CHAPTER 14

Gases

What You'll Learn

- You will use gas laws to calculate how pressure, temperature, volume, and number of moles of a gas will change when one or more of these variables is altered.
- You will compare properties of real and ideal gases.
- You will apply the gas laws and Avogadro's principle to chemical equations.

Why It's Important

From barbecuing on a gas grill to taking a ride in a hot-air balloon, many activities involve gases. It is important to be able to predict what effect changes in pressure, temperature, volume, or amount, will have on the properties and behavior of a gas.

Chemistry Online

Visit the Glencoe Chemistry Web site at **chemistrymc.com** to find links about gases.

Firefighters breathe air that has been compressed into tanks that they can wear on their backs.

DISCOVERY LAB

More Than Just Hot Air

How does a temperature change affect the air in a balloon?

Safety Precautions

Always wear goggles to protect eyes from broken balloons.

Procedure

1. Inflate a round balloon and tie it closed.
2. Fill the bucket about half full of cold water and add ice.
3. Use a string to measure the circumference of the balloon.
4. Stir the water in the bucket to equalize the temperature. Submerge the balloon in the ice water for 15 minutes.
5. Remove the balloon from the water. Measure the circumference.

Materials

5-gal bucket
round balloon
ice
string

Analysis

What happens to the size of the balloon when its temperature is lowered? What might you expect to happen to its size if the temperature is raised?

Section 14.1 The Gas Laws

Objectives

- **State** Boyle's law, Charles's law, and Gay-Lussac's law.
- **Apply** the three gas laws to problems involving the pressure, temperature, and volume of a gas.

Vocabulary

Boyle's law
Charles's law
Gay-Lussac's law

The manufacturer of the air tank in the photo on the opposite page had to understand the nature of the gases the tank contains. Understanding gases did not happen accidentally. The work of many scientists over many years has contributed to our present knowledge of the nature of gases. The work of three scientists in particular was valuable enough that laws describing gas behavior were named in their honor. In this section, you'll study three important gas laws: Boyle's law, Charles's law, and Gay-Lussac's law. Each of these laws relates two of the variables that determine the behavior of gases—pressure, temperature, volume, and amount of gas present.

Kinetic Theory

You can't understand gases without understanding the movement of gas particles. Remember from your study of the kinetic-molecular theory in Chapter 13 that gas particles behave differently than those of liquids and solids. The kinetic theory provides a model that is used to explain the properties of solids, liquids, and gases in terms of particles that are always in motion and the forces that exist between them. The kinetic theory assumes the following concepts about gases are true.

- *Gas particles do not attract or repel each other.* Gases are free to move within their containers without interference from other particles.
- *Gas particles are much smaller than the distances between them.* You saw in the **DISCOVERY LAB** that gas has volume. However, the kinetic theory assumes that gas particles themselves have virtually no volume.

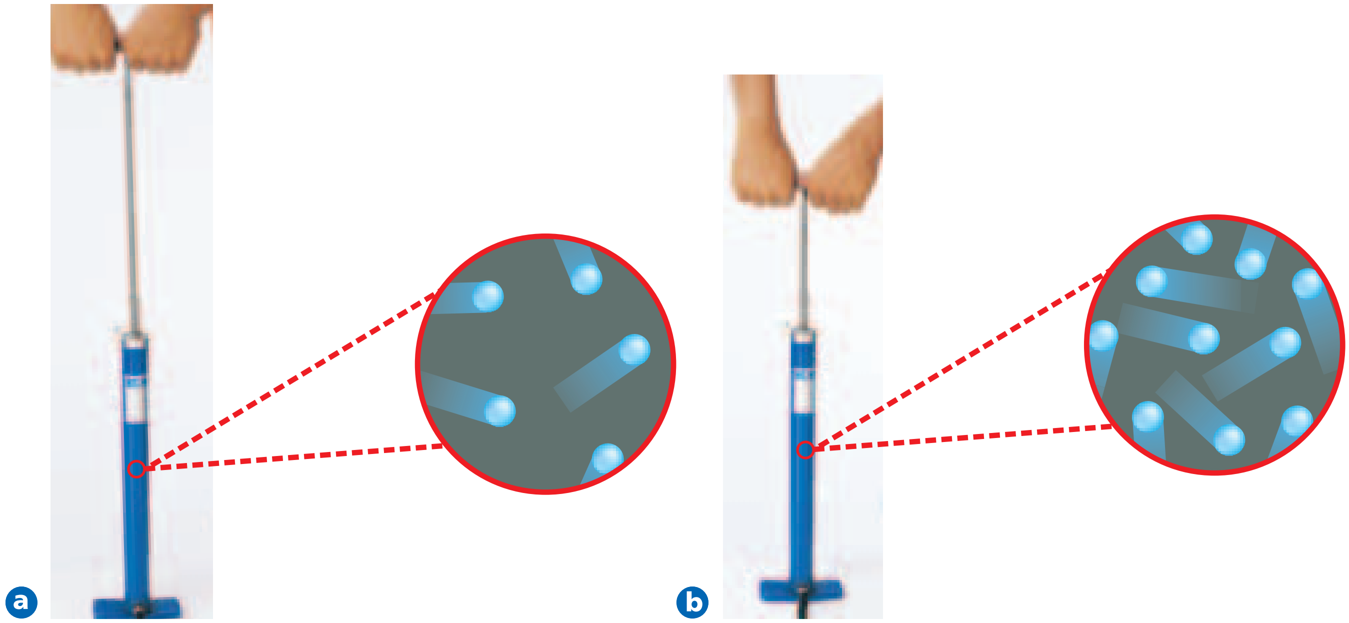


Figure 14-1

The kinetic theory relates pressure and the number of collisions per unit time for a gas. **a** When the bicycle pump is pulled out as far as it will go, the pressure of the air inside the pump equals that of the atmosphere. **b** If the piston is pushed down half the length of the pump, the air particles are squeezed into a space half the original size. Pressure doubles because the frequency of collisions between the gas particles and the inner wall of the pump has doubled.

Try at Home LAB

See page 958 in Appendix E for
Under Pressure

Almost all the volume of a gas is empty space. Gases can be compressed by moving gas particles closer together because of this low density of particles.

- *Gas particles are in constant, random motion.* Gas particles spread out and mix with each other because of this motion. The particles move in straight lines until they collide with each other or with the walls of their container.
- *No kinetic energy is lost when gas particles collide with each other or with the walls of their container.* Such collisions are completely elastic. As long as the temperature stays the same, the total kinetic energy of the system remains constant.
- *All gases have the same average kinetic energy at a given temperature.* As temperature increases, the total energy of the gas system increases. As temperature decreases, the total energy of the gas system decreases.

The nature of gases Actual gases don't obey all the assumptions made by the kinetic theory. But for many gases, their behavior approximates the behavior assumed by the kinetic theory. You will learn more about real gases and how they vary from these assumptions in Section 14.3.

Notice how all the assumptions of the kinetic theory are based on the four factors previously mentioned—the number of gas particles present and the temperature, the pressure, and the volume of the gas sample. These four variables all work together to determine the behavior of gases. When one variable changes, it affects the other three. Look at the following example of how a change in one variable affects at least one other variable.

What happens to the gas in a plastic balloon if you squeeze it, decreasing its volume? Because the balloon is closed, the amount of gas is constant. Assume the temperature is held constant. Decreasing the volume pushes the gas particles closer together. Recall from the kinetic-molecular theory that as gas particles are pushed closer together, the number of collisions between particles themselves and between the particles and the walls of their container increases. As the number of collisions per unit time increases, so does the observed pressure. Therefore, as the volume of a gas decreases, its pressure increases. Similarly, if the balloon is no longer squeezed, the volume increases and the pressure decreases. You can see another example of this principle in **Figure 14-1**. The interdependence of the variables of volume, pressure, temperature, and amount of gas is the basis for the following gas laws.

Boyle's Law

Robert Boyle (1627–1691), an Irish chemist, did experiments like the one shown in **Figure 14-2** to study the relationship between the pressure and the volume of a gas. By taking careful quantitative measurements, he showed that if the temperature is constant, doubling the pressure of a fixed amount of gas decreases its volume by one-half. On the other hand, reducing the pressure by half results in a doubling of the volume. A relationship in which one variable increases as the other variable decreases is referred to as an inversely proportional relationship. For help with understanding inverse relationships, see the **Math Handbook** page 905.

Boyle's law states that the volume of a given amount of gas held at a constant temperature varies inversely with the pressure. Look at the graph in **Figure 14-2** in which pressure versus volume is plotted for a gas. The plot of an inversely proportional relationship results in a downward curve. If you choose any two points along the curve and multiply the pressure times the volume at each point, how do your two answers compare? Note that the product of the pressure and the volume for each of points 1, 2, and 3 is 10 atm·L. From the graph, what would the volume be if the pressure is 2.5 atm? What would the pressure be if the volume is 2 L?

The products of pressure times volume for any two sets of conditions are equal, so Boyle's law can be expressed mathematically as follows.

$$P_1V_1 = P_2V_2$$

P_1 and V_1 represent a set of initial conditions for a gas and P_2 and V_2 represent a set of new conditions. If you know any three of these four values for a gas at constant temperature, you can solve for the fourth by rearranging the equation. For example, if P_1, V_1, and P_2 are known, dividing both sides of the equation by P_2 will isolate the unknown variable V_2.

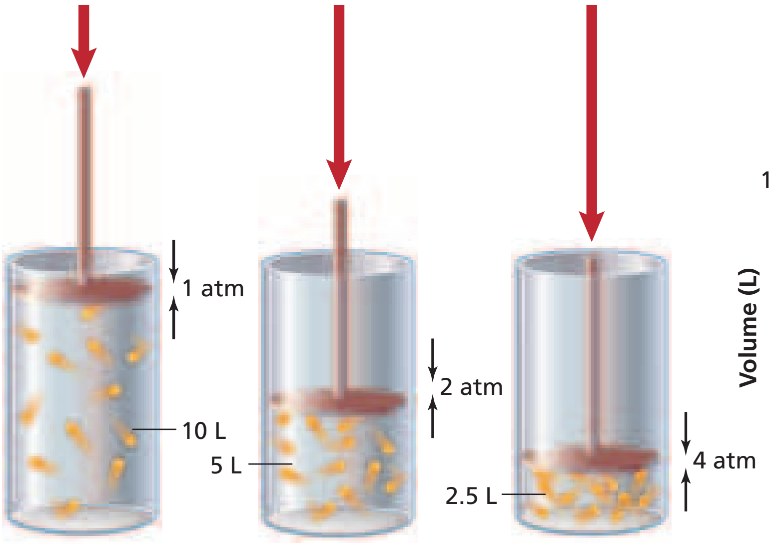
Use the equation for Boyle's law to calculate the volume that corresponds to a pressure of 2.5 atm, assuming that the amount of gas and temperature are constant. Then find what pressure corresponds to a volume of 2.0 L. Use 2.0 atm for P_1 and 5 L for V_1. How do these answers compare to those you found using the graph in **Figure 14-2?**

Careers Using Chemistry

Meteorologist

Would you like to be able to plan your days better because you know what the weather will be? Then consider a career as a meteorologist.

Most weather is caused by the interaction of energy and air. Meteorologists study how these interactions affect and are affected by changes in temperature and pressure of the air. For example, winds and fronts are direct results of pressure changes caused by uneven heating of Earth's atmosphere by the sun.



Condition 1:

$k = P_1V_1$
$k = (1\text{ atm})(10\text{ L})$
$k = 10\text{ atm•L}$

Condition 2:

$k = P_2V_2$
$k = (2\text{ atm})(5\text{ L})$
$k = 10\text{ atm•L}$

Condition 3:

$k = P_3V_3$
$k = (4\text{ atm})(2.5\text{ L})$
$k = 10\text{ atm•L}$

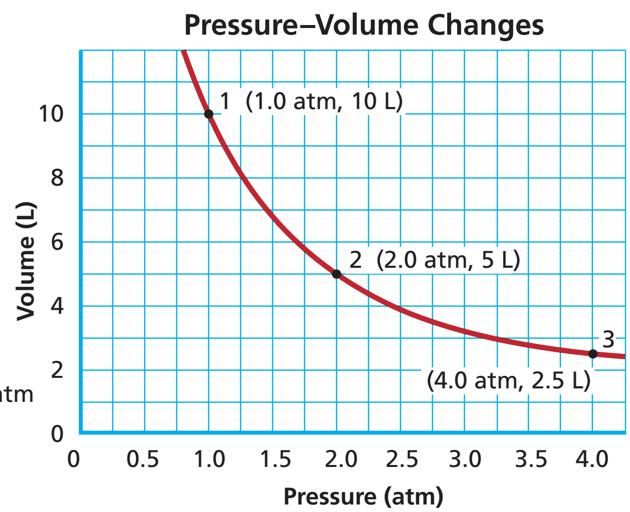


Figure 14-2

The gas particles in this cylinder take up a given volume at a given pressure. As pressure increases, volume decreases. The graph shows that pressure and volume have an inverse relationship, which means that as pressure increases, volume decreases. This relationship is illustrated by a downward curve in the line from condition 1 to condition 2 to condition 3.

