CHAPTER 12

Exploring Gas Laws

You are probably used to sharing scientific ideas and observations with your classmates as you work in pairs or in small groups. If you are familiar with Internet chat rooms and e-mail, you may even share your ideas with people around the world. Sharing scientific ideas is an essential part of scientific discovery. Back in the early nineteenth century, the ideas that led to the complete gas laws were shared between colleagues in much the same way that you share scientific ideas with your classmates.

During the nineteenth century, scientists across Europe organized *academies*, or science societies. (This practice still exists today.) Belonging to an academy allowed much more communication among scientists. They read the papers and reports that other scientists had published, wrote letters, and held meetings to discuss ideas. They worked together to develop many important theories and laws, including the gas laws.

In this chapter, you will study more of the gas laws. Although these laws were discovered almost 200 years ago, they are still accepted today. You will learn how the gas laws allow you to solve many problems involving gas behaviour. You will also learn about some of the modern applications and technological advances associated with the gas laws. You will have a chance to test the gas laws through your own experiments and to compare your results with those of early scientists. Finally, you will learn about some of the gas reactions that occur in our atmosphere.

How did sharing scientific observations lead to the gas laws?

Chapter Preview

Concepts and Skills You Will Need

Before you begin this chapter, review the following concepts and skills:

- manipulating equations algebraically, and applying them to solve problems (Chapter 11, section 11.4)
- using correct significant digits in problem solving (Chapter 1, section 1.2)
- converting an amount in grams to an amount in moles, and then to a number of atoms or molecules (Chapter 5, section 5.3)
- solving stoichiometry problems, including writing balanced equations (Chapter 4, section 4.1 and Chapter 7, section 7.1)

12.1 The Ideal Gas Law

Section Preview/ Specific Expectations

In this section, you will

- **state** Avogadro's hypothesis, and **explain** how it contributes to our understanding of the reactions of gases
- **describe** the quantitative relationships that exist among pressure, volume, temperature, and amount of substance, based on the ideal gas law
- **solve** quantitative problems involving the ideal gas law
- **use** and **convert** appropriate units to express pressure and temperature
- **communicate** your understanding of the following terms: *law of combining volumes, law of multiple proportions, Avogadro's hypothesis, molar volume, ideal gas law*

Near the beginning of the nineteenth century, Joseph Gay-Lussac experimented with the volumes of gases. He found that adding two volumes of gas to one volume of gas produced only two volumes of gas. Puzzled, Gay-Lussac tried adding three volumes of gas to one volume. The result was still two! When he tried adding one volume of gas to a second volume of gas, again the result was two. What was going on?

In England, around the same time, John Dalton studied the masses of compounds as they reacted to produce products. After Dalton read about the similar work of other scientists, such as Lavoisier and the British scientist Joseph Priestley, he contacted Gay-Lussac. He described his results and hypotheses to Gay-Lussac. In 1808, both men published their theories. After examining the theories of Dalton and Gay-Lussac, an Italian scientist named Amedeo Avogadro formulated a hypothesis that combined their theories.

\Figure **The information that was shared by these three scientists led to the gas laws that we use today.**

The Molar Volume of Gases

Gay-Lussac measured the *volumes* of gases before and after a reaction. His research led him to devise the **law of combining volumes**: When gases react, the volumes of the reactants and the products, measured at equal temperatures and pressures, are always in whole number ratios. For example, 2 volumes of hydrogen gas react with 1 volume of oxygen gas to produce 2 volumes of water vapour.

John Dalton examined the *masses* of compounds before and after a reaction. Dalton's research led him to propose the **law of multiple proportions**: The masses of the elements that combine can be expressed in small whole number ratios.

By combining these ideas, Avogadro related the *volume* of a gas to the *amount* that is present (calculated from the mass). Avogadro divided Dalton's mass ratios by the molar masses of the elements to obtain the mole ratios. He realized that these mole ratios were the same as the volume ratios that Gay-Lussac had obtained. For example, 1 L of hydrogen gas reacts with 1 L of chlorine gas. Avogadro decided that there must be the same number of molecules in each litre of gas. Thus, **Avogadro's hypothesis** was formulated: Equal volumes of all ideal gases at the same temperature and pressure contain the same number of molecules.

Figures 12.2, 12.3, and 12.4 show the three reactions that produced Gay-Lussac's confusing observations. You can see that the mole ratios are the same as the volume ratios. Today our knowledge of atoms and molecules helps us understand Gay-Lussac's results. We know that gases are made of molecules that may contain more than one atom.

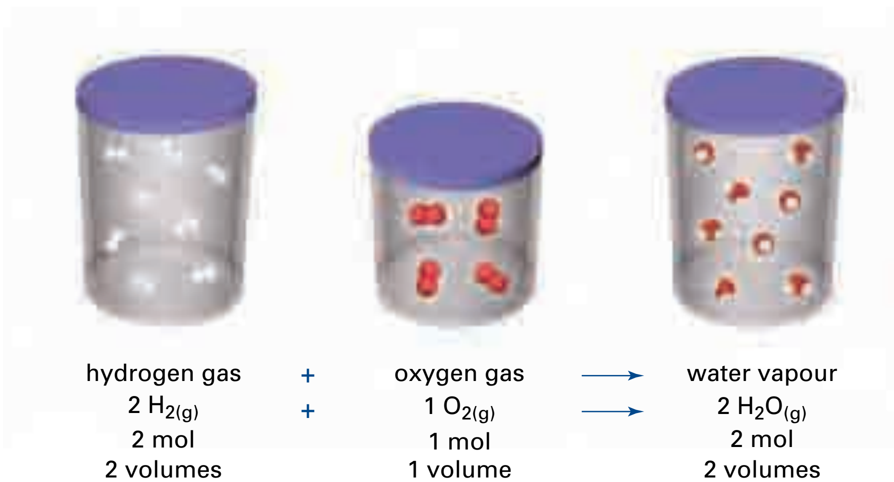

Figure 12.2 **Hydrogen and oxygen gases combine to form water vapour.**

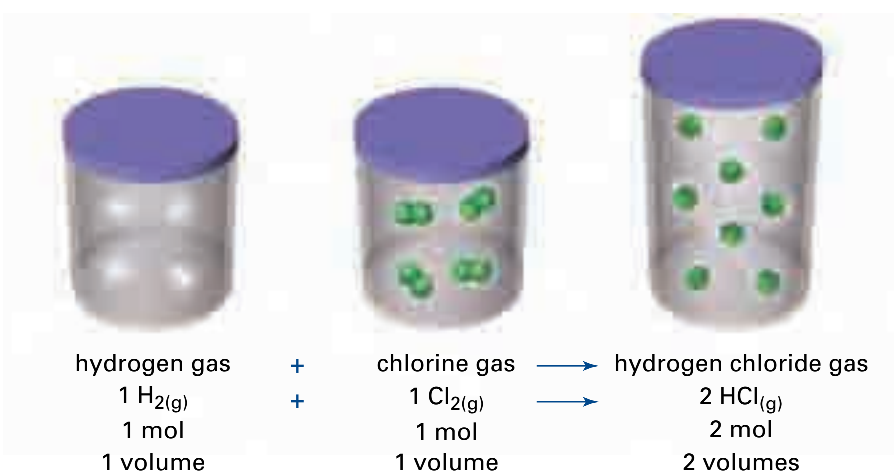

Figure 12.3 **Hydrogen and chlorine gases combine to form hydrogen chloride gas.**

Language LINK

The word *symposium* comes from the ancient Greeks. A symposium was a gathering of intellectuals to drink, feast, and talk. All sorts of new ideas arose from these meetings. The participants took away the new ideas to work on them further. Then they brought their findings back to the next symposium.

Many scientific laws are given the name of one person. It is important to remember, however, that most laws are the culmination of many scientists' work over a long time. Science would never move forward without ideas being shared.

Look up the word "symposium" in a dictionary. What meaning does it have today? Do a quick Internet search of the word "symposium." What modern symposiums can you find? What subjects do they cover?

mind STRETCH

As a class, hold a symposium about gas balloons. Research hot air balloons and helium balloons. Also research any other information about balloons, the history of ballons, and balloon travel that interests you. Prepare papers and posters to share your information.

Figure 12.4 **Hydrogen and nitrogen combine to form ammonia gas.**

Avogadro's hypothesis can be written as a mathematical law, shown here.

Avogadro's Law

Avogadro's law gives us a mathematical relationship between the volume of a gas (V) and the number of moles of gas present (n).

$$n \propto V \quad \text{or} \quad n = kV \quad \text{or} \quad \frac{n_1}{V_1} = \frac{n_2}{V_2}$$

where n = number of moles
V = volume
k = a constant

CHECKPOINT

The next Sample Problem involves the combined gas law. Review this law in section 11.4 before continuing.

Based on Avogadro's law, one mole of a gas occupies the same volume as one mole of another gas at the same temperature and pressure. The molar volume of a gas is the space that is occupied by one mole of the gas. **Molar volume** is measured in units of L/mol. You can find the molar volume of a gas by dividing its volume by the number of moles that are present ($\frac{V}{n}$). Look at the Sample Problem below to find out how to calculate molar volume. Then complete the following Thought Lab to find the molar volumes of carbon dioxide gas, oxygen gas, and methane gas at STP.

Sample Problem

The Molar Volume of Nitrogen

Problem

A resealable 1.30 L container has a mass of 4.73 g. Nitrogen gas, $N_{2(g)}$, is added to the container until the pressure is 98.0 kPa at 22.0°C. Together, the container and the gas have a mass of 6.18 g. Calculate the molar volume of nitrogen gas at STP.

What Is Required?

You need to find the volume of one mole of nitrogen (the molar volume) at STP.

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What Is Given?

Set out all the data in a table like the one shown below.

Situation 1: in the container	Situation 2: at STP
P_i = 98.0 kPa	P_f = 101.3 kPa
V_i = 1.30 L	V_f = ?
T_i = 22.0°C, or 295 K	T_f = 0°C, or 273 K
m_i = 6.18 g − 4.73 g = 1.45 g (Subtract the mass of the container.)	m_f = 1.45 g (The mass remains the same.)
n_i = ?	$n_f = n_i$ = ?

Plan Your Strategy

Algebraic method

Step 1 Calculate the number of moles of nitrogen gas (n_i) by dividing the mass of the nitrogen in the container by the molar mass of nitrogen gas (28.02 g/mol).

$$n = \frac{m}{M}$$

Step 2 Use the combined gas law from Chapter 11 to find the volume of nitrogen at STP (V_f).

Step 3 Use the volume (V_f) and the number of moles ($n_f = n_i$) to find the molar volume ($\frac{V}{n}$). The molar volume is the volume of one mole of gas.

$$\text{Molar volumes} = \frac{V}{n}$$

Ratio method

Step 1 Calculate the number of moles of nitrogen gas (n_i) by dividing the mass of the nitrogen in the container by the molar mass of nitrogen gas (28.02 g/mol).

$$n = \frac{m}{M}$$

Step 2 Since the pressure increases from 98.0 kPa to 101.3 kPa, the volume will decrease. Multiply the initial volume by a pressure ratio that is less than 1. Since the temperature decreases from 295 K to 273 K, the volume will decrease further. Multiply the initial volume by a temperature ratio that is less than 1.

Step 3 Since there is less than 1 mol of nitrogen gas present, the volume of 1 mol of nitrogen gas (the molar volume) will be greater than the volume you calculated in step 2. To find the molar volume, multiply by a mole ratio that is greater than 1.

PROBLEM TIP

In the Sample Problem, you will see two different methods of solving the problem: the algebraic method and the ratio method. Choose the method you prefer to solve this type of problem.

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Act on Your Strategy

Algebraic method

Step 1 Calculate the number of moles.

$$n = \frac{m}{M}$$
$$= \frac{1.45\ \cancel{\text{g}}}{28.02\ \cancel{\text{g}}/\text{mol}}$$
$$= 0.0517\ \text{mol}$$

Step 2 Find the volume of nitrogen at STP.

$$\frac{P_iV_i}{T_i} = \frac{P_fV_f}{T_f}$$
$$\therefore V_f = \frac{P_iV_iT_f}{T_iP_f}$$
$$= \frac{98.0\ \cancel{\text{kPa}} \times 1.30\ \text{L} \times 273\ \cancel{\text{K}}}{295\ \cancel{\text{K}} \times 101.3\ \cancel{\text{kPa}}}$$
$$= 1.16\ \text{L}$$

Step 3 Find the molar volume.

$$\text{Molar volume} = V/n$$
$$= \frac{1.16\ \text{L}}{0.0517\ \text{mol}}$$
$$= 22.4\ \text{L/mol}$$

Therefore, the molar volume of nitrogen at STP is 22.4 L/mol.

Ratio method

Step 1 Calculate the number of moles.

$$n = \frac{m}{M}$$
$$= \frac{1.45\ \cancel{\text{g}}}{28.02\ \cancel{\text{g}}/\text{mol}}$$
$$= 0.0517\ \text{mol}$$

Step 2 V_f is the final volume of the nitrogen gas.

$$V_f = 1.30\ \text{L} \times \frac{98.0\ \cancel{\text{kPa}}}{101.3\ \cancel{\text{kPa}}} \times \frac{273\ \cancel{\text{K}}}{295\ \cancel{\text{K}}}$$
$$= 1.16\ \text{L}$$

Step 3 V_m is the volume of 1 mol of the nitrogen gas.

$$V_m = 1.16\ \text{L} \times \frac{1.00\ \cancel{\text{mol}}}{0.0517\ \cancel{\text{mol}}}$$
$$= 22.4\ \text{L}$$

The molar volume is equal to the volume V_m divided by 1.00 mol. Thus the molar volume is 22.4 L/mol.

Check Your Solution

The answer is expressed in the correct units. It agrees with the accepted value. There are three significant digits in the answer. This is consistent with the least number of significant digits in the question.

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Practice Problems

1. At 19°C and 100 kPa, 0.021 mol of oxygen gas, $O_{2(g)}$, occupy a volume of 0.50 L. What is the molar volume of oxygen gas at this temperature and pressure?
2. What is the molar volume of hydrogen gas, $H_{2(g)}$, at 255°C and 102 kPa, if a 1.09 L volume of the gas has a mass of 0.0513 g?
3. A sample of helium gas, $He_{(g)}$, has a mass of 11.28 g. At STP, the sample has a volume of 63.2 L. What is the molar volume of this gas at 32.2°C and 98.1 kPa?
4. In the Sample Problem, you discovered that the molar volume of nitrogen gas is 22.4 L at STP.
   (a) How many moles of nitrogen are present in 10.0 L at STP?
   (b) What is the mass of this gas sample?

ThoughtLab Molar Volume of Gases

Two students decided to calculate the molar volumes of carbon dioxide, oxygen, and methane gas. First they measured the mass of an empty 150 mL syringe under vacuum conditions. This ensured that the syringe did not contain any air. Next they filled the syringe with 150 mL of carbon dioxide gas. They measured and recorded the mass of the syringe plus the gas. The students repeated their procedure for oxygen gas and for methane gas.

Finally, the students found the temperature of the room to be 23.0°C (296 K). They found the pressure to be 98.7 kPa. They took these values to be the temperature and pressure of the three gases. The students' results are given in the table.

Procedure

1. Copy the table into your notebook.
2. Calculate the molar volume of carbon dioxide gas at the given temperature and pressure, and at STP. Write your calculations and answers in the table.
3. Do the same calculations for oxygen and methane gas. Write your calculations and answers in the table.

Analysis

1. Compare the three molar volumes at STP. What do you observe?
2. The accepted molar volume of a gas at STP is 22.4 L/mol. Use this value to calculate the percent error in your experimental data for each gas.

Three Gases at 296 K and 98.7 kPa

Gas	carbon dioxide	oxygen	methane
Volume of gas (V)	150 mL	150 mL	150 mL
Mass of empty syringe	25.08 g	25.08 g	25.08 g
Mass of gas + syringe	25.34 g	25.27 g	25.18 g
Mass of gas (m)			
Molar mass (M)			
Number of moles of gas ($n = m/M$)			
Volume of gas at STP (273 K and 101.3 kPa)			
Molar volume at STP ($MV = V/n$)			

In the Thought Lab, you found that the molar volume of one gas is roughly the same as the molar volume of another gas at the same temperature and pressure. In fact, the molar volume of an *ideal* gas at STP is 22.4 L/mol. Figure 12.5 shows a balloon with a volume of 22.4 L compared to some common objects. This is a fairly large volume of gas. For example, a basketball has a volume of only 7.5 L.

Figure 12.5 **One mole of any gas at STP occupies 22.4 L (22.4 dm^3). How large is 22.4 L? The other objects are shown for comparison.**

CHECKPOINT

What are the temperature and pressure at STP? Go back to section 11.4 to remind yourself.

