 10.0 L
n of N_2 = 0.500 mol
T of N_2 = 298 K

Unknown: P of N_2 in atm

2 PLAN

$$n, V, T \longrightarrow P$$

The gas sample undergoes no change in conditions. Therefore, the ideal gas law can be rearranged and used to find the pressure as follows.

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

3 COMPUTE

$$P = \frac{(0.500\ \cancel{\text{mol}})\left(\dfrac{0.0821\ \cancel{\text{L}} \cdot \text{atm}}{\cancel{\text{mol}} \cdot \cancel{\text{K}}}\right)(298\ \cancel{\text{K}})}{10.0\ \cancel{\text{L}}} = 1.22\ \text{atm}$$

4 EVALUATE

All units cancel correctly to give the result in atmospheres. The answer is properly limited to three significant figures. It is also close to an estimated value of 1.5, computed as $(0.5 \times 0.1 \times 300)/10$.

PRACTICE

1. What pressure, in atmospheres, is exerted by 0.325 mol of hydrogen gas in a 4.08 L container at 35°C? — *Answer* 2.01 atm
2. A gas sample occupies 8.77 L at 20°C. What is the pressure, in atmospheres, given that there are 1.45 mol of gas in the sample? — *Answer* 3.98 atm

SAMPLE PROBLEM 11-4

What is the volume, in liters, of 0.250 mol of oxygen gas at 20.0°C and 0.974 atm pressure?

SOLUTION

1 ANALYZE

Given: P of O_2 = 0.974 atm
n of O_2 = 0.250 mol
To use the value 0.0821 L•atm/(mol•K) for R, the temperature (°C) must be converted to kelvins.
T of O_2 = 20.0°C + 273.2 = 293.2 K

Unknown: V of O_2 in L

2 *PLAN*

$$P, n, T \longrightarrow V$$

The ideal gas law can be rearranged to solve for V for this sample, which undergoes no change in conditions.

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

3 *COMPUTE*

$$V = \frac{(0.250\ \cancel{\text{mol}}\ O_2)\left(\dfrac{0.0821\ \text{L} \cdot \cancel{\text{atm}}}{\cancel{\text{mol}} \cdot \cancel{\text{K}}}\right)(293.2\ \cancel{\text{K}})}{0.974\ \cancel{\text{atm}}} = 6.17\ \text{L}\ O_2$$

4 *EVALUATE* Units cancel to give liters, as desired. The answer is correctly limited to three significant figures. It is also close to an estimated value of 6, calculated as $0.2 \times 0.1 \times 300$.

PRACTICE

1. A sample that contains 4.38 mol of a gas at 250 K has a pressure of 0.857 atm. What is the volume? *Answer* 105 L
2. How many liters are occupied by 0.909 mol of nitrogen at 125°C and 0.901 atm pressure? *Answer* 33.0 L N_2

SAMPLE PROBLEM 11-5

What mass of chlorine gas, Cl_2, in grams, is contained in a 10.0 L tank at 27°C and 3.50 atm of pressure?

SOLUTION

1 *ANALYZE*

Given: P of Cl_2 = 3.50 atm
V of Cl_2 = 10.0 L
T of Cl_2 = 27°C + 273 = 300. K

Unknown: mass of Cl_2 in g

2 *PLAN* The ideal gas law can be rearranged to solve for n after temperature is converted to kelvins.

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

The number of moles is then converted to grams.

$$m\ (\text{in g}) = n \times \frac{\text{g}}{\text{mol}}$$

3 *COMPUTE*

$$n = \frac{(3.50\ \cancel{\text{atm}})(10.0\ \cancel{\text{L}}\ Cl_2)}{\left(\dfrac{0.0821\ \cancel{\text{L}} \cdot \cancel{\text{atm}}}{\text{mol} \cdot \cancel{\text{K}}}\right)(300.\ \cancel{\text{K}})} = 1.42\ \text{mol}\ Cl_2$$

$$\text{mass of } Cl_2 = 1.42\ \cancel{\text{mol}} \times \frac{70.90\ \text{g}\ Cl_2}{\cancel{\text{mol}}} = 101\ \text{g}\ Cl_2$$

4 *EVALUATE* Units cancel to leave the desired unit. The result is correctly given to three significant figures. The answer is close to the estimated value.

PRACTICE

1. How many grams of carbon dioxide gas are there in a 45.1 L container at 34°C and 1.04 atm? *Answer* 81.9 g CO_2

2. What is the mass, in grams, of oxygen gas in a 12.5 L container at 45°C and 7.22 atm? *Answer* 111 g O_2

3. A sample of carbon dioxide with a mass of 0.30 g was placed in a 250 mL container at 400. K. What is the pressure exerted by the gas? *Answer* 0.90 atm

Finding Molar Mass or Density from the Ideal Gas Law

Suppose that the pressure, volume, temperature, and mass are known for a gas sample. You can calculate the number of moles (n) in the sample, using the ideal gas law. Then you can calculate the molar mass (grams per mole) by dividing the known mass by the number of moles.

An equation showing the relationship between density, pressure, temperature, and molar mass can be derived from the ideal gas law. The number of moles (n) is equal to mass (m) divided by molar mass (M). Substituting m/M for n into $PV = nRT$ gives the following.

$$PV = \frac{mRT}{M} \quad \text{or} \quad M = \frac{mRT}{PV}$$

Density (D) is mass (m) per unit volume (V). Writing this definition in the form of an equation gives $D = m/V$. You can see that m/V appears in the right-hand equation above. Introducing density into that equation gives the following.

$$M = \frac{mRT}{PV} = \frac{DRT}{P}$$

Solving for density gives this equation.

$$D = \frac{MP}{RT}$$

You can see that the density of a gas varies directly with molar mass and pressure and inversely with Kelvin temperature.

SAMPLE PROBLEM 11-6

At 28°C and 0.974 atm, 1.00 L of gas has a mass of 5.16 g. What is the molar mass of this gas?