Math Handbook

Review using inverse relationships in the **Math Handbook** on page 905 of your textbook.

EXAMPLE PROBLEM 14-1

Boyle's Law

A sample of helium gas in a balloon is compressed from 4.0 L to 2.5 L at a constant temperature. If the pressure of the gas in the 4.0-L volume is 210 kPa, what will the pressure be at 2.5 L?

1. Analyze the Problem

You are given the initial and final volumes and the initial pressure of a sample of helium. Boyle's law states that as volume decreases, pressure increases if temperature remains constant. Because the volume in this problem is decreasing, the pressure will increase. So the initial pressure should be multiplied by a volume ratio greater than one.

Known	Unknown
$V_1 = 4.0$ L	$P_2 = ?$ kPa
$V_2 = 2.5$ L	
$P_1 = 210$ kPa	

2. Solve for the Unknown

Divide both sides of the equation for Boyle's law by V_2 to solve for P_2.

$$P_1V_1 = P_2V_2$$

$$P_2 = P_1\left(\frac{V_1}{V_2}\right)$$

Substitute the known values into the rearranged equation.

$$P_2 = 210\text{ kPa}\left(\frac{4.0\text{ L}}{2.5\text{ L}}\right)$$

Multiply and divide numbers and units to solve for P_2.

$$P_2 = 210\text{ kPa}\left(\frac{4.0\cancel{\text{L}}}{2.5\cancel{\text{L}}}\right) = 340\text{ kPa}$$

3. Evaluate the Answer

When the volume is decreased by almost half, the pressure is expected to almost double. The calculated value of 340 kPa is reasonable. The unit in the answer is kPa, a pressure unit.

Practice!

For more practice with Boyle's law problems, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Assume that the temperature and the amount of gas present are constant in the following problems.

1. The volume of a gas at 99.0 kPa is 300.0 mL. If the pressure is increased to 188 kPa, what will be the new volume?
2. The pressure of a sample of helium in a 1.00-L container is 0.988 atm. What is the new pressure if the sample is placed in a 2.00-L container?
3. Air trapped in a cylinder fitted with a piston occupies 145.7 mL at 1.08 atm pressure. What is the new volume of air when the pressure is increased to 1.43 atm by applying force to the piston?
4. If it takes 0.0500 L of oxygen gas kept in a cylinder under pressure to fill an evacuated 4.00-L reaction vessel in which the pressure is 0.980 atm, what was the initial pressure of the gas in the cylinder?
5. A sample of neon gas occupies 0.220 L at 0.860 atm. What will be its volume at 29.2 kPa pressure?

Charles's Law

Have you ever noticed that on a cold day, a tire on a car might look as if it's low on air? However, after driving the car for awhile, the tire warms up and looks less flat. What made the difference in the tire? When canning vegetables at home, why are they often packed hot in the jars and then sealed? These questions can be answered by applying another of the gas laws, Charles's law. Another example of how gases are affected by temperature is shown in the **problem-solving LAB.** If kelvin temperature is doubled, so is the volume.

How are gas temperature and volume related? The French physicist Jacques Charles (1746–1823) studied the relationship between volume and temperature. In his experiments, he observed that as temperature increases, so does the volume of a gas sample when the pressure is held constant. This property can be explained by the kinetic-molecular theory; at a higher temperature, gas particles move faster, striking each other and the walls of their container more frequently and with greater force. For the pressure to stay constant, volume must increase so that the particles have farther to travel before striking the walls. Having to travel farther decreases the frequency with which the particles strike the walls of the container.

Look at **Figure 14-3**, which includes a graph of volume versus temperature for a gas sample kept at a constant pressure. Note that the resulting plot is a straight line. Note also that you can predict the temperature at which the volume will reach a value of zero liters by extrapolating the line at temperatures below which values were actually measured. The temperature that corresponds to zero volume is −273.15°C, or 0 on the kelvin (K) temperature scale. This temperature is referred to as absolute zero, and it is the lowest possible theoretical temperature. Theoretically, at absolute zero, the kinetic energy of particles is zero, so all motion of gas particles at that point ceases.

Examine the relationship between temperature and volume shown by the cylinders in **Figure 14-3**. In the graph, note that 0°C does not correspond to zero volume. Although the relationship is linear, it is not direct. For example, you can see from the graph that increasing the temperature from 25°C to 50°C does not double the volume of the gas. If the kelvin temperature is plotted instead, a direct proportion is the result. See **Figure 14-4** on the next page.

History CONNECTION

Jacques Charles's interest in the behavior of gases was sparked by his involvement in ballooning. He built the first balloon that was filled with hydrogen instead of hot air. Charles's investigations attracted the attention of King Louis XVI, who allowed Charles to establish a laboratory in the Louvre, which is a museum in Paris.

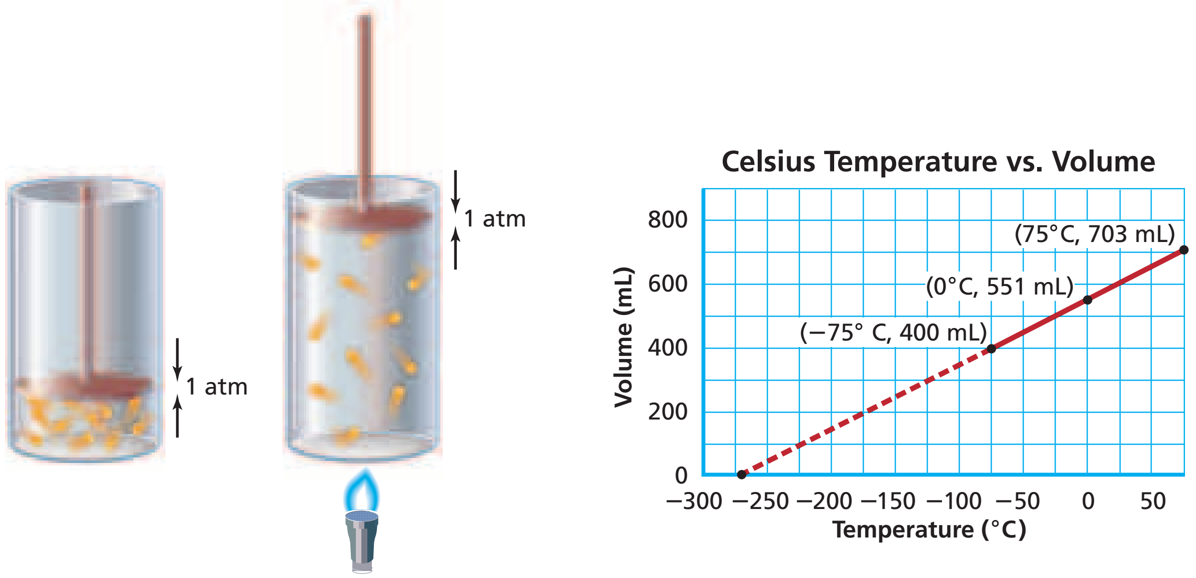


Figure 14-3

The gas particles in this cylinder take up a given volume at a given temperature. When the cylinder is heated, the kinetic energy of the particles increases. The volume of the gas increases, pushing the piston outward. Thus the distance that the piston moves is a measure of the increase in volume of the gas as it is heated. Note that the graph of volume versus temperature extrapolates to −273.15°C, or 0 K.

CONTENTS

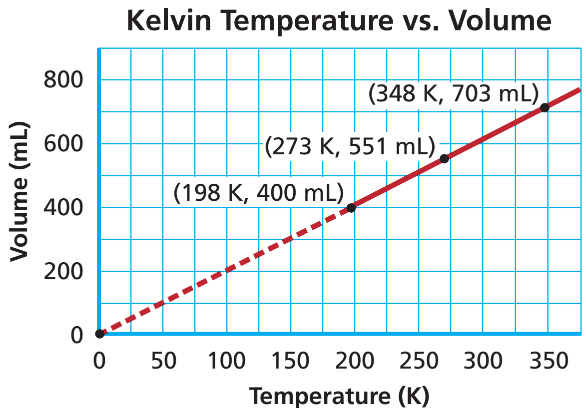


Figure 14-4

This graph illustrates the directly proportional relationship between the volume and the kelvin temperature of a gas held at constant pressure. When the kelvin temperature doubles, the volume doubles.

Charles's law states that the volume of a given mass of gas is directly proportional to its kelvin temperature at constant pressure. For help with understanding direct relationships, see the **Math Handbook,** page 905. So for any two sets of conditions, Charles's law can be expressed as

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

Here V_1 and T_1 represent any initial pair of conditions, while V_2 and T_2 are any new set of conditions. As with Boyle's law, if you know any three of the four values, you can calculate the fourth using the equation.

The temperature must be expressed in kelvin units when using the equation for Charles's law. The kelvin scale starts at absolute zero, which corresponds to –273.15°C and is 0 K. Because a Celsius degree and a kelvin unit are the same size, it is easy to convert a temperature in Celsius to kelvin units. Round 273.15 to 273, and add it to the Celsius temperature.

$$T_K = 273 + T_C$$

problem-solving LAB

How is turbocharging in a car engine maximized?

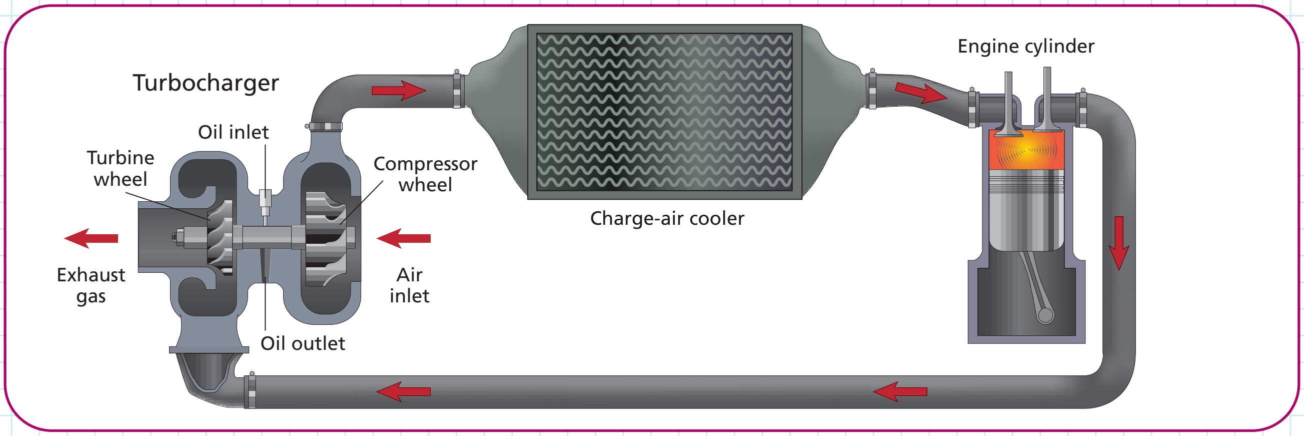
Interpreting Scientific Illustrations After gasoline and air are burned in the combustion chamber of an automobile, the resulting hot gases are exhausted out the tailpipe. The horsepower of an automobile engine can be significantly improved if the energy of these exhaust gases is used to operate a compressor that forces additional air into the combustion chamber. Outside air is then blown over this compressed air to cool it before it enters the engine. Increasing the power of an engine in this manner is known as turbocharging.

Analysis

Examine the illustration of an engine fitted with a turbocharging system. The paths of the exhaust gas, entering combustion air, and the cooling air are shown.

Thinking Critically

1. What property of the exhaust gas is being used to turn the turbine that runs the compressor? Explain.
2. If more power is to be gained from this design, what must also accompany the extra supply of oxygen to the combustion chamber?
3. What property does the compressor alter so that more air can be injected into the combustion chamber? Explain.
4. Why does the air in the compressor get hot, and why does cooling help to improve the power of the engine?



EXAMPLE PROBLEM 14-2

Charles's Law

A gas sample at 40.0°C occupies a volume of 2.32 L. If the temperature is raised to 75.0°C, what will the volume be, assuming the pressure remains constant?

1. Analyze the Problem

You are given the initial temperature and volume of a sample of gas. Charles's law states that as the temperature increases, so does the volume, assuming the pressure is constant. Because the temperature in this problem is increasing, the volume will increase. So the initial volume should be multiplied by a volume ratio greater than one.