In Chapter 11, you learned that temperature, pressure, and volume are related. Based on Avogadro's law, the number of moles is related to the temperature, pressure, and volume of a gas. Therefore, Avogadro's law can be applied to solve gas problems involving moles and volume, when the temperature and pressure remain constant. Figure 12.6 explains the relationship among temperature, pressure, volume, and number of moles of a gas.

The following Sample Problems show you how to do gas calculations using Avogadro's law.

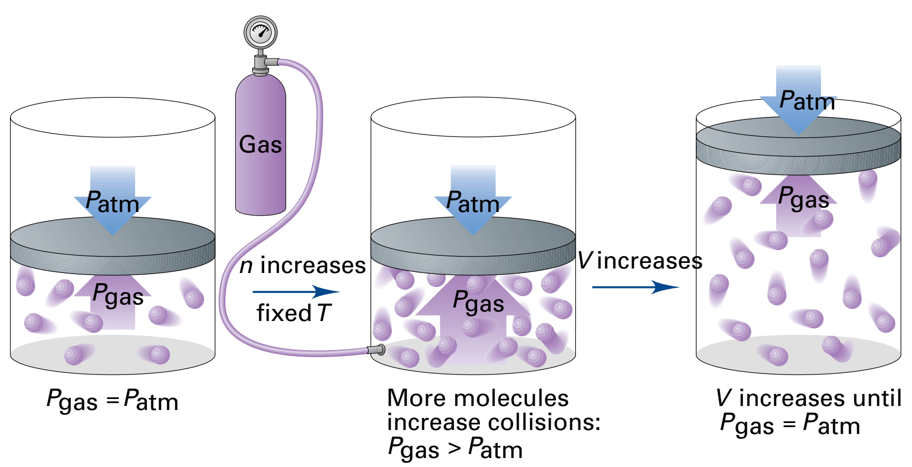

Figure 12.6 **At a temperature (*T*), a given amount of a gas (*n*) produces a pressure (*P*). When more gas is added, the number of moles (*n*) increases. This results in more molecular collisions with the container walls, thus increasing the pressure of the gas. The volume (*V*) of the gas increases until the pressure is again equal to the pressure of the surroundings.**

Sample Problem

Volumes of Gases

Problem

What is the volume of 3.0 mol of nitrous oxide, $NO_{2(g)}$, at STP?

What Is Required?

You need to find the volume of nitrous oxide at STP (V_f).

What Is Given?

You know that 1.00 mol of a gas occupies 22.4 L at STP.

$\therefore n_i = 1.0$ mol

$V_i = 22.4$ L

$n_f = 3.0$ mol of NO_2

Plan Your Strategy

Algebraic method

Use Avogadro's law: $\frac{n_i}{V_i} = \frac{n_f}{V_f}$

Cross multiply to solve for V_f, the unknown volume of NO_2.

Ratio method

There are 3 mol of nitrous oxide. Thus, the volume of nitrous oxide at STP must be larger than the volume of 1 mol of gas at STP. Multiply by a mole ratio that is greater than 1.

Act on Your Strategy

Algebraic method

$$
\begin{aligned}
V_f &= \frac{n_f V_i}{n_i} \\
&= \frac{3.0 \cancel{\text{mol}} \times 22.4 \text{ L}}{1.00 \cancel{\text{mol}}} \\
&= 67 \text{ L}
\end{aligned}
$$

Ratio method

$$
\begin{aligned}
V_f &= 22.4 \text{ L} \times \frac{3.0 \cancel{\text{mol}}}{1.0 \cancel{\text{mol}}} \\
&= 67 \text{ L}
\end{aligned}
$$

Therefore, there are 67 L of nitrous oxide.

Check Your Solution

The significant digits and the units are all correct.
The volume of nitrous oxide is three times the volume of 1 mol of gas at STP. This makes sense, since there are 3 mol of nitrous oxide.

PROBLEM TIP

In these Sample Problems, you will see two different methods of solving the problem: the algebraic method and the ratio method. Choose the method you prefer to solve this type of problem.

Sample Problem

Moles of Gas

Problem

Suppose that you have 44.8 L of methane gas at STP.

(a) How many moles are present?

(b) What is the mass (in g) of the gas?

(c) How many molecules of gas are present?

What Is Required?

(a) You need to calculate the number of moles.

(b) You need to calculate the mass of the gas.

(c) You need to calculate the number of molecules.

What Is Given?

The gas is at STP. Thus 1.00 mol of gas has a volume of 22.4 L. You know that one mole contains 6.02×10^{23} molecules. There are 44.8 L of gas.

Plan Your Strategy

Algebraic method

(a) Use Avogadro's law. Solve for the number of moles by cross multiplying.

(b) Multiply the number of moles (n) by the molar mass (M) to find the mass of the gas (m).

$$m = n \times M$$

(c) Multiply the number of moles (n) by the Avogadro constant (6.02×10^{23}) to find the number of molecules.

$$\text{\# of molecules} = \text{n} \times 6.02 \times 10^{23}\ \text{molecules/mol}$$

Ratio method

(a) The volume of the unknown gas is 44.8 L. Since the volume is greater than 22.4 L, there is more than 1 mol of gas. To find the unknown number of moles (n), multiply by a volume ratio that is greater than 1.

(b) Multiply the number of moles (n) by the molar mass (M) to find the mass of the gas (m).

$$m = n \times M$$

(c) Multiply the number of moles (n) by the Avogadro constant (6.02×10^{23}) to find the number of molecules.

$$\text{\# of molecules} = \text{n} \times 6.02 \times 10^{23}\ \text{molecules/mol}$$

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Act on Your Strategy

Algebraic method

(a) $\frac{n_i}{V_i} = \frac{n_f}{V_f}$

$n_f = \frac{n_i V_f}{V_i}$

$= \frac{1.00 \text{ mol} \times 44.8 \cancel{\text{L}}}{22.4 \cancel{\text{L}}}$

$= 2.00 \text{ mol}$

(b) Find the molar mass of methane, CH_4.

$1\text{C} = 1 \times 12.01 \text{ g/mol}$

$\underline{4\text{H} = 4 \times 1.01 \text{ g/mol}}$

$M_{CH_4} = 16.05 \text{ g/mol}$

$m = n \times M$

$= 2.00 \cancel{\text{mol}} \times 16.05 \text{ g/}\cancel{\text{mol}}$

$= 32.1 \text{ g}$

(c) $\# \text{ molecules} = 2.00 \cancel{\text{mol}} \times 6.02 \times 10^{23} \frac{\text{molecules}}{\cancel{\text{mol}}}$

$= 1.20 \times 10^{24} \text{ molecules}$

Ratio method

(a) $n = 1.00 \text{ mol} \times \frac{44.8 \cancel{\text{L}}}{22.4 \cancel{\text{L}}}$

$= 2.00 \text{ mol}$

(b) Find the molar mass of methane, CH_4.

$1\text{C} = 1 \times 12.01 \text{ g/mol}$

$\underline{4\text{H} = 4 \times 4.04 \text{ g/mol}}$

$M_{CH_4} = 16.05 \text{ g/mol}$

$m = n \times M$

$= 2.00 \text{ mol} \times 16.05 \text{ g/mol}$

$= 32.1 \text{ g}$

(c) $\# \text{ molecules} = 2.00 \cancel{\text{mol}} \times 6.02 \times 10^{23} \frac{\text{molecules}}{\cancel{\text{mol}}}$

$= 1.20 \times 10^{24} \text{ molecules}$

Therefore, 2.00 mol of methane are present. The mass of the gas is 32.1 g. 1.20×10^{24} molecules are present.

Check Your Solution

The significant digits are correct.
The volume of methane is double the volume of 1 mol of gas. It makes sense that 2 mol of methane are present. It also makes sense that the number of molecules present is double the Avogadro constant.

PROBLEM TIP

To solve many of these problems, try setting up a proportion and solving by cross multiplication.

Practice Problems

5. A balloon contains 2.0 L of helium gas at STP. How many moles of helium are present?

6. How many moles of gas are present in 11.2 L at STP? How many molecules?

7. What is the volume, at STP, of 3.45 mol of argon gas?

8. A certain set of conditions allows 4.0 mol of gas to be held in a 70 L container. What volume do 6.0 mol of gas need under the same conditions of temperature and pressure?

9. At STP, a container holds 14.01 g of nitrogen gas, 16.00 g of oxygen gas, 66.00 g of carbon dioxide gas, and 17.04 g of ammonia gas. What is the volume of the container?

10. **(a)** What volume do 2.50 mol of oxygen occupy at STP?

(b) How many molecules are present in this volume of oxygen?

(c) How many oxygen atoms are present in this volume of oxygen?

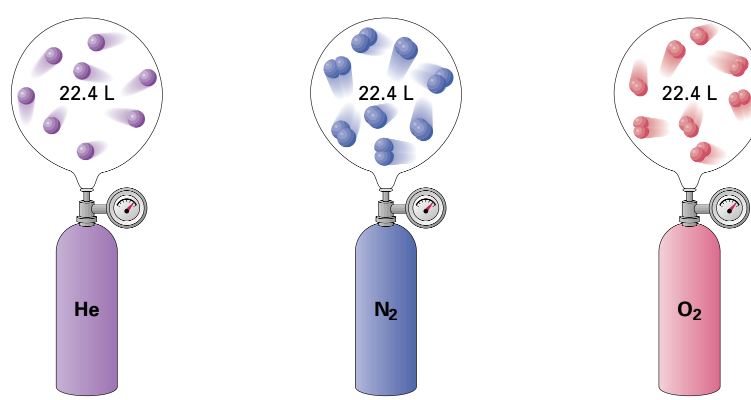
11. What volume do 2.00×10^{24} atoms of neon occupy at STP?

Table 12.1 Molar Volume of Several Real Gases at STP

Gas	Molar volume (L/mol)
helium, He	22.398
neon, Ne	22.401
argon, Ar	22.410
hydrogen, H_2	22.430
nitrogen, N_2	22.413
oxygen, O_2	22.414
carbon dioxide, CO_2	22.414
ammonia, NH_3	22.350

Volumes of Real Gases

You know that ideal gases have a volume of 22.4 L at STP. Do *real* gases have the same volume? The volumes of several real gases at STP are given in Table 12.1. All the volumes are very close to 22.4 L/mol, the molar volume of an ideal gas. Scientists have decided that 22.4 L/mol is an acceptable approximation for *any* gas at STP when using gas laws. Although the volumes are the same, one mole of a gas will have a different mass and density than one mole of another gas. (See Figure 12.7.)



n = 1 mol	n = 1 mol	n = 1 mol
P = 1 atm (760 torr)	P = 1 atm (760 torr)	P = 1 atm (760 torr)
T = 0°C (273 K)	T = 0°C (273 K)	T = 0°C (273 K)
V = 22.4 L	V = 22.4 L	V = 22.4 L
Number of gas particles = 6.022×10^{23}	Number of gas particles = 6.022×10^{23}	Number of gas particles = 6.022×10^{23}
Mass = 4.003 g	Mass = 28.02 g	Mass = 32.00 g
d = 0.179 g/L	d = 1.25 g/L	d = 1.43 g/L

Figure 12.7 **At STP, gases such as helium, nitrogen, and oxygen behave as ideal gases. They have a molar volume of 22.4 L.**

A Deeper Look: Real Gas Deviations

Do real gases *always* behave like ideal gases? At STP, you have seen that most real gases behave like ideal gases. At high pressures and/or low temperatures, however, gases no longer behave ideally. To understand why, recall the characteristics of an ideal gas, according to the kinetic molecular theory. (You learned this in Chapter 11.)

- Gas molecules have insignificant volume of their own.
- Molecules have no attractive forces between each other or with their container.
- Molecules move in perfectly straight lines.
- Collisions are completely elastic and, thus, do not use up energy.

These characteristics led to the ideal gas law. They are accurate enough for most applications. They are not perfect, however. In fact, most gases fall short of these characteristics in many ways:

- The particles of a real gas have a significant volume of their own.
- Molecules do attract each other.
- Molecules do not necessarily move in straight lines.
- Collisions are not completely elastic.

At high pressures and/or low temperatures, gases have smaller volumes. This means that the gas molecules are closer together. Since they are closer together, the molecules interact more than when they are far apart. It is no longer true to say that the molecules have no attractive forces between each other or with their container. Part A of Figure 12.8 illustrates how this affects the pressure of a gas. Also, since the total volume is smaller, the amount of space taken up by the gas molecules is more important. You can no longer ignore the volume of the gas molecules. Part B of Figure 12.8 illustrates that gas molecules do occupy part of the volume of the container.

CHECKPOINT

Look back at Table 12.1. It gives more accurate molar volumes for several gases. Using what you have learned about the way real gases behave, explain why these molar volumes are slightly different from the molar volume of an ideal gas.

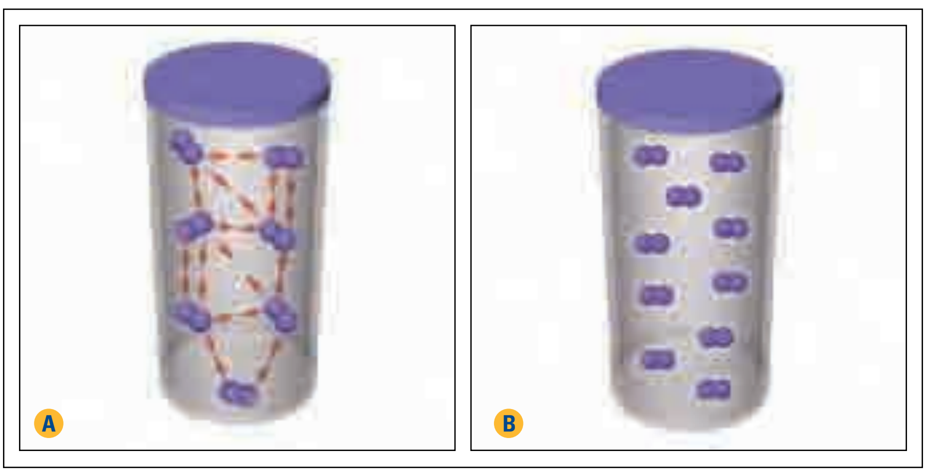


Figure 12.8 **(A) Because particles are attracted to each other, the pressure is reduced. (B) Since particles take up space, the total volume of empty space is smaller than the volume of the container.**

Scientists who want more accuracy in their experiments have adapted the ideal gas law to reflect the behaviour of real gases. In later chemistry courses, you will learn more about this corrected version of the ideal gas law, called the *Van der Waals equation*.

To summarize, at low pressures and high temperatures, most gases behave as ideal gases. Under any conditions that allow attractive forces between molecules to occur, gases no longer behave as an ideal gas. They behave as real gases.

History LINK

The French chemist, Antoine Lavoisier (1743–1794), was the first person to notice the gas-volume relationship. He observed water decomposing to give two volumes of hydrogen and one volume of oxygen. He mentioned this relationship in his 1789 textbook *Eléments de la Chemie*.

Arriving at the Ideal Gas Law

In the Sample Problem earlier, you used three steps to find the molar volume of nitrogen gas. After calculating the number of moles, you calculated the final volume using the combined gas law. Then you used Avogadro's law to find the volume of one mole of nitrogen. There is an easier way to do problems like this—by combining the two gas laws. After Avogadro's work, it did not take scientists long to connect $V \alpha n$ (Avogadro's law) with $V \alpha \frac{T}{P}$ (the combined gas law).

$$V \alpha n$$

$$V \alpha \frac{T}{P}$$

$$\therefore V \alpha \frac{nT}{P}$$

$$\therefore \frac{PV}{nT} = R, \text{ where } R \text{ is a constant}$$

As you know, all gases behave in a similar way. The *universal gas constant* (R), which applies to all gases, was derived for the final equation given above. Examine the calculation below to see how R was derived.

For one mole of gas at STP,

$$P = 101.3 \text{ kPa}$$
$$T = 273 \text{ K}$$
$$V = 22.4 \text{ L}$$
$$n = 1.00 \text{ mol}$$

$$R = \frac{PV}{nT}$$
$$= \frac{101.3 \text{ kPa} \times 22.4 \text{ L}}{1.00 \text{ mol} \times 273.15 \text{ K}}$$
$$= 8.31 \frac{\text{kPa·L}}{\text{mol·K}}$$

Electronic Learning Partner

Go to the Chemistry 11 Electronic Learning Partner for an interactive simulation on the ideal gas law.

When it is measured more accurately, the universal gas constant (R) has a value of 8.314 $\frac{\text{kPa·L}}{\text{mol·K}}$. This was the final piece of the equation. The resulting equation is an efficient tool for solving many problems that involve gases.