SOLUTION

1 *ANALYZE*

Given: P of gas = 0.974 atm
V of gas = 1.00 L
T of gas = 28°C + 273 = 301 K
m of gas = 5.16 g

Unknown: M of gas in g/mol

2 PLAN

$$P, V, T, m \longrightarrow M$$

You can use the rearranged ideal gas law provided earlier in this section to find the answer.

$$M = \frac{mRT}{PV}$$

3 COMPUTE

$$M = \frac{(5.16\text{ g})\left(\dfrac{0.0821\ \cancel{\text{L}} \cdot \cancel{\text{atm}}}{\text{mol} \cdot \cancel{\text{K}}}\right)(301\ \cancel{\text{K}})}{(0.974\ \cancel{\text{atm}})(1.00\ \cancel{\text{L}})} = 131\text{ g/mol}$$

4 EVALUATE

Units cancel as needed. The answer is correctly given to three significant figures. It is also close to an estimated value of 150, calculated as $(5 \times 0.1 \times 300)/1$.

PRACTICE

1. What is the molar mass of a gas if 0.427 g of the gas occupies a volume of 125 mL at 20.0°C and 0.980 atm? *Answer* 83.8 g/mol
2. What is the density of a sample of ammonia gas, NH_3, if the pressure is 0.928 atm and the temperature is 63.0°C? *Answer* 0.572 g/L NH_3
3. The density of a gas was found to be 2.0 g/L at 1.50 atm and 27°C. What is the molar mass of the gas? *Answer* 33 g/mol
4. What is the density of argon gas, Ar, at a pressure of 551 torr and a temperature of 25°C? *Answer* 1.18 g/L Ar

SECTION REVIEW

1. How does the ideal gas law reduce to the following:
   a. Boyle's law
   b. Charles's law
   c. Gay-Lussac's law
   d. Avogadro's law
2. What is the volume, in liters, of 0.100 g of $C_2H_2F_4$ vapor at 0.0928 atm and 22.3°C?
3. Why must the units P, V, T, and n match up to those of the ideal gas constant in solving problems? What would be the units for R if P is in pascals, T is in kelvins, V is in liters, and n is in moles?
4. What is the molar mass of a 1.25 g sample of gas that occupies a volume of 1.00 L at a pressure of 0.961 atm and a temperature of 27.0°C?
5. Name two quantities besides pressure, mass, volume, and number of moles that can be calculated using the ideal gas law.

Stoichiometry of Gases

SECTION 11-3

Objectives

- Explain how Gay-Lussac's law and Avogadro's law apply to the volumes of gases in chemical reactions.
- Use a chemical equation to specify volume ratios for gaseous reactants or products, or both.
- Use volume ratios and the gas laws to calculate volumes, masses, or molar amounts of gaseous reactants or products.

You can apply the discoveries of Gay-Lussac and Avogadro to calculate the stoichiometry of reactions involving gases. For gaseous reactants or products, the coefficients in chemical equations not only indicate molar amounts and mole ratios but also reveal volume ratios. For example, consider the reaction of carbon monoxide with oxygen to give carbon dioxide.

$2CO(g)$	+	$O_2(g)$	$\longrightarrow$	$2CO_2(g)$
2 molecules		1 molecule		2 molecules
2 mol		1 mol		2 mol
2 volumes		1 volume		2 volumes

The possible volume ratios can be expressed in the following ways.

$$\textbf{a.}\quad \frac{2 \text{ volumes CO}}{1 \text{ volume } O_2} \text{ or } \frac{1 \text{ volume } O_2}{2 \text{ volumes CO}}$$

$$\textbf{b.}\quad \frac{2 \text{ volumes CO}}{2 \text{ volumes } CO_2} \text{ or } \frac{2 \text{ volumes } CO_2}{2 \text{ volumes CO}} \qquad \textbf{c.}\quad \frac{1 \text{ volume } O_2}{2 \text{ volumes } CO_2} \text{ or } \frac{2 \text{ volumes } CO_2}{1 \text{ volume } O_2}$$

Volumes can be compared in this way only if all are measured at the same temperature and pressure.

Volume-Volume Calculations

Suppose the volume of a gas involved in a reaction is known and you need to find the volume of another gaseous reactant or product, assuming both reactant and product exist under the same conditions. Use volume ratios like those given above in exactly the same way you would use mole ratios.

SAMPLE PROBLEM 11-7

Propane, C_3H_8, is a gas that is sometimes used as a fuel for cooking and heating. The complete combustion of propane occurs according to the following equation.

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

(a) What will be the volume, in liters, of oxygen required for the complete combustion of 0.350 L of propane? (b) What will be the volume of carbon dioxide produced in the reaction? Assume that all volume measurements are made at the same temperature and pressure.

SOLUTION

1 ANALYZE

Given: balanced chemical equation
V of propane = 0.350 L

Unknown: **a.** V of O_2 in L; **b.** V of CO_2 in L

2 PLAN

a. V of $C_3H_8 \longrightarrow V$ of O_2; **b.** V of $C_3H_8 \longrightarrow V$ of CO_2

All volumes are to be compared at the same temperature and pressure. Therefore, volume ratios can be used like mole ratios to find the unknowns.

3 COMPUTE

$$\textbf{a. } 0.350\ \cancel{\text{L } C_3H_8} \times \frac{5\text{ L } O_2}{1\ \cancel{\text{L } C_3H_8}} = 1.75\text{ L } O_2$$

$$\textbf{b. } 0.350\ \cancel{\text{L } C_3H_8} \times \frac{3\text{ L } CO_2}{1\ \cancel{\text{L } C_3H_8}} = 1.05\text{ L } CO_2$$

4 EVALUATE

Each result is correctly given to three significant figures. The answers are reasonably close to estimated values of 2, calculated as 0.4 × 5, and 1.2, calculated as 0.4 × 3, respectively.

PRACTICE

1. Assuming all volume measurements are made at the same temperature and pressure, what volume of hydrogen gas is needed to react completely with 4.55 L of oxygen gas to produce water vapor? *Answer* 9.10 L H_2

2. What volume of oxygen gas is needed to react completely with 0.626 L of carbon monoxide gas, CO, to form gaseous carbon dioxide? Assume all volume measurements are made at the same temperature and pressure. *Answer* 0.313 L O_2

CHEMISTRY INTERACTIVE TUTOR

Module 5: Equations and Stoichiometry

Volume-Mass and Mass-Volume Calculations

Stoichiometric calculations may involve both masses and gas volumes. Sometimes the volume of a reactant or product is given and the mass of a second gaseous substance is unknown. In other cases, a mass amount may be known and a volume may be the unknown. The calculations require routes such as the following.

gas volume A ⟶ moles A ⟶ moles B ⟶ mass B

or

mass A ⟶ moles A ⟶ moles B ⟶ gas volume B

To find the unknown in cases like these, you must know the conditions under which both the known and unknown gas volumes have been measured. The ideal gas law is useful for calculating values at standard and nonstandard conditions.