Known	Unknown
$T_1 = 40.0°C$	$V_2 = ?\ L$
$V_1 = 2.32\ L$	
$T_2 = 75.0°C$	

2. Solve for the Unknown

Add 273 to the Celsius temperature to obtain the kelvin temperature.

$$T_K = 273 + T_C$$

Substitute the known Celsius temperatures for T_1 and T_2 to convert them to kelvin units.

$$T_1 = 273 + 40.0°C = 313\ K$$

$$T_2 = 273 + 75.0°C = 348\ K$$

Multiply both sides of the equation for Charles's law by T_2 to solve for V_2.

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

$$V_2 = V_1\left(\frac{T_2}{T_1}\right)$$

Substitute the known values into the rearranged equation.

$$V_2 = 2.32\ L\left(\frac{348\ K}{313\ K}\right)$$

Multiply and divide numbers and units to solve for V_2.

$$V_2 = 2.32\ L\left(\frac{348\ \cancel{K}}{313\ \cancel{K}}\right) = 2.58\ L$$

3. Evaluate the Answer

The increase in kelvin units is relatively small. Therefore, you expect the volume to show a small increase, which agrees with the answer. The unit of the answer is liters, a volume unit.

Math Handbook

Review using direct relationships in the **Math Handbook** on page 905 of this textbook.

PRACTICE PROBLEMS

Assume that the pressure and the amount of gas present remain constant in the following problems.

6. A gas at 89°C occupies a volume of 0.67 L. At what Celsius temperature will the volume increase to 1.12 L?

7. The Celsius temperature of a 3.00-L sample of gas is lowered from 80.0°C to 30.0°C. What will be the resulting volume of this gas?

8. What is the volume of the air in a balloon that occupies 0.620 L at 25°C if the temperature is lowered to 0.00°C?

Practice!

For more practice with Charles's law problems, go to **Supplemental Practice Problems** in Appendix A.

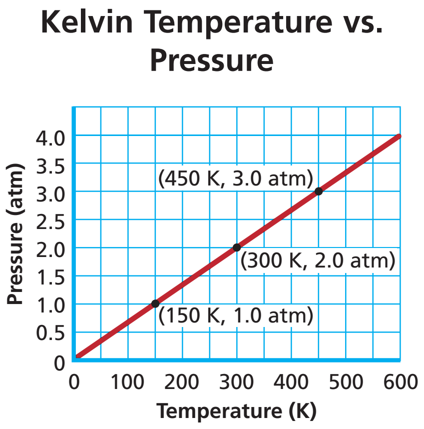


Figure 14-5

Compare the relationship between pressure and kelvin temperature as shown in this graph and the relationship between volume and kelvin temperature as shown in **Figure 14-4**. Notice that both show direct relationships.

Gay-Lussac's Law

Boyle's law relates pressure and volume of a gas, and Charles's law states the relationship between a gas's temperature and volume. What is the relationship between pressure and temperature? Pressure is a result of collisions between gas particles and the walls of their container. An increase in temperature increases collision frequency and energy, so raising the temperature should also raise the pressure if the volume is not changed.

How are temperature and pressure of a gas related? Joseph Gay-Lussac explored the relationship between temperature and pressure of a contained gas at a fixed volume. He found that a direct proportion exists between the kelvin temperature and the pressure, such as that illustrated in **Figure 14-5**. **Gay-Lussac's law** states that the pressure of a given mass of gas varies directly with the kelvin temperature when the volume remains constant. It can be expressed mathematically.

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

As with Boyle's and Charles's laws, if you know any three of the four variables, you can calculate the fourth using this equation. Remember that temperature must be in kelvin units whenever it is used in a gas law equation.

EXAMPLE PROBLEM 14-3

Gay-Lussac's Law

The pressure of a gas in a tank is 3.20 atm at 22.0°C. If the temperature rises to 60.0°C, what will be the gas pressure in the tank?

1. Analyze the Problem

You are given the initial pressure and the initial and final temperatures of a gas sample. Gay-Lussac's law states that if the temperature of a gas increases, so does its pressure. Because the temperature in this problem is increasing, the pressure will increase. So the initial pressure should be multiplied by a volume ratio greater than one.

Known

$P_1 = 3.20$ atm
$T_1 = 22.0°C$
$T_2 = 60.0°C$

Unknown

$P_2 = ?$ atm

2. Solve for the Unknown

Add 273 to the Celsius temperature to obtain the kelvin temperature.

$T_K = 273 + T_C$

Substitute the known Celsius temperatures for T_1 and T_2 to convert them to kelvin units.

$T_1 = 273 + 22.0°C = 295$ K

$T_2 = 273 + 60.0°C = 333$ K

Multiply both sides of the equation for Gay-Lussac's law by T_2 to solve for P_2.

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

$$P_2 = P_1\left(\frac{T_2}{T_1}\right)$$

This cooker is sealed so that the volume is constant. Pressure increases in the cooker as temperature increases.

Substitute the known values into the rearranged equation.

$$P_2 = 3.20\ \text{atm}\left(\frac{333\ \text{K}}{295\ \text{K}}\right)$$

Multiply and divide numbers and units to solve for P_2.

$$P_2 = 3.20\ \text{atm}\left(\frac{333\ \cancel{\text{K}}}{295\ \cancel{\text{K}}}\right) = 3.61\ \text{atm}$$

3. Evaluate the Answer

Gay-Lussac's law states that pressure and temperature are directly proportional. Kelvin temperature shows a small increase, so you expect the pressure to show a small increase, which agrees with the answer calculated. The unit is atm, a pressure unit.

PRACTICE PROBLEMS

Assume that the volume and the amount of gas are constant in the following problems.

9. A gas in a sealed container has a pressure of 125 kPa at a temperature of 30.0°C. If the pressure in the container is increased to 201 kPa, what is the new temperature?

10. The pressure in an automobile tire is 1.88 atm at 25.0°C. What will be the pressure if the temperature warms up to 37.0°C?

11. Helium gas in a 2.00-L cylinder is under 1.12 atm pressure. At 36.5°C, that same gas sample has a pressure of 2.56 atm. What was the initial temperature of the gas in the cylinder?

12. If a gas sample has a pressure of 30.7 kPa at 0.00°C, by how much does the temperature have to decrease to lower the pressure to 28.4 kPa?

13. A rigid plastic container holds 1.00 L methane gas at 660 torr pressure when the temperature is 22.0°C. How much more pressure will the gas exert if the temperature is raised to 44.6°C?

Practice!

For more practice with Gay-Lussac's law problems, go to **Supplemental Practice Problems** in Appendix A.

You have seen how the variables of temperature, pressure, and volume affect a gas sample. The gas laws covered in this section each relate two of these three variables if the other variable remains constant. What happens when all three of these variables change? You'll investigate this situation in the next section.

Section 14.1 Assessment

14. State Boyle's, Charles's, and Gay-Lussac's laws using sentences, then equations.

15. A weather balloon of known initial volume is released. The air pressures at its initial and final altitudes are known. Why can't you find its new volume by using these known values and Boyle's law?

16. Which of the three variables that apply to equal amounts of gases are directly proportional? Which are inversely proportional?

17. **Thinking Critically** Explain why gases such as the oxygen found in tanks used at hospitals are compressed. Why must care be taken to prevent compressed gases from reaching a high temperature?

18. **Concept Mapping** Draw a concept map that shows the relationship among pressure, volume, and temperature variables for gases and Boyle's, Charles's, and Gay-Lussac's laws.

Section 14.2

The Combined Gas Law and Avogadro's Principle

Español

Objectives

- **State** the relationship among temperature, volume, and pressure as the combined gas law.
- **Apply** the combined gas law to problems involving the pressure, temperature, and volume of a gas.
- **Relate** numbers of particles and volumes by using Avogadro's principle.

Vocabulary

combined gas law
Avogadro's principle
molar volume

In the previous section, you applied the three gas laws covered so far to problems in which either pressure, volume, or temperature of a gas sample was held constant as the other two changed. As illustrated in **Figure 14-6**, in a number of applications involving gases, all three variables change. If all three variables change, can you calculate what their new values will be? In this section you will see that Boyle's, Charles's, and Gay-Lussac's laws can be combined into a single equation that can be used for just that purpose.

The Combined Gas Law

Boyle's, Charles's, and Gay-Lussac's laws can be combined into a single law. This **combined gas law** states the relationship among pressure, volume, and temperature of a fixed amount of gas. All three variables have the same relationship to each other as they have in the other gas laws: Pressure is inversely proportional to volume and directly proportional to temperature, and volume is directly proportional to temperature. The equation for the combined gas law can be expressed as

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

As with the other gas laws, this equation allows you to use known values for the variables under one set of conditions to find a value for a missing variable under another set of conditions. Whenever five of the six values from the two sets of conditions are known, the sixth can be calculated using this expression for the combined gas law.

This combined law lets you work out problems involving more variables that change, and it also provides a way for you to remember the other three laws without memorizing each equation. If you can write out the combined gas law equation, equations for the other laws can be derived from it by remembering which variable is held constant in each case.

For example, if temperature remains constant as pressure and volume vary, then $T_1 = T_2$. After simplifying the combined gas law under these conditions, you are left with

$$P_1V_1 = P_2V_2$$

You should recognize this equation as the equation for Boyle's law. See whether you can derive Charles's and Gay-Lussac's laws from the combined gas law.

Figure 14-6

Constructing an apparatus that uses gases must take into account the changes in gas variables such as pressure, volume, and temperature that can take place. As a hot-air balloonist ascends in the sky, pressure and temperature both decrease, and the volume of the gas in the balloon is affected by those changes.

EXAMPLE PROBLEM 14-4

The Combined Gas Law

A gas at 110 kPa and 30.0°C fills a flexible container with an initial volume of 2.00 L. If the temperature is raised to 80.0°C and the pressure increased to 440 kPa, what is the new volume?

1. Analyze the Problem

You are given the initial pressure, temperature, and volume of a gas sample as well as the final pressure and temperature. The volume of a gas is directly proportional to kelvin temperature, so volume increases as temperature increases. Therefore the volume should be multiplied by a temperature factor greater than one. Volume is inversely proportional to pressure, so as pressure increases, volume decreases. Therefore the volume should be multiplied by a pressure factor that is less than one.

Known

P_1 = 110 kPa
T_1 = 30.0°C
V_1 = 2.00 L
T_2 = 80.0°C
P_2 = 440 kPa

Unknown

V_2 = ? L

2. Solve for the Unknown

Add 273 to the Celsius temperature to obtain the kelvin temperature.

$$T_K = 273 + T_C$$

Substitute the known Celsius temperatures for T_1 and T_2 to convert them to kelvin units.

$$T_1 = 273 + 30.0°\text{C} = 303 \text{ K}$$

$$T_2 = 273 + 80.0°\text{C} = 353 \text{ K}$$

Multiply both sides of the equation for the combined gas law by T_2 and divide it by P_2 to solve for V_2.

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

$$V_2 = V_1\left(\frac{P_1}{P_2}\right)\left(\frac{T_2}{T_1}\right)$$

Substitute the known values into the rearranged equation.

$$V_2 = 2.00 \text{ L}\left(\frac{110 \text{ kPa}}{440 \text{ kPa}}\right)\left(\frac{353 \text{ K}}{303 \text{ K}}\right)$$

Multiply and divide numbers and units to solve for V_2.

$$V_2 = 2.00 \text{ L}\left(\frac{110 \cancel{\text{kPa}}}{440 \cancel{\text{kPa}}}\right)\left(\frac{353 \cancel{\text{K}}}{303 \cancel{\text{K}}}\right) = 0.58 \text{ L}$$

3. Evaluate the Answer

Increasing the temperature causes the volume to increase, but increasing the pressure causes the volume to decrease. Because the pressure change is much greater than the temperature change, the volume undergoes a net decrease. The calculated answer agrees with this. The unit is L, a volume unit.

Math Handbook

Review rearranging algebraic equations in the **Math Handbook** on page 897 of this textbook.

For more practice with combined gas law problems, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

Assume that the amount of gas is constant in the following problems.

19. A helium-filled balloon at sea level has a volume of 2.1 L at 0.998 atm and 36°C. If it is released and rises to an elevation at which the pressure is 0.900 atm and the temperature is 28°C, what will be the new volume of the balloon?

20. At 0.00°C and 1.00 atm pressure, a sample of gas occupies 30.0 mL. If the temperature is increased to 30.0°C and the entire gas sample is transferred to a 20.0-mL container, what will be the gas pressure inside the container?

21. A sample of air in a syringe exerts a pressure of 1.02 atm at a temperature of 22.0°C. The syringe is placed in a boiling water bath at 100.0°C. The pressure of the air is increased to 1.23 atm by pushing the plunger in, which reduces the volume to 0.224 mL. What was the original volume of the air?