The **ideal gas law** states that the pressure multiplied by the volume is equal to the number of moles multiplied by the universal gas constant and the temperature.

$$PV = nRT$$

Guidelines for Using the Ideal Gas Law

- Always convert the temperature to kelvins (K).
- Always convert the masses to moles (mol).
- Always convert the volumes to litres (L).
- Using the ideal gas law will be easier if you always convert the pressures to kilopascals (kPa). Then you can memorize the value of R (8.314 $\frac{\text{kPa·L}}{\text{mol·K}}$) and use it for every calculation. If you happen to forget the value of R, you can calculate it by finding R for 1 mol of gas at STP.

Converting the Units of the Universal Gas Constant

CHECKPOINT

Go back to section 11.2 to review the units for pressure.

Some people prefer to use a converted value of R, rather than converting all the pressures to kilopascals, as suggested in the guidelines above. What are the units for R if the pressure is given in atmospheres?

$$\frac{8.314\ \cancel{\text{kPa}}\cdot\text{L}}{\text{mol}\cdot\text{K}} \times \frac{1\text{atm}}{101.3\ \cancel{\text{kPa}}} = 0.08206\ \frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}}$$

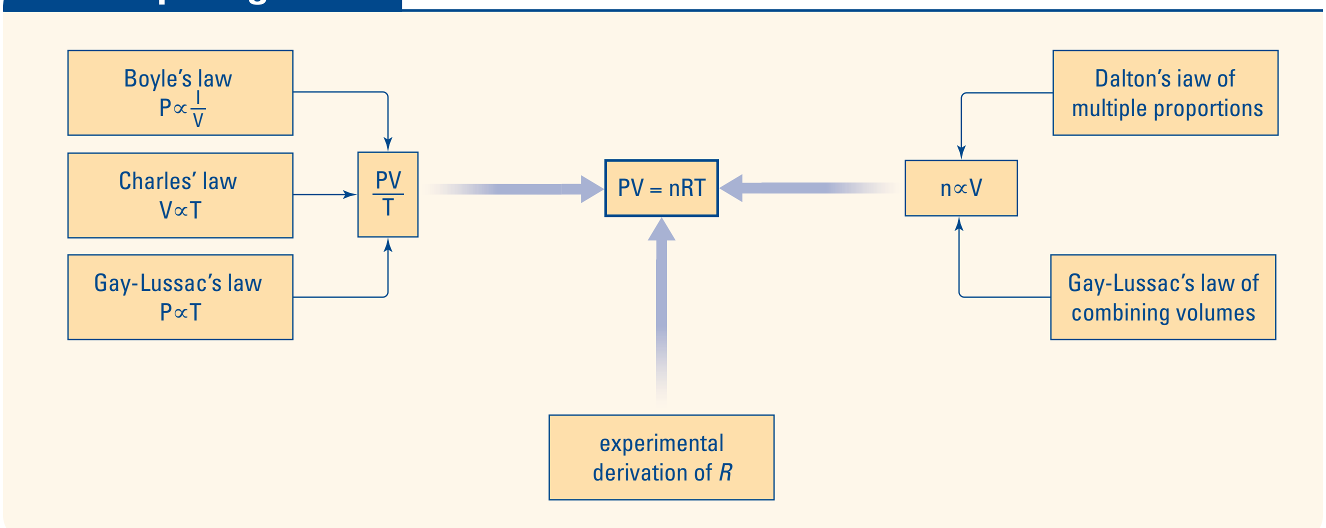
What are the units for R if the pressure is given in torr (mm Hg)?

$$8.314\ \frac{\cancel{\text{kPa}}\cdot\text{L}}{\text{mol}\cdot\text{K}} \times \frac{760\ \text{mmHg}}{101.3\ \cancel{\text{kPa}}} = 62.37\ \frac{\text{mmHg}\cdot\text{L}}{\text{mol}\cdot\text{K}}$$

Since R is an exact constant, always express it with *four* significant digits. Do not worry about changing the significant digits of R in your calculations. Always keep four digits, and then fix the number of significant digits in your answer.

The diagram below illustrates the evolution of the ideal gas law. It is a brief summary of what you have learned so far in this section. The following Sample Problems show you how to do calculations using the ideal gas law.

Concept Organizer — The Evolution of the Ideal Gas Law



Sample Problem

Calculating Molar Volume Using $PV = nRT$

Problem

Use the ideal gas law to calculate the molar volume of a gas at standard ambient temperature and pressure (SATP). The conditions for SATP are 298 K and 100 kPa.

What Is Required?

You need to find the molar volume, $MV = \frac{V}{n}$, at SATP. The units for your answer will be in mol/L.

Continued ...

Continued ...

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What Is Given?

$P = 100 \text{ kPa}$

$n = 1.00 \text{ mol}$

$R = 8.314 \text{ kPa·L/mol·K}$

$T = 298 \text{ K}$

PROBLEM TIP

The molar volume is in L/1.00 mol. If you let $n = 1.00$ mol, you only need to find V.

Plan Your Strategy

Let $PV = nRT$, and solve for V. Then substitute the known values into the equation.

Act on Your Strategy

$$PV = nRT$$

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

$$= \frac{1.00 \cancel{\text{mol}} \times 8.314 \cancel{\text{kPa}}\text{·L/}\cancel{\text{mol·K}} \times 298 \cancel{\text{K}}}{100 \cancel{\text{kPa}}}$$

$$= 24.8 \text{ L}$$

The molar volume of a gas at SATP is 24.8 L/mol.

Check Your Solution

The molar volume is in units of L/mol. 24.8 L/mol is slightly larger than the 22.4 L/mol molar volume at STP. This is the expected result, since the pressure has decreased and the temperature has increased.

Sample Problem

A Laughing Matter

Problem

Dentists sometimes use laughing gas (dinitrogen oxide, N_2O) to keep patients relaxed during dental procedures. A cylinder of laughing gas has a diameter of 23.0 cm and a height of 140 cm. The pressure is 108 kPa, and the temperature is 294 K. How many grams of laughing gas are in the cylinder?

What Is Required?

You need to find the mass (in g) of laughing gas in a cylinder at 108 kPa and 294 K.

What Is Given?

$P = 108 \text{ kPa}$

$V = \pi r^2 h$, where $r = 11.5$ cm and $h = 140$ cm

Continued ...

Continued ...
FROM PAGE 486

$R = 8.314$ kPa·L/mol·K

$T = 294$ K

Plan Your Strategy

Step 1 Use the formula for the volume of a cylinder ($\pi r^2 h$) to approximate the volume of the gas inside.

Step 2 Solve the ideal gas equation for n. Put the known values into the equation to find the number of moles present (n).

Step 3 Calculate the molar mass (M) of dinitrogen oxide. Use it to find the mass (m) of the gas present.

Act on Your Strategy

Step 1 Find the volume of the cylinder.

$$\begin{aligned} V &= \pi r^2 h \\ &= \pi \times 11.5^2 \text{ cm}^2 \times 140 \text{ cm} \\ &= 5.73 \times 10^4 \text{ cm}^3 \\ &= 57.3 \text{ L} \end{aligned}$$

Step 2 Use the gas law to find the number of moles.

$$\begin{aligned} PV &= nRT \\ \therefore n &= \frac{PV}{RT} \\ &= \frac{108 \cancel{\text{kPa}} \times 57.3 \cancel{\text{L}}}{8.314 \cancel{\text{kPa}}\cdot\cancel{\text{L}}/\text{mol}\cdot\cancel{\text{K}} \times 294 \cancel{\text{K}}} \\ &= 2.53 \text{ mol} \end{aligned}$$

Step 3 Calculate the mass of the gas.

For N_2O, the molar mass is

$$\begin{aligned} 2\text{N} &= 2 \times 14.01 \text{ g/mol} \\ 1\text{O} &= 1 \times 16.00 \text{ g/mol} \\ \hline M_{N_2O} &= 44.02 \text{ g/mol} \end{aligned}$$

For 2.53 mol, the mass (m) is

$$\begin{aligned} m &= n \times M \\ &= 2.53 \cancel{\text{mol}} \times 44.02 \text{ g/}\cancel{\text{mol}} \\ &= 111 \text{ g} \end{aligned}$$

Therefore, 111 g of laughing gas is in the cylinder.

Check Your Solution

The answer is in grams, which is the unit you were asked for. There are three significant digits. This is consistent with the least number of significant digits in the question.

Practice Problems

12. 4.00 L of ammonia gas in a container holds 2.17 mol at 206 kPa. What is the temperature inside the container?

Continued ...

Continued ...
FROM PAGE 487

13. How many kilograms of chlorine gas are contained in 87.6 m^3 at 290 K and 2.40 atm? **Hint:** 1 m^3 = 1000 L
14. Calculate the volume of 3.03 g of hydrogen gas at a pressure of 560 torr and a temperature of 139 K.
15. A 6.0×10^2 L reaction tank contains 5.0 mol of oxygen gas and 28 mol of nitrogen gas. If the temperature is 83°C, what is the pressure of the oxygen, in kPa?

Section Wrap-up

In this section, you learned that the ideal gas law was derived from Avogadro's law and the combined gas law. You can use the ideal gas law to calculate the pressure, temperature, volume, or number of moles of a gas. In the next section, you will learn how to apply the ideal gas law to identify unknown gases on the basis of their densities and/or molar masses.

Remember that you can manipulate the ideal gas law to solve for any variable.

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

The universal gas constant (R) always has the same value: 8.314 kPa·L/mol·K. (You need to use an alternative value of R if pressure is given in atmospheres or in torr.)

Section Review

1. I How many grams of sulfur dioxide are in 36.2 L at STP?
2. I At certain conditions, 20 mol of a gas occupy 498 L. What volume do 35 mol occupy?
3. I What is the pressure of 0.76 mol of gas at 48°C in 8.0 L?
4. I How many grams of bromine gas, Br_2, are in 3.12 L at a pressure of 2.4 atm and a temperature of –20°C? How many molecules of bromine gas are there?
5. I What is the temperature of 6.02 mol of hydrogen sulfide in a 132 L tank at 0.95 atm?
6. K/U What was Avogadro's hypothesis? Explain how Avogadro contributed to our understanding of gases and the relationships among the properties of gases.
7. K/U 1.00 m^3 of regular air at 1.0 atm is compressed to one eighth of the volume at a constant temperature. What is the pressure contribution of the nitrogen? **Hint:** See the Table 11.2 in Chapter 11, section 11.4.

PROBEWARE

If you have access to probeware, do the Gas Laws lab, or a similar lab available from a probeware company.

Applications of the Ideal Gas Law

12.2

Section Preview/ Specific Expectations

In this section, you will

- **solve** more quantitative problems using the ideal gas law
- **determine** the molar mass of an unknown gas through experimentation

Studying gases and their properties becomes more interesting when you realize that your own body is a container for gases. Your lungs hold air, an important solution of gases that you need to live. As well, there is the embarrassing type of gas, called methane, that results from the digestive process. How many moles of oxygen do your lungs hold? Using the laws and properties you have learned, you can now calculate this fact about your own body. (See Figure 12.9.)

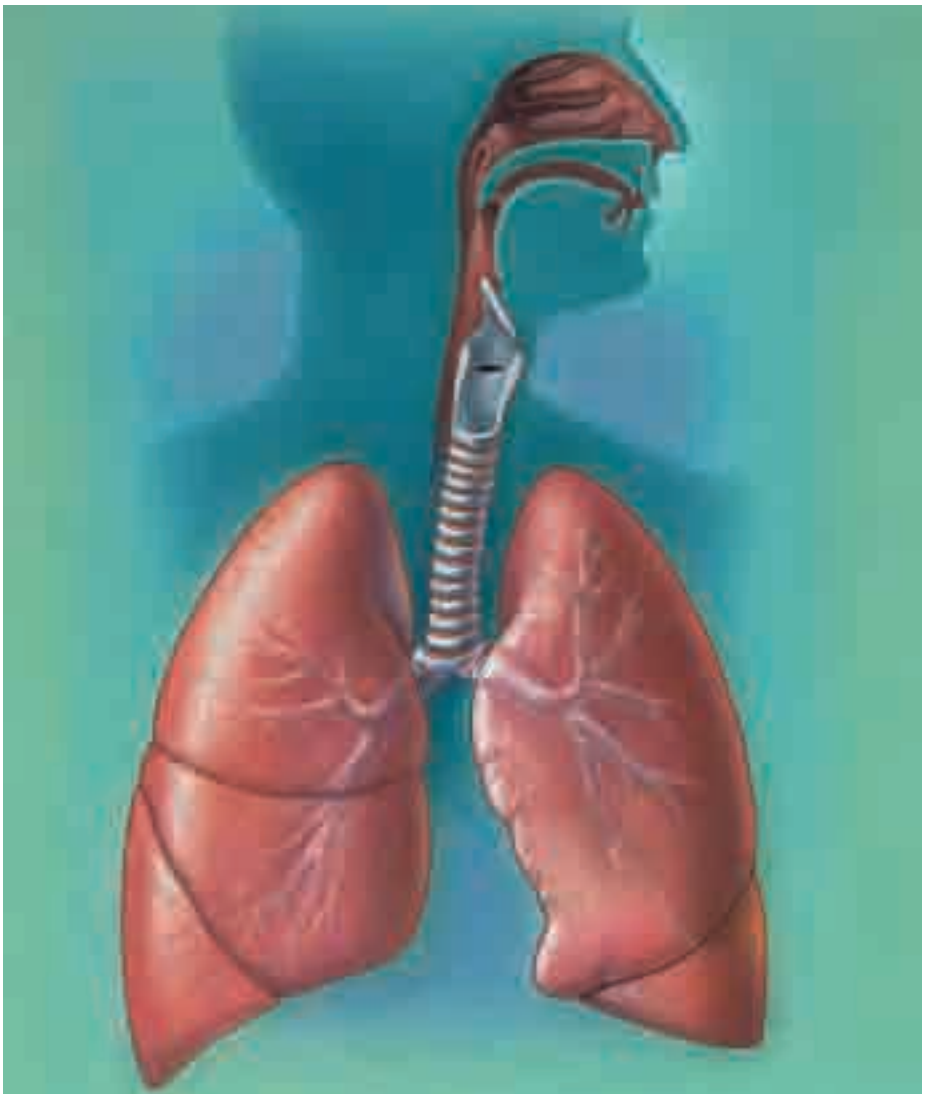

Figure 12.9 **Your lungs hold about 4 L of air. When you breathe out, only 500 mL of air is expelled. The same amount of air is taken in when you inhale one breath. If air contains 20% oxygen gas, how many moles of oxygen gas do your lungs contain at 37°C and 100 kPa?**

Biology LINK

As you breathe in, your diaphragm moves down and your rib cage moves out to increase the volume of your lungs. As the volume increases, the air pressure in your lungs decreases. Air from the outside rushes in to fill the expanded volume. This is an application of Boyle's law. In the same way, your lungs decrease in volume when you exhale. The resulting increase in pressure pushes air out of your lungs.

In this section, you will learn more about two properties of a gas that are closely related to molar volume: density and molar mass. You have already encountered these properties, but now you will use them to help you with your gas calculations.

CHECKPOINT

In this book, we use a capital *M* to symbolize molar mass. When working with solutions, you may see M used to express the molar concentration of a solution (for example, 6 M $HCl_{(aq)}$. Make sure that you do not become confused! Always check to make sure that you understand what *M* is referring to.

Density and Molar Mass

As you learned in the last section, the *molar volume* of a gas is defined as the space that is occupied by one mole of the gas. It is always given in units of L/mol.

The *density* of a gas is similar to the density of a solid or a liquid. Density is found by dividing mass by volume. The density of a gas is usually reported in units of g/L.

The *molar mass* of a gas refers to the mass (in g) of one mole of the gas. You can calculate molar mass by adding the masses of atoms in the periodic table. You can also calculate molar mass by dividing the mass of a sample by the number of moles that are present. Molar mass is always expressed in the units g/mol. Table 12.2 summarizes molar volume, density, and molar mass.

Table 12.2 Molar Volume, Density, and Molar Mass

	Molar volume	Density	Molar mass
Unit	L/mol	g/L	g/mol
Meaning	volume/amount	mass/volume	mass/amount
Calculations	Molar volume = $\frac{\text{Volume}}{\text{Number of moles}}$ $MV = V/n$	Density = $\frac{\text{Mass}}{\text{Volume}}$ $D = m/V$	Molar mass = Sum of molar masses of the atoms in the compound *or* Molar mass = $\frac{\text{Mass}}{\text{Number of moles}}$ $M = m/n$

mind STRETCH

Molar volume (L/mol), density (g/L), and molar mass (g/mol) are closely related. You can calculate each property using the other two. Analyze the units to discover the exact relationship.

