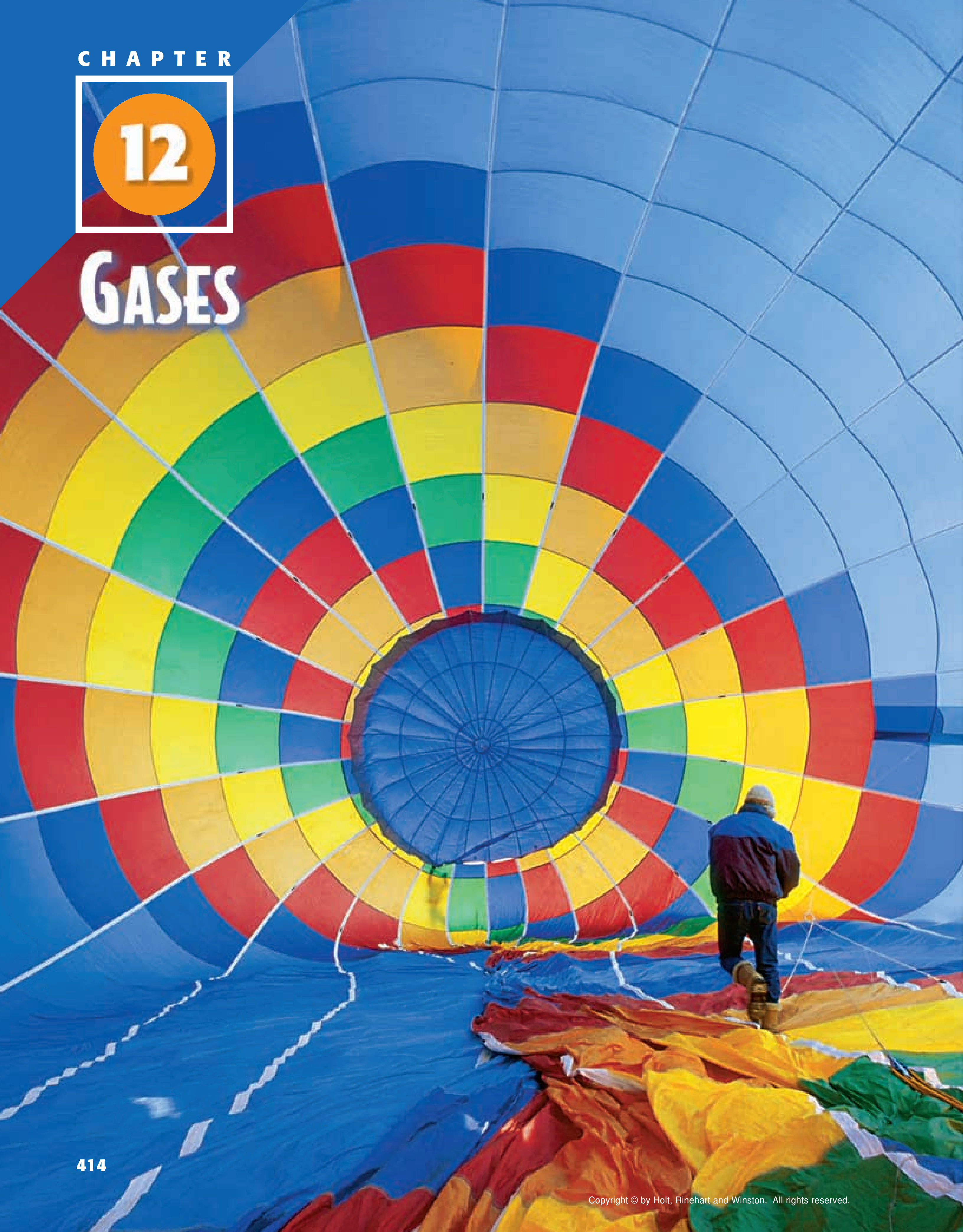
CHAPTER 12

GASES



The hot-air balloon pictured here is being filled with hot air. Hot air expands, so the air in the balloon will be less dense. The balloon will be lifted up by the cooler, denser air outside it. The hot-air balloon demonstrates some important facts about gases: gases expand when heated, they have weight, and they have mass.

START-UP ACTIVITY

SAFETY PRECAUTIONS

Pressure Relief

PROCEDURE

1. Blow up a **round latex balloon,** and let the air out.
2. Put the round part of the balloon inside a **PET bottle.** Roll the neck of the balloon over the mouth of the bottle, and secure the neck of the balloon with a **rubber band** so that it dangles into the bottle.
3. Try to blow up the balloon. Record the results.
4. Answer the first two analysis questions below.
5. Design a modification to the balloon-in-a-bottle apparatus that will allow the balloon to inflate. Answer the third analysis question below.
6. If your teacher approves of your design, try it out.

ANALYSIS

1. What causes the balloon to expand?
2. What happens to air that is trapped in the bottle when you blow into the balloon?
3. Describe your design modification, and include a sketch, if applicable. Explain why your design works.

Pre-Reading Questions

1. **Among the states of matter, what is unique about gases?**
2. **Do gases have mass and weight? How can you tell?**
3. **Gases are considered fluids. Why?**

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SECTION

1 Characteristics of Gases

KEY TERMS

- **pressure**
- **newton**
- **pascal**
- **standard temperature and pressure**
- **kinetic-molecular theory**

OBJECTIVES

1. **Describe** the general properties of gases.
2. **Define** pressure, give the SI unit for pressure, and convert between standard units of pressure.
3. **Relate** the kinetic-molecular theory to the properties of an ideal gas.