22. An unopened, cold 2.00-L bottle of soda contains 46.0 mL of gas confined at a pressure of 1.30 atm at a temperature of 5.0°C. If the bottle is dropped into a lake and sinks to a depth at which the pressure is 1.52 atm and the temperature is 2.09°C, what will be the volume of gas in the bottle?

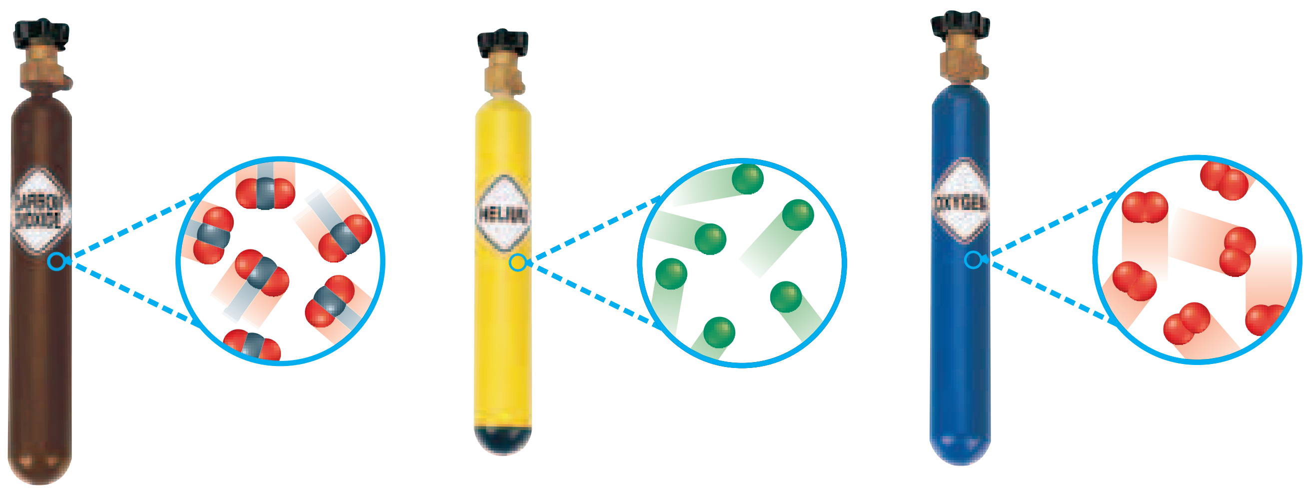
23. A sample of gas of unknown pressure occupies 0.766 L at a temperature of 298 K. The same sample of gas is then tested under known conditions and has a pressure of 32.6 kPa and occupies 0.644 L at 303 K. What was the original pressure of the gas?

Avogadro's Principle

The particles making up different gases can vary greatly in size. However, according to the kinetic-molecular theory, the particles in a gas sample are usually far enough apart that size has a negligible influence on the volume occupied by a fixed number of particles, as shown in **Figure 14-7**. For example, 1000 relatively large krypton gas particles occupy the same volume as 1000 much smaller helium gas particles at the same temperature and pressure. It was Avogadro who first proposed this idea in 1811. Today, it is known as **Avogadro's principle**, which states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles.

Figure 14-7

Compressed gas tanks of equal volume that are at the same pressure and temperature contain equal numbers of gas particles, regardless of which gas they contain. Refer to **Table C-1** in Appendix C for a key to atom color conventions.



CONTENTS

Remember from Chapter 11 that the most convenient unit for counting numbers of atoms or molecules is the mole. One mole contains 6.02×10^{23} particles. The **molar volume** for a gas is the volume that one mole occupies at 0.00°C and 1.00 atm pressure. These conditions of temperature and pressure are known as standard temperature and pressure (STP). Avogadro showed experimentally that one mole of any gas will occupy a volume of 22.4 L at STP. The fact that this value is the same for all gases greatly simplifies many gas law calculations. Because the volume of one mole of a gas at STP is 22.4 L, you can use the following conversion factor to find the number of moles, the mass, and even the number of particles in a gas sample.

$$\text{Conversion factor: } \frac{22.4 \text{ L}}{1 \text{ mol}}$$

Suppose you want to find the number of particles in a sample of gas that has a volume of 3.72 L at STP. First, find the number of moles of gas in the sample.

$$3.72 \cancel{\text{L}} \times \frac{1 \text{ mol}}{22.4 \cancel{\text{L}}} = 0.166 \text{ mol}$$

A mole of gas contains 6.02×10^{23} gas particles. Use this definition to convert the number of moles to the number of particles.

$$0.166 \cancel{\text{mol}} \times \frac{6.02 \times 10^{23} \text{ particles}}{1 \cancel{\text{mol}}} = 9.99 \times 10^{22} \text{ particles}$$

The following example problems show you how to use molar volume in other ways.

EXAMPLE PROBLEM 14-5

Avogadro's Principle—Using Moles

Calculate the volume that 0.881 mol of gas at standard temperature and pressure (STP) will occupy.

1. Analyze the Problem

You are given the temperature and pressure of a gas sample and the amount of gas the sample contains. According to Avogadro's principle, 1 mol of gas occupies 22.4 L at STP. The number of moles of gas should be multiplied by the conversion factor $\frac{22.4 \text{ L}}{1 \text{ mol}}$ to find the volume.

Known	Unknown
n = 0.881 moles	V = ? L
T = 0.00°C	
P = 1.00 atm	

2. Solve for the Unknown

Because conditions are already at STP, multiply the known number of moles by the conversion factor that relates liters to moles to solve for the unknown volume.

$$V = 0.881 \cancel{\text{mol}} \left(\frac{22.4 \text{ L}}{1 \cancel{\text{mol}}}\right) = 19.7 \text{ L}$$

3. Evaluate the Answer

Because the amount of gas present is slightly less than one mole, you expect the answer to be slightly less than 22.4 L. The answer agrees with that prediction. The unit in the answer is L, a volume unit.

Math Handbook

Review unit conversion in the **Math Handbook** on page 901 of this textbook.

For more practice with Avogadro's principle problems that use moles, go to **Supplemental Practice Problems** in Appendix A.

PRACTICE PROBLEMS

24. Determine the volume of a container that holds 2.4 mol of gas at STP.

25. What size container do you need to hold 0.0459 mol N_2 gas at STP?

26. What volume will 1.02 mol of carbon monoxide gas occupy at STP?

27. How many moles of nitrogen gas will be contained in a 2.00-L flask at STP?

28. If a balloon will rise off the ground when it contains 0.0226 mol of helium in a volume of 0.460 L, how many moles of helium are needed to make the balloon rise when its volume is 0.865 L? Assume that temperature and pressure stay constant.

EXAMPLE PROBLEM 14-6

Avogadro's Principle—Using Mass

Calculate the volume that 2.0 kg of methane gas (CH_4) will occupy at STP.

1. Analyze the Problem

You are given the temperature and pressure of a gas sample and the mass of gas the sample contains. One mole of a gas occupies 22.4 L at STP. The number of moles can be calculated by dividing the mass of the sample, *m*, by its molar mass, M.

Known	Unknown
m = 2.00 kg	V = ? L
T = 0.00°C	
P = 1.00 atm	

2. Solve for the Unknown

Use atomic masses and numbers of each type of atom to determine molecular mass for methane. Express that molecular mass in grams per mole to determine molar mass.

$$\mathrm{M} = \frac{1\ \cancel{\text{C atom}} \times 12.01\ \text{amu}}{\cancel{\text{C atom}}} + \frac{4\ \cancel{\text{H atoms}} \times 1.01\ \text{amu}}{\cancel{\text{H atom}}}$$

$$= 12.01\ \text{amu} + 4.04\ \text{amu} = 16.05\ \text{amu};\ 16.05\ \text{g/mol}$$

Multiply the mass of methane by a conversion factor to change it from kg to g.

$$2.00\ \cancel{\text{kg}}\left(\frac{1000\ \text{g}}{1\ \cancel{\text{kg}}}\right) = 2.00 \times 10^3\ \text{g}$$

Divide the mass in grams of methane by its molar mass to find the number of moles.

$$\frac{m}{\mathrm{M}} = \frac{2.00 \times 10^3\ \cancel{\text{g}}}{16.05\ \cancel{\text{g}}/\text{mol}} = 125\ \text{mol}$$

Because conditions are already at STP, multiply the known number of moles by the conversion factor of 22.4 L/1 mol to solve for the unknown volume.

$$V = 125\ \cancel{\text{mol}}\left(\frac{22.4\ \text{L}}{1\ \cancel{\text{mol}}}\right) = 2.80 \times 10^3\ \text{L}$$

3. Evaluate the Answer

The mass of methane present is much more than 1 mol, so you expect a large volume, which is in agreement with the answer. The units are L, a volume unit.

Avogadro's principle is essential to manufacturers that use gases. This factory produces ammonia.

PRACTICE PROBLEMS

29. How many grams of carbon dioxide gas are in a 1.0-L balloon at STP?

30. What volume in milliliters will 0.00922 g H_2 gas occupy at STP?

31. What volume will 0.416 g of krypton gas occupy at STP?

32. A flexible plastic container contains 0.860 g of helium gas in a volume of 19.2 L. If 0.205 g of helium is removed without changing the pressure or temperature, what will be the new volume?

33. Calculate the volume that 4.5 kg of ethylene gas (C_2H_4) will occupy at STP.

Practice!

For more practice with Avogadro's principle problems that use mass, go to **Supplemental Practice Problems** in Appendix A.

Why aren't weather balloons completely inflated when they are released? How strong must the walls of a scuba tank be? Look at the example of the combined gas law and Avogadro's principle shown in **Figure 14-8.** Using the combined gas law and Avogadro's principle together will help you understand how gases are affected by pressure, temperature, and volume.

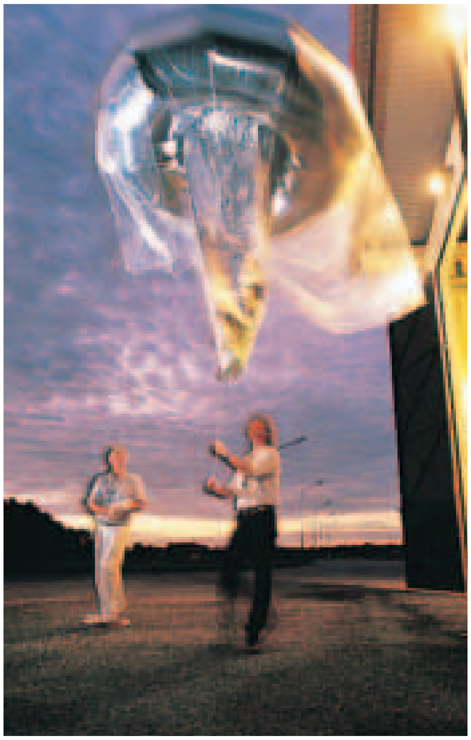


Figure 14-8

The combined gas law and Avogadro's principle have many practical applications that scientists and manufacturers must consider.

Section 14.2 Assessment

34. State the combined gas law using a sentence and then an equation.

35. What variable is assumed to be constant when using the combined gas law?

36. What three laws are used to make the combined gas law?

37. Explain why Avogadro's principle holds true for gases that have large particles and also for gases that have small particles.

38. Why must conditions of temperature and pressure be stated to do calculations involving molar volume?

39. **Thinking Critically** Think about what happens when a bottle of carbonated soft drink is shaken before being opened. Use the gas laws to explain whether the effect will be greater when the liquid is warm or cold.

40. **Applying Concepts** Imagine that you are going on an airplane trip in an unpressurized plane. You are bringing aboard an air-filled pillow that you have inflated fully. Predict what will happen when you try to use the pillow while the plane is at its cruising altitude.

Section 14.3

The Ideal Gas Law

Objectives

- **Relate** the amount of gas present to its pressure, temperature, and volume by using the ideal gas law.
- **Compare** the properties of real and ideal gases.

Vocabulary

ideal gas constant (R)
ideal gas law

In the last section, you learned that Avogadro noted the importance of being able to calculate the number of moles of a gas present under a given set of conditions. The laws of Avogadro, Boyle, Charles, and Gay-Lussac can be combined into a single mathematical statement that describes the relationship among pressure, volume, temperature, and number of moles of a gas. This formula is called the ideal gas law because it works best when applied to problems involving gases contained under certain conditions. The particles in an ideal gas are far enough apart that they exert minimal attractive or repulsive forces on one another and occupy a negligible volume.

The Ideal Gas Law

The number of moles is a fourth variable that can be added to pressure, volume, and temperature as a way to describe a gas sample. Recall that as the other gas laws were presented, care was taken to state that the relationships hold true for a "fixed mass" or a "given amount" of a gas sample. Changing the number of gas particles present will affect at least one of the other three variables.