Dense Gases Can Be Deadly

You have learned that the volume of a gas at a certain temperature and pressure is the same as the volume of any other gas at the same temperature and pressure. For example, all gases have the approximate volume of 22.4 L at STP. The *molecular masses* of different gases, however, are all different. This means that each gas has a different density, or mass per unit of volume. (Look back at Figure 12.7 to see three examples.)

Understanding the densities of gases can be useful in the everyday world. For example, miners who drill deep into the ground must know which gases are present, and which have the highest densities. (See Figure 12.10.) They must take appropriate safety precautions to avoid explosions, poisoning, or suffocation.

On December 3, 1984, an industrial accident in Bhopal, India, released a large quantity of methylisocyanate, CH_3NCO, a dense gas, into the air. This highly irritating and toxic gas caused the death of more than 3000 people living nearby. The Chemistry Bulletin on the next page gives an example of a natural disaster involving a dense gas. Figure 12.11, on page 492, illustrates a popular use for gases that are less dense than air.

Figure 12.10 Dense gases sit at the bottoms of pits, such as mines and wells. Miners must understand the behaviour of these gases to avoid accidents. For example, dense carbon dioxide gas can cause suffocation.

Chemistry Bulletin

Science | Technology | Society | Environment

The Killing Lakes of Cameroon

On August 15, 1984, a cloud of deadly gas burst from Lake Monoun in Cameroon, a country in western Africa. Thirty-seven people died from suffocation. Two years later, on August 21, 1986, Lake Nyos, a larger and deeper lake, ejected a full cubic kilometre of the same gas. The gas travelled silently into neighbouring villages, killing 1700 people and thousands of livestock. What was this toxic gas?

Lake Nyos and Lake Monoun both sit in volcanic craters. The lakes are hazardous because of their volcanic origin, even though both volcanoes are dormant. Volcanoes are vents through Earth's crust. They carry *magma*, a mixture of molten rock and dissolved gases, to the surface of Earth. When magma rises to Earth's surface in volcanoes, the pressure is decreased. The gases come out of solution and expand.

Magma contains a large quantity of dissolved carbon dioxide and minerals. These are released from the magma into the ground water under a volcano's crater. Lake Nyos and Lake Monoun are both fed at the bottom by volcanic springs of mineral-rich carbonated ground water. In some crater lakes, there is enough water circulation for the carbon dioxide to bubble up from the bottom and be released at the surface. The surface waters of Lake Nyos and Lake Monoun, however, do not mix with deeper waters. The carbon dioxide remains at the bottom. Volcanic springs continue to supply the lakes with carbon dioxide, which remains trapped at the bottom. Fresher surface water sits on this dense lower layer.

Scientists believe that the tragedies at Lake Nyos and Lake Monoun were caused by a disruption of the water layers, perhaps triggered by a landslide, an earthquake, or even a strong wind. Lower carbonated water was suddenly released into the upper water. As it moved up to the surface, the pressure on the gas decreased while the temperature increased. As a result, the carbon dioxide gas rapidly bubbled out of solution. At Lake Nyos, the sudden release of the lower lake water into the upper layers caused a plume of water and gas to rise high into the air.

How did this massive release of carbon dioxide cause so many deaths? Carbon dioxide gas is one-and-a-half times as dense as air. It sinks to the ground and displaces the oxygenated air we need to breathe. The invisible, odourless carbon dioxide that was released from Lake Nyos and Lake Monoun settled on the ground. It travelled rapidly down the slopes into populated regions. People quickly became unconscious and died of suffocation.

Carbon dioxide continues to accumulate at the bottoms of both lakes. An international team of scientists is developing a plan to release this carbon dioxide by controlled degassing. They plan to insert long pipes deep into both lakes and suck up some of the dense bottom water. This will create a pressure difference and cause a fountain of gas-rich water to jet from the pipes.

Making Connections

1. Calculate and compare the density of one mole of $CO_{2(g)}$ with the density of one mole of $N_{2(g)}$ at STP. (Nitrogen gas is the main constituent of air.) Explain the significance of your calculations.
2. How is carbonated water produced artificially, and what is it used for? Do research to find out.

Web LINK

www.school.mcgrawhill.ca/resources/
Many accidents have occurred because of dense gases in mines. Which gases are dangerous? What safety precautions should be taken? To learn more, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next.

History LINK

In World War I, a highly irritating gas called phosgene, $COCl_2$, was used against the Allied troops. Phosgene is about 3.4 times more dense than air. In concentrations above 50 ppm, it causes the lungs to fill up with fluid. This results in diminished lung capacity, and subsequent collapse of the heart. Phosgene can bring about death within hours.

Figure 12.11 **Helium balloons and hot air balloons also depend on the density of gases. These objects are filled with gases that are less dense than air. Therefore, they are able to float.**

In the following Sample Problem, you will use the ideal gas law to find the density of nitrogen gas.

Sample Problem

Finding the Density of Nitrogen Gas

Problem

Nitrogen gas makes up almost 80% of our atmosphere. What is the density of pure nitrogen gas, in g/L, at 12.50°C and 126.63 kPa?

What Is Required?

Find the density of nitrogen gas, in g/L, at 12.50°C and 126.63 kPa pressure.

What Is Given?

P = 126.63 kPa

T = 12.50°C

Plan Your Strategy

Step 1 Change the temperature to kelvins. The temperature in this question is given to two decimal places. Therefore, use a conversion factor with two decimal places: + 273.15.

Step 2 Calculate the molar mass (M) of nitrogen gas, N_2, using the molar mass in the periodic table.

Step 3 Since the volume is not given, set it as 1.00 L. Since the pressure is given in kilopascals, use R = 8.314 kPa·L/mol·K. Substitute the numbers and units for P, T, R, and V into the ideal gas law equation. Solve for n.

Continued ...

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FROM PAGE 492

Step 4 Convert n to the mass of nitrogen (m) by multiplying the number of moles by the molar mass (M) of N_2.

Step 5 Find the density by dividing the mass by the volume (1.00 L).

Act on Your Strategy

Step 1 $T = (12.50°\text{C} + 273.15)$
$= 285.65\text{ K}$

Step 2 $M_{N_2} = 2 \times 14.01\text{ g/mol}$
$= 28.02\text{ g/mol}$

Step 3 $PV = nRT$

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

$$= \frac{(126.63\ \cancel{\text{kPa}})(1.00\ \cancel{\text{L}})}{(8.314\ \cancel{\text{kPa}}\cdot\cancel{\text{L}}/\text{mol}\cdot\cancel{\text{K}})(285.65\ \cancel{\text{K}})}$$

$$= 5.3320 \times 10^{-2}\text{ mol}$$

Step 4 The mass of N_2 is

$$m = n \times M$$
$$= (5.3348 \times 10^{-2}\text{ mol})(28.02\text{ g/mol})$$
$$= 1.4940\text{ g}$$

Step 5 The mass of 1.00 L of nitrogen gas is 1.494 g. By dividing this by the volume, 1.00 L, we obtain a density of 1.494 g/L.

Check Your Solution

When the units cancel out in the ideal gas equation, mol remains.
When units cancel in the density equation, g/L remain.
The least number of digits in the question is four. Therefore, the answer should have four significant digits, which it does.

Practice Problems

16. Oxygen makes up about 20% of our atmosphere. Find the density of pure oxygen gas, in g/L, for the conditions in the Sample Problem: 12.50°C and 126.63 kPa.

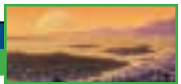
17. Find the density of butane gas, C_4H_{10}, (in g/L) at SATP conditions: 298 K and 100 kPa.

18. The atmosphere of the imaginary planet Xylo is made up entirely of poisonous chlorine gas, Cl_2. The atmospheric pressure of this inhospitable planet is 155.0 kPa, and the temperature is 89°C. What is the density of the atmosphere?

19. The atmosphere of planet Yaza, from the same star system as Xylo, is made of fluorine gas, F_2. The density of the atmosphere on Yaza is twice the density of the atmosphere on Xylo. The temperature of both planets is the same. What is the atmospheric pressure of Yaza?

Molar Mass of a Gas

You can find the molar mass of a gaseous element or compound in the same way that you find the molar mass of any other element or compound: by adding up the masses of the atoms. You can also find the molar mass by dividing the mass by the number of moles. (See Figure 12.12.)



COURSE CHALLENGE

How can you use what you learn in this Sample Problem to help you identify the gas given off by water plants? You will apply your learning later on, in the Chemistry Course Challenge.

Figure 12.12 How could you determine the molar mass of vaporized iodine gas on paper? How could you determine it in a laboratory?

In the laboratory, calculating the molar mass of an unknown gas can help you identify it. The next Sample Problem will demonstrate this.

Sample Problem

Using Molar Mass to Identify an Unknown Gas

Problem

A scientist isolates 2.366 g of a gas. The sample occupies a volume of 800 mL at 78.0°C and 103 kPa. Use these data to calculate the molar mass of the gas. Is the gas most likely to be bromine, krypton, neon, or fluorine?

What Is Given?

P = 103 kPa

V = 800 mL

m = 2.366 g

R = 8.314 kPa·L/mol·K

T = 78.0°C

What Is Required?

You need to find the number of moles. Then you need to use the mass and the number of moles to find the molar mass of the gas.

Continued ...

Continued ...
FROM PAGE 494

Plan Your Strategy

Convert the temperature to kelvins. Solve $PV = nRT$ for n. Then substitute in the known values to find the number of moles. Finally, set up a proportion to find the number of grams that would be in one mole, using the equation $M = \frac{m}{n}$, where M is the molar mass, m is the mass, and n is the number of moles.

Act on Your Strategy

$$\begin{aligned} T(\text{K}) &= T(^\circ\text{C}) + 273 \\ &= 78.0^\circ\text{C} + 273 \\ &= 351 \text{ K} \end{aligned}$$

Using the ideal gas law,

$PV = nRT$

$$\begin{aligned} \therefore\ n &= \frac{PV}{RT} \\ &= \frac{103\ \cancel{\text{kPa}} \times 0.800\ \cancel{\text{L}}}{8.314\ \cancel{\text{kPa·L}}/\text{mol·}\cancel{\text{K}} \times 351\ \cancel{\text{K}}} \\ &= 0.0282 \text{ mol} \end{aligned}$$

$$\begin{aligned} M &= \frac{m}{n} \\ &= \frac{2.366 \text{ g}}{0.0282 \text{ mol}} \\ &= 83.9 \text{ g/mol} \end{aligned}$$

The molar mass of the gas is 83.9 g/mol.

To identify the gas, compare the molar masses of the four gases mentioned.

Bromine, Br_2, has a molar mass of 2×79.9 g/mol $= 141.8$ g/mol.

Krypton, Kr, has a molar mass of 83.8 g/mol.

Neon, Ne, has a molar mass of 20.2 g/mol.

Fluorine, F_2, has a molar mass of 2×18.9 g/mol $= 38.0$ g/mol.

Therefore, the gas must be krypton.

Check Your Solution

The units of the answer are g/mol, the correct units for molar mass. The answer has three significant digits, equal to the least number of digits in the question.

The answer is probably correct, since it is so close to the molar mass of one of the given gases.

In the following investigation, you will find the molar mass of an unknown gas. You will use this mass to identify the gas.

The ideal gas law gives great flexibility for solving many different types of problems. After the investigation, you will find another Sample Problem. It illustrates how you can use the ideal gas law with methods you have previously learned, to identify an unknown gas. Practice problems are located at the end of this Sample Problem.

SKILL FOCUS

Performing and recording

Analyzing and interpreting

Communicating results

Investigation 12-A

Calculating the Molar Mass of an Unknown Gas: Teacher Demonstration

Cigarette lighters contain a gaseous fuel that burns quickly. It produces a large amount of heat using only a small amount of gas. Your teacher will measure the volume and mass of a sample of this gas. Then you will use these data to calculate the molar mass of the gas.

Materials

4 L beaker or plastic pail
500 mL graduated cylinder
disposable cigarette lighter
needle nose pliers
plastic wrap
balance or scale
tap water
thermometer
barometer
hair dryer

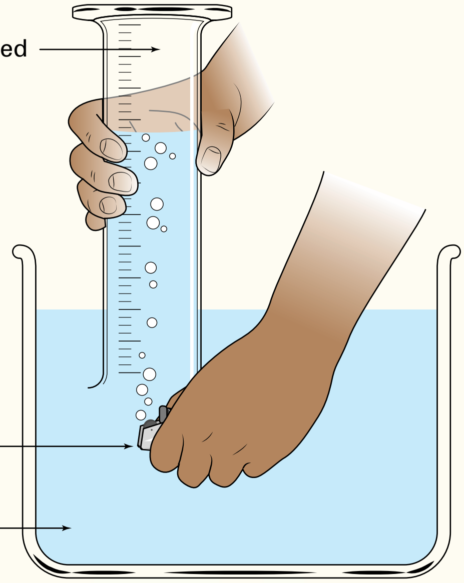
Your teacher will take the following Safety Precautions and perform the following steps.

Safety Precautions

- Remember that this gas is flammable. Do not try to produce flames with the lighters. Before beginning, check that there are no flames (such as lit Bunsen burners) in the laboratory.
- If water is spilled on the floor, wipe it up immediately so that no one steps in it.
- Release all the gas collected into an operating fume hood after the investigation. Expel the remaining gas in the lighter outside before disposing of the lighter.

Procedure

1. Use pliers to remove the striker, flint, and spring from the disposable lighter.
2. Fill the 4 L beaker (or pail) about two-thirds full of tap water. Determine the temperature of the tap water. Use this measurement to approximate the temperature of the gas.
3. Briefly immerse the lighter in the water, then shake it dry. Use the hair dryer on a low or cool setting to dry the lighter as much as possible. This is to set a standard for drying the empty lighter later. CAUTION Do not overheat the lighter.
4. Determine the mass of the lighter.
5. Fill the graduated cylinder with water. Cover the cylinder tightly with a piece of plastic wrap. With your hand over the plastic wrap, place the cylinder upside down into the beaker. Make sure that no air bubbles are trapped in the cylinder. Slide the plastic wrap away. A water-filled measuring tube to collect gas that has been created. The gas will displace the water as it rises, giving an accurate measurement.



6. As shown in the diagram, hold the lighter underwater, below the graduated cylinder in the beaker. Carefully depress the button on the lighter to release gas into the cylinder. The entire lighter does not need to be emptied. Just gather enough gas for an accurate measurement.
7. Add tap water to the beaker (or pail), or lift the cylinder, so that *the water inside the cylinder is at exactly the same level as the water in the beaker*. This equalizes the pressure in the cylinder with the pressure of the atmosphere. Record the volume of the gas collected when the water levels are equal inside and outside the cylinder.
8. Dry the lighter with the hair dryer. Measure its mass.
9. Record the air pressure in the room.
10. Wash your hands.

Analysis

1. (a) Subtract the final mass of the lighter from its initial mass. This will give you the mass of the gas used.
   (b) Use the volume, the mass, the temperature of the water, and the air pressure to calculate the number of moles of gas.
2. Use the mass of the gas and the number of moles to calculate the molar mass.

Conclusions

3. The gas in the lighter has the formula C_nH_{2n+2}.
   (a) Use the periodic table to calculate the molar masses of the compounds with $n = 1$ to $n = 5$.
   (b) Identify the gas in the lighter. You will learn how to name this gas in Chapter 13.
4. How do your results compare with the theoretical molar mass calculated from the periodic table?
5. Calculate the percent error for your results.
6. What were the sources of error?

Figure 12.13 Methane gas, CH_4, is used as a fuel for Bunsen burners. Suppose that you are given a container of an unknown gas. What methods could you use to find out if it is methane?

Sample Problem

Identifying a Compound Using Percent Composition and the Ideal Gas Law

Problem

As geologists study the area where an ancient marsh was located, they discover an unknown gas seeping from the ground. They collect a sample of the gas, and take it to a lab for analysis. Lab technicians find that the gas is made up of 80.0% carbon and 20.0% hydrogen. They also find that a 4.60 g sample occupies a volume of 2.50 L at 1.50 atm and 25.0°C. What is the molecular formula of the gas?

What Is Required?

You need to find the molecular formula (and thus the identity) of an unknown gas.

What Is Given?

$P = 1.50$ atm

$V = 2.50$ L

$m = 4.60$ g

$R = 8.314$ kPa·L/mol·K

$T = 25.0$°C

The percentage composition of the gas is 80.0% carbon and 20.0% hydrogen.

Plan Your Strategy

Step 1 Find the empirical formula of the gas, using the molar masses of carbon and hydrogen and the percent compositions.

Step 2 Solve the ideal gas law for the number of moles. First change the temperature to kelvins. Then change the pressure to kilopascals and use $R = 8.314$ kPa·L/mol·K. (You could also use $R = 0.08206$ atm·L/mol·K.)

Step 3 Find the molar mass (M) of the compound by dividing the mass of the sample (m) by the number of moles (n) in the sample.