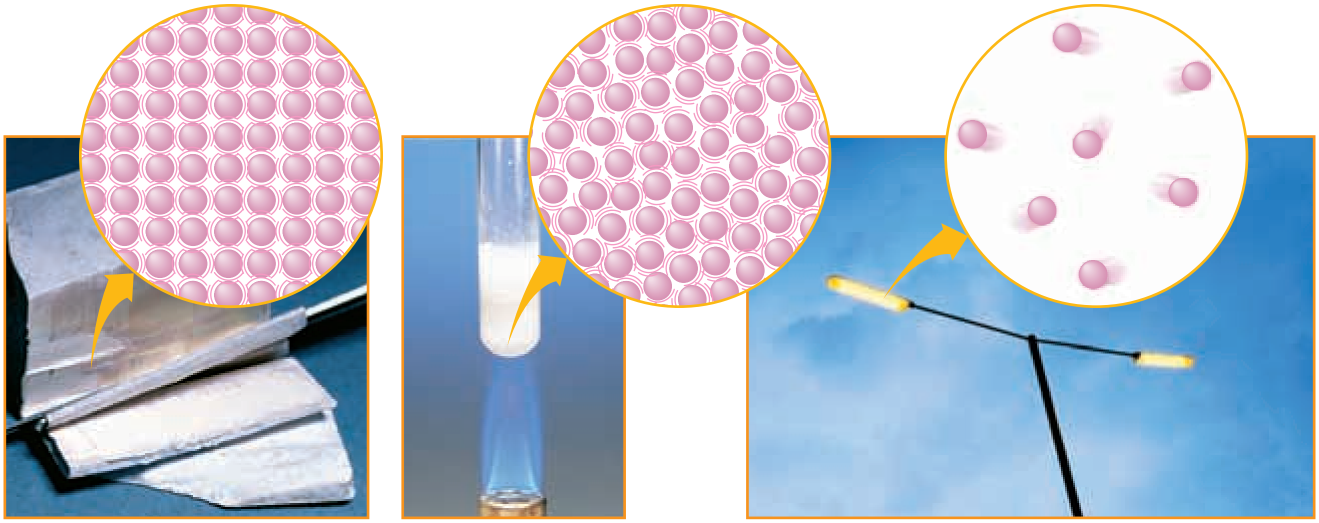
Properties of Gases

Each state of matter has its own properties. Gases have unique properties because the distance between the particles of a gas is much greater than the distance between the particles of a liquid or a solid. Although liquids and solids seem very different from each other, both have small intermolecular distances. Gas particles, however, are much farther apart from each other than liquid and solid particles are. In some ways, gases behave like liquids; in other ways, they have unique properties.

Gases Are Fluids

Gases are considered *fluids*. People often use the word *fluid* to mean "liquid." However, the word *fluid* actually means "any substance that can flow." Gases are fluids because they are able to flow. Gas particles can flow because they are relatively far apart and therefore are able to move past each other easily. In **Figure 1,** a strip of copper is reacting with nitric acid to form nitrogen dioxide, a brown gas. Like all gases, nitrogen dioxide is a fluid. The gas flows over the sides of the beaker.

Figure 1
The reaction in the beaker in this photo has formed NO_2, a brown gas, which flows out of the container.



Gases Have Low Density

Gases have much lower densities than liquids and solids do. Because of the relatively large distances between gas particles, most of the volume occupied by a gas is empty space. As shown in **Figure 2,** particles in solids and liquids are almost in contact with each other, but gas particles are much farther apart. This distance between particles shows why a substance in the liquid or solid state always has a much greater density than the same substance in the gaseous state does. The low density of gases also means that gas particles travel relatively long distances before colliding with each other.

Figure 2
Particles of sodium in the solid, liquid, and gas phases are shown. The atoms of gaseous sodium are much farther apart, as in a sodium vapor street lamp.

Gases Are Highly Compressible

Suppose you completely fill a syringe with liquid and try to push the plunger in when the opening is plugged. You cannot make the space the liquid takes up become smaller. It takes very great pressure to reduce the volume of a liquid or solid. However, if only gas is in the syringe, with a little effort you can move the plunger down and compress the gas. As shown in **Figure 3,** gas particles can be pushed closer together. The space occupied by the gas particles themselves is very small compared with the total volume of the gas. Therefore, applying a small pressure will move the gas particles closer together and will decrease the volume.

Figure 3
The volume occupied by a gas can be reduced because gas molecules can move closer together.

Figure 4
When the partition between the gas and the empty container (vacuum) is removed, the gas flows into the empty container and fills the entire volume of both containers.

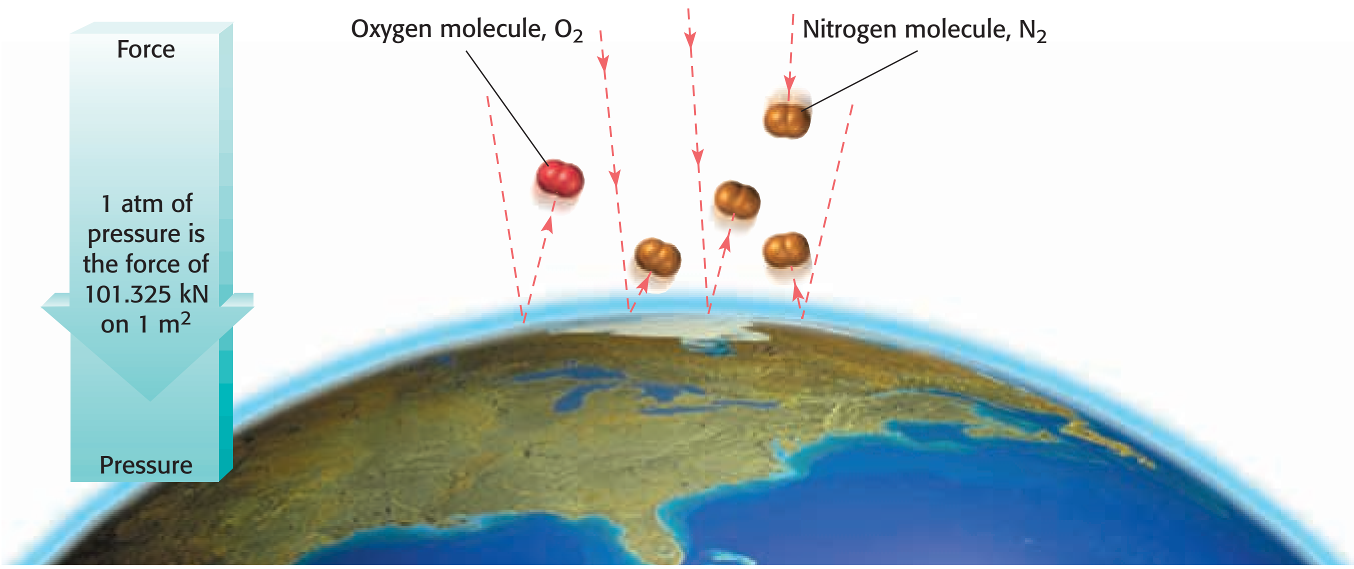
Gases Completely Fill a Container

A solid has a certain shape and volume. A liquid has a certain volume but takes the shape of the lower part of its container. In contrast, a gas completely fills its container. This principle is shown in **Figure 4.** One of the containers contains a gas, and the other container is empty. When the barrier between the two containers is removed, the gas rushes into the empty container until it fills both containers equally. Gas particles are constantly moving at high speeds and are far apart enough that they do not attract each other as much as particles of solids and liquids do. Therefore, a gas expands to fill the entire volume available.

Gas Pressure

Gases are all around you. Earth's atmosphere, commonly known as *air,* is a mixture of gases: mainly nitrogen and oxygen. Because you cannot always feel air, you may have thought of gases as being weightless, but all gases have mass; therefore, they have weight in a gravitational field. As gas molecules are pulled toward the surface of Earth, they collide with each other and with the surface of Earth more often, as shown in **Figure 5.** Collisions of gas molecules are what cause *air pressure.*

Figure 5
Gas molecules in the atmosphere collide with Earth's surface, creating atmospheric pressure.