As **Figure 14-9** illustrates, increasing the number of particles present in a sample will raise the pressure if the volume and temperature are kept constant. If the pressure and temperature are constant and more gas particles are added, the volume will increase.

Because pressure, volume, temperature, and the number of moles present are all interrelated, it would be helpful if one equation could describe their relationship. Remember that the combined gas law relates volume, temperature, and pressure of a sample of gas.

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

For a specific sample of gas, you can see that this relationship of pressure, volume, and temperature is always the same. You could say that

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

where k is a constant based on the amount of gas present, n. Experiments using known values of P, T, V, and n show that

$$\mathrm{k} = n\mathrm{R}$$

where **R** represents an experimentally determined constant that is referred to as the **ideal gas constant.** Therefore, the **ideal gas law**,

$$PV = n\mathrm{R}T$$

describes the physical behavior of an ideal gas in terms of the pressure, volume, temperature, and number of moles of gas present. The ideal gas law is used in the **CHEMLAB** in this chapter.

The ideal gas constant In the ideal gas equation, the value of R depends on the units used for pressure. **Table 14-1** shows the numerical value of R for different units of pressure.

Figure 14-9

The volume and temperature of this tire stay the same as air is added. However, the pressure in the tire increases as the amount of air present increases.

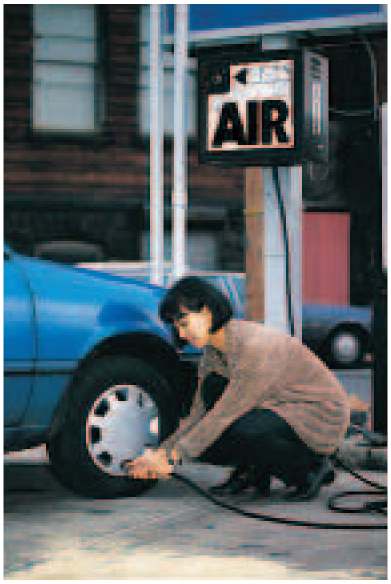


Table 14-1

Numerical Values of the Gas Constant, R					
Units of R	**Numerical value of R**	**Units of *P***	**Units of *V***	**Units of *T***	**Units of *n***
$\frac{L \cdot atm}{mol \cdot K}$	0.0821	atm	L	K	mol
$\frac{L \cdot kPa}{mol \cdot K}$	8.314	kPa	L	K	mol
$\frac{L \cdot mm\ Hg}{mol \cdot K}$	62.4	mm Hg	L	K	mol

The R value you will probably find most useful is the first one listed in the table, 0.0821 $\frac{L \cdot atm}{mol \cdot K}$. Use this R in problems in which the unit of volume is liters, the pressure is in atmospheres, and the temperature is in kelvins.

Real versus ideal gases What does the term *ideal gas* mean? An ideal gas is one whose particles take up no space and have no intermolecular attractive forces. An ideal gas follows the gas laws under all conditions of temperature and pressure.

In the real world, no gas is truly ideal. All gas particles have some volume, however small it may be, because of the sizes of their atoms and the lengths of their bonds. All gas particles also are subject to intermolecular interactions. Despite that, most gases will behave like ideal gases at many temperature and pressure levels. Under the right conditions of temperature and pressure, calculations made using the ideal gas law closely approximate actual experimental measurements.

When is the ideal gas law not likely to work for a real gas? Real gases deviate most from ideal gas behavior at extremely high pressures and low temperatures. As the amount of space between particles and the speed at which the particles move decrease, the effects of the volume of gas particles and intermolecular attractive forces become increasingly important. The gas behaves as a real gas in **Figure 14-10a**. Lowering the temperature of nitrogen gas results in less kinetic energy of the gas particles, which means their intermolecular attractive forces are strong enough to bond them more closely together. When the temperature is low enough, this real gas condenses to form a liquid. The gas in **Figure 14-10b** also behaves as a real gas. Increasing the pressure on a gas such as propane lowers the volume and forces the gas particles closer together until their volume is no longer negligible compared to the volume of the tank. Real gases such as propane will liquefy if enough pressure is applied.

Figure 14-10

Real gases deviate most from ideal behavior at low temperatures and high pressures. **a** Liquid nitrogen is used to store biological tissue samples at low temperatures. **b** Increased pressure allows a larger mass of propane to fit into a smaller volume for easier transport. Propane is sold as LP (liquid propane) for this reason, although it is actually burned for fuel as a gas.

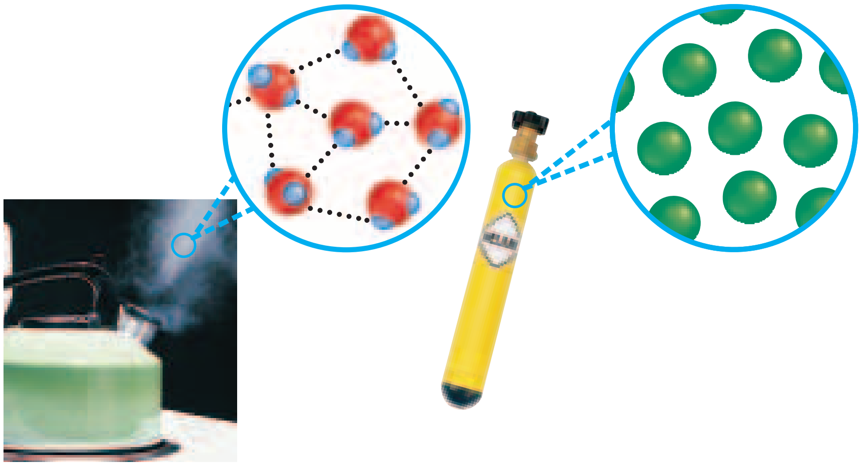


Figure 14-11

In polar gas molecules, such as water vapor, oppositely charged poles attract each other through electrostatic forces.
A nonpolar gas, such as helium, is uncharged. Thus nonpolar gases are more likely to behave like ideal gases than are polar gases.

The nature of the particles making up a gas also affects how ideally the gas behaves. For example, polar gas molecules such as water vapor generally have larger attractive forces between their particles than nonpolar molecules such as chlorine gas. The oppositely charged ends of polar molecules are pulled together through electrostatic forces, as shown in **Figure 14-11**. Therefore polar gases do not behave as ideal gases. Also, the particles of gases composed of molecules such as butane (C_4H_{10}) occupy more actual volume than an equal number of gas particles of smaller molecules such as helium (He). Therefore, larger gas molecules tend to cause a greater departure from ideal behavior than do smaller gas molecules.

Applying the Ideal Gas Law

Look again at the combined gas law on page 428. Notice that it cannot be used to find n, the number of moles of a gas. However, the ideal gas law can be used to solve for the value of any one of the four variables P, V, T, or n if the values of the other three are known. Rearranging the $PV = n\mathrm{R}T$ equation allows you to also calculate the molar mass and density of a gas sample if the mass of the sample is known.

To find the molar mass of a gas sample, the mass, temperature, pressure, and volume of the gas must be known. Remember from Chapter 12 that the number of moles of a gas (n) is equal to the mass (m) divided by the molar mass (M). Therefore, the n in the equation can be replaced by m/M.

$$PV = \frac{m\mathrm{R}T}{\mathrm{M}} \text{ or } \mathrm{M} = \frac{m\mathrm{R}T}{PV}$$

Math Handbook

Review fractions in the **Math Handbook** on page 907 of this textbook.

EXAMPLE PROBLEM 14-7

The Ideal Gas Law—Using Moles

Calculate the number of moles of gas contained in a 3.0-L vessel at 3.00×10^2 K with a pressure of 1.50 atm.

1. Analyze the Problem

You are given the volume, temperature, and pressure of a gas sample. When using the ideal gas law to solve for n, choose the value of R that contains the pressure and temperature units given in the problem.

Known	Unknown
V = 3.0 L	n = ? mol
$T = 3.00 \times 10^2$ K	
P = 1.50 atm	
$\mathrm{R} = 0.0821 \frac{\mathrm{L \cdot atm}}{\mathrm{mol \cdot K}}$	

2. Solve for the Unknown

Divide both sides of the ideal gas law equation by RT to solve for n.

$PV = nRT$

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

Substitute the known values into the rearranged equation.

$$n = \frac{(1.50\ \text{atm})(3.0\ \text{L})}{\left(0.0821\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(3.00 \times 10^2\ \text{K})}$$

Multiply and divide numbers and units to solve for n.

$$n = \frac{(1.50\ \cancel{\text{atm}})(3.0\ \cancel{\text{L}})}{\left(0.0821\ \frac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\text{mol}\cdot\cancel{\text{K}}}\right)(3.00 \times 10^2\ \cancel{\text{K}})} = 0.18\ \text{mol}$$

3. Evaluate the Answer

One mole of a gas at STP occupies 22.4 L. In this problem, the volume is much smaller than 22.4 L while the temperature and pressure values are not too different from those at STP. The answer agrees with the prediction that the number of moles present will be significantly less than one mole. The unit of the answer is the mole.

PRACTICE PROBLEMS

41. If the pressure exerted by a gas at 25°C in a volume of 0.044 L is 3.81 atm, how many moles of gas are present?

42. Determine the Celsius temperature of 2.49 moles of gas contained in a 1.00-L vessel at a pressure of 143 kPa.

43. Calculate the volume that a 0.323-mol sample of a gas will occupy at 265 K and a pressure of 0.900 atm.

44. What is the pressure in atmospheres of a 0.108-mol sample of helium gas at a temperature of 20.0°C if its volume is 0.505 L?

45. Determine the kelvin temperature required for 0.0470 mol of gas to fill a balloon to 1.20 L under 0.988 atm pressure.

For more practice with ideal gas law problems that use moles, go to **Supplemental Practice Problems** in Appendix A.

Recall from Chapter 2 that density (D) is defined as mass (m) per unit volume (V). After rearranging the ideal gas equation to solve for molar mass, D can be substituted for m/V.

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

This equation can be rearranged to solve for the density of a gas.

$$D = \frac{\text{M}P}{RT}$$

Why might you need to know the density of a gas? Consider what requirements are necessary to fight a fire. One way to put out a fire is to remove its oxygen source by covering it with another gas that will neither burn nor support combustion. This gas must have a greater density than oxygen so that it will fall to the level of the fire. You can observe applications of density when you do the **miniLAB** later in this section and read the **Chemistry and Technology** feature.

The density of a gas can be used to identify it. This gas is denser than air and can be poured from one container to another.

EXAMPLE PROBLEM 14-8

The Ideal Gas Law—Using Molar Mass

What is the molar mass of a pure gas that has a density of 1.40 g/L at STP?

1. Analyze the Problem

You are given the density, temperature, and pressure of a sample of gas. Because density is known and mass and volume are not, use the form of the ideal gas equation that involves density.

Known

$D = 1.40\ \frac{\text{g}}{\text{L}}$

$T = 0.00°\text{C}$

$P = 1.00\ \text{atm}$

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

Unknown

$\text{M} = ?\ \frac{\text{g}}{\text{mol}}$

2. Solve for the Unknown

Convert the standard T to kelvin units.

$T_K = 273 + T_C$

$T_K = 273 + 0.00°\text{C} = 273\ \text{K}$

Use the form of the ideal gas law that includes density (D) and solves for M.

$$\text{M} = \frac{DRT}{P}$$

Substitute the known values into the equation.

$$\text{M} = \frac{\left(1.40\ \frac{\text{g}}{\text{L}}\right)\left(0.0821\ \frac{\text{L·atm}}{\text{mol·K}}\right)(273\ \text{K})}{1\ \text{atm}}$$

Multiply and divide numbers and units to solve for M.

$$\text{M} = \frac{\left(1.40\ \frac{\text{g}}{\cancel{\text{L}}}\right)\left(0.0821\ \frac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\text{mol}\cdot\cancel{\text{K}}}\right)(273\ \cancel{\text{K}})}{1\ \cancel{\text{atm}}} = 31.4\ \text{g/mol}$$

3. Evaluate the Answer

You would expect the molar mass of a gas to fall somewhere between that of one of the lightest gases under normal conditions, such as 2 g/mol for H_2, and that of a relatively heavy gas, such as 222 g/mol for Rn. The answer seems reasonable. The unit is g/mol, which is the molar mass unit.

PRACTICE PROBLEMS

46. How many grams of gas are present in a sample that has a molar mass of 70.0 g/mol and occupies a 2.00-L container at 117 kPa and 35.1°C?

47. Calculate the grams of N_2 gas present in a 0.600-L sample kept at 1.00 atm pressure and a temperature of 22.0°C.

48. What is the density of a gas at STP that has a molar mass of 44.0 g/mol?

49. What is the molar mass of a sample of gas that has a density of 1.09 g/L at 1.02 atm pressure and 25.0°C?

50. Calculate the density a gas will have at STP if its molar mass is 39.9 g/mol.

Practice!