SAMPLE PROBLEM 11-8

Calcium carbonate, $CaCO_3$, also known as limestone, can be heated to produce calcium oxide (lime), an industrial chemical with a wide variety of uses. The balanced equation for the reaction follows.

$$\mathbf{CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)}$$

How many grams of calcium carbonate must be decomposed to produce 5.00 L of carbon dioxide gas at STP?

SOLUTION

1 ANALYZE

Given: balanced chemical equation
desired volume of CO_2 produced at STP = 5.00 L

Unknown: mass of $CaCO_3$ in grams

2 PLAN

The known volume is given at STP. This tells us the pressure and temperature. The ideal gas law can be used to find the moles of CO_2. The mole ratios from the balanced equation can then be used to calculate the moles of $CaCO_3$ needed. (Note that volume ratios do not apply here because calcium carbonate is a solid.)

3 COMPUTE

$$n = \frac{PV}{RT} = \frac{(1\ \cancel{\text{atm}})(5.00\ \cancel{\text{L}}\ CO_2)}{\left(\frac{0.0821\ \cancel{\text{L}} \cdot \cancel{\text{atm}}}{\text{mol} \cdot \cancel{\text{K}}}\right)(273\ \cancel{\text{K}})} = 0.223 \text{ mol } CO_2$$

$$0.223\ \cancel{\text{mol } CO_2} \times \frac{1\ \cancel{\text{mol } CaCO_3}}{1\ \cancel{\text{mol } CO_2}} \times \frac{100.09 \text{ g } CaCO_3}{1\ \cancel{\text{mol } CaCO_3}} = 22.3 \text{ g } CaCO_3$$

4 EVALUATE

Units all cancel correctly. The answer is properly given to three significant figures. It is close to an estimated value of 20, computed as $(0.2 \times 100)/25$.

PRACTICE

1. What mass of sulfur must be used to produce 12.61 L of gaseous sulfur dioxide at STP according to the following equation? *Answer* 18.0 g S_8

$$S_8(s) + 8O_2(g) \longrightarrow 8SO_2(g)$$

2. How many grams of water can be produced from the complete reaction of 3.44 L of oxygen gas, at STP, with hydrogen gas? *Answer* 5.53 g H_2O

SAMPLE PROBLEM 11-9

Tungsten, W, a metal used in light-bulb filaments, is produced industrially by the reaction of tungsten oxide with hydrogen.

$$\mathbf{WO_3(s) + 3H_2(g) \longrightarrow W(s) + 3H_2O(l)}$$

How many liters of hydrogen gas at 35°C and 0.980 atm are needed to react completely with 875 g of tungsten oxide?

SOLUTION

1 ANALYZE

Given: balanced chemical equation
reactant mass of WO_3 = 875 g
P of H_2 = 0.980 atm
T of H_2 = 35°C + 273 = 308 K

Unknown: V of H_2 in L at known nonstandard conditions

2 PLAN

Moles of H_2 are found by converting the mass of WO_3 to moles and then using the mole ratio. The ideal gas law is used to find the volume from the calculated number of moles of H_2.

3 COMPUTE

$$875\ \cancel{\text{g WO}_3} \times \frac{1\ \cancel{\text{mol WO}_3}}{231.84\ \cancel{\text{g WO}_3}} \times \frac{3\ \text{mol H}_2}{1\ \cancel{\text{mol WO}_3}} = 11.3\ \text{mol H}_2$$

$$V = \frac{nRT}{P} = \frac{(11.3\ \cancel{\text{mol}}\ \text{H}_2)\left(\dfrac{0.0821\ \text{L} \cdot \cancel{\text{atm}}}{\cancel{\text{mol}} \cdot \cancel{\text{K}}}\right)(308\ \cancel{\text{K}})}{0.980\ \cancel{\text{atm}}} = 292\ \text{L H}_2$$

4 EVALUATE

Unit cancellations are correct, as is the use of three significant figures for each result. The answer is reasonably close to an estimated value of 330, computed as (11 × 0.1 × 300)/1.

PRACTICE

1. What volume of chlorine gas at 38°C and 1.63 atm is needed to react completely with 10.4 g of sodium to form NaCl? *Answer* 3.54 L Cl_2

2. How many liters of gaseous carbon monoxide at 27°C and 0.247 atm can be produced from the burning of 65.5 g of carbon according to the following equation? *Answer* 544 L CO

$$2C(s) + O_2(g) \longrightarrow 2CO(g)$$

SECTION REVIEW

1. How many liters of ammonia gas can be formed from the reaction of 150. L of hydrogen gas? Assume that there is complete reaction of hydrogen with excess nitrogen gas and that all measurements are made at the same temperature and pressure.

2. How many liters of H_2 gas at STP can be produced by the reaction of 4.60 g of Na and excess water, according to the following equation?

$$2Na(s) + 2H_2O(l) \longrightarrow H_2(g) + 2NaOH(aq)$$

3. How many grams of Na are needed to react with H_2O to liberate 4.00×10^2 mL of H_2 gas at STP?

4. What volume of oxygen gas in liters can be collected at 0.987 atm pressure and 25.0°C when 30.6 g of $KClO_3$ decompose by heating, according to the following equation?

$$2KClO_3(s) \xrightarrow[MnO_2]{\Delta} 2KCl(s) + 3O_2(g)$$

Effusion and Diffusion

SECTION 11-4

OBJECTIVES

- State Graham's law of effusion.
- Determine the relative rates of effusion of two gases of known molar masses.
- State the relationship between the molecular velocities of two gases and their molar masses.

The constant motion of gas molecules causes them to spread out to fill any container in which they are placed. As you learned in Chapter 10, the gradual mixing of two gases due to their spontaneous, random motion is known as *diffusion,* illustrated in Figure 11-6. *Effusion* is the process whereby the molecules of a gas confined in a container randomly pass through a tiny opening in the container. In this section, you will learn how effusion can be used to estimate the molar mass of a gas.