You may have noticed that your ears sometimes "pop" when you ascend to high altitudes or fly in an airplane. This popping happens because the density of the air changes when you change altitudes. The atmosphere is denser as you move closer to Earth's surface because the weight of atmospheric gases at any elevation compresses the gases below. Less-dense air exerts less pressure. Your ears pop when the air inside your ears changes to the same pressure as the outside air.

Measuring Pressure

The scientific definition of **pressure** is "force divided by area." To find pressure, you need to know the force and the area over which that force is exerted.

The unit of force in SI units is the **newton.** One newton is the force that gives an acceleration of 1 m/s^2 to an object whose mass is 1 kg.

$$1 \text{ newton} = 1 \text{ kg} \times 1 \text{ m/s}^2 = 1 \text{ N}$$

The SI unit of pressure is the **pascal,** Pa, which is the force of one newton applied over an area of one square meter.

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

One pascal is a small unit of pressure. It is the pressure exerted by a layer of water that is 0.102 mm deep over an area of one square meter.

Atmospheric pressure can be measured by a barometer, as shown in **Figure 6.** The atmosphere exerts pressure on the surface of mercury in the dish. This pressure goes through the fluid and up the column of mercury. The mercury settles at a point where the pressure exerted downward by its weight equals the pressure exerted by the atmosphere.

pressure
the amount of force exerted per unit area of surface

newton
the SI unit for force; the force that will increase the speed of a 1 kg mass by 1 m/s each second that the force is applied (abbreviation, N)

pascal
the SI unit of pressure; equal to the force of 1 N exerted over an area of 1 m^2 (abbreviation, Pa)

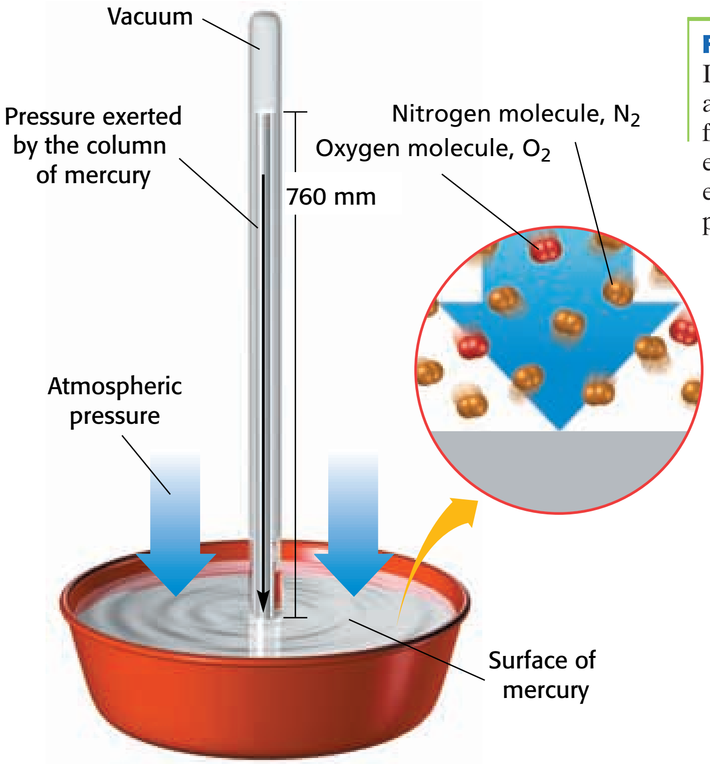


Figure 6
In the sealed column of a barometer, the mercury falls until the pressure exerted by its weight equals the atmospheric pressure.

Table 1 Pressure Units

Unit	Abbreviation	Equivalent number of pascals
Atmosphere	atm	1 atm = 101 325 Pa
Bar	bar	1 bar = 100 025 Pa
Millimeter of mercury	mm Hg	1 mm Hg = 133.322 Pa
Pascal	Pa	1
Pounds per square inch	psi	1 psi = $6.892\ 86 \times 10^3$ Pa
Torr	torr	1 torr = 133.322 Pa

At sea level, the atmosphere keeps the mercury in a barometer at an average height of 760 mm, which is 1 atmosphere. One millimeter of mercury is also called a *torr,* after Evangelista Torricelli, the Italian physicist who invented the barometer. Other units of pressure are listed in **Table 1.**

In studying the effects of changing temperature and pressure on a gas, one will find a standard for comparison useful. Scientists have specified a set of standard conditions called **standard temperature and pressure,** or STP, which is equal to 0°C and 1 atm.

standard temperature and pressure

for a gas, the temperature of 0°C and the pressure 1.00 atmosphere

SAMPLE PROBLEM A

Converting Pressure Units

Convert the pressure of 1.000 atm to millimeters of mercury.

1 Gather information.

From **Table 1,** 1 atmosphere = 101 325 Pa, 1 mm Hg = 133.322 Pa

2 Plan your work.

Both units of pressure can be converted to pascals, so use pascals to convert atmospheres to millimeters of mercury.

The conversion factors are $\frac{101\ 325 \text{ Pa}}{1 \text{ atm}}$ and $\frac{1 \text{ mm Hg}}{133.322 \text{ Pa}}$.

3 Calculate.

$$1.000 \text{ atm} \times \frac{101\ 325 \text{ Pa}}{1 \text{ atm}} \times \frac{1 \text{ mm Hg}}{133.322 \text{ Pa}} = 760.0 \text{ mm Hg}$$

4 Verify your results.

By looking at **Table 1,** you can see that about 100 000 pascals equal one atmosphere, and the number of pascals that equal 1 mm Hg is just over 100. Therefore, the number of millimeters of mercury equivalent to 1 atm is just below 1000 (100 000 divided by 100). The answer is therefore reasonable.

PRACTICE HINT

Remember that millimeters of mercury is actually a unit of pressure, so it cannot be canceled by a length measurement.

PRACTICE

1. The critical pressure of carbon dioxide is 72.7 atm. What is this value in units of pascals?
2. The vapor pressure of water at 50.0°C is 12.33 kPa. What is this value in millimeters of mercury?
3. In thermodynamics, the standard pressure is 100.0 kPa. What is this value in units of atmospheres?
4. A tire is inflated to a gauge pressure of 30.0 psi (Which must be added to the atmospheric pressure of 14.7 psi to find the total pressure in the tire). Calculate the total pressure in the tire in pascals.

PROBLEM SOLVING SKILL

The Kinetic-Molecular Theory

The properties of gases stated earlier are explained on the molecular level in terms of the **kinetic-molecular theory.** The kinetic-molecular theory is a model that is used to predict gas behavior.

The kinetic-molecular theory states that gas particles are in constant rapid, random motion. The theory also states that the particles of a gas are very far apart relative to their size. This idea explains the fluidity and compressibility of gases. Gas particles can easily move past one another or move closer together because they are farther apart than liquid or solid particles.