For more practice with ideal gas law problems that use molar mass, go to **Supplemental Practice Problems** in Appendix A.

miniLAB

The Density of Carbon Dioxide

Hypothesizing Air is a mixture of mostly nitrogen and oxygen. Use observations to form a hypothesis about which has greater density, air or carbon dioxide.

Materials masking tape, aluminum foil, metric ruler, 1-L beaker, candle, matches, thermometer, barometer or weather radio, baking soda ($NaHCO_3$), vinegar (5% CH_3COOH)

Procedure

1. Record the temperature and the barometric pressure of the air in the classroom.
2. Roll a 23-cm × 30-cm piece of aluminum foil into a cylinder that is 6 cm × 30 cm. Tape the edges with masking tape.
3. Use matches to light a candle. **CAUTION:** *Run water over the extinguished match before throwing it away. Keep all hair and loose clothing away from the flame.*
4. Place 30 g of baking soda in the bottom of a large beaker. Add 40 mL of vinegar.
5. Quickly position the foil cylinder at approximately 45° up and away from the top of the candle flame.
6. While the reaction in the beaker is actively producing CO_2 gas, carefully pour the gas, but not the liquid, out of the beaker and into the top of the foil tube. Record your observations.

Analysis

1. Based on your observations, state a hypothesis about whether CO_2 is heavier or lighter than air.
2. Use the combined gas law to calculate molar volume at room temperature and atmospheric pressure.
3. Carbon dioxide gas (CO_2) has a molar mass of 44 g/mol. The two major components of air, which are oxygen and nitrogen, have molar masses of 32 g/mol and 28 g/mol, respectively. Calculate the room-temperature densities in g/L of nitrogen (N_2), oxygen (O_2), and carbon dioxide (CO_2) gases.
4. Do these calculations confirm your hypothesis? Explain.

Section 14.3 Assessment

51. Write the equation for the ideal gas law.

52. Use the kinetic-molecular theory to analyze and evaluate the ideal gas law's applicability to real gases.

53. List common units for each variable in the ideal gas law.

54. **Thinking Critically** Which of the following gases would you expect to behave most like an ideal gas at room temperature and atmospheric pressure: water vapor, carbon dioxide, helium, or hydrogen? Explain.

55. **Making and Using Graphs** The accompanying data show the volume of hydrogen gas collected at a number of different temperatures. Illustrate these data with a graph and use them to determine the temperature at which the volume will reach a value of 0 mL. What is this temperature called? For more help, refer to **Drawing Line Graphs** in the **Math Handbook** on page 903 of this text.

Volume of H_2 Collected at Different Temperatures

Trial	1	2	3	4	5	6
T (°C)	300	175	110	0	−100	−150
V (mL)	48	37	32	22	15	11

Section 14.4 Gas Stoichiometry

Objectives

- **Determine** volume ratios for gaseous reactants and products by using coefficients from a chemical equation.
- **Calculate** amounts of gaseous reactants and products in a chemical reaction using the gas laws.

All the laws you have learned so far involving gases can be applied to calculate the stoichiometry of reactions in which gases are reactants or products. Recall that the coefficients in chemical equations represent molar amounts of substances taking part in the reaction. For example, when butane gas burns, the reaction is represented by the following chemical equation.

$$2C_4H_{10}(g) + 13O_2(g) \longrightarrow 8CO_2(g) + 10H_2O(g)$$

From the balanced chemical equation, you know that 2 mol of butane reacts with 13 mol of oxygen, producing 8 mol of carbon dioxide and 10 mol of water vapor. By examining this balanced equation, you are able to find mole ratios of substances in this reaction. Avogadro's principle states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles. Thus, when gases are involved, the coefficients in a balanced chemical equation represent not only molar amounts but also relative volumes. For example, if 2 L of butane reacts, the reaction involves 13 L of oxygen and produces 8 L of carbon dioxide and 10 L of water vapor.

Calculations Involving Only Volume

To find the volume of a gaseous reactant or product in a reaction, you must know the balanced chemical equation for the reaction and the volume of at least one other gas involved in the reaction. Examine the reaction showing the combustion of methane, which takes place every time you light a Bunsen burner.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(g)$$

Because the coefficients represent volume ratios for gases taking part in the reaction, you can determine that it takes 2 L of oxygen to react completely with 1 L of methane. The complete combustion of 1 L of methane will produce 1 L of carbon dioxide and 2 L of water vapor, as shown in **Figure 14-12**.

What volume of methane is needed to produce 26 L of water vapor? Because the volume ratio of methane and water vapor is 1:2, the volume of methane needed is half that of the water vapor. Thus, 13 L of methane is needed to produce 26 L of water vapor. What volume of oxygen is needed to produce 6.0 L of carbon dioxide?

Note that no conditions of temperature and pressure are listed. They are not needed as part of the calculation because after mixing, each gas is at the same temperature and pressure. The same temperature and pressure affect all gases in the same way, so these conditions don't need to be considered.

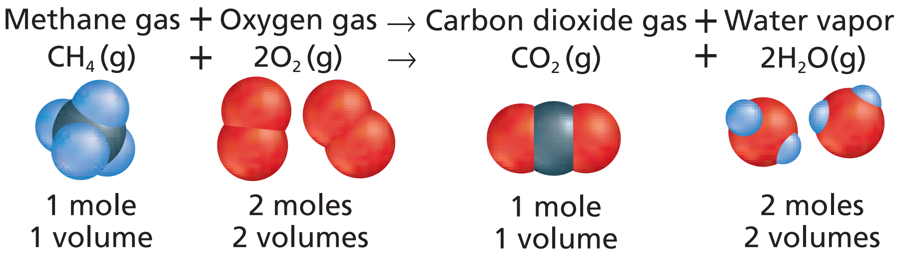


Figure 14-12

The coefficients in a balanced equation show the relationships among numbers of moles of all reactants and products. The coefficients also show the relationships among volumes of any gaseous reactants or products. From these coefficients, volume ratios can be set up for any pair of gases in the reaction.

EXAMPLE PROBLEM 14-9

Volume-Volume Problems

What volume of oxygen gas is needed for the complete combustion of 4.00 L of propane gas (C_3H_8)? Assume constant pressure and temperature.

1 Analyze the Problem

You are given the volume of a gaseous reactant in a chemical reaction. Remember that the coefficients in a balanced chemical equation provide the volume relationships of gaseous reactants and products.

Known	Unknown
$V_{C_3H_8} = 4.00$ L	$V_{O_2} = ?$ L

2. Solve for the Unknown

Write the balanced equation for the combustion of C_3H_8.

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

Use the balanced equation to find the volume ratio for O_2 and C_3H_8.

$$\frac{5 \text{ volumes } O_2}{1 \text{ volume } C_3H_8}$$

Multiply the known volume of C_3H_8 by the volume ratio to find the volume of O_2.

$$(4.00 \text{ L } \cancel{C_3H_8}) \times \frac{5 \cancel{\text{ volumes}} O_2}{1 \cancel{\text{ volume } C_3H_8}} = 20.0 \text{ L } O_2$$

3. Evaluate the Answer

The coefficients in the combustion equation show that a much larger volume of O_2 than C_3H_8 is used up in the reaction, which is in agreement with the calculated answer. The unit of the answer is L, a unit of volume.

Correct proportions of gases are needed for many chemical reactions. The combustion of propane heats the air that inflates the balloon.

PRACTICE PROBLEMS

56. What volume of oxygen is needed to react with solid sulfur to form 3.5 L SO_2?

57. Determine the volume of hydrogen gas needed to react completely with 5.00 L of oxygen gas to form water.

58. How many liters of propane gas (C_3H_8) will undergo complete combustion with 34.0 L of oxygen gas?

59. What volume of oxygen is needed to completely combust 2.36 L of methane gas (CH_4)?

Practice!

For more practice with volume-volume problems, go to **Supplemental Practice Problems** in Appendix A.

Calculations Involving Volume and Mass

To do stoichiometric calculations that involve both gas volumes and masses, you must know the balanced equation for the reaction involved, at least one mass or volume value for a reactant or product, and the conditions under which the gas volumes have been measured. Then the ideal gas law can be used along with volume or mole ratios to complete the calculation.

In doing this type of problem, remember that the balanced chemical equation allows you to find ratios for moles and gas volumes only—not for masses. All masses given must be converted to moles or volumes before being used as part of a ratio. Also remember that the temperature units used must be kelvin.

EXAMPLE PROBLEM 14-10

Volume-Mass Problems

Ammonia is synthesized from hydrogen and nitrogen gases.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

If 5.00 L of nitrogen reacts completely by this reaction at a constant pressure and temperature of 3.00 atm and 298 K, how many grams of ammonia are produced?

Ammonia is essential in the production of chemical fertilizers.

1. Analyze the Problem

You are given the volume, pressure, and temperature of a gas sample. The mole and volume ratios of gaseous reactants and products are given by the coefficients in the balanced chemical equation. Volume can be converted to moles and thus related to mass by using molar mass and the ideal gas law.

Known	Unknown
$V_{N_2} = 5.00$ L	$m_{NH_3} = ?$ g
$P = 3.00$ atm	
$T = 298$ K	

2. Solve for the Unknown

Determine volume ratios from the balanced chemical equation.

$$\frac{1 \text{ volume } N_2}{2 \text{ volumes } NH_3}$$

Use this ratio to determine how many liters of gaseous ammonia will be made from 5.00 L of nitrogen gas.

$$5.00 \text{ L } \cancel{N_2}\left(\frac{2 \text{ } \cancel{\text{volumes}} \text{ } \cancel{NH_3}}{1 \text{ } \cancel{\text{volume}} \text{ } N_2}\right) = 10.0 \text{ L } NH_3$$

Rearrange the equation for the ideal gas law to solve for n.

$$PV = nRT$$

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

Substitute the known values into the rearranged equation using the volume of NH_3 for V. Multiply and divide numbers and units to solve for the number of moles of NH_3.

$$n = \frac{(3.00 \text{ } \cancel{\text{atm}})(10.0 \text{ } \cancel{\text{L}})}{\left(0.0821 \frac{\cancel{\text{L}}\cdot\cancel{\text{atm}}}{\text{mol}\cdot\cancel{\text{K}}}\right)(298 \text{ } \cancel{\text{K}})} = 1.23 \text{ mol } NH_3$$

Find the molar mass, M, of NH_3 by finding the molecular mass and expressing it in units of g/mol.

$$M = \left(\frac{1 \text{ } \cancel{\text{N atom}} \times 14.01 \text{ amu}}{\text{N } \cancel{\text{atom}}}\right) + \left(\frac{3 \text{ } \cancel{\text{H atoms}} \times 1.01 \text{ amu}}{\cancel{\text{H atom}}}\right)$$

$$= 17.04 \text{ amu}$$

$$M = 17.04 \frac{\text{g}}{\text{mol}}$$

Convert moles of ammonia to grams of ammonia using molar mass of ammonia as a conversion factor.

$$1.23 \text{ } \cancel{\text{mol}} \text{ } NH_3 \times 17.04 \frac{\text{g}}{\cancel{\text{mol}}} = 21.0 \text{ g } NH_3$$

3. Evaluate the Answer

To check your answer, calculate the volume of reactant nitrogen at STP. Then, use molar volume and the mole ratio between N_2 and NH_3 to determine how many moles of NH_3 were produced. You can convert the answer to grams using the molar mass of NH_3. All data was given in three significant digits as is the answer.

PRACTICE PROBLEMS

60. Ammonium nitrate is a common ingredient in chemical fertilizers. Use the reaction shown to calculate the mass of solid ammonium nitrate that must be used to obtain 0.100 L of dinitrogen oxide gas at STP.

$NH_4NO_3(s) \rightarrow N_2O(g) + 2H_2O(g)$

61. Calcium carbonate forms limestone, one of the most common rocks on Earth. It also forms stalactites, stalagmites, and many other types of formations found in caves. When calcium carbonate is heated, it decomposes to form solid calcium oxide and carbon dioxide gas.

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

How many liters of carbon dioxide will be produced at STP if 2.38 kg of calcium carbonate reacts completely?

62. Determine how many moles of water vapor will be produced at 1.00 atm and 200°C by the complete combustion of 10.5 L of methane gas (CH_4).

63. When iron rusts, it undergoes a reaction with oxygen to form iron(III) oxide.

$4\,Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

Calculate the volume of oxygen gas at STP that is required to completely react with 52.0 g of iron.

64. Solid potassium metal will react with Cl_2 gas to form ionic potassium chloride. How many liters of Cl_2 gas are needed to completely react with 0.204 g of potassium at STP?

Practice!
For more practice with volume-mass problems, go to **Supplemental Practice Problems** in Appendix A.