Graham's Law of Effusion

The rates of effusion and diffusion depend on the relative velocities of gas molecules. The velocity of a gas varies inversely with its mass. Lighter molecules move faster than heavier molecules at the same temperature.

Recall that the average kinetic energy of the molecules in any gas depends only on the temperature and equals $\frac{1}{2}mv^2$. For two different gases, A and B, at the same temperature, the following relationship is true.

$$\frac{1}{2}M_A v_A{}^2 = \frac{1}{2}M_B v_B{}^2$$

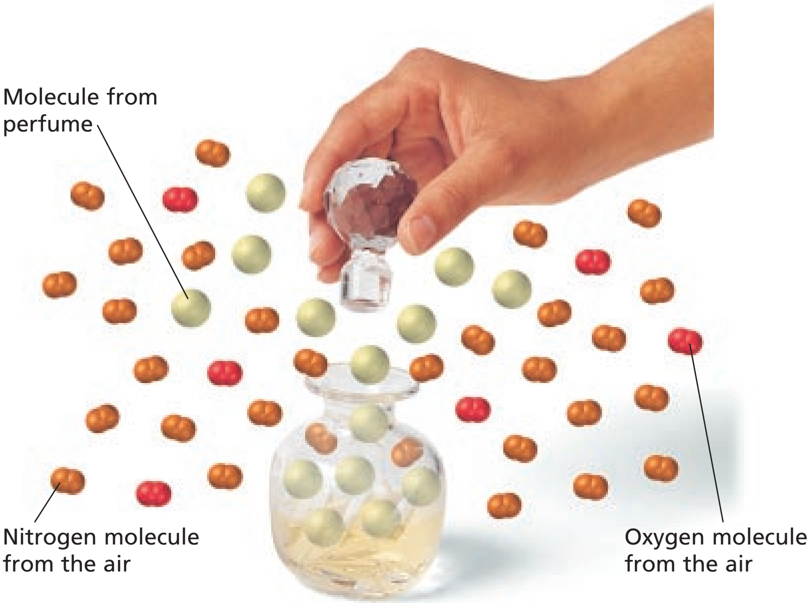
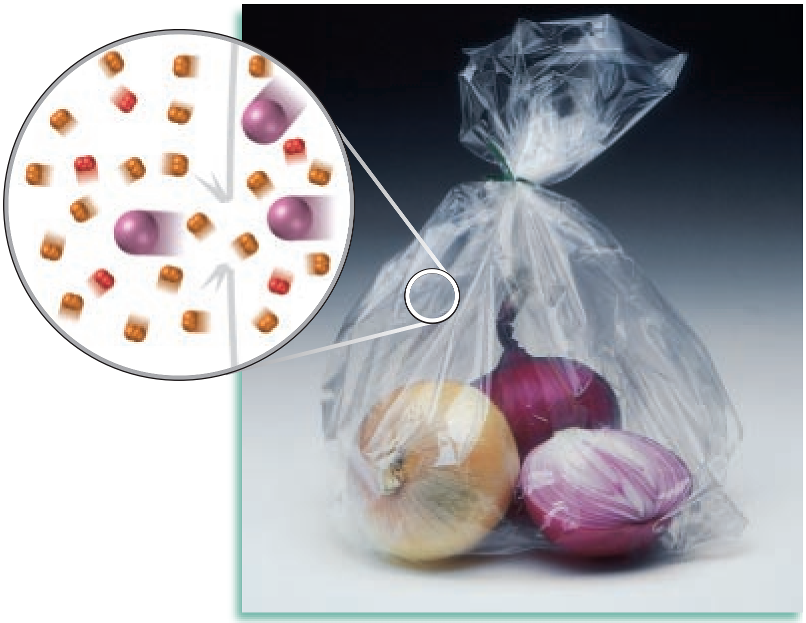


FIGURE 11-6 When a bottle of perfume is opened, some of its molecules diffuse into the air and mix with the molecules in the air. At the same time, molecules from the air, such as nitrogen and oxygen, diffuse into the bottle and mix with the gaseous scent molecules.

FIGURE 11-7 If you can smell the onions even when the bag is tightly sealed, molecules from the onions may be effusing through a small hole in the bag.



M_A and M_B represent the molar masses of gases A and B, and v_A and v_B represent their molecular velocities. Multiplying the equation by 2 gives the following.

$$M_A v_A^2 = M_B v_B^2$$

Suppose you wanted to compare the velocities of the two gases. You would first rearrange the equation above to give the velocities as a ratio.

$$\frac{v_A^2}{v_B^2} = \frac{M_B}{M_A}$$

The square root of each side of the equation is then taken.

$$\frac{v_A}{v_B} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$$

This equation shows that the molecular velocities of two different gases are inversely proportional to the square roots of their molar masses. Because the rates of effusion are directly proportional to molecular velocities, the equation can be written as follows.

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

In the mid-1800s, the Scottish chemist Thomas Graham studied the effusion and diffusion of gases. Figure 11-7 illustrates the process of effusion. Compare this with the diffusion process. The equation derived above is a mathematical statement of some of Graham's discoveries. It describes the rates of effusion. **Graham's law of effusion** *states that the rates of effusion of gases at the same temperature and pressure are inversely proportional to the square roots of their molar masses.*

Quick LAB

Diffusion

Wear Safety Goggles and an Apron.

Materials

- household ammonia
- perfume or cologne
- two 250 mL beakers
- Two watch glasses
- 10 mL graduated cylinder
- clock or watch with second hand

Question

Do different gases diffuse at different rates?

Procedure

Record all of your results in a data table.

1. Outdoors or in a room separate from the one in which you will carry out the rest of the investigation, pour approximately 10 mL of the household ammonia into one of the 250 mL beakers, and cover it with a watch glass. Pour roughly the same amount of perfume or cologne into the second beaker. Cover it with a watch glass also.
2. Take the two samples you just prepared into a large, draft-free room. Place the samples about 12 to 15 feet apart and at the same height. Position someone as the observer midway between the two beakers. Remove both watch-glass covers at the same time.
3. Note whether the observer smells the ammonia or the perfume first. Record how long this takes. Also, record how long it takes the vapor of the other substance to reach them. Air the room after you have finished.

Discussion

1. What do the times that the two vapors took to reach the observer show about the two gases?
2. What factors other than molecular mass (which determines diffusion rate) could affect how quickly the observer smells each vapor?