Gas particles in constant motion collide with each other and with the walls of their container. The kinetic-molecular theory states that the pressure exerted by a gas is a result of collisions of the molecules against the walls of the container, as shown in **Figure 7.** The kinetic-molecular theory considers collisions of gas particles to be perfectly elastic; that is, energy is completely transferred during collisions. The total energy of the system, however, remains constant.

kinetic-molecular theory
a theory that explains that the behavior of physical systems depends on the combined actions of the molecules constituting the system

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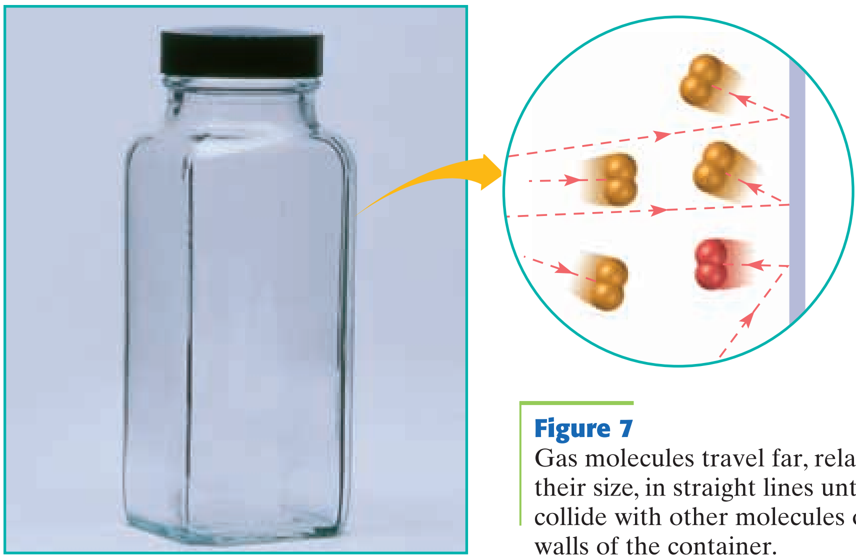
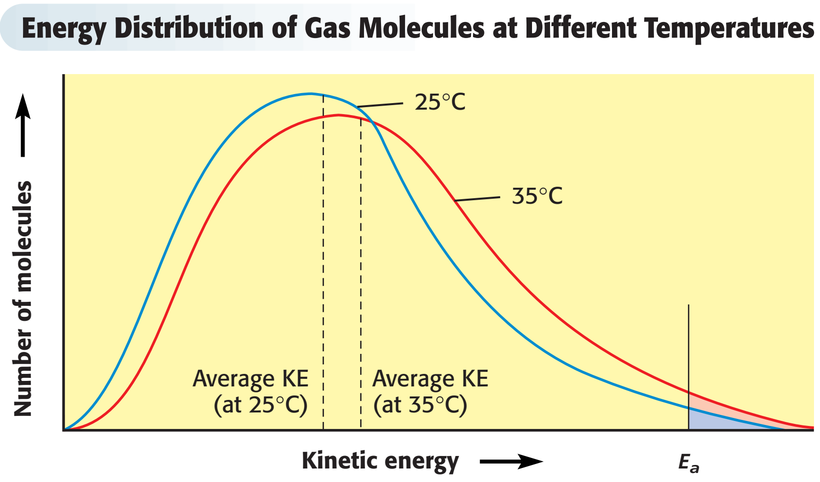


Figure 7
Gas molecules travel far, relative to their size, in straight lines until they collide with other molecules or the walls of the container.

Figure 8
Increasing the temperature of a gas shifts the energy distribution in the direction of greater average kinetic energy.



Gas Temperature Is Proportional to Average Kinetic Energy

Molecules are always in random motion. The average kinetic energy of random motion is proportional to the absolute temperature, or temperature in kelvins. Heat increases the energy of random motion of a gas.

But not all molecules are traveling at the same speed. As a result of multiple collisions, the molecules have a range of speeds. **Figure 8** shows the numbers of molecules at various energies. For a 10°C rise in temperature from STP, the average energy increases about 3%, while the number of very high-energy molecules approximately doubles or triples.

1 Section Review

UNDERSTANDING KEY IDEAS

1. What characteristic of gases makes them different from liquids or solids?

2. Why are gases considered fluids?

3. What happens to gas particles when a gas is compressed?

4. What is the difference between force and pressure?

5. What is the SI unit of pressure, and how is it defined?

6. How does the kinetic-molecular theory explain the pressure exerted by gases?

7. How is a gas's ability to fill a container different from that of a liquid or a solid?

PRACTICE PROBLEMS

8. The atmospheric pressure on top of Mount Everest is 58 kPa. What is this pressure in atmospheres?

9. The vapor pressure of water at 0°C is 4.579 mm Hg. What is this pressure in pascals?

10. A laboratory high-vacuum system may operate at 1.0×10^{-5} mm Hg. What is this pressure in pascals?

CRITICAL THINKING

11. How does the kinetic-molecular theory explain why atmospheric pressure is greater at lower altitudes than at higher altitudes?

12. Molecules of hydrogen escape from Earth, but molecules of oxygen and nitrogen are held to the surface. Why?

SECTION 2

The Gas Laws

KEY TERMS

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

OBJECTIVES

1. **State** Boyle's law, and use it to solve problems involving pressure and volume.
2. **State** Charles's law, and use it to solve problems involving volume and temperature.
3. **State** Gay-Lussac's law, and use it to solve problems involving pressure and temperature.
4. **State** Avogadro's law, and explain its importance in determining the formulas of chemical compounds.

Measurable Properties of Gases

In this section, you will study the relationships between the measurable properties of a gas, represented by the variables shown below.

P = pressure exerted by the gas

T = temperature in kelvins of the gas

V = total volume occupied by the gas

n = number of moles of the gas

Pressure-Volume Relationships

As you read in the last section, gases have pressure, and the space that they take up can be made smaller. In 1662, the English scientist Robert Boyle studied the relationship between the volume and the pressure of a gas. He found that as pressure on a gas increases in a closed container, the volume of the gas decreases. In fact, the product of the pressure and volume, PV, remains almost constant if the temperature remains the same. **Table 2** shows data for experiments similar to Boyle's.

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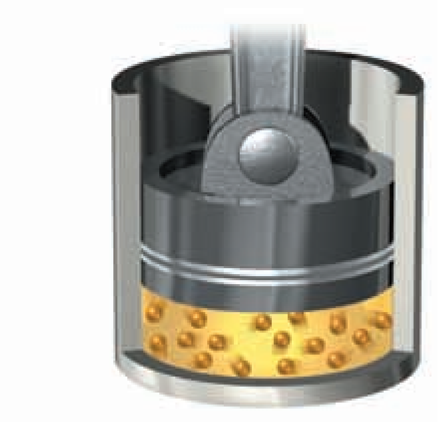
Table 2 Pressure-Volume Data for a Sample of Gas at Constant Temperature

Pressure (kPa)	Volume (L)	PV (kPa × L)
150	0.334	50.1
200	0.250	50.0
250	0.200	50.0
300	0.167	50.1

Figure 9

a Gas molecules in a car-engine cylinder spread apart to fill the cylinder.



b As the cylinder's volume decreases, there are more molecular collisions, and the pressure of the gas increases.