Chemistry Online

Topic: Gases
To learn more about gases, visit the Chemistry Web site at **chemistrymc.com**

Activity: Research how the gas laws are important to fish and scuba divers. Explain your answers using equations when possible.

Stoichiometric problems such as these are considered in industrial processes that involve gases. How much of a reactant should be purchased? How much of a product will be produced? What conditions of temperature and presssure are necessary? Answers to these questions are essential to effective production of a product.

Section 14.4 Assessment

65. How do mole ratios compare to volume ratios for gaseous reactants and products in a balanced chemical equation?

66. Is the volume of a gas directly or inversely proportional to the number of moles of a gas at constant temperature and pressure? Explain.

67. Determine the volume ratio of ammonia to nitrogen in the reaction shown.

$3H_2 + N_2 \rightarrow 2NH_3$

Which will occupy a larger volume at a given temperature and pressure: one mole of H_2 or one mole of NH_3?

68. **Thinking Critically** One mole of a gas occupies a volume of 22.4 L at STP. Calculate the temperature and pressure conditions needed to fit two moles of a gas into a volume of 22.4 L.

69. **Predicting** Using what you have learned about gases, predict what will happen to the size of the reaction vessel you need to carry out a reaction involving gases if the temperature is doubled and the pressure is held constant.

Using the Ideal Gas Law

The ideal gas law is a powerful tool that the chemist—and now you—can use to determine the molar mass of an unknown gas. By measuring the temperature, pressure, volume, and mass of a gas sample, you can calculate the molar mass of the gas.

Problem

How can the equation for the ideal gas law be used to calculate the molar mass of a gas?

Objectives

- **Measure** the mass, volume, temperature, and pressure of an insoluble gas collected over water.
- **Calculate** the molar mass of an unknown gas using the ideal gas equation.

Materials

aerosol can of duster
600-mL graduated beaker
bucket or bowl
thermometer (°C)
barometer or weather radio
plastic microtip pipette
latex tubing
glass tubing
scissors
electrical or duct tape
balance

Safety Precautions

- **Read and observe all cautions listed on the aerosol can of office equipment duster.**
- **Do not have any open flames in the room.**

Pre-Lab

1. Read the entire **CHEMLAB.**
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.
3. Because you will collect the aerosol gas over water, the beaker contains both the aerosol gas and water vapor. Form a hypothesis about how the presence of water vapor will affect the calculated value of the molar mass of the gas. Explain.
4. The following gases are or have been used in aerosol cans, some as propellants. Use the gases' molecular formulas to calculate their molar masses.
   a. propane, C_3H_8
   b. butane, C_4H_{10}
   c. dichlorodifluoromethane, CCl_2F_2
   d. tetrafluoroethane, $C_2H_2F_4$
5. Given the following data for a gas, use the equation for the ideal gas law to calculate the molar mass.
   a. mass = 0.810 g
   b. pressure = 0.954 atm
   c. volume = 0.461 L
   d. temperature = 291 K

Data and Calculations	
Mass of can before release of gas (g)	
Mass of can after release of gas (g)	
Mass of gas released (g)	
Air temperature (°C)	
Air temperature (K)	
Air pressure (list what unit was used)	
Air pressure (atm)	
Volume of gas collected (L)	

Procedure

1. Place the bucket in the sink and fill it with water.
2. Submerge the beaker in the water. Then, invert it in the bucket, being careful to keep it completely filled with water.
3. Measure the mass of an aerosol can of office equipment duster. Record the mass in the data table.

4. Use scissors to cut the stem from a plastic microtip pipette.
5. Fit the pipette stem over the long plastic spray tip that comes with the aerosol can to extend the length of the tip and enlarge the diameter.
6. Connect one end of 30 cm of latex tubing to glass tubing that is 8 cm long.
7. Connect the other end to the pipette stem that is attached to the aerosol can. If necessary, tape any connections so that they don't leak.
8. Place the end of the glass tubing under the pour spout of the inverted beaker as shown in the photo.

9. Hold the beaker down while you slowly release the gas from the aerosol can. Collect between 400 and 500 mL of the gas by water displacement.
10. To equalize the air pressure, lift the beaker so that the water level inside and outside the beaker is the same.
11. Carefully read the volume of the gas collected using the graduations on the beaker.
12. Record this volume of the gas collected in the data table.
13. Remove the tubing from the aerosol can.
14. Measure the mass of the can and record it in the data table.
15. Using a barometer or weather radio, record the atmospheric pressure in the data table.
16. Using a thermometer, determine air temperature. Record it in the data table.

Cleanup and Disposal

1. Dispose of the empty can according to the instructions on its label.
2. Pour the water down the drain.
3. Discard any tape and the pipettes in the trash can.
4. Return all lab equipment to its proper place.

Analyze and Conclude

1. **Using Numbers** Fill in the remainder of the data table by calculating the mass of the gas that was released from the aerosol can, converting the atmospheric pressure from the units measured into atmospheres, and converting the air temperature into kelvins. Substitute your data from the table into the form of the ideal gas equation that solves for M. Calculate the molar mass of the gas in the can using the appropriate value for R.
2. **Using Numbers** Read the contents of the can and determine which of the gases from step 4 in the Pre-Lab is the most likely propellant.
3. **Error Analysis** Remember that you are collecting the gas after it has bubbled through water. What might happen to some of the gas as it goes through the water? What might be present in the gas in the beaker in addition to the gas from the can? Calculate the percent error using your calculated molar mass compared to the molar mass of the gas in the aerosol can.
4. **Interpreting Data** Were your data consistent with the ideal gas law? Evaluate the pressure and temperature at which your experiment was done, and the polarity of the gas. Would you expect the gas in your experiment to behave as an ideal gas or a real gas?

Real-World Chemistry

1. Explain why the label on an aerosol can warns against exposing the can to high heat.
2. Use the ideal gas law to explain why the wind blows.

CHEMISTRY and Technology

Giving a Lift to Cargo

Have you ever been caught in a traffic jam caused by a truck carrying an oversized piece of industrial equipment? The need to transport heavy loads more efficiently is one factor leading to the revival of a technology that most people considered to be dead—the airship, sometimes called a dirigible or zeppelin. The burning of the zeppelin *Hindenburg* in May 1937, followed closely by World War II, ended nearly all commercial use of these "lighter-than-air" ships. But by the 1990s, chemists and engineers had developed strong, lightweight alloys and tough, fiber-reinforced composite plastics. These materials, along with computerized control and satellite navigation systems are making commercial airships practical again.

Modern airships

Modern airships use helium to provide lift. The *Hindenburg* used hydrogen gas, which provides about twice the lift of helium but is no longer used because it is extremely flammable. The helium is contained in bags made of a space-age, leakproof fabric called Tedlar, a type of plastic. The bags are loosely inflated so that the helium pressure is about the same as atmospheric pressure. The quantity of helium determines the lifting ability of the ship.

Airships and Boyle's law

As the airship rises, atmospheric pressure decreases and the helium expands, as Boyle's law predicts. As the ship reaches the desired altitude, air is pumped into another bag called a *ballonet.* The pressure of the ballonet prevents the helium bags from expanding further, thus keeping the ship at that altitude. To descend, more air is pumped into the ballonet. This added pressure squeezes the helium bags, causing the volume to decrease and the density of the helium to increase. Also, the compressed air adds weight to the ship. The lifting power of the helium is reduced and the airship moves downward.

Uses for modern airships

The old airships of 1900 to 1940 were used mostly for luxury passenger service. Among modern airships, the German Zeppelin-NT and a ship being developed by the Hamilton Airship Company in South Africa are designed to carry passengers.

Probably the most interesting new airship is the huge CargoLifter from a German-American company. Its length is slightly less than the length of three football fields. CargoLifter's skeleton is constructed of a strong carbon fiber composite material that is much less dense than metals. At the mooring mast, it is as tall as a 27-story building and contains 450 000 m^3 of helium. It is designed to carry loads up to 160 metric tons (352 000 pounds). It can pick up and deliver objects such as large turbines and entire locomotives. Because modern roads are not needed, equipment can be delivered to locations in developing countries that would otherwise be impossible to reach.

Investigating the Technology

1. **Thinking Critically** No plastics, fabrics, or metals that are "lighter-than-air" exist. Yet these materials are used to make airships. Why is it possible to describe an airship as a lighter-than-air craft?

Chemistry Online

Visit the Chemistry Web site at **chemistrymc.com** to find links to more information about airships.

Summary

14.1 The Gas Laws

- Boyle's law states that the pressure and volume of a contained gas are inversely proportional if temperature is constant.
- Charles's law states that the volume and kelvin temperature of a contained gas are directly proportional if pressure is constant.
- Gay-Lussac's law states that the pressure and kelvin temperature of a contained gas are directly proportional if volume is constant.

14.2 The Combined Gas Law and Avogadro's Principle

- Boyle's, Charles's, and Gay-Lussac's laws are brought together in the combined gas law, which permits calculations involving changes in the three gas variables of pressure, volume, and temperature.
- Avogadro's principle states that equal volumes of gases at the same temperature and pressure contain equal numbers of particles. The volume of one mole of a gas at STP is 22.4 L.

14.3 The Ideal Gas Law

- The combined gas law and Avogadro's principle are used together to form the ideal gas law. In the equation for the ideal gas law, R is the ideal gas constant.
- The ideal gas law allows you to determine the number of moles of a gas when its pressure, temperature, and volume are known.
- Real gases deviate from behavior predicted for ideal gases because the particles of a real gas occupy volume and are subject to intermolecular forces.
- The ideal gas law can be used to find molar mass if the mass of the gas is known, or the density of the gas if its molar mass is known.

14.4 Gas Stoichiometry

- The coefficients in a balanced chemical equation specify volume ratios for gaseous reactants and products.
- The gas laws can be used along with balanced chemical equations to calculate the amount of a gaseous reactant or product in a reaction.

Key Equations and Relationships

- Boyle's law: $P_1V_1 = P_2V_2$, constant temperature (p. 421)
- Charles's law: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, constant pressure (p. 424)
- Gay-Lussac's law: $\frac{P_1}{T_1} = \frac{P_2}{T_2}$, constant volume (p. 426)
- Combined gas law: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$ (p. 428)
- Ideal gas law: $PV = nRT$ (p. 434)
- Finding molar mass: $\mathrm{M} = \frac{m\mathrm{R}T}{PV}$ (p. 437)
- Finding density: $D = \frac{\mathrm{M}P}{\mathrm{R}T}$ (p. 437)

Vocabulary

- Avogadro's principle (p. 430)
- Boyle's law (p. 421)
- Charles's law (p. 424)
- combined gas law (p. 428)
- Gay-Lussac's law (p. 426)
- ideal gas constant (R) (p. 434)
- ideal gas law (p. 434)
- molar volume (p. 431)

chemistrymc.com/vocabulary_puzzlemaker

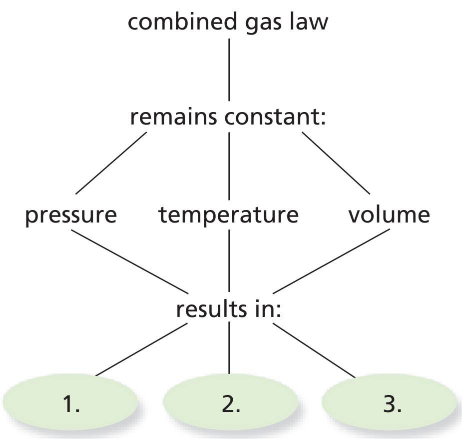
CHAPTER 14 ASSESSMENT

ChemistryOnline

Go to the Glencoe Chemistry Web site at **chemistrymc.com** *for additional Chapter 14 Assessment.*

Concept Mapping

70. Complete the following concept map that shows how Boyle's law, Charles's law, and Gay-Lussac's law are derived from the combined gas law.