Step 4 Compare the molar mass of the unknown gas with the molar mass of the empirical formula. To find the molecular formula, multiply the empirical formula by the ratio of the two molar masses.

Act on Your Strategy

Step 1 Find the empirical formula. Assume that the total mass of the sample is 100.0 g. Thus the mass of the carbon in the sample is 80.0% × 100.0 g = 80.0 g. The mass of the hydrogen is 20.0% × 100.0 g = 20.0 g.

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Now find the number of moles of carbon and hydrogen using the formula $n = \frac{m}{M}$

For carbon,

$$n = \frac{80.0\ \cancel{g}}{12.01\ \cancel{g}/\text{mol}} = 6.67\ \text{mol}$$

For hydrogen,

$$n = \frac{20.0\ \cancel{g}}{1.01\ \cancel{g}/\text{mol}} = 20.0\ \text{mol}$$

Finally, find the simplest mole ratio of the two elements in the compound. This will be the empirical, or simplest, formula of the gas.

The ratio of the elements in the compound is

$\frac{6.67}{6.67}$ mol of C to $\frac{20.0}{6.67}$ mol of H

or 1.0 mol of C to 3.0 mol of H

The empirical formula of the unknown gas is CH_3.

Step 2 Use the ideal gas law.

$$T = (25.0^\circ\text{C} + 273) = 298\ \text{K}$$

$$PV = nRT$$

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

If you convert the units of R ...	If you convert the pressure to kPa ...
$R = 0.08206\ \text{atm·L/mol·K}$	$P = 1.50\ \text{atm} \times 101.3\ \text{kPa/atm}$ $= 152\ \text{kPa}$
$n = \frac{(1.50\ \text{atm})(2.50\ \text{L})}{(0.08206\ \text{atm·L/mol·K})(298\ \text{K})}$ $= 0.153\ \text{mol}$	$n = \frac{(1.52\ \text{kPa})(2.50\ \text{L})}{(8.314\ \text{kPa·L/mol·K})(298\ \text{K})}$ $= 0.153\ \text{mol}$

Step 3 Find the molar mass.

$$M = \frac{m}{n} = \frac{4.60\ \text{g}}{0.153\ \text{mol}} = 30.1\ \text{g/mol}$$

Step 4 Find the molecular formula.

The molar mass of the compound is 30.1 g/mol.

The molar mass of the empirical formula, CH_3, is 15.04 g/mol.

$$\text{Ratio of molar masses} = \frac{30.07\ \cancel{\text{g/mol}}}{15.04\ \cancel{\text{g/mol}}} = 2.00$$

Thus, the molecular formula of the unknown gas is twice the empirical formula.

$$CH_3 \times 2 = C_2H_6$$

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Check Your Solution

In step 2, the units in the ideal gas equation cancel out to give moles. C_2H_6 is a reasonable answer for the molecular formula of the gas, since it is a simple integer ratio of the two molar masses.

Practice Problems

20. A 1.56 L gas sample has a mass of 3.22 g at 100 kPa and 281 K. What is the molar mass of the gas?

21. 2.0 L of haloethane has a mass of 14.1 g at 344 K and 1.01 atm. What is the molar mass of haloethane?

22. A vapour has a mass of 0.548 g in 237 mL, at 373 K and 755 torr. What is the molar mass of the vapour?

23. The mass of a 5.00 L evacuated container is 125.00 g. When the container is filled with argon gas at 298 K and 105.0 kPa, it has a mass of 133.47 g.

(a) Calculate the density of argon under these conditions.

(b) What is the density of argon at STP?

24. A gaseous compound contains 92.31% carbon and 7.69% hydrogen by mass. 4.35 g of the gas occupies 4.16 L at 22.0°C and 738 torr. Determine the molecular formula of the gas.

Section Wrap-up

In this section, you learned how density and molar mass are related to the ideal gas law. You also learned how to identify an unknown substance by calculating its molar mass, both theoretically and in the laboratory. Before you continue, take the time to complete the following Section Review questions. They will help you remember what you have learned.

Section Review

Unit Issue Prep

You will be debating a question related to gas pollution in the Unit Issue. How does the density of a gas determine whether the gas pollutes Earth's surface or the atmosphere?

1 I What is the density of methane, $CH_{4(g)}$, if 4.5 mol are in 100 L?

2 I 8.1 g of a gas occupy 12.3 L of space at 27°C and 8 atm.

(a) What is the molar mass of the gas?

(b) What might the gas be?

3 I A gas that consists of only nitrogen and oxygen atoms is found to contain 30% nitrogen. A 9.23 g sample of the gas occupies 2.2 L at STP. What is the gas?

4 I You are given a sample of an unknown gas. Describe how you can identify the gas in the laboratory. What measurements will you take? What apparatus might you need?

5 K/U How can the densities of two gases at STP be different, even though their volumes are the same?

Gas Law Stoichiometry

12.3

Section Preview/ Specific Expectations

In this section, you will

- **perform** stoichiometric calculations involving the number of moles, number of atoms, number of molecules, mass, and volume of substances in a balanced chemical reaction
- **determine** the molar volume of hydrogen in an investigation

Many chemical reactions in everyday life involve gases. Figure 12.14 shows a common reaction that has a gas as a reactant. Other reactions, such as the electrolyzation of salt to give chlorine gas, have gases as products. To carry out an accurate and efficient reaction, scientists must know the number of moles of all the reactants. When one or more of the reactants is a gas, this means using the ideal gas law.

Figure 12.14 **In this photograph, oxygen gas reacts with calcium to produce calcium oxide.**

You have already learned that the ideal gas law can be used to solve for different variables in several different types of situations. As you may recall, *the term "stoichiometry" refers to the relationship between the number of moles of the reactants and the number of moles of the products in a chemical reaction*. In this section, you will learn how to use Gay-Lussac's law of combining volumes and the ideal gas law to solve stoichiometric problems that involve gases.

Volume to Volume Stoichiometry

At the beginning of this chapter, you were introduced to Gay-Lussac's law of combining volumes: *When gases react, the volumes of the reactants and the products, measured at equal temperatures and pressures, are always in whole number ratios*. As well, you learned that the mole ratios from a chemical equation are the same as the ratios of the volumes of the gases.

This information will help you with a certain type of gas stoichiometry problem. When a gas reacts to produce another gas, you can use Gay-Lussac's law of combining volumes to find the volumes of the gases. The following Sample Problem shows you how.

CHECKPOINT

Go back to the beginning of this chapter. Make sure that you understand why the mole ratios are the same as the volume ratios.

Sample Problem

Gay-Lussac's Law of Combining Volumes

Problem

Ammonia is produced by a reaction of nitrogen gas and hydrogen gas. The chemical equation for the reaction is

$$N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$$

Suppose that 12.0 L of nitrogen gas reacts with hydrogen gas at the same temperature and pressure.

(a) What volume of ammonia gas is produced?

(b) What volume of hydrogen is consumed?

What Is Required?

(a) Calculate the volume of ammonia gas produced when 12.0 L of nitrogen gas reacts.

(b) Calculate the volume of hydrogen gas used up by the reaction.

What Is Given?

From the equation, you know that 12.0 L of nitrogen gas is used. The mole ratios from the equation are

$$\frac{2 \text{ mol } NH_{3(g)}}{1 \text{ mol } N_{2(g)}}$$

$$\frac{3 \text{ mol } H_{2(g)}}{1 \text{ mol } N_{2(g)}}$$

Plan Your Strategy

You know that the mole ratios of the volumes of gases are the same as the ratios of the volumes. Therefore, you can use the mole ratios to find the volumes of ammonia gas and hydrogen gas. You do not need to use the temperature and pressure, since they remain the same in this problem.

Act on Your Strategy

(a) Let x be the volume of ammonia gas.

$$\frac{2 \text{ mol } NH_{3(g)}}{1 \text{ mol } N_{2(g)}} = \frac{x \text{ L } NH_{3(g)}}{12.0 \text{ L } N_{2(g)}}$$

$$(12.0 \text{ L } N_{2(g)}) \frac{2 \cancel{\text{mol}} NH_{3(g)}}{1 \cancel{\text{mol}} N_{2(g)}} = \frac{x \text{ L } NH_{3(g)}}{12.0 \text{ L } N_{2(g)}} (12.0 \text{ L } N_{2(g)})$$

$$x = 24.0 \text{ L } NH_{3(g)}$$

Therefore, 24.0 L of ammonia gas is produced.

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(b) Let y be the volume of hydrogen gas.

$$\frac{3 \text{ mol } H_{2(g)}}{1 \text{ mol } N_{2(g)}} = \frac{y \text{ L } H_{2(g)}}{12.0 \text{ L } N_{2(g)}}$$

$$(12.0 \text{ L } N_{2(g)})\frac{3 \text{ mol } H_{2(g)}}{1 \text{ mol } N_{2(g)}} = \frac{y \; H_{2(g)}}{12.0 \text{ L } N_{2(g)}}(12.0 \text{ L } N_{2(g)})$$

$y = 36.0 \text{ L } H_{2(g)}$

Therefore, 36.0 L of hydrogen gas is consumed.

Check Your Solution

The number of significant digits in the answer is the same as the number of significant digits in the question.

The mole ratio of ammonia to nitrogen is 2:1. Thus, it makes sense that the volume of ammonia gas is twice the volume of nitrogen gas.

The mole ratio of hydrogen to nitrogen is 3:1. It makes sense that the volume of hydrogen gas is three times the volume of nitrogen gas.

Practice Problems

25. Use the following balanced equation to answer the questions below.

$$2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$$

(a) What is the mole ratio of oxygen gas to water vapour?

(b) What is the volume ratio of oxygen gas to water vapour?

(c) What is the volume ratio of hydrogen gas to oxygen gas?

(d) What is the volume ratio of water vapour to hydrogen gas?

26. 1.5 L of propane gas are burned in a barbecue. The following equation shows the reaction. Assume all gases are at STP.

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

(a) What volume of carbon dioxide gas is produced?

(b) What volume of oxygen is consumed?

27. Use the following equation to answer the questions below.

$$SO_{2(g)} + O_{2(g)} \rightarrow SO_{3(g)}$$

(a) Balance the equation.

(b) 12.0 L of sulfur trioxide, $SO_{3(g)}$, are produced at 100°C. What volume of oxygen is consumed?

(c) What assumption must you make to answer part (b)?

28. 2.0 L of gas A react with 1.0 L of gas B to produce 1.0 L of gas C. All gases are at STP.

(a) Write the balanced chemical equation for this reaction.

(b) Each molecule of gas A is made of two identical “a” atoms. That is, gas A is really $a_{2(g)}$. In the same way, each molecule of gas B is made of two identical “b” atoms. What is the chemical formula of gas C in terms of “a” and “b” atoms?

Solving Gas Stoichiometry Problems

Earlier in this course, you learned how to do stoichiometry calculations. To solve gas stoichiometry problems, you will incorporate the ideal gas law into what you learned previously. The following steps will help you do this.

How to Solve Gas Stoichiometry Problems

1. Write a balanced equation for the reaction.
2. Write the given information under the appropriate reactants and products. Put a question mark under the reactant or product for which information is needed.
3. Convert all amounts to moles.
4. Compare molar amounts using stoichiometry ratios from the balanced equation. Solve for the unknown molar amount.
5. Convert the new molar amount into the units required. You may multiply by a conversion factor, or use a set of conditions with the ideal gas law, $PV = nRT$.

Using the Ideal Gas Law for the Gaseous Product of a Reaction

The best way to find out how to do a stoichiometry problem using the ideal gas law is to study an example. In the following Sample Problem, you will use a balanced equation and the ideal gas law to find the volume of a gas produced. (Refer to Chapter 4, section 4.1, if you want to review how to write balanced equations.)

Sample Problem

Mass to Volume Stoichiometry

Problem

Ancient alchemists liked to use strong sulfuric acid to produce dramatically dangerous effects. One interesting reaction occurs when sulfuric acid reacts with iron metal to produce gas and an iron(II) compound. What volume of gas is produced when excess sulfuric acid reacts with 40.0 g of iron at 18.0°C and 100.3 kPa?

What Is Required?

Calculate the volume of gas that is produced when sulfuric acid reacts with iron under specific temperature and pressure conditions.

What Is Given?

Reactants: sulfuric acid and iron

Products: an iron(II) compound and a gas

Mass of iron = 40.0 g

Temperature = 18.0°C

Pressure = 100.3 kPa

Continued ...

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Plan Your Strategy

Step 1 Write a balanced equation for the chemical reaction.

Step 2 Find the number of moles of iron present. Use this value, along with the mole ratios from the balanced equation, to find the number of moles of gas produced.

Step 3 Use the ideal gas law. You know the number of moles of gas, the temperature, and the pressure. (Do not forget to change the temperature to kelvins.) Solve for the volume of the gas.

Act on Your Strategy

Step 1 Write the balanced equation. (This reaction is a single displacement reaction.)

$$Fe_{(s)} + H_2SO_{4(aq)} \rightarrow H_{2(g)} + FeSO_{4(aq)}$$

40.0 g ? L

Step 2 Find the number of moles of iron, and the number of moles of gas.

To find the number of moles of iron, divide the mass by the molar mass. You can find the molar mass of iron in the periodic table: 55.85 g/mol. **Note:** If the reactant was a compound, such as $FeCl_2$, you would need to calculate the molar mass by adding the molar masses of all the atoms.

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{40.0 \text{ g Fe}}{55.85 \text{ g/mol}} \\ &= 0.716 \text{ mol Fe} \end{aligned}$$

From the balanced equation, the mole ratio is

$$\frac{1 \text{ mol } H_2}{1 \text{ mol Fe}}$$

Use this ratio to find the number of moles of hydrogen gas formed by the reaction.

$$\frac{n \text{ mol } H_2}{0.716 \text{ mol Fe}} = \frac{1 \text{ mol } H_2}{1 \text{ mol Fe}}$$

$$(0.716 \text{ mol Fe})\frac{n \text{ mol } H_2}{0.716 \text{ mol Fe}} = \frac{1 \text{ mol } H_2}{1 \text{ mol Fe}}(0.716 \text{ mol Fe})$$

$n = 0.716 \text{ mol } H_2$

Step 3 Use the ideal gas law to solve for the volume, since all the other quantities are now known.

First change the temperature to kelvins.

$18.0°C + 273 = 291$ K

You now have all the values you need to solve for volume.

$P = 100.3$ kPa

$n = 0.716$ mol

$R = 8.314$ kPa·L/mol·K

$T = 291$ K

Continued ...

Web LINK

www.school.mcgrawhill/resources

One of the problems with air bags (see Figure 12.15 on the next page) is that they can harm a child or small adult because too much gas is produced. To find out what volume of nitrogen gas in an air bag is safe for a child, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next. Use the volume you find to calculate the mass of $NaN_{3(s)}$ that is needed to produce this volume at 22.0°C and 105 kPa.

Figure 12.15 Air bags must be tested thoroughly before being manufactured for public use.

CHEM FACT

An automobile air bag fills up with about 65 L of nitrogen gas in approximately 27 ms. This can prevent a driver from being seriously injured. The sodium that is produced is extremely caustic, however. It reacts with iron(III) oxide as follows:

$$6Na_{(s)} + Fe_2O_{3(s)} \rightarrow 3Na_2O_{(s)} + 2Fe_{(s)}$$

The sodium oxide then reacts with carbon dioxide and water vapour.

$$Na_2O(s) + 2CO_{2(g)} + H_2O_{(g)} \rightarrow 2NaHCO_{3(s)}$$

The sodium hydrogen carbonate that is produced is a harmless substance. It is better known as baking soda.

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$$PV = nRT$$

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

$$= \frac{0.716\ \cancel{\text{mol}} \times 8.314\ \cancel{\text{kPa}}\cdot\text{L}/\cancel{\text{mol}}\cdot\cancel{\text{K}} \times 291\ \cancel{\text{K}}}{100.3\ \cancel{\text{kPa}}}$$

$$= 17.3\ \text{L}$$

Therefore, 17.3 L of hydrogen gas is produced by this reaction at 18.0°C and 100.3 kPa.

Check Your Solution

The answer is slightly less than the molar volume of hydrogen gas at STP. Since less than one mole of hydrogen gas was formed, this seems reasonable.

Practice Problems

29. Engineers design automobile air bags that deploy almost instantly on impact. To do this, an air bag must provide a large amount of gas in a very short time. Many automobile manufacturers use solid sodium azide, NaN_3, along with suitable catalysts, to provide the gas that is needed to inflate the air bag. The balanced equation for this reaction is

$$2NaN_{3(s)} \rightarrow 2Na_{(s)} + 3N_{2(g)}$$

(a) What volume of nitrogen gas will be produced if 117.0 g of sodium azide are stored in the steering wheel at 20.2°C and 101.2 kPa?