Boyle's Law

Figure 9 shows what happens to the gas molecules in a car cylinder as it compresses. As the volume decreases, the concentration, and therefore pressure, increases. This concept is shown in graphical form in **Figure 10.** The inverse relationship between pressure and volume is known as **Boyle's law.** The third column of **Table 2** shows that at constant temperature, the product of the pressure and volume of a gas is constant.

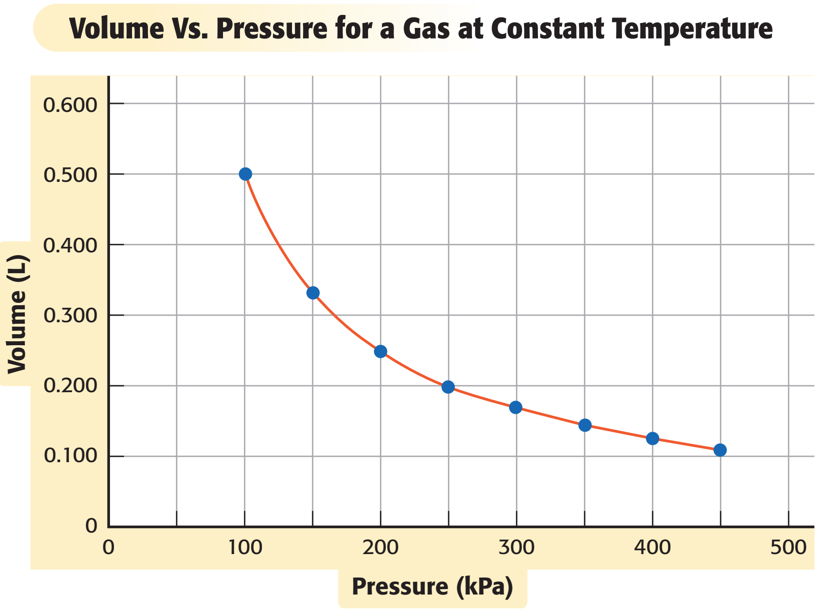
Boyle's law
the law that states that for a fixed amount of gas at a constant temperature, the volume of the gas increases as the pressure of the gas decreases and the volume of the gas decreases as the pressure of the gas increases

$$PV = k$$

If the temperature and number of particles are not changed, the PV product remains the same, as shown in the equation below.

$$P_1V_1 = P_2V_2$$

Figure 10
This pressure-volume graph shows an inverse relationship: as pressure increases, volume decreases.



SAMPLE PROBLEM B

Solving Pressure-Volume Problems

A given sample of gas occupies 523 mL at 1.00 atm. The pressure is increased to 1.97 atm, while the temperature remains the same. What is the new volume of the gas?

1 Gather information.

The initial volume and pressure and the final pressure are given. Determine the final volume.

$P_1 = 1.00$ atm $\quad\quad V_1 = 523$ mL

$P_2 = 1.97$ atm $\quad\quad V_2 = ?$

2 Plan your work.

Place the known quantities into the correct places in the equation relating pressure and volume.

$$P_1V_1 = P_2V_2$$

$$(1.00 \text{ atm})(523 \text{ mL}) = (1.97 \text{ atm})V_2$$

3 Calculate.

$$V_2 = \frac{(1.00 \text{ atm})(523 \text{ mL})}{1.97 \text{ atm}} = 265 \text{ mL}$$

4 Verify your results.

The pressure was almost doubled, so the new volume should be about one-half the initial volume. The answer is therefore reasonable.

PRACTICE HINT

It does not matter what units you use for pressure and volume when using Boyle's law as long as they are the same on both sides of the equation.

PRACTICE

PROBLEM SOLVING SKILL

1. A sample of oxygen gas has a volume of 150.0 mL at a pressure of 0.947 atm. What will the volume of the gas be at a pressure of 1.000 atm if the temperature remains constant?
2. A sample of gas in a syringe has a volume of 9.66 mL at a pressure of 64.4 kPa. The plunger is depressed until the pressure is 94.6 kPa. What is the new volume, assuming constant temperature?
3. An air mass of volume 6.5×10^5 L starts at sea level, where the pressure is 775 mm Hg. It rises up a mountain where the pressure is 622 mm Hg. Assuming no change in temperature, what is the volume of the air mass?
4. A balloon has a volume of 456 mL at a pressure of 1.0 atm. It is taken under water in a submarine to a depth where the air pressure in the submarine is 3.3 atm. What is the volume of the balloon? Assume constant temperature.

Temperature-Volume Relationships

Heating a gas makes it expand. Cooling a gas makes it contract. This principle is shown in **Figure 11.** As balloons are dipped into liquid nitrogen, the great decrease in temperature makes them shrink. When they are removed from the liquid nitrogen, the gas inside the balloons warms up, and the balloons expand to their original volume.

In 1787, the French physicist Jacques Charles discovered that a gas's volume is directly proportional to the temperature on the Kelvin scale if the pressure remains the same.

Charles's Law

Charles's law

the law that states that for a fixed amount of gas at a constant pressure, the volume of the gas increases as the temperature of the gas increases and the volume of the gas decreases as the temperature of the gas decreases

The direct relationship between temperature and volume is known as **Charles's law.** The kinetic-molecular theory states that gas particles move faster on average at higher temperatures, causing them to hit the walls of their container with more force. Repeated strong collisions cause the volume of a flexible container, such as a balloon, to increase. Likewise, gas volume decreases when the gas is cooled, because of the lower average kinetic energy of the gas particles at the lower temperature. If the absolute temperature is reduced by half, then the average kinetic energy is reduced by half, and the particles will strike the walls with half of the energy they had at the higher temperature. In that case, the volume of the gas will be reduced to half of the original volume if the pressure remains the same.

The direct relationship between volume and temperature is shown in **Figure 12,** in which volume-temperature data are graphed using the Kelvin scale. If you read the line in **Figure 12** all the way down to 0 K, it looks as though the gas's volume becomes zero. Does a gas volume really become zero at absolute zero? No. Before this temperature is reached, the gas becomes a liquid and then freezes to a solid, each of which has a certain volume.

Figure 11

a Air-filled balloons are dipped into liquid nitrogen.

b The extremely low temperature of the nitrogen causes them to shrink in volume.

c When the balloons are removed from the liquid nitrogen, the air inside them quickly warms, and the balloons expand to their original volume.