Applications of Graham's Law

Graham's experiments dealt with the densities of gases. The density of a gas varies directly with its molar mass. Therefore, the square roots of the molar masses in the equation on page 352 can be replaced by the square roots of the gas densities, giving the following relationship.

$$\frac{\text{rate of effusion of A}}{\text{rate of effusion of B}} = \frac{\sqrt{M_B}}{\sqrt{M_A}} = \frac{\sqrt{\text{density}_B}}{\sqrt{\text{density}_A}}$$

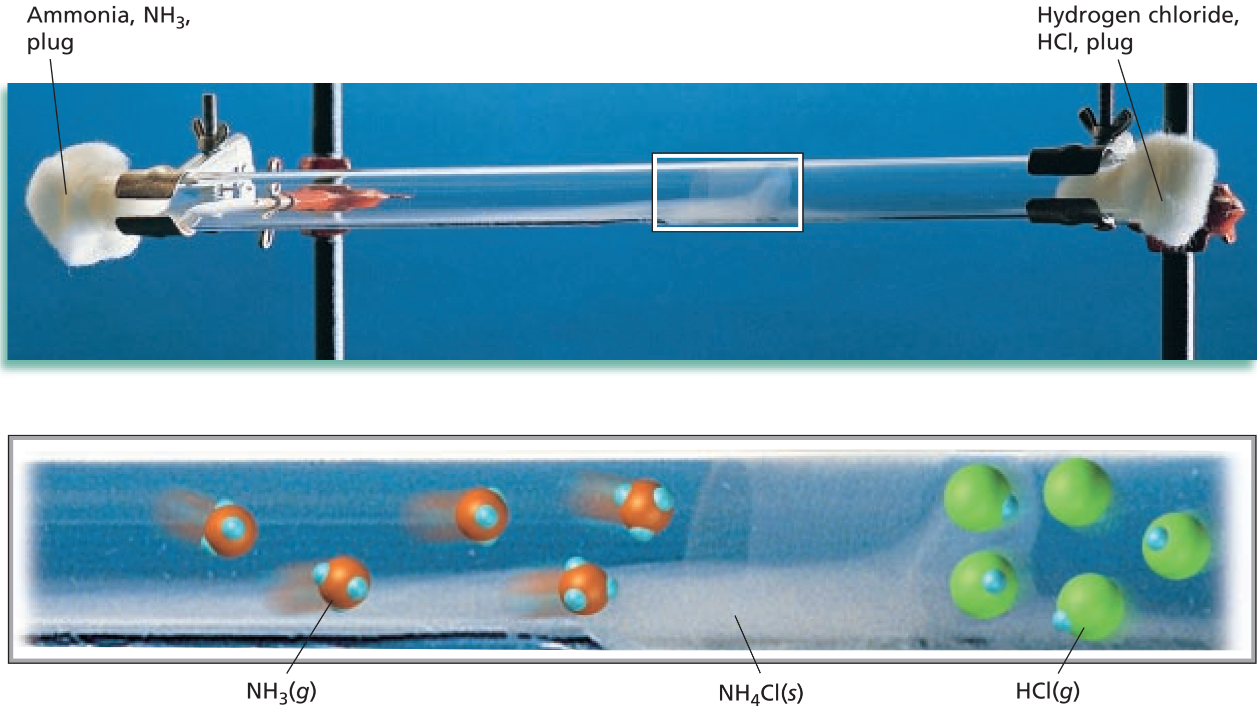
In the experiment shown in Figure 11-8, ammonia gas, NH_3, and hydrogen chloride gas, HCl, diffuse toward each other from opposite ends of a glass tube. A white ring forms at the point where the two gases meet and combine chemically. The white ring is solid ammonium chloride, NH_4Cl. Notice that it forms closer to the HCl end of the tube than to the NH_3 end. The NH_3 has diffused faster than the HCl.

If the two gases had equal vapor pressures (resulting in equal concentrations), this result could be attributed completely to differences in molar mass. The lighter NH_3 molecules (molar mass 17.04 g) diffuse faster than the heavier HCl molecules (molar mass 36.46 g). However, diffusion rates depend on both molar mass and concentration.

The diffusion of the two gases in the tube takes longer than would be predicted from the known velocities of NH_3 and HCl. This is because the experiment does not take place in a vacuum. The oxygen and nitrogen molecules present in the air collide with the diffusing ammonia and hydrogen chloride molecules, slowing them down. To describe all these random collisions taking place in this system would be very complicated. This is why we use these equations to represent the *relative* rates of effusion based on the *relative* velocities of the molecules.

Graham's law also provides a method for determining molar masses. The rates of effusion of gases of known and unknown molar mass can be compared at the same temperature and pressure. The unknown molar mass can then be calculated using Graham's law. One application of Graham's law was used in order to separate the heavier $^{238}_{92}U$ isotope from the lighter $^{235}_{92}U$ isotope. The uranium was converted to a gaseous compound and passed through porous membranes, where the isotopes diffused at different rates due to their different densities and were thereby separated.

FIGURE 11-8 Cotton plugs moistened with solutions of ammonia and hydrogen chloride were placed at opposite ends of the glass tube several minutes before this photograph was taken. Why does the white ring of ammonium chloride form closer to the right end than the left end?



SAMPLE PROBLEM 11-10

Compare the rates of effusion of hydrogen and oxygen at the same temperature and pressure.

SOLUTION

1 ANALYZE

Given: identities of two gases, H_2 and O_2

Unknown: relative rates of effusion

2 PLAN

molar mass ratio ⟶ ratio of rates of effusion

The ratio of the rates of effusion of two gases at the same temperature and pressure can be found from Graham's law.

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

3 COMPUTE

$$\frac{\text{rate of effusion of } H_2}{\text{rate of effusion of } O_2} = \frac{\sqrt{M_{O_2}}}{\sqrt{M_{H_2}}} = \frac{\sqrt{32.00 \text{ g/mol}}}{\sqrt{2.02 \text{ g/mol}}} = \sqrt{\frac{32.00 \text{ g/mol}}{2.02 \text{ g/mol}}} = 3.98$$

Hydrogen effuses 3.98 times faster than oxygen.