(b) How many molecules of nitrogen are present in this volume?

(c) How many atoms are present in this volume?

30. 0.72 g of hydrogen gas, H_2, reacts with 8.0 L of chlorine gas, Cl_2, at STP. How many litres of hydrogen chloride gas, HCl, are produced?

31. How many grams of baking soda (sodium hydrogen carbonate, $NaHCO_3$), must be used to produce 45 mL of carbon dioxide gas at 190°C and 101.3 kPa in a pan of muffins? (The mole ratio of $NaHCO_3$ to CO_2 is 2:1.)

32. How much zinc (in grams) must react with hydrochloric acid to produce 18 mL of gas at SATP? (**Hint:** Zinc chloride, $ZnC\ell_{2(s)}$ is a product.)

33. 35 g of propane gas burned in a barbecue, according to the following equation:

$$C_3H_{8(g)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$$

All the gases are measured at SATP.

(a) What volume of water vapour is produced?

(b) What volume of oxygen is consumed?

34. What mass of oxygen is reacted to produce 0.62 L of water vapour at 100°C and 101.3 KPa? Start by balancing the following equation:

$$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{(g)}$$

Including Water Vapour Pressure in Gas Calculations

You can collect many gases by allowing them to bubble up through water into a container that is filled with water. (See Figure 12.16). This is the method your teacher used to collect the gas in Investigation 12-A. Unfortunately, molecules of water vapour mix with the gas sample. To avoid error, the pressure that was contributed by the water vapour must be subtracted when finding the pressure of the gas.

As an example, consider hydrogen gas, which is often collected over water. The hydrogen that is collected is a mixture of hydrogen and water vapour. As you learned from Dalton's law of partial pressures in Chapter 11, the pressure of this mixture is

$$P_{\text{total}} = P_{\text{hydrogen}} + P_{\text{water vapour}}$$

To find the partial pressure of dry hydrogen, subtract the pressure of the water vapour from the total pressure.

$$P_{\text{hydrogen}} = P_{\text{total}} - P_{\text{water vapour}}$$

The pressure of the water vapour is the same for any gas that is collected at a particular temperature. For example, the pressure of water vapour at 25°C is 3.17 kPa. Table 12.3 gives the pressure of water vapour at different temperatures.

When using the ideal gas law for a gas collected over water, you must correct the pressure before you substitute it into the gas law. The following Sample Problem shows you how to do this.

Figure 12.16 This is an efficient and convenient method for collecting hydrogen gas. Unfortunately molecules of water vapour mix with the gas sample.

Table 12.3
Pressure of Water Vapour

Temperature (°C)	Pressure (kPa)
17	1.94
18	2.06
19	2.20
20	2.34
21	2.49
22	2.64
23	2.81
24	2.98
25	3.17
26	3.36
27	3.56
28	3.78
29	4.00
30	4.24

Sample Problem

Calculating the Volume of a Gas Collected Over Water

Problem

A student reacts magnesium with excess dilute hydrochloric acid to produce hydrogen gas. She uses 0.15 g of magnesium metal. What volume of dry hydrogen does she collect over water at 28°C and 101.8 kPa?

What Is Required?

You need to find the volume of hydrogen collected over water in this reaction.

What Is Given?

T = 28.0°C

P = 101.8 kPa

Mass of magnesium (m) = 0.15 g

Pressure of water vapour at 28°C = 3.78 kPa

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Plan Your Strategy

Step 1 Write a balanced chemical equation for the reaction.

Step 2 Calculate the number of moles of magnesium by dividing the mass given (m) by the molar mass of magnesium (M). Use the number of moles of magnesium, along with the mole ratio from the equation, to calculate the number of moles of hydrogen gas produced by the reaction.

Step 3 Convert the temperature to kelvins. Since the hydrogen is collected over water, subtract the pressure of the water vapour at 28°C from the atmospheric pressure.

Step 4 Use the ideal gas law to find the unknown volume (V) of hydrogen gas.

Act on Your Strategy

Step 1 The balanced chemical equation is

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

0.15 g ? L

Step 2 Find the number of moles of magnesium.

From the periodic table, the molar mass of magnesium is 24.31 g/mol.

$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{0.15\ \cancel{g}}{24.31\ \cancel{g}/\text{mol}} \\ &= 6.2 \times 10^{-3}\ \text{mol} \end{aligned}$$

The mole ratio of hydrogen gas to magnesium in this reaction is

$$\frac{1\ \text{mol}\ H_2}{1\ \text{mol Mg}}$$

Using the mole ratio,

$$\frac{n\ \text{mol}\ H_2}{6.2 \times 10^{-3}\ \text{mol}} = \frac{1\ \text{mol}\ H_2}{1\ \text{mol Mg}}$$

Cross multiply to get

$$(6.2 \times 10^{-3}\ \text{mol Mg})\frac{n\ \text{mol}\ H_2}{6.2 \times 10^{-3}\ \text{mol Mg}} = \frac{1\ \text{mol}\ H_2}{1\ \text{mol Mg}}(6.2 \times 10^{-3}\ \text{mol Mg})$$

$$n = 6.2 \times 10^{-3}\ \text{mol}$$

Step 3 Convert the temperature and pressure to kelvins.

$$\begin{aligned} T &= (28^\circ\text{C} + 273) \\ &= 301\ \text{K} \end{aligned}$$

The pressure that is exerted by the hydrogen gas is

$$\begin{aligned} P_{\text{hydrogen}} &= 101.8\ \text{kPa} - 3.78\ \text{kPa} \\ &= 98.0\ \text{kPa} \end{aligned}$$

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Step 4 Use the ideal gas law.

$$PV = nRT$$

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

$$= \frac{(6.2 \times 10^{-3}\ \cancel{\text{mol}})(8.314\ \cancel{\text{kPa}}\cdot\text{L}/\cancel{\text{mol}\cdot\text{K}})(301\ \cancel{\text{K}})}{(98.0\ \cancel{\text{kPa}})}$$

$$= 0.16\ \text{L}$$

The student collects 0.16 L of dry hydrogen.

Check Your Solution

The final answer is rounded to two significant digits. This is the least number of significant digits in the question. The mass of the magnesium is a small number. Therefore, the volume of the hydrogen produced is also a small number.

Using the Ideal Gas Law for the Gaseous Reactant of a Reaction

So far, you have used the ideal gas law for reactions with a gas *product*. You can also use the ideal gas law for reactions with a gas as a *reactant*. The following Sample Problem shows you how to do this.

Sample Problem

Space Shuttle Science: Gas as a Reactant

Problem

When astronauts travel in a space shuttle (Figure 12.17), carbon dioxide must be removed from the air they breathe. One method is to bubble the air in the shuttle though a solution of lithium hydroxide. The lithium hydroxide converts any carbon dioxide into lithium carbonate.

$$2LiOH_{(aq)} + CO_{2(g)} \rightarrow Li_2CO_{3(aq)} + H_2O_{(\ell)}$$

(You learned about a different method in Chapter 7 using solid LiOH.) Air containing 25.0 L of carbon dioxide is passed through 1.5 mol/L LiOH solution over a 20 min period. The atmospheric pressure in the shuttle is 0.85 atm, and the temperature is 28.3˚C. What mass of lithium carbonate is produced?

Note: Although the air is bubbled through an aqueous solution, you do not need to consider the pressure of the water vapour. This is because you are dealing with the *reactant* not the *product*.

Figure 12.17

Continued ...

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What Is Required?

Find the mass of lithium carbonate that is produced when 25.0 L of carbon dioxide is bubbled through an aqueous solution of lithium hydroxide.

What Is Given?

$2LiOH_{(aq)} + CO_{2(g)} \rightarrow Li_2CO_{3(aq)} + H_2O_{(\ell)}$

1.0 L $\quad V = 25.0$ L
1.5 mol/L $\quad T = 28.3°C$
$P = 0.85$ atm

Plan Your Strategy

Step 1 Convert the temperature to kelvins. Convert the pressure from atm to kPa, or use $R = 0.8206$ atm·L/mol·K.

Step 2 Use the ideal gas law to find the number of moles of carbon dioxide that reacts.

Step 3 Use the stoichiometry of the equation to determine the number of moles of lithium carbonate produced.

Step 4 Use the periodic table to determine the molar mass of lithium carbonate. To find the mass of lithium carbonate produced, multiply the number of moles by the molar mass.

Act on Your Strategy

Step 1 Convert the temperature and pressure.

$$T = 28.3°C + 273$$
$$= 301 \text{ K}$$

$$0.85 \cancel{\text{atm}} \times \frac{101.3 \text{ kPa}}{1 \cancel{\text{atm}}} = 86 \text{ kPa}$$

Step 2 Find the number of moles of $CO_{2(g)}$.

$$PV = nRT$$

$$\therefore n = \frac{PV}{RT}$$
$$= \frac{(0.85 \cancel{\text{atm}})(25.0 \cancel{\text{L}})}{(0.08206 \cancel{\text{atm}}{\cdot}\cancel{\text{L}}/\text{mol}{\cdot}\cancel{\text{K}})(301 \cancel{\text{K}})}$$
$$= 0.86 \text{ mol}$$

Therefore, 0.86 mol of $CO_{2(g)}$ passes through the LiOH solution.

Step 3 Find the number of moles of Li_2CO_3 produced.

From the balanced equation, we know that 1 mol of carbon dioxide produces 1 mol of lithium carbonate.

$$\frac{n \text{ mol } Li_2CO_3}{0.86 \text{ mol } CO_2} = \frac{1 \text{ mol } Li_2CO_3}{1 \text{ mol } CO_2}$$

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$$(0.86 \text{ mol } CO_2)\ \frac{n \text{ mol } Li_2CO_3}{0.86 \text{ mol } CO_2} = \frac{1\text{mol } Li_2CO_3}{1\text{mol } CO_2}\ (0.86 \text{ mol } CO_2)$$

n = 0.86 mol of Li_2CO_3

Step 4 Find the molar mass of Li_2CO_3. Then find the mass of Li_2CO_3 produced.

$$M_{Li_2CO_3} = (2 \times 6.94 + 12.01 + 3 \times 16.00) \text{ g/mol}$$
$$= 73.89 \text{ g/mol}$$

Thus, the mass of Li_2CO_3 produced is

$$m = n \times M$$
$$= 0.86 \text{ mol} \times 73.89 \text{ g/mol}$$
$$= 64 \text{ g}$$

Therefore, 64 g of Li_2CO_3 is produced in this reaction.

Check Your Solution.

The answer has two significant digits. This is the least number of significant digits in the question.

When the units cancel in the ideal gas equation, mol remains. When the units cancel in the final calculation, g remains.

The answer seems reasonable.

Practice Problems

35. Oxygen, O_2, reacts with magnesium, Mg, to produce 243 g of magnesium oxide, MgO, at 101.3 kPa and 45°C. How many litres of oxygen are consumed? Start by writing the balanced equation.

36. Zinc reacts with nitric acid to produce 34 L of dry hydrogen gas at 900 torr and 20°C. How many grams of zinc are consumed?

37. 0.75 L of hydrogen gas is collected over water at 25.0°C and 101.6 kPa. What volume will the dry hydrogen occupy at 103.3 kPa and 25.0°C?

38. 3070 kg of coal burns to produce carbon dioxide. Assume that the coal is 95% pure carbon and the combustion is 80% efficient. (**Hint:** The mole ratio of $C_{(s)}$ to $CO_{2(g)}$ is 5:4.) How many litres of carbon dioxide are produced at SATP?

39. When 7.48 g of iron reacts with chlorine gas, 21.73 g of product is formed.

(a) How many moles of chlorine are used?

(b) What is the formula for the product?

(c) Write the equation for the reaction that occurs.

Now you have a chance to do an exciting investigation: the reaction of magnesium metal with strong acid. Remember to take appropriate safety precautions and follow your teacher's directions when working with the acid.

Investigation 12-B

SKILL FOCUS

Predicting

Performing and recording

Analyzing and interpreting

The Production of Hydrogen Gas

In this investigation, you will produce hydrogen gas by reacting strong acid with magnesium metal.

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

You will collect the hydrogen gas over water in a graduated cylinder.

Question

What is the molar volume of dry hydrogen gas at STP? Calculate it using the volume of hydrogen gas that you produce and the mass of one reactant.

Prediction

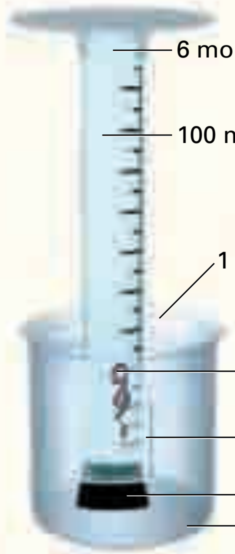
Predict the volume of hydrogen gas that will be produced. Use the mass of the magnesium and the balanced chemical equation given above. Assume 100% yield, and use regular stoichiometry. Organize your calculations clearly.

Safety Precautions

- Before beginning this investigation, check that there are no open flames (such as lit Bunsen burners) in the laboratory.
- The acid that you are using in this investigation is strong enough to burn. Wear your safety glasses and lab apron at all times. Handle the acid carefully. Wipe up any spills of water or acid immediately. If you accidentally spill any acid on your skin, wash it off immediately with large amounts of cool water.
- When you have finished the investigation, you can safely wash the products down the sink. You must dilute them, however, by running water down afterward.

Materials

scale or balance
100 mL graduated cylinder
stopper with two holes to fit graduated cylinder
1 L beaker or bowl
water at room temperature
6.0 mol/L hydrochloric acid, HCl
6 to 7 cm piece of magnesium ribbon
10 to 15 cm piece of copper wire
steel wool
barometer and thermometer
clamp and ring stand (optional)



Procedure

1. Prepare a table, in your notebook. Show your calculations.

Observations and Results

Observations	Trial 1	Trial 2 (if time permits)
mass of magnesium ribbon (g)		
temperature of water (°C)		
barometric pressure (kPa)		
volume of hydrogen collected (mL → L)		
vapour pressure of water at this temperature (kPa)*		

Results	Trial 1	Trial 2 (if time permits)
number of moles of magnesium (mol)		
volume of collected dry hydrogen at STP (L)		
molar volume of hydrogen at STP (L/mol)		

*Find the pressure of the water vapour in Table 12.3, or ask your teacher.

2. Obtain a piece of magnesium ribbon that is about 6 to 7 cm long. Use steel wool to clean the outside of the ribbon. Measure the mass of the ribbon.

3. Use the mass of the magnesium and the balanced equation to predict the volume of hydrogen gas that will be produced. Show your calculations in your notebook.

4. Fill the beaker (or bowl) about half full of water at room temperature. Measure the temperature of the water. (You will use this temperature to approximate the temperature of the gas produced.)

5. Measure and record the barometric pressure.

6. Add 15 mL of water to the graduated cylinder. Then, *very carefully*, pour 10 to 15 mL of 6 mol/L HCl into the graduated cylinder. *Very slowly and carefully*, pour water at room temperature down the *sides* of the cylinder until the cylinder is completely filled. Your objective in pouring the water this way is to avoid mixing it with the acid at the bottom of the cylinder. **CAUTION** Normally you should avoid adding water to an acid. Be particularly careful during this step.

7. Attach the magnesium ribbon to the copper wire. Dangle the magnesium in the graduated cylinder. The magnesium should hang 1 to 2 cm below the stopper. Put the stopper in the cylinder. Do not worry if a small amount of water overflows out of the cylinder.

8. Hold your gloved finger over the holes in the stopper. Tip the cylinder upside down into the 1 L beaker. Be careful that no air bubbles get into the cylinder. Hold or clamp the tube into place. Watch the reaction proceed. Record your observations in your notebook.

9. Add water, at room temperature, to the beaker until the level of the water inside the cylinder is exactly the same as the level of the water in the beaker. This equalizes the pressure of the hydrogen gas with the air pressure outside the tube. **Note:** Another way to equalize the pressure is to raise the graduated cylinder slightly to align the water levels.

10. Record the volume of the trapped gas.

11. All the magnesium should be used up by the reaction, since it is the limiting reagent. If any magnesium ribbon does remain after the reaction, rinse it with water, dry it with a paper towel, and measure its mass. To find the mass of the magnesium used up by the reaction, subtract the final mass from the initial mass.

12. Empty and clean all your apparatus. Clean your work space. Wash your hands.

Analysis

1. Calculate the molar volume of hydrogen. Use the volume of the H_2 gas, and the water temperature. Also, use the barometric pressure minus the pressure of the water vapour.

2. Use the combined or ideal gas law to translate the conditions to STP. Redo the calculations.

Conclusions

3. What was the class average for the molar volume of dry hydrogen gas at STP? How close was your molar volume to the class average?

4. How close was your molar volume to the accepted molar volume of a gas at STP? Calculate the percent error for your molar volume.