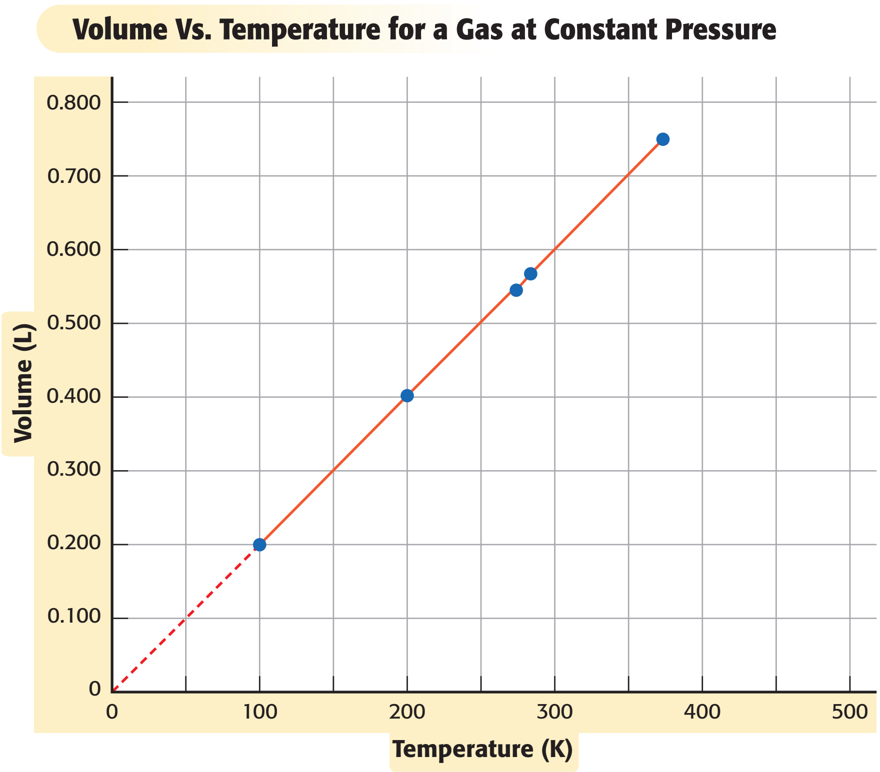


Figure 12
When the temperature scale is in kelvins, the graph shows a direct proportion between volume of a sample of gas and the temperature.

Data for an experiment of the type carried out by Charles are given in **Table 3.** At constant pressure, the volume of a sample of gas divided by its absolute temperature is a constant, k. Charles's law can be stated as the following equation.

$$\frac{V}{T} = k$$

If all other conditions are kept constant, V/T will remain the same. Therefore, Charles's law can also be expressed as follows.

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

Table 3 Volume-Temperature Data for a Sample of Gas at a Constant Pressure

Volume (mL)	Temperature (K)	V/T (mL/K)
748	373	2.01
567	283	2.00
545	274	1.99
545	273	2.00
546	272	2.01
402	200	2.01
199	100	1.99

SAMPLE PROBLEM C

Solving Volume-Temperature Problems

A balloon is inflated to 665 mL volume at 27°C. It is immersed in a dry-ice bath at −78.5°C. What is its volume, assuming the pressure remains constant?

1 Gather information.

The initial volume and temperature and the final temperature are given. Determine the final volume.

$V_1 = 665$ mL $\qquad T_1 = 27°C$

$V_2 = ?$ $\qquad T_2 = -78.5°C$

2 Plan your work.

Convert the temperatures from degrees Celsius to kelvins:

$T_1 = 27°C + 273 = 300$ K $\qquad T_2 = -78.5°C + 273 = 194.5$ K

Place the known quantities into the correct places in the equation relating volume and temperature.

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

$$\frac{665 \text{ mL}}{300 \text{ K}} = \frac{V_2}{194.5 \text{ K}}$$

3 Calculate.

$$V_2 = \frac{(665 \text{ mL})(194.5 \text{ K})}{300 \text{ K}} = 431 \text{ mL}$$

4 Verify your results.

Charles's law tells you that volume decreases as temperature decreases. The temperature decreased by about one-third, and according to the calculation, so did the volume. The answer is therefore reasonable.

PRACTICE HINT

In gas law problems, always convert temperatures to kelvins. The gas law equations do not work for temperatures expressed in the Celsius or Fahrenheit scales.

PRACTICE

PROBLEM SOLVING SKILL

1. Helium gas in a balloon occupies 2.5 L at 300.0 K. The balloon is dipped into liquid nitrogen that is at a temperature of 80.0 K. What will the volume of the helium in the balloon at the lower temperature be?
2. A sample of neon gas has a volume of 752 mL at 25.0°C. What will the volume at 50.0°C be if pressure is constant?
3. A helium-filled balloon has a volume of 2.75 L at 20.0°C. The volume of the balloon changes to 2.46 L when placed outside on a cold day. What is the temperature outside in degrees Celsius?
4. When 1.50×10^3 L of air at 5.00°C is injected into a household furnace, it comes out at 30.0°C. Assuming the pressure is constant, what is the volume of the heated air?

Figure 13
As the temperature of a gas increases, the average kinetic energy of the molecules increases. This means that at constant volume, pressure increases with temperature.

Temperature-Pressure Relationships

As you have learned, pressure is the result of collisions of particles with the walls of the container. You also know that the average kinetic energy of particles is proportional to the sample's average absolute temperature. Therefore, if the absolute temperature of the gas particles is doubled, their average kinetic energy is doubled. If there is a fixed amount of gas in a container of fixed volume, the collisions will have twice the energy, so the pressure will double, as shown in **Figure 13.** The French scientist Joseph Gay-Lussac is given credit for discovering this relationship in 1802. **Figure 14** shows a graph of data for the change of pressure with temperature in a gas of constant volume. The graph is a straight line with a positive slope, which indicates that temperature and pressure have a directly proportional relationship.

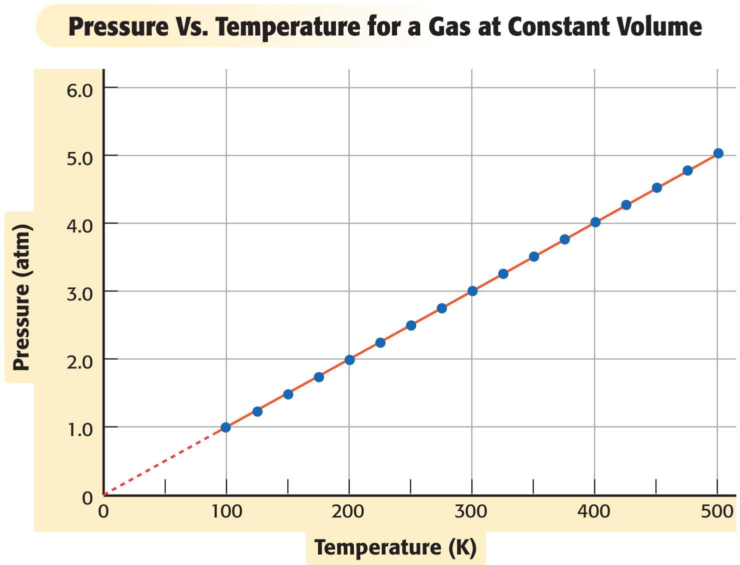


Figure 14
This graph shows that gas pressure is directly proportional to Kelvin temperature, at constant volume.