4 EVALUATE

The result is correctly reported to three significant figures. It is also approximately equivalent to an estimated value of 4, calculated as $\sqrt{32}/\sqrt{2}$.

PRACTICE

1. A sample of hydrogen effuses through a porous container about 9 times faster than an unknown gas. Estimate the molar mass of the unknown gas.

   Answer 160 g/mol

2. Compare the rate of effusion of carbon dioxide with that of hydrogen chloride at the same temperature and pressure.

   Answer CO_2 will effuse about 0.9 times as fast as HCl.

3. If a molecule of neon gas travels at an average of 400 m/s at a given temperature, estimate the average speed of a molecule of butane gas, C_4H_{10}, at the same temperature.

   Answer about 235 m/s

SECTION REVIEW

1. Compare diffusion with effusion.
2. Estimate the molar mass of a gas that effuses at 1.6 times the effusion rate of carbon dioxide.
3. List the following gases in order of increasing average molecular velocity at 25°C: H_2O, He, HCl, BrF, and NO_2.

CHAPTER 11 REVIEW

CHAPTER SUMMARY

11-1

- Gay-Lussac's law of combining volumes of gases states that the volumes of reacting gases and their products at the same temperature and pressure can be expressed as ratios of small whole numbers.
- Avogadro's law states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules. At constant temperature and pressure, gas volume is thus directly proportional to the number of moles.
- Gay-Lussac's law and Avogadro's law can be used to show that molecules of the reactive elemental gases are diatomic.
- The volume occupied by one mole of an ideal gas at STP is called the standard molar volume. The standard molar volume of an ideal gas is 22.414 10 L at standard temperature and pressure.

Vocabulary

Avogadro's law (334)
Gay-Lussac's law of combining volumes of gases (333)
standard molar volume of a gas (335)

11-2

- Charles's law, Boyle's law, and Avogadro's law can be combined to create the ideal gas law. The ideal gas law is stated mathematically as follows.

$$PV = nRT$$

- The value and units of the ideal gas constant depend on the units of the variables used in the ideal gas law.
- The ideal gas law can be used to calculate gas pressure, volume, temperature, or number of moles when three of these four variables are known and the gas sample does not undergo a change in conditions.
- The ideal gas law can also be used to calculate the density or molar mass of a gas sample.

Vocabulary

ideal gas constant (342)
ideal gas law (340)

11-3

- Given a balanced equation and the volume of a reacting gas, the volume of another gaseous product or reactant can be calculated using their volume ratios, as long as reactants and products exist under the same conditions.
- Given the volume of a reactant or product gas, the mass of a second reactant or product can be calculated using the ideal gas law and mole-to-mass conversion factors.
- If the mass of a substance is known, the ideal gas law and the proper mass-to-mole conversion factors can be used to calculate the volume of a gas.

11-4

- Graham's law of effusion states that the relative rates of effusion of gases at the same temperature and pressure are inversely proportional to the square roots of their molar masses.
- Graham's law reflects the fact that less massive molecules effuse faster than do more massive ones.
- Graham's law can be used to compare the rates of effusion of gases at the same temperature and pressure.
- Given the relative rates of effusion of two gases and the identity of one of them, Graham's law can be used to estimate the molar mass of the other.

Vocabulary

Graham's law of effusion (352)

REVIEWING CONCEPTS

1. a. What restrictions are there on the use of Gay-Lussac's law of combining volumes?
b. At the same temperature and pressure, what is the relationship between the volume of a gas and the number of molecules present? (11-1)

2. According to Avogadro,
a. what is the relationship between gas volume and number of moles at constant temperature and pressure?
b. what is the mathematical expression denoting this relationship? (11-1)

3. What is the relationship between the number of molecules and the mass of 22.4 L of different gases at STP? (11-1)

4. Why must the temperature and pressure be specified when stating gas density values? (11-1)

5. a. Write the equation for the ideal gas law.
b. What relationship is expressed in the ideal gas law? (11-2)

6. a. In what situation does the ideal gas law apply?
b. Why do you have to pay particular attention to units when using this law? (11-2)

7. a. In a balanced chemical equation, what is the relationship between the molar ratios and the volume ratios of gaseous reactants and products?
b. What restriction applies to the use of the volume ratios in solving stoichiometry problems? (11-3)

8. a. Distinguish between diffusion and effusion.
b. At a given temperature, what factor determines the rates at which different molecules undergo these processes? (11-4)

PROBLEMS

Molar Volume and Gas Density

9. Suppose a 5.00 L sample of O_2 at a given temperature and pressure contains 1.08×10^{23} molecules. How many molecules would be contained in each of the following at the same temperature and pressure?
a. 5.00 L H_2
b. 5.00 L CO_2
c. 10.00 L NH_3

10. How many molecules are contained in each of the following?
a. 1.00 mol O_2
b. 2.50 mol He
c. 0.0650 mol NH_3
d. 11.5 g NO_2

11. Find the mass of each of the following.
a. 2.25 mol Cl_2
b. 3.01×10^{23} molecules H_2S
c. 25.0 molecules SO_2

12. What is the volume, in liters, of each of the following at STP? (Hint: See Sample Problem 11-1.)
a. 1.00 mol O_2
b. 3.50 mol F_2
c. 0.0400 mol CO_2
d. 1.20×10^{-6} mol He

13. How many moles are contained in each of the following at STP?
a. 22.4 L N_2
b. 5.60 L Cl_2
c. 0.125 L Ne
d. 70.0 mL NH_3

14. Find the mass, in grams, of each of the following at STP. (Hint: See Sample Problem 11-2.)
a. 11.2 L H_2
b. 2.80 L CO_2
c. 15.0 mL SO_2
d. 3.40 cm^3 F_2

15. Find the volume, in liters, of each of the following at STP.
a. 8.00 g O_2
b. 3.50 g CO
c. 0.0170 g H_2S
d. 2.25×10^5 kg NH_3

Ideal Gas Law

16. Calculate the pressure, in atmospheres, exerted by each of the following. (Hint: See Sample Problem 11-3.)
a. 2.50 L of HF containing 1.35 mol at 320. K
b. 4.75 L of NO_2 containing 0.86 mol at 300. K
c. 7.50×10^2 mL of CO_2 containing 2.15 mol at 57°C