Mastering Concepts

71. State the laws of Boyle, Charles, and Gay-Lussac as equations. (14.1)

72. What will happen to the pressure of a contained gas if its temperature is lowered? (14.1)

73. Why is it important not to puncture an aerosol can? (14.1)

74. Explain why an unopened bag of potato chips left in a hot car appears to become larger. (14.1)

75. If two variables have an inverse relationship, what happens to the value of one as the value of the other increases? (14.1)

76. If two variables have a direct relationship, what happens to the value of one as the value of the other is increased? (14.1)

77. Label the following examples as being generally representative of direct or inverse relationship. (14.1)

- **a.** popularity of a musical group versus how hard it is to get tickets for its concert
- **b.** number of carats a diamond weighs versus its cost
- **c.** number of people helping versus how long it takes to clean up after a party

78. Label the following examples as being representative of a direct or inverse relationship. (14.1)

- **a.** pressure versus volume of a gas
- **b.** volume versus temperature of a gas
- **c.** pressure versus temperature of a gas

79. What four variables are used to describe gases? (14.1)

80. List the standard conditions for gas measurements. (14.2)

81. Write the equation for the combined gas law. Identify the units most commonly used with each variable. (14.2)

82. State Avogadro's principle. (14.2)

83. What volume is occupied by one mole of a gas at STP? What volume do two moles occupy at STP? (14.2)

84. What units must be used to express the temperature in the equation for the ideal gas law? Explain. (14.3)

85. List two conditions under which a gas is least likely to behave ideally. (14.3)

86. Write the value and units for the gas constant R in two common forms. (14.3)

87. What information is needed to solve a volume-mass problem that involves gases? (14.4)

Mastering Problems

The Gas Laws (14.1)

88. Use Boyle's, Charles's, or Gay-Lussac's law to calculate the missing value in each of the following.

- **a.** $V_1 = 2.0$ L, $P_1 = 0.82$ atm, $V_2 = 1.0$ L, $P_2 = ?$
- **b.** $V_1 = 250$ mL, $T_1 = ?$, $V_2 = 400$ mL, $T_2 = 298$ K
- **c.** $V_1 = 0.55$ L, $P_1 = 740$ mm Hg, $V_2 = 0.80$ L, $P_2 = ?$
- **d.** $T_1 = 25°C$, $P_1 = ?$, $T_2 = 37°C$, $P_2 = 1.0$ atm

89. What is the pressure of a fixed volume of a gas at 30.0°C if it has a pressure of 1.11 atm at 15.0°C?

90. A fixed amount of oxygen gas is held in a 1.00-L tank at a pressure of 3.50 atm. The tank is connected to an empty 2.00-L tank by a tube with a valve. After this valve has been opened and the oxygen is allowed to flow freely between the two tanks at a constant temperature, what is the final pressure in the system?

91. Hot-air balloons rise because the hot air inside the balloon is less dense than the cooler air outside. Calculate the volume an air sample will occupy inside a balloon at 43.0°C if it occupies 2.50 L at the outside air temperature of 22.0°C, assuming the pressure is the same at both locations.

The Combined Gas Law (14.2)

92. A sample of nitrogen gas is stored in a 500.0-mL flask at 108 kPa and 10.0°C. The gas is transferred to a 750.0-mL flask at 21.0°C. What is the pressure of nitrogen in the second flask?

93. The air in a dry, sealed 2-L soda bottle has a pressure of 0.998 atm at sea level at a temperature of 34.0°C. What will be its pressure if it is brought to a higher altitude where the temperature is only 23.0°C?

94. A weather balloon is filled with helium that occupies a volume of 5.00×10^4 L at 0.995 atm and 32.0°C. After it is released, it rises to a location where the pressure is 0.720 atm and the temperature is −12.0°C. What is the volume of the balloon at that new location?

Avogadro's Principle (14.2)

95. Propane, C_3H_8, is a gas commonly used as a home fuel for cooking and heating.

a. Calculate the volume that 0.540 mol of propane occupies at STP.

b. Think about the size of this volume compared to the amount of propane that it contains. Why do you think propane is usually liquefied before it is transported?

96. Carbon monoxide, CO, is a product of incomplete combustion of fuels. Find the volume that 42 g of carbon monoxide gas occupies at STP.

The Ideal Gas Law (14.3)

97. The lowest pressure achieved in a laboratory is about 1.0×10^{-15} mm Hg. How many molecules of gas are present in a 1.00-L sample at that pressure and a temperature of 22.0ºC?

98. Determine the density of chlorine gas at 22.0°C and 1.00 atm pressure.

99. Geraniol is a compound found in rose oil that is used in perfumes. What is the molar mass of geraniol if its vapor has a density of 0.480 $\frac{g}{L}$ at a temperature of 260.0°C and a pressure of 0.140 atm?

100. A 2.00-L flask is filled with propane gas (C_3H_8) at 1.00 atm and −15.0°C. What is the mass of the propane in the flask?

Gas Stoichiometry (14.4)

101. Ammonia is formed industrially by reacting nitrogen and hydrogen gases. How many liters of ammonia gas can be formed from 13.7 L of hydrogen gas at 93.0°C and a pressure of 40.0 kPa?

102. When 3.00 L of propane gas is completely combusted to form water vapor and carbon dioxide at a temperature of 350°C and a pressure of 0.990 atm, what mass of water vapor will result?

103. When heated, solid potassium chlorate ($KClO_3$) decomposes to form solid potassium chloride and oxygen gas. If 20.8 g of potassium chlorate decomposes, how many liters of oxygen gas will form at STP?

104. Use the reaction shown below to answer these questions.

$$CO(g) + NO(g) \longrightarrow N_2(g) + CO_2(g)$$

a. Balance the equation.

b. What is the volume ratio of carbon monoxide to carbon dioxide in the balanced equation?

c. If 42.7 g CO is reacted completely at STP, what volume of N_2 gas will be produced?

Mixed Review

Sharpen your problem-solving skills by answering the following.

105. Gaseous methane (CH_4) undergoes complete combustion by reacting with oxygen gas to form carbon dioxide and water vapor.

a. Write a balanced equation for this reaction.

b. What is the volume ratio of methane to water in this reaction?

106. If 2.33 L of propane at 24°C and 67.2 kPa is completely burned in excess oxygen, how many moles of carbon dioxide will be produced?

107. Use Boyle's, Charles's, or Gay-Lussac's law to calculate the missing value in each of the following.

a. V_1 = 1.4. L, P_1 = ?, V_2 = 3.0 L, P_2 = 1.2 atm

b. V_1 = 705 mL, T_1 = 273 K, V_2 = ?, T_2 = 323 K

c. V_1 = 0.540 L, P_1 = ?, V_2 = 0.990 L, P_2 = 775 mm Hg

d. T_1 = 37°C, P_1 = 5.0 atm, P_2 = 2.5 atm, T_2 = ?

108. Determine the pressure inside a television picture tube with a volume of 3.50 L that contains 2.00×10^{-5} g of nitrogen gas at 22.0°C.

109. Determine how many liters 8.80 g of carbon dioxide gas would occupy at:

a. STP

b. 160°C and 3.00 atm

c. 288 K and 118 kPa

110. If 5.00 L of hydrogen gas, measured at 20.0ºC and 80.1 kPa is burned in excess oxygen to form water, what mass of oxygen (measured at the same temperature and pressure) will be consumed?

Thinking Critically

111. Making and Using Graphs Automobile tires become underinflated as temperatures drop during the winter months if no additional air is added to the tires at the start of the cold season. For every 10°F drop in temperature, the air pressure in a car's tires goes down by about 1 psi (14.7 psi equals 1.00 atm). Complete the following table. Then make a graph illustrating how the air pressure in a tire changes over the temperature range from 40°F to −10°F, assuming you start with a pressure of 30.0 psi at 40°F.

Tire Inflation Based on Temperature

Temperature (°F)	Pressure (psi)
40	
30	
20	
10	
0	
−10	

112. Applying Concepts When nitroglycerin ($C_3H_5N_3O_9$) explodes, it decomposes into the following gases: CO_2, N_2, NO, and H_2O. If 239 g of nitroglycerin explodes, what volume will the mixture of gaseous products occupy at 1.00 atm pressure and 2678°C?

113. Analyze and Conclude What is the numerical value of the ideal gas constant (R) in $\frac{cm^3 \cdot Pa}{K \cdot mol}$

114. Applying Concepts Calculate the pressure of a mixture of two gases that contains 4.67×10^{22} molecules CO and 2.87×10^{24} molecules of N_2 in a 6.00-L container at 34.8°C.

Writing in Chemistry

115. It was the dream of many early balloonists to complete a trip around the world in a hot-air balloon, a goal not achieved until 1999. Write about what you imagine a trip in a balloon would be like, including a description of how manipulating air temperature would allow you to control altitude.

116. Investigate and explain the function of the regulators on the air tanks used by scuba divers.

Cumulative Review

Refresh your understanding of previous chapters by answering the following.

117. Convert each of the following mass measurements to its equivalent in kilograms. (Chapter 2)

a. 247 g
b. 53 Mg
c. 7.23 μg
d. 975 mg

118. How many atoms of each element are present in five formula units of calcium permanganate? (Chapter 8)

119. Terephthalic acid is an organic compound used in the formation of polyesters. It contains 57.8 percent C, 3.64 percent H, and 38.5 percent O. The molar mass is known to be approximately 166 g/mol. What is the molecular formula of terephthalic acid? (Chapter 11)

120. The particles of which of the following gases have the highest average speed? The lowest average speed? (Chapter 13)

a. carbon monoxide at 90°C
b. nitrogen trifluoride at 30°C
c. methane at 90°C
d. carbon monoxide at 30°C

STANDARDIZED TEST PRACTICE
CHAPTER 14

Use these questions and the test-taking tip to prepare for your standardized test.

1. The kinetic-molecular theory describes the microscopic behavior of gases. One main point of the theory is that within a sample of gas, the frequency of collisions between individual gas particles and between the particles and the walls of their container increases if the sample is compressed. The gas law that states this relationship in mathematical terms is

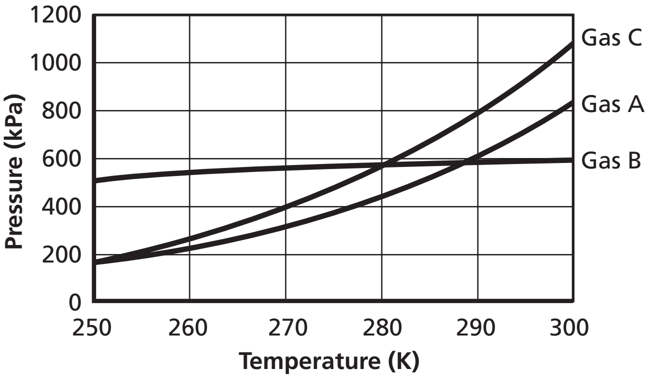
a. Gay-Lussac's Law.
b. Charles's Law.
c. Boyle's Law.
d. Avogadro's Law.

2. Three 2.0-L containers are placed in a 50°C room. Samples of 0.5 mol N_2, 0.5 mol Xe, and 0.5 mol ethene (C_2H_4) are pumped into Containers 1, 2, and 3, respectively. Inside which container will the pressure be greatest?

a. Container 2
b. Container 3
c. Containers 2 and 3 have the same, higher pressure
d. Containers 1, 2, and 3 have equal pressures

Interpreting Graphs Use the graph to answer questions 3–5.

Pressures of Three Gases at Different Temperatures



3. It can be seen from the graph that

a. as temperature increases, pressure decreases.
b. as pressure increases, volume decreases.
c. as temperature decreases, moles decrease.
d. as pressure decreases, temperature decreases.

4. Which of these gases is an ideal gas?

a. Gas A
b. Gas B
c. Gas C
d. none of the above

5. What is the predicted pressure of Gas B at 310 K?

a. 260 kPa
b. 620 kPa
c. 1000 kPa
d. 1200 kPa

6. What volume will 0.875 moles of SF_4 occupy at STP?

a. 19.6 L
b. 21.4 L
c. 22.4 L
d. 32.7 L

7. While it is on the ground, a blimp is filled with 5.66×10^6 L of He gas. The pressure inside the grounded blimp, where the temperature is 25°C, is 1.10 atm. Modern blimps are non-rigid, which means that their volume is changeable. If the pressure inside the blimp remains the same, what will be the volume of the blimp at a height of 2300 m, where the temperature is 12°C?

a. 5.66×10^6 L
b. 2.72×10^6 L
c. 5.4×10^6 L
d. 5.92×10^6 L

8. The reaction that provides blowtorches with their intense flame is the combustion of acetylene (C_2H_2) to form carbon dioxide and water vapor. Assuming that the pressure and temperature of the reactants are the same, what volume of oxygen gas is required to completely burn 5.60 L of acetylene?

a. 2.24 L
b. 5.60 L
c. 11.2 L
d. 14.0 L

9. A sample of argon gas is compressed into a volume of 0.712 L by a piston exerting 3.92 atm of pressure. The piston is slowly released until the pressure of the gas is 1.50 atm. What is the new volume of the gas?

a. 0.272 L
b. 3.67 L
c. 1.86 L
d. 4.19L

10. Assuming ideal behavior, how much pressure will 0.0468 g of ammonia (NH_3) gas exert on the walls of a 4.00-L container at 35.0°C?

a. 0.0174 atm
b. 0.296 atm
c. 0.00198 atm
d. 0.278 atm

TEST-TAKING TIP

Ask Questions If you've got a question about what will be on the test, the way the test is scored, the time limits placed on each section, or anything else. . .by all means ask! Will you be required to know the specific names of the gas laws, such as Boyle's law and Charles's law?