Gay-Lussac's Law

Gay-Lussac's law

the law that states that the pressure of a gas at a constant volume is directly proportional to the absolute temperature

The direct relationship between temperature and pressure is known as **Gay-Lussac's law.** Because the pressure of a gas is proportional to its absolute temperature, the following equation is true for a sample of constant volume.

$$P = kT$$

This equation can be rearranged to the following form.

$$\frac{P}{T} = k$$

At constant volume, the following equation applies.

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

If any three of the variables in the above equation are known, then the unknown fourth variable can be calculated.

SAMPLE PROBLEM D

Solving Pressure-Temperature Problems

An aerosol can containing gas at 101 kPa and 22°C is heated to 55°C. Calculate the pressure in the heated can.

1 Gather information.

The initial pressure and temperature and the final temperature are given. Determine the final pressure.

$P_1 = 101$ kPa $\qquad T_1 = 22°C$

$P_2 = ?$ $\qquad T_2 = 55°C$

PRACTICE HINT

When solving gas problems, be sure to identify which variables (P, V, n, T) remain constant and which change. That way you can pick the right equation to use.

2 Plan your work.

Convert the temperatures from degrees Celsius to kelvins:

$T_1 = 22°C + 273 = 295$ K $\qquad T_2 = 55°C + 273 = 328$ K

Use the equation relating temperature and pressure.

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

$$\frac{101 \text{ kPa}}{295 \text{ K}} = \frac{P_2}{328 \text{ K}}$$

3 Calculate.

$$P_2 = \frac{(101 \text{ kPa})(328 \text{ K})}{295 \text{ K}} = 113 \text{ kPa}$$

4 Verify your results.

The temperature of the gas increases by about 10%, so the pressure should increase by the same proportion.

PRACTICE

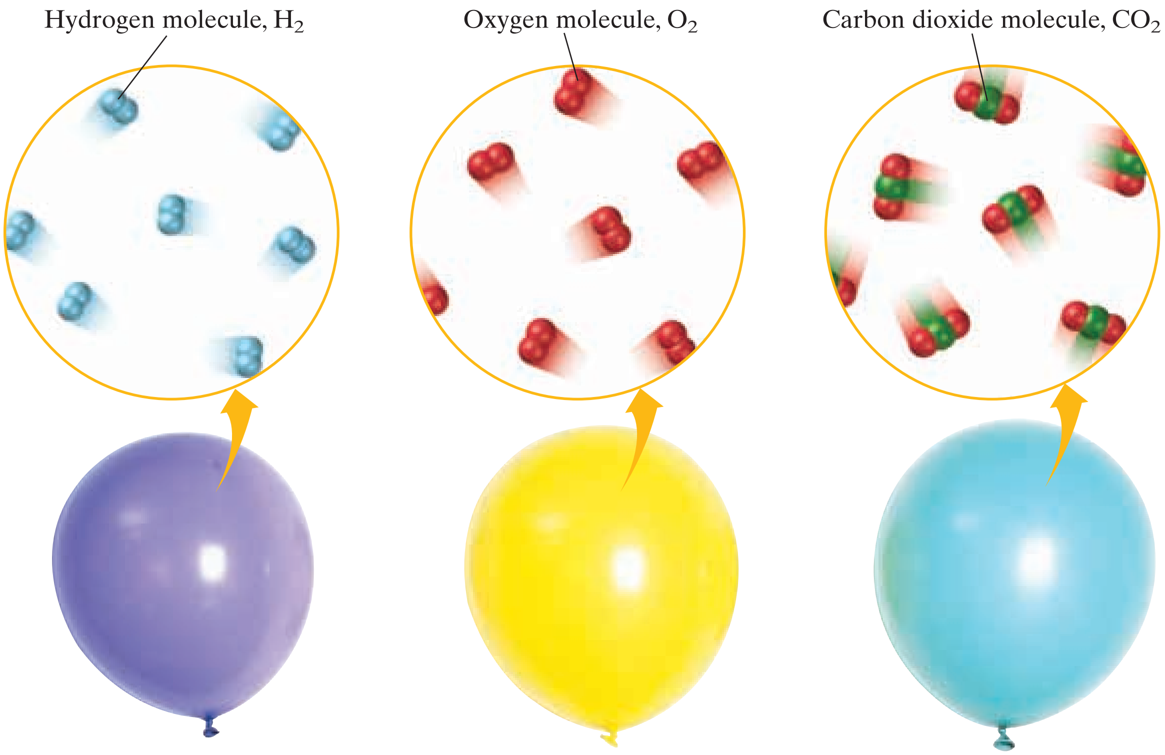
PROBLEM SOLVING SKILL

1. At 122°C the pressure of a sample of nitrogen is 1.07 atm. What will the pressure be at 205°C, assuming constant volume?
2. The same sample of nitrogen as in item 1 starts at 122°C and 1.07 atm. After cooling, the pressure is measured to be 0.880 atm. What is the new temperature?
3. A sample of helium gas is at 122 kPa and 22°C. Assuming constant volume, what will the temperature be when the pressure is 203 kPa?
4. The air in a steel-belted tire is at a gauge pressure of 29.8 psi at a temperature of 20°C. After the tire is driven fast on a hot road, the temperature in the tire is 48°C. What is the tire's new gauge pressure?

Volume-Molar Relationships

In 1811, the Italian scientist Amadeo Avogadro proposed the idea that equal volumes of all gases, under the same conditions, have the same number of particles. This idea is shown in **Figure 15,** which shows equal numbers of molecules of the gases H_2, O_2, and CO_2, each having the same volume. A result of this relationship is that molecular masses can be easily determined. Unfortunately, Avogadro's insight was not recognized right away, mainly because scientists at the time did not know the difference between atoms and molecules. Later, the Italian chemist Stanislao Cannizzaro used Avogadro's principle to determine the true formulas of several gaseous compounds.

Figure 15
At the same temperature and pressure, balloons of equal volume contain equal numbers of molecules, regardless of which gas they contain.



Avogadro's Law

Avogadro's law

the law that states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules

Avogadro's idea turned out to be correct and is now known as **Avogadro's law.** With this knowledge, chemists gained insight into the formulas of chemical compounds for the first time. In 1858, Cannizzaro used Avogadro's law to deduce that the correct formula for water is H_2O. This important discovery will be discussed in more detail in the next section.