5. How would your results have been different had you not cleaned the magnesium ribbon before you used it?

6. What were some possible sources of error in your investigation?

Section Wrap-up

In this section, you learned how to use Gay-Lussac's law to calculate volumes of gases in a gas reaction. You also learned how to use the ideal gas law to find the volumes of gases used or produced in reactions. Building on Dalton's law of partial pressures, from Chapter 11, you learned how to calculate the molar volume of a gas collected over water. Finally, you learned, first-hand, how to produce hydrogen gas in a laboratory.

In the next section, you will see how gases in the atmosphere interact with the Sun's light. You will also find out about the dangers of gas pollution.

Web LINK

www.school.mcgrawhill.ca/resources/
What kind of company or industry uses, produces, or sells hydrogen gas? What is hydrogen gas used for? What safety precautions must be taken when handling it? To answer these questions, go to the web site above. Go to **Science Resources**, then to **Chemistry 11** to find out where to go next.

Section Review

1. K/U Gay-Lussac's law of combining volumes provides a short-cut for some gas calculations. What type of gas reaction lets you use this law?

2. K/U What values should you measure in the laboratory in order to calculate the molar volume of a gas?

3. C Why is it necessary to correct the pressure of a gas that is collected over water? How would you do this?

4. K/U Which will occupy a greater volume: 2.0 mol of nitrogen gas at STP or 1.9 mol of oxygen gas at SATP? Explain.

5. I 2.0 L of hydrogen gas reacts with oxygen at 5°C and 99 kPa. How many litres of water vapour are produced?

6. I 13 g of lead reacts with hydrofluoric acid to produce hydrogen gas at 22°C and 88.3 kPa.
   - **(a)** Describe the apparatus you would need to carry out this reaction in the laboratory.
   - **(b)** How many litres of *dry* hydrogen gas are collected?

7. I Plants consume carbon dioxide gas as they produce sugar during photosynthesis.
$$6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + 6O_2$$
   To produce 50 g of sugar, $C_6H_{12}O_6$, how many litres of gas at SATP does a sugar beet need to consume?

8. I 28 L of oxygen gas reacts with 52 L of hydrogen gas in an 80 L vessel at 25°C and 3.0 atm.
   - **(a)** How many grams of water are produced? **Hint:** Find out which reagent will be used up by the reaction.
   - **(b)** What will the pressure be in the cylinder after the reaction if all of the water that is formed condenses and is removed?

9. MC After completing the Web Link task on this page, prepare a brochure for an imaginary hydrogen manufacturing company. Your brochure should advertise the many uses of your product. It must also contain safety information for the customer.

Atmospheric Reactions and Pollution

12.4

Section Preview/ Specific Expectations

In this section, you will

- **explain** Canadian initiatives to improve air quality
- **communicate** your understanding of the following terms: *chlorofluorocarbons (CFCs), Montréal Protocol*

It is a warm summer weekend, so you decide to visit a national park. It is nice to be away from all that pollution. But are you bringing the pollution with you? Take a look at Figure 12.18. How did these campers get to the park? Cars and buses release gaseous pollution into the air. On the other hand, what would it be like to live without vehicles? Can we compromise in a way that protects the environment while maintaining a reasonable standard of living?

Figure 12.18 **These campers are enjoying the environment, but they are harming it at the same time. How?**

Fortunately many people are now aware of issues such as pollution from car exhaust and ozone depletion. At the same time, Canadian chemical and related industries employ about 250 000 people and generate tens of billions of dollars annually. These industries make the products that entertain, feed, clothe, and keep Canadians comfortable. When discussing atmospheric chemistry and pollution, we must consider two important influences: the economy and the environment.

In this section, you will look at the chemistry that takes place among the gases in the atmosphere. You will examine an important gas called ozone and the pollutants that affect this gas.

Gas Chemistry in the Atmosphere

Many important chemical reactions take place in the atmosphere. All these reactions involve gases. There are two main types of gas chemistry in the atmosphere:

- interactions between gases already present in the atmosphere
- interactions between atmospheric gases and gases produced by processes on Earth

The Ozone Cycle

High in the atmosphere, a gas called ozone, O_3, absorbs ultraviolet (UV) radiation from the Sun. The radiation separates the ozone into oxygen gas, O_2, and an oxygen atom. After passing through a few more steps, ozone is re-formed when molecules of oxygen gas combine with oxygen atoms. By absorbing energy from the Sun in this way, ozone prevents harmful UV radiation from reaching Earth's surface. Figure 12.19 illustrates the cycle that occurs as ozone is formed, absorbs UV radiation, and breaks up. Figure 12.20 shows where ozone is located in the atmosphere.

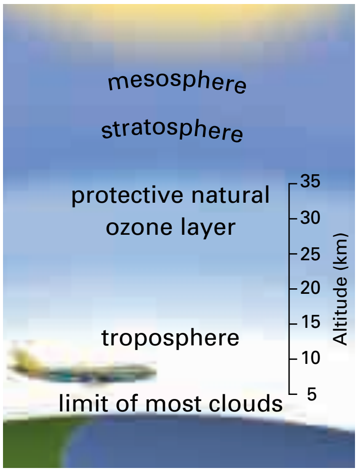


Figure 12.20 Most ozone is located about 25 km above Earth's surface. This is a very thin level, however. If all the ozone in the atmosphere were compressed, it would be only a few millimetres thick.

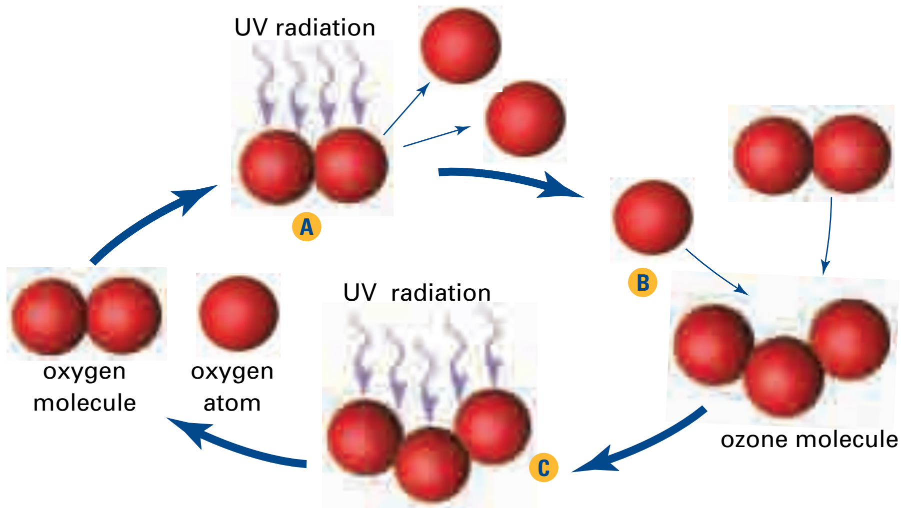


Figure 12.19 The ozone cycle: (A) Oxygen gas in the atmosphere absorbs energy from the Sun. Each oxygen molecule breaks up into two oxygen atoms. (B) An oxygen atom combines with a molecule of oxygen gas to form ozone. Another molecule is needed to absorb extra energy. (C) Ozone absorbs ultraviolet radiation and breaks up again into oxygen gas and an oxygen atom.

Pollutants in the Atmosphere

Gases from living and non-living processes on Earth's surface interact with the gases in the atmosphere. For example, oxygen and carbon dioxide in the air are involved in animal and plant respiration. A forest fire burns plants to produce carbon dioxide gas. On a more damaging level, human technology and industrial processes produce many polluting gases. As you will learn in Chapter 14, burning fossil fuels is a major source of gas pollution in the atmosphere.

Chlorofluorocarbons (CFCs) are an important class of polluting gases that are not usually caused by burning fossil fuels. CFCs are stable and harmless near the ground. When they make their way up into the atmosphere, however, they interact and interfere with atmospheric processes. In particular, these gases interfere with the production and reactions of ozone, O_3. You will learn more about CFCs later in this section.

Figure 12.21 Smog was originally defined as a mixture of smoke and fog. Today it can also be photochemical, caused by sunlight breaking down air pollutants.

Ozone Near the Ground

Ozone does not only exist high up in the stratosphere. It is also present much closer to us, in the *troposphere*: the layer of the atmosphere that lies directly over Earth. Ozone near the ground is largely produced when nitrogen oxide gas from car and truck exhaust fumes reacts with oxygen gas in sunlight. Ozone is a major component of smog in cities. (See Figure 12.21.)

At this level, ozone is a pollutant with a harsh odour. In humans and animals, it causes respiratory problems, including coughing, wheezing, and eye irritation. It retards plant growth, reduces the productivity of crops, and damages forests. Concentrations of ozone as low as 0.1 ppm (parts per million) can decrease photosynthesis by 50%. In addition, it damages plastics, breaks down rubber, and corrodes metals.

Careers in Chemistry

Environmental Technician

Gases in the atmosphere (such as carbon dioxide and methane) allow heat from the Sun to enter the atmosphere and prevent it from leaving again. This is called the *greenhouse effect*. Thanks to the greenhouse effect, Earth remains fairly warm, with an average temperature of about 15.5°C. Without this effect, Earth's temperature would be about −18°C.

Human activities over the last hundred years have caused the level of carbon dioxide and other gases in the atmosphere to increase. As a result, more heat is trapped in Earth's atmosphere. According to Environment Canada, Canada's average temperature has risen by about 1°C over the last century. This is causing more frequent and more intense winter storms.

Change from the Ground Up

How can we decrease the production of greenhouse gases? One way is to focus on the sources of these gases. As global warming increases, chemists who study atmospheric processes will be more in demand.

Change begins with accurate measurements. Environmental technicians assess and monitor pollution levels in air, water, and soil. To become an environmental technician, you need a high school diploma, with advanced-level credits in mathematics, English, and science (preferably chemistry or physics). You also need a two- or three-year community college program in environmental technology.

Environmental technicians gather gas samples from smokestacks and PCBs from transformers. They also set up equipment in the field to create baseline studies and monitor changes in the environment.

Make Career Connections

Human Resource Development Canada has offices in every province. It also has a web site where you can access descriptions and requirements for many careers, including that of environmental technician. What other environmental career opportunities can you locate? Prepare a brief report of your findings, and present it to the class.

CFCs and Ozone Depletion

As you learned earlier, **chlorofluorocarbons (CFCs)** are chemicals that interfere with the ozone cycle high up in the atmosphere. CFCs are non-toxic, nonflammable compounds that contain atoms of chlorine, carbon, and fluorine. These gases are human-made compounds that were released into the atmosphere primarily from refrigeration and aerosol devices.

In 1928, Thomas Midgley invented the first CFC compound. Because they were useful but safe, they were referred to as "miracle compounds." In particular, dichlorodifluoromethane, CCl_2F_2, also known as Freon, was discovered to be an efficient refrigerant.

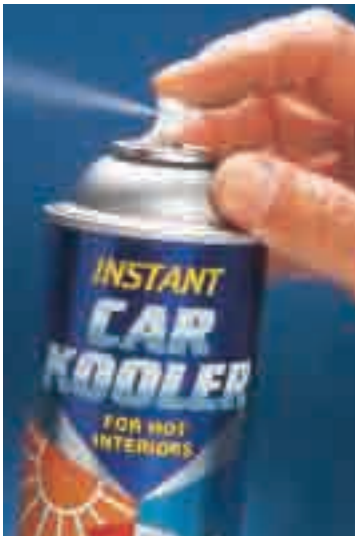


Figure 12.22 **CFCs have been used as refrigerants, coolants in home and automobile air conditioners, and propellants in aerosol containers such as hair sprays.**

Because of Freon, household refrigeration became common. In fact, much of our modern-day life style first became possible because of CFCs. By 1974, millions of tonnes of CFCs had been produced (see Figure 12.22). At the University of California, chemists F. Sherwood Rowland and Mario Molina began to wonder where all of these CFCs ended up. They realized that CFCs are chemically very stable. However, they began to calculate what happens when CFCs are exposed to high levels of radiation far up in the atmosphere. As it turned out, their fears were well-founded. In 1985, British scientists in the Antarctic noticed a large decrease in the ozone layer above the Antarctic. A "hole" in the ozone layer was beginning to form. In 1995, Rowland and Molina, along with a third ozone scientist, won the Nobel Prize for their work with CFCs.

How CFCs Attack Ozone

Today we know that CFCs high in the atmosphere break apart under ultraviolet radiation to produce chlorine atoms. These chlorine atoms destroy ozone molecules.

$$Cl + O_3 \rightarrow ClO + O_2$$

The product, ClO, reacts with an oxygen atom and releases the chlorine atom. The chlorine atom attacks another ozone molecule. Over time, one chlorine atom can destroy thousands of ozone molecules. Figure 12.23 illustrates this process for the CFC trichlorofluoromethane, CCl_3F. Eventually the chlorine atom reacts with a different compound in the atmosphere to form a stable, less harmful product.

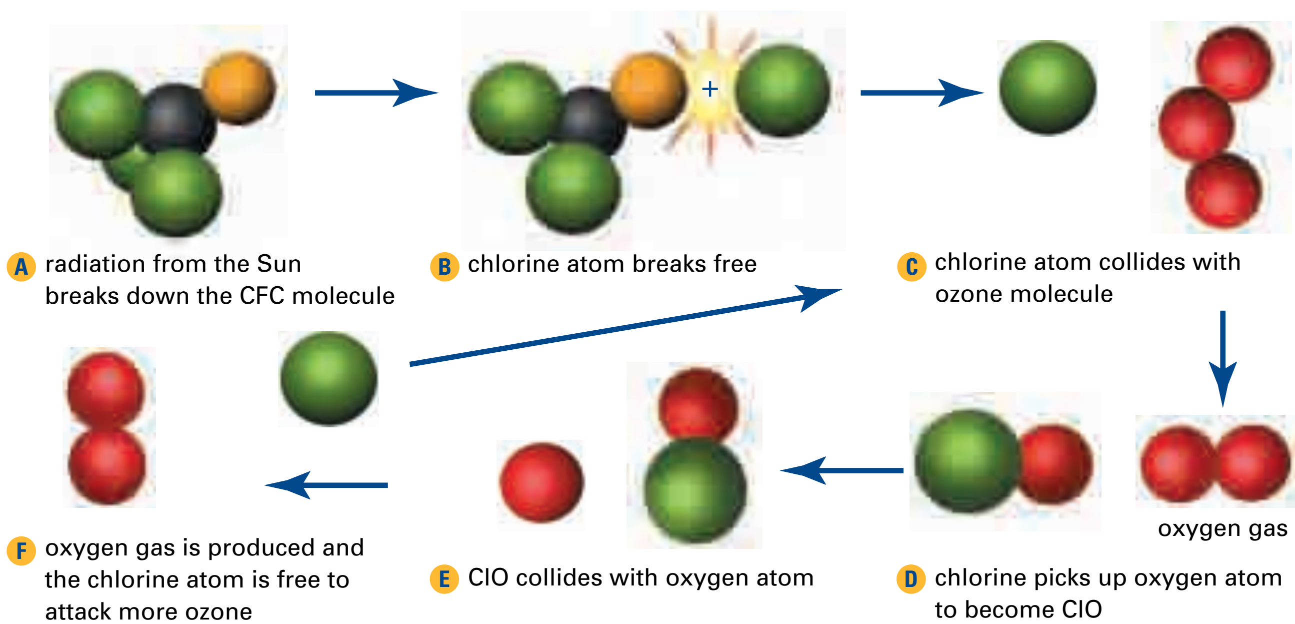


Figure 12.23 **The breakdown of a CFC by ultraviolet radiation produces single chlorine atoms that attack and destroy ozone. Because the chlorine atoms are released, they attack ozone again and again in a chain reaction.**

Although they are the most abundant ozone-depleting substance, CFCs are not the only culprits. Other chemicals that damage the ozone layer include methyl bromide, CH_3Br, carbon tetrachloride, CCl_4, and halons such as carbon trifluorobromide, CF_3Br.

The Effects of Ozone Depletion

How do we know that CFCs and other ozone-destroying molecules are having an effect on the atmosphere? As Figure 12.24 illustrates, ozone levels over Canada and other parts of the world have decreased significantly since the late 1970s.

With reduced ozone levels, more ultraviolet radiation from the Sun reaches Earth. Among humans, UV-induced skin cancer and eye damage are becoming a serious threat. The increased levels of radiation also damage phytoplankton in fresh and marine ecosystems. Since phytoplankton are the base of the aquatic food chain, this damage affects all other water species. As you learned earlier, the *presence* of ozone close to Earth damages crops and forests. A *lack* of ozone in the atmosphere, however, also reduces the yield of crops, such as barley and canola, and harms forests.