17. Calculate the volume, in liters, occupied by each of the following. (Hint: See Sample Problem 11-4.)
a. 2.00 mol of H_2 at 300. K and 1.25 atm
b. 0.425 mol of NH_3 at 37°C and 0.724 atm
c. 4.00 g of O_2 at 57°C and 0.888 atm

18. Determine the number of moles of gas contained in each of the following.
a. 1.25 L at 250. K and 1.06 atm
b. 0.80 L at 27°C and 0.925 atm
c. 7.50×10^2 mL at –50.°C and 0.921 atm

19. Find the mass of each of the following. (Hint: See Sample Problem 11-5.)
a. 5.60 L of O_2 at 1.75 atm and 250. K
b. 3.50 L of NH_3 at 0.921 atm and 27°C
c. 125 mL of SO_2 at 0.822 atm and –53°C

20. Find the molar mass of each gas measured at the specified conditions. (Hint: See Sample Problem 11-6.)
a. 0.650 g occupying 1.12 L at 280. K and 1.14 atm
b. 1.05 g occupying 2.35 L at 37°C and 0.840 atm
c. 0.432 g occupying 7.50×10^2 mL at –23°C and 1.03 atm

21. If the density of an unknown gas is 3.20 g/L at –18°C and 2.17 atm, what is the molar mass of this gas?

22. One method of estimating the temperature of the center of the sun is based on the assumption that the center consists of gases that have an average molar mass of 2.00 g/mol. If the density of the center of the sun is 1.40 g/cm^3 at a pressure of 1.30×10^9 atm, calculate the temperature in degrees Celsius.

Gas Stoichiometry

23. Carbon monoxide reacts with oxygen to produce carbon dioxide. If 1.0 L of carbon monoxide reacts with oxygen,
a. how many liters of oxygen are required? (Hint: See Sample Problem 11-7.)
b. how many liters of carbon dioxide are produced?

24. Acetylene gas, C_2H_2, undergoes combustion to produce carbon dioxide and water vapor. If 75.0 L of CO_2 are produced,
a. how many liters of C_2H_2 are required?
b. what volume of H_2O vapor is produced?
c. what volume of O_2 is required?

25. If liquid carbon disulfide reacts with 4.50×10^2 mL of oxygen to produce the gases carbon dioxide and sulfur dioxide, what volume of each product is produced?

26. Assume that 5.60 L of H_2 at STP react with CuO according to the following equation:
$CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(g)$
Make sure the equation is balanced before beginning your calculations.
a. How many moles of H_2 react? (Hint: See Sample Problem 11-8.)
b. How many moles of Cu are produced?
c. How many grams of Cu are produced?

27. Solid iron(III) hydroxide decomposes to produce iron(III) oxide and water vapor. If 0.75 L of water vapor is produced at STP,
a. how many grams of iron(III) hydroxide were used?
b. how many grams of iron(III) oxide are produced?

28. If 29.0 L of methane, CH_4, undergoes complete combustion at 0.961 atm and 20°C, how many liters of each product are formed?

29. If air is 20.9% oxygen by volume,
a. how many liters of air are needed for complete combustion of 25.0 L of octane vapor, C_8H_{18}?
b. what volume of each product is produced?

30. A modified Haber process for making ammonia is conducted at 550.°C and 2.50×10^2 atm. If 10.0 kg of nitrogen (the limiting reactant) is used and the process goes to completion, what volume of ammonia is produced?

31. When liquid nitroglycerin, $C_3H_5(NO_3)_3$, explodes, the products are carbon dioxide, nitrogen, oxygen, and water vapor. If 5.00×10^2 g of nitroglycerin explode at STP, what is the total volume, at STP, for all gases produced?

32. The principal source of sulfur on Earth is deposits of free sulfur occurring mainly in volcanically active regions. The sulfur was initially formed by the reaction between the two volcanic vapors SO_2 and H_2S to form $H_2O(l)$ and $S_8(s)$. What volume of each gas, at 0.961 atm and 22°C, was needed to form a sulfur deposit of 4.50×10^5 kg on the slopes of a volcano in Hawaii?

33. A 3.25 g sample of solid calcium carbide, CaC_2, reacted with water to produce acetylene gas, C_2H_2, and aqueous calcium hydroxide. If the acetylene was collected over water at 17°C and 0.974 atm, how many milliliters of acetylene were produced?

34. Balance the following chemical equation.
$Mg(s) + O_2(g) \longrightarrow MgO(s)$
Then, based on the quantity of reactant or product given, determine the corresponding quantities of the specified reactants or products, assuming that the system is at STP.
a. 22.4 L O_2 = ___ mol O_2 $\longrightarrow$ ___ mol MgO
b. 11.2 L O_2 = ___ mol O_2 $\longrightarrow$ ___ mol MgO
c. 1.40 L O_2 = ___ mol O_2 $\longrightarrow$ ___ mol MgO

35. Assume that 8.50 L of I_2 are produced using the following reaction that takes place at STP:
$KI(aq) + Cl_2(g) \longrightarrow KCl(aq) + I_2(g)$
Balance the equation before beginning your calculations.
a. How many moles of I_2 are produced?
b. How many moles of KI were used?
c. How many grams of KI were used?

36. Suppose that 6.50×10^2 mL of hydrogen gas are produced through a replacement reaction involving solid iron and sulfuric acid, H_2SO_4, at STP. How many grams of iron(II) sulfate are also produced?

37. Methanol, CH_3OH, is made by causing carbon monoxide and hydrogen gases to react at high temperature and pressure. If 4.50×10^2 mL of CO and 825 mL of H_2 are mixed,
a. which reactant is present in excess?
b. how much of that reactant remains after the reaction?
c. what volume of CH_3OH is produced, assuming the same pressure?

38. Assume that 13.5 g of Al react with HCl according to the following equation, at STP:
$Al(s) + HCl(aq) \longrightarrow AlCl_3(aq) + H_2(g)$
Remember to balance the equation first.
a. How many moles of Al react?
b. How many moles of H_2 are produced?
c. How many liters of H_2 at STP are produced? (Hint: See Sample Problem 11-9.)

Effusion and Diffusion

39. Quantitatively compare the rates of effusion for the following pairs of gases at the same temperature and pressure.
a. hydrogen and nitrogen (Hint: See Sample Problem 11-10.)
b. fluorine and chlorine

40. What is the ratio of the average velocity of hydrogen molecules to that of neon atoms at the same temperature and pressure?