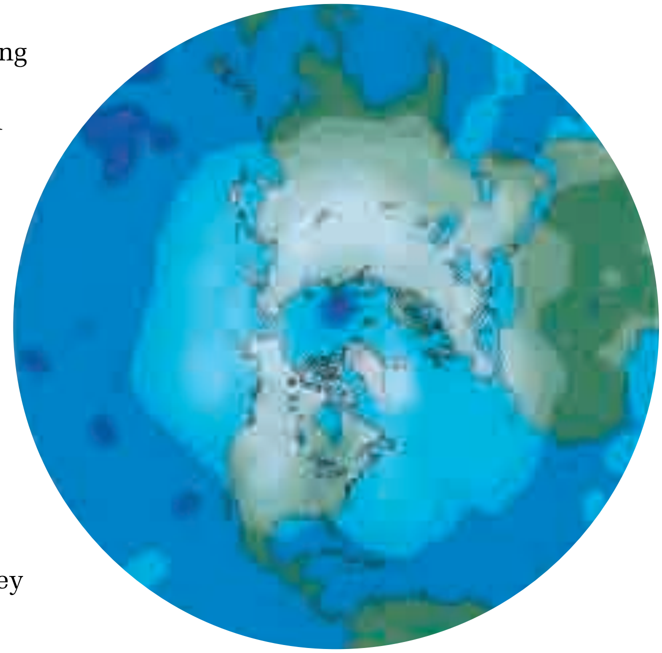


Figure 12.24 **Ozone thinning is still occurring over the northern hemisphere, as shown here. The dark blue in the centre indicates the presence of an ozone "hole."**

Improving Air Quality

What has happened since the discovery that CFCs and other chemicals were harming the ozone layer? The chemical industry has invented and produced environmentally friendly refrigerants. All new cars produced in North America have air conditioners that contain ozone-friendly refrigerants. Used CFCs from older refrigeration units are being recycled so that they do not escape into the atmosphere. A lot has already been accomplished, but more remains to be done.

Much of the positive activity comes from an international agreement called the **Montréal Protocol**. This agreement was signed in Montréal, Canada, on September 16, 1987. It has been successful in drastically reducing the use of CFCs worldwide. The Montréal Protocol is particularly significant because, for the first time, individual countries put a common planetary goal ahead of their own economic interests.

The Montréal Protocol stated that the production and consumption of all substances that deplete the ozone layer would be phased out by the year 2000 in developed countries. (Methyl chloroform would be phased out by 2005.) The chemicals that are named in the agreement include CFCs, halons, carbon tetrachloride, methyl chloroform, and methyl bromide. Once CFC production and consumption are stopped, scientists hope that the ozone layer will recover within 50 or 60 years. The success of the Montréal Protocol depends however, on the co-operation of both developed and developing countries.

CHEM FACT

In 1993, a scientific link was established between ozone depletion and increases in ultraviolet radiation. It was found that increased exposure to UV-B radiation causes skin cancer, the formation of cataracts, and the suppression of the human immune system. Research has shown skin cancer to be as common as all other types of cancer combined. Sunscreens can protect humans from the risk of some skin cancers. Unfortunately they do not appear to provide protection against damage to the immune system.

Section Wrap-up

In this section, you learned how ozone high in the atmosphere interacts with UV light from the Sun. In addition, you learned where CFCs come from, and how they damage the ozone layer. Finally, you saw that the Montréal Protocol is striving to prevent further ozone damage.

Canadians in Chemistry

Parisa Ariya was born in Tehran, the capital of Iran. She chose atmospheric chemistry for a career. Scientists in this field study the transformation of molecules in the atmosphere (the layer of gases surrounding Earth). They also study the atmosphere's interactions with oceans, land, and living things. After studying in several countries, Dr. Ariya became a professor at McGill University in Montréal.

Dr. Parisa Ariya

One of Dr. Ariya's particular areas of interest is halogen chemistry. Halogens such as chlorine, Cl, and bromine, Br, occur naturally in ocean waters. As well, they enter ocean waters as run-off from human activities. The oceans emit these halogens into the atmosphere. There they react with, and destroy, ozone, O_3.

Ariya and her students are trying to determine what kinds of halogens exist in Earth's atmosphere, how quickly they are produced and degraded, and what their major sources are. As they find answers to these questions, they may be able to recommend ways to reduce halogens in the oceans and atmosphere. They may also develop ways to modify halogens' reactions with ozone and other gases so that the halogens do less harm.

Another of Ariya's research interests is sulfur, S. She and her students are studying its atmospheric reactions with ozone and hydrogen peroxide, H_2O_2. Through field studies, laboratory experiments, and modelling, they are trying to determine the impacts of such reactions.

"You can enjoy nature through sports," says Dr. Ariya, who is an avid soccer player and swimmer. "But as a scientist you also enjoy nature intellectually, methodically. Science keeps the mind alive because you're constantly learning. Science is fun!"

Section Review

Unit Issue Prep

Research the Ontario Drive-Clean program. How is pollution from car exhausts being regulated by the government? Search for information in preparation for the Unit Issue.

1. C Describe the cycle that ozone goes through as it absorbs ultraviolet radiation. Use a diagram.
2. C Compare the effects of ozone near the ground with its effects high in the atmosphere.
3. K/U What are chlorofluorocarbons? How do they affect the ozone layer?
4. K/U Describe the Montréal Protocol. Why is it so significant?
5. MC Canadian scientists developed the Brewer Ozone Spectrophotometer, a state-of-the-art ozone-measuring device. It is the most accurate ozone-measuring device in the world. Use the Internet or reference books to find out how the Brewer Ozone Spectrophotometer works. Report your findings to the class.

CHAPTER 12 Review

Reflecting on Chapter 12

Summarize this chapter in the format of your choice. Here are a few ideas to use as guidelines:

- Describe Avogadro's hypothesis, and explain how it relates to the properties of gases.
- Explain how the ideal gas law came about. Describe how to use it to calculate gas properties.
- Explain how to use the ideal gas law for stoichiometric problems involving gases as reactants or products in a chemical reaction.
- Describe how to identify an unknown gas using the ideal gas law.
- Explain the importance of ozone, and describe the action of CFCs on ozone. Explain what the Montréal Protocol has done for the ozone issue.

Reviewing Key Terms

For each of the following terms, write a sentence that shows your understanding of its meaning.

law of combining volumes	molar volume
	ideal gas law
law of multiple proportions	chlorofluorocarbons (CFCs)
Avogadro's hypothesis	Montréal Protocol

Knowledge/Understanding

1. (a) How did the work of John Dalton and Joseph Gay-Lussac lead to Avogadro's hypothesis?
   (b) How does Avogadro's hypothesis help us to understand gas behaviour and gas reactions?
2. (a) What is the law of multiple proportions?
   (b) State Gay-Lussac's law of combining volumes.
   (c) How are the two laws related? What useful problem-solving gas law did they lead to?
3. What is the molar volume of an ideal gas at STP?
4. What are the characteristics of an ideal gas? How is it different from a real gas?
5. Is the density of one mole of hydrogen gas at STP the same as the density of one mole of oxygen gas at STP? Explain your answer.
6. (a) Additional gas is added to a container with a fixed volume. What happens to the pressure and temperature of the gas?
   (b) A balloon filled with gas experiences a drop in pressure. What happens to the volume of the gas?
7. A student reacts a sample of zinc with hydrochloric acid. He collects the gas as it bubbles through water. How will this student need to correct his measurements? Explain your answer.
8. What are the effects of ozone pollution near the ground?
9. Why are high levels of ultraviolet radiation potentially dangerous?
10. How is the Montréal Protocol attempting to improve future air quality?

Inquiry

11. As you work on gas reactions in the laboratory, your barometer shatters. Now it is impossible to measure the pressure of the room.
    (a) Design a simple investigation that allows you to calculate the pressure of the room. Assume you have a thermometer.
    (b) Identify the materials and equipment you will need to carry out the investigation.
    (c) Which variables will you hold constant? Which will you change? Which will you measure?
12. What volume does each amount of helium occupy at STP?
    (a) 1.00 mol
    (b) 12.5 mol
    (c) 100.0 g
13. How many molecules are in 0.250 m^3 of oxygen at STP?
14. What volume does 2.00 mol of oxygen occupy at 750 torr and 30.0°C?
15. What volume does 1.50 g of nitrogen gas occupy at 100.0°C and 5.00 atm?
16. Oxygen that is needed in a school laboratory is stored in a pressurized 2.00 L cylinder. 25.0 g of oxygen is contained in the cylinder at 20.0°C. Under what pressure is the oxygen stored?
17. Propane, C_3H_8, that is needed for a barbecue is stored in a 40.0 dm^3 metal cylinder. The pressure gauge on the cylinder reads 25.0 atm

at 20.0°C. What mass of propane is in the cylinder? (**Hint:** 1 dm^3 = 1 L)

18. Find the density (in g/L) of each atmospheric gas.
(a) oxygen at 1000 torr and 30.0°C
(b) helium at 10.0 atm and 20°C

19. **(a)** Pressurized CO_2 is used in the soft-drink manufacturing industry. How many grams of carbon dioxide are in a 500.0 cm^3 tank at –50.0°C and 2.00 atm?
(b) How many grams of oxygen does this tank hold at the same temperature and pressure?

20. 9.0 g of an unknown gas is stored in a 5.00 L metal tank at 0.0°C and 202.0 kPa. To identify the gas, investigators decide to find out its molecular mass.
(a) What is the molecular mass of the gas?
(b) What is the gas?

21. A 25.0 L tank, stored at –20.0°C, contains 10.0 g of helium and 10.0 g of hydrogen gas.
(a) What is the total number of moles of gas in the tank?
(b) What is the total pressure (in kPa) in the tank?
(c) What is the partial pressure of helium in the tank?

22. A 60.0 g sample of nitrogen gas is stored in a 5.0 L tank at a pressure of 10.0 atm. At what temperature (in °C) is the gas stored?

23. A 13.4 g sample of an unknown liquid is vapourized at 85.0°C and 100.0 kPa. The vapour has a volume of 4.32 L. The percentage composition of the liquid is found to be 52.1% carbon, 13.2% hydrogen, and 34.7% oxygen. What is the molecular formula of the liquid?

24. An unlabelled bottle of an unknown liquid is found on a shelf in a laboratory storeroom. 10.0 g of the liquid is vaporized at 120.0°C and 5.0 atm. The volume of the vapour is found to be 568.0 cm^3. The liquid is found to be made up of 84.2% carbon and 15.8% hydrogen. What is the molecular formula of the liquid?

25. A 4.2 g sample of a volatile liquid contains 1.0 g of carbon and 0.25 g of hydrogen. The rest of the liquid is chlorine. When the sample is vaporized at 101.0 kPa and 60.0°C, it occupies a volume of 2.2 L. What is the molecular formula of the liquid?

26. Methanol has potential to be used as an alternative fuel. It burns in the presence of oxygen to produce carbon dioxide and water.
$CH_3OH_{(\ell)} + O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$
(a) Balance this equation.
(b) 10 L of oxygen is completely consumed at STP. What volume of CO_2 is produced?
(c) What mass of methanol is consumed in this reaction?

27. A student wants to prepare carbon dioxide using sodium carbonate and dilute hydrochloric acid.
$Na_2CO_{3(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + CO_{2(g)} + H_2O_{(\ell)}$
How much sodium carbonate should the student react with excess hydrochloric acid to produce 1.00 L of carbon dioxide at 24.0°C and 760 torr?

28. A scientist makes hydrogen gas in the laboratory by reacting calcium metal with an excess of hydrochloric acid.
$Ca_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_{2(g)}$
The scientist reacts 5.00 g of calcium and collects the hydrogen over water at 25.0°C and 103.0 kPa. What volume of dry hydrogen is produced?

29. A chemist collects oxygen over water at 22.0°C and 105.0 kPa using the following reaction:
$2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$
What volume of dry oxygen is obtained if the chemist heats 25.0 g of potassium chlorate?

30. Ammonia, a useful fertilizer, is produced by the following reaction:
$CH_{4(g)} + H_2O_{(\ell)} + N_2O_{(g)} \rightarrow 2NH_{3(g)} + CO_{2(g)}$
500.0 g of methane reacts with excess H_2O and N_2O. At 27.0°C and 1.20 atm, what volume of ammonia gas is produced?

31. Hydrochloric acid dissolves limestone, as shown in the following chemical equation:
$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(\ell)}$
12.0 g of $CaCO_3$ reacts with 110 mL of 1.25 mol/L HCl. At 22.0°C and 99.0 kPa, what volume of carbon dioxide is produced?

32. Butane from a disposable lighter burns according to the following equation:

$C_4H_{10(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(\ell)}$

(a) Balance this chemical equation.

(b) How many grams of butane are needed to produce 300.0 mL of carbon dioxide at 50.0˚C and 1.25 atm? How many grams of oxygen are needed?

Communication

33. Draw a concept map showing the connections between the following gas laws:
- Boyle's law
- Charles' law
- Gay-Lussac's law
- the combined gas law
- the ideal gas law
- Gay-Lussac's law of combining volumes
- Dalton's law of partial pressures

34. Prepare a poster that will explain the relationships between the pressure, volume, and temperature of a gas to students in a younger grade.

35. Express the ideal gas constant in each of the following units.

(a) kPa·L/mol·K

(b) atm·L/mol·K

(c) torr·L/mol·K

Making Connections

36. Complex carbohydrates are starches that your body can convert to glucose, a type of sugar. Simple carbohydrate foods contain glucose, ready for immediate use by the human body. Breathing and burning glucose, $C_6H_{12}O_6$, produces energy in a jogger's muscles, according to the following unbalanced equation:

$C_6H_{12}O_{6(aq)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(g)}$

Just before running, Myri eats two oranges. The oranges give her body 25 g of glucose to make energy. The temperature outside is 27˚C, and the atmospheric pressure is 102.3 kPa. Although 21% of the air Myri breathes in is oxygen, she breathes out about 16% of this oxygen. (In other words, she only uses about 5%.)

(a) How many litres of air does Myri breathe while running to burn up the glucose she consumed?

(b) How many litres of water vapour does she produce? How many litres of carbon dioxide gas does she produce?

(c) Suppose that the water vapour Myri breathes out is condensed to its liquid form. If the density of the water is 1.0 g/mL, what is its volume?

37. Sulfur dioxide reacts with oxygen in air to produce sulfur trioxide. The sulfur trioxide then reacts with water to produce sulfuric acid. Write balanced chemical equations for these reactions.

38. Write a story to describe what your life would be like if you did not participate in any polluting activities (such as riding in petroleum-powered cars) or use any products (such as plastic) that cause pollution. Include as many pollution-causing products as possible.

Answers to Practice Problems and Short Answers to Section Review Questions:

Practice Problems: **1.** 24 L/mol **2.** 42.8 L/mol **3.** 25.9 L/mol **4.(a)** 0.446 mol **(b)** 12.5 g **5.** 0.089 mol **6.** 0.500 mol, 3.01×10^{23} molecules **7.** 77.3 L **8.** 1.0×10^2 L **9.** 78.38 L **10.(a)** 56. 0 L **(b)** 1.51×10^{24} molecules **(c)** 3.01×10^{24} atoms **11.** 74.4 L **12.** 45.7 K **13.** 626 kg **14.** 23.3 L **15.** 25 kPa **16.** 1.71 g/L **17.** 2.35 g/L **18.** 3.65 g/L **19.** 578 kPa **20.** 48.2 g/mol **21.** 2.0×10^2 g/mol **22.** 71.3 g/mol **23.(a)** 1.69 g/L **(b)** 1.78 g/L **24.** C_2H_2 **25.(a)** 1:2 **(b)** 1:2 **(c)** 2:1 **(d)** 1:1 **26.(a)** 4.5 L **(b)** 7.5 L **27.(b)** 6.00 L **28.(b)** a_4b_2 **29.(a)** 65.0 L **(b)** 1.63×10^{24} molecules **(c)** 3.25×10^{24} atoms **30.** 16 L **31.** 0.20 g **32.** 0.047 g **33.(a)** 79 L **(b)** 98 L **34.** 0.32 g **35.** 79 L **36.** 1.1×10^2 g **37.** 0.71 L **38.** 4.8×10^6 L **39.(a)** 0.201 mol **(b)** $FeCl_3$

Section Review: **12.1:** **1.** 104 g **2.** 8.7×10^2 L **3.** 2.5×10^2 kPa **4.** 58 g; 2.2×10^{23} molecules **5.** –19˚C **7.** 6.2 atm **12.2:** **1.** 0.72 g/L **2.(a)** 2.0 g/mol **(b)** hydrogen gas **3.** N_2O_4 **12.3:** **5.** 2.0 L **6.(b)** 1.7 L **7.** 41 L **8.(a)** 1.1×10^2 g **(b)** 0.075 atm