41. At a certain temperature and pressure, chlorine molecules have an average velocity of 0.0380 m/s. What is the average velocity of sulfur dioxide molecules under the same conditions?

42. A sample of helium effuses through a porous container 6.50 times faster than does unknown gas X. What is the molar mass of the unknown gas?

MIXED REVIEW

43. An unknown gas effuses at 0.850 times the effusion rate of nitrogen dioxide, NO_2. Estimate the molar mass of the unknown gas.

44. Use the ideal gas law, $PV = nRT$, to derive Boyle's law and Charles's law.

45. A container holds 265 mL of chlorine gas, Cl_2. Assuming that the gas sample is at STP, what is its mass?

46. Suppose that 3.11 mol of carbon dioxide is at a pressure of 0.820 atm and a temperature of 39°C. What is the volume of the sample, in liters?

47. Compare the rates of diffusion of carbon monoxide, CO, and sulfur trioxide, SO_3.

48. A gas sample that has a mass of 0.993 g occupies 0.570 L. Given that the temperature is 281 K and the pressure is 1.44 atm, what is the molar mass of the gas?

49. The density of a gas is 3.07 g/L at STP. Calculate the gas's molar mass.

50. How many moles of helium gas would it take to fill a gas balloon with a volume of 1000. cm^3 when the temperature is 32°C and the atmospheric pressure is 752 mm Hg?

51. A gas sample is collected at 16°C and 0.982 atm. If the sample has a mass of 7.40 g and a volume of 3.96 L, find the volume of the gas at STP and the molar mass.

CRITICAL THINKING

52. Evaluating Methods In solving a problem, what types of conditions involving temperature, pressure, volume, or number of moles would allow you to use

a. the combined gas law?

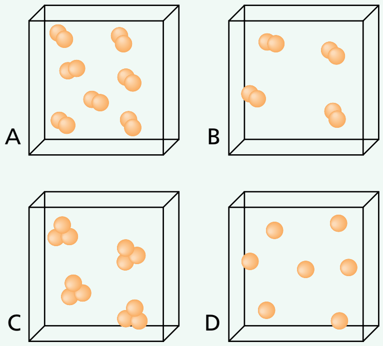
b. the ideal gas law?

53. Relating Ideas Write expressions relating the rates of effusion, molar masses, and densities of two different gases, A and B.

54. Evaluating Ideas Gay-Lussac's law of combining volumes holds true for relative volumes at any proportionate size. Use Avogadro's law to explain why this proportionality exists.

55. Designing Experiments Design an experiment to prove that the proportionality described in item 54 exists.

56. Interpreting Concepts The diagrams that follow represent equal volumes of four different gases.



Use the diagrams to answer the following questions:

a. Are these gases at the same temperature and pressure? How do you know?

b. If the molar mass of gas B is 38 g/mol and that of gas C is 46 g/mol, which gas sample is more dense?

c. To make the densities of gas samples B and C equal, which gas should expand in volume?

d. If the densities of gas samples A and C are equal, what is the relationship between their molecular masses?

TECHNOLOGY & LEARNING

57. Graphing Calculator Calculating Pressure Using the Ideal Gas Law

The graphing calculator can run a program that calculates the pressure in atmospheres, given the number of moles of a gas (n), volume (V), and temperature (T). Given a 0.50 mol gas sample with a volume of 10. L at 298 K, you can calculate the pressure according to the ideal gas law. Begin by using the program to carry out the calculation. Next, use it to make calculations.

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

Note: Answers are written with five significant figures.

a. What is the pressure for a gas with an amount of 1.3 mol, volume of 8.0 L, and temperature of 293 K?

b. What is the pressure for a gas with an amount of 2.7 mol, volume of 8.5 L, and temperature of 310 K?

c. A gas with an amount of 0.75 mol and a volume of 6.0 L is measured at two different temperatures: 300 K and 275 K. At which temperature is the pressure greater?

HANDBOOK SEARCH

58. Most elements from Groups 1, 2, and 13 will react with water, acids, or bases to produce hydrogen gas. Review the common reactions information in the *Elements Handbook* and answer the following:

a. Write the equation for the reaction of barium with water.

b. Write the equation for the reaction between cesium and hydrochloric acid.
c. Write the equation for the reaction of gallium with hydrofluoric acid.
d. What mass of barium would be needed to react with excess water to produce 10.1 L of H_2 at STP?
e. What masses of cesium and hydrochloric acid would be required to produce 10.1 L of H_2 at STP?

59. Group 1 metals react with oxygen to produce oxides, peroxides, or superoxides. Review the equations for these common reactions in the *Elements Handbook* and answer the following:
a. How do oxides, peroxides, and superoxides differ?
b. What mass of product will be formed from a reaction of 5.00 L of oxygen with excess sodium? The reaction occurs at 27°C and 1 atm.

60. Most metals react with chlorine from Group 17 to produce salts. Review these common reactions in the *Elements Handbook* and answer the following:
a. Write the equation for the reaction of iron and chlorine gas.
b. What mass of iron salt is produced if an excess of iron reacts with 450. mL of chlorine gas at 27°C and 1 atm?

RESEARCH & WRITING

61. How do scuba divers use the laws and principles that describe the behavior of gases to their advantage? What precautions do they take to prevent the bends?

62. Explain the processes involved in the liquefaction of gases. What substances that are gases under normal room conditions are typically used in the liquid form? Why?

63. Research the relationship between explosives and the establishment of Nobel Prizes. Prepare a report that describes your findings.

64. Write a summary describing how Gay-Lussac's work on combining volumes relates to Avogadro's study of gases. Explain how certain conclusions about gases followed logically from consideration of the work of both scientists.

ALTERNATIVE ASSESSMENT

65. During a typical day, record every instance in which you encounter the diffusion or effusion of gases (for example, smelling perfume).

66. **Performance** Qualitatively compare the molecular masses of various gases by noting how long it takes you to smell them from a fixed distance. Work only with materials that are not dangerous, such as flavor extracts, fruit peels, and onions.

67. **Performance** Design an experiment to gather data to verify the Ideal Gas Law. If your teacher approves of your plan, carry it out. Illustrate your data with a graph, and determine if the data are consistent with the Ideal Gas Law.