Avogadro's law also means that gas volume is directly proportional to the number of moles of gas at the same temperature and pressure. This relationship is expressed by the equation below, in which k is a proportionality constant.

$$V = kn$$

But volumes of gases change with changes in temperature and pressure. A set of conditions has been defined for measuring volumes of gases. For example, we know that argon exists as single atoms, and that its atomic mass is 39.95 g/mol. It has been determined that 22.41 L of argon at 0°C and 1 atm have a mass of 39.95 g. *Therefore, 22.41 L is the volume of 1 mol of any gas at STP.* The mass of 22.41 L of a gas at 0°C and a pressure of 1 atm will be equal to the gas's molecular mass.

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Topic : Avogadro's Law

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2 Section Review

UNDERSTANDING KEY IDEAS

1. What is the name of the gas law relating pressure and volume, and what does it state?
2. What is the name of the gas law relating volume and absolute temperature, and what does it state?
3. What is the name of the gas law relating pressure and absolute temperature, and what does it state?
4. What relationship does Avogadro's law express?

PRACTICE PROBLEMS

5. A sample of gas occupies 1.55 L at 27.0°C and 1.00 atm pressure. What will the volume be if the pressure is increased to 50.0 atm, but the temperature is kept constant?
6. A sample of nitrogen gas occupies 1.55 L at 27°C and 1.00 atm pressure. What will the volume be at –100°C and the same pressure?
7. A 1.0 L volume of gas at 27.0°C exerts a pressure of 85.5 kPa. What will the pressure be at 127°C? Assume constant volume.
8. A sample of nitrogen has a volume of 275 mL at 273 K. The sample is heated and the volume becomes 325 mL. What is the new temperature in kelvins?
9. A small cylinder of oxygen contains 300.0 mL of gas at 15 atm. What will the volume of this gas be when released into the atmosphere at 0.900 atm?

CRITICAL THINKING

10. A student has the following data: V_1 = 822 mL, T_1 = 75°C, T_2 = –25°C. He calculates V_2 and gets –274 mL. Is this correct? Explain why or why not.
11. Aerosol cans have a warning not to dispose of them in fires. Why?
12. What volume of carbon dioxide contains the same number of molecules as 20.0 mL of oxygen at the same conditions?

SECTION

3 Molecular Composition of Gases

Key Terms

- ideal gas
- ideal gas law
- diffusion
- effusion
- Graham's law of diffusion
- Gay-Lussac's law of combining volumes
- partial pressure
- Dalton's law of partial pressure

Objectives

1. **Solve** problems using the ideal gas law.
2. **Describe** the relationships between gas behavior and chemical formulas, such as those expressed by Graham's law of diffusion, Gay-Lussac's law of combining volumes, and Dalton's law of partial pressures.
3. **Apply** your knowledge of reaction stoichiometry to solve gas stoichiometry problems.

The Ideal Gas Law

You have studied four different gas laws, which are summarized in **Table 4.** Boyle's law states the relationship between the pressure and the volume of a sample of gas. Charles's law states the relationship between the volume and the absolute temperature of a gas. Gay-Lussac's law states the relationship between the pressure and the temperature of a gas. Avogadro's law relates volume to the number of moles of gas.

No gas perfectly obeys all four of these laws under all conditions. Nevertheless, these assumptions work well for most gases and most conditions. As a result, one way to model a gas's behavior is to assume that the gas is an **ideal gas** that perfectly follows these laws. An ideal gas, unlike a real gas, does not condense to a liquid at low temperatures, does not have forces of attraction or repulsion between the particles, and is composed of particles that have no volume.

ideal gas

an imaginary gas whose particles are infinitely small and do not interact with each other

Table 4 Summary of the Basic Gas Laws

Boyle's law	$P_1V_1 = P_2V_2$
Charles's law	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$
Gay-Lussac's law	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$
Avogadro's law	$V = kn$

STUDY TIP

Organizing the Gas Laws

You can use **Table 4** as a reference to help you learn and understand the gas laws.

- You may want to make a more elaborate table in which you write down each law's name, a brief explanation of its meaning, its mathematical representation, and the variable that is kept constant.

The Ideal Gas Law Relates All Four Gas Variables

ideal gas law

the law that states the mathematical relationship of pressure (P), volume (V), temperature (T), the gas constant (R), and the number of moles of a gas (n); $PV = nRT$

In using the basic gas laws, we have made use of four variables: pressure, P, volume, V, absolute temperature, T, and number of moles, n. Boyle's law, Charles's law, Gay-Lussac's law, and Avogadro's law can be combined into one equation that gives the relationship between all four variables, P, V, T, and n, for any sample of gas. This relationship is called the **ideal gas law.** When any three variables are given, the fourth can be calculated. The ideal gas law is represented mathematically below.

$$PV = nRT$$

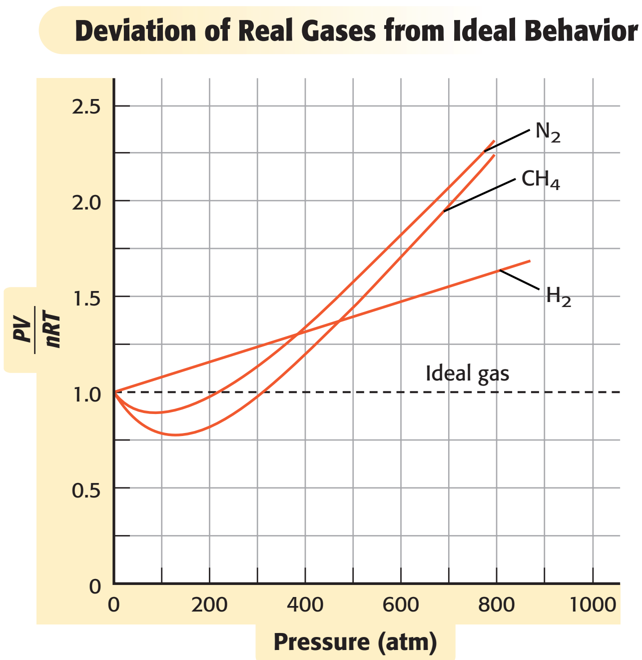
R is a proportionality constant. The value for R used in any calculation depends on the units used for pressure and volume. In this text, we will normally use units of kilopascals and liters when using the ideal gas law, so the value you will use for R is as follows.

$$R = \frac{8.314\ \text{L}\bullet\text{kPa}}{\text{mol}\bullet\text{K}}$$

If the pressure is expressed in atmospheres, then the value of R is 0.0821 (L•atm)/(mol•K).

The ideal gas law describes the behavior of real gases quite well at room temperature and atmospheric pressure. Under those conditions, the volume of the particles themselves and the forces of attraction between them can be ignored. The particles behave in ways that are similar enough to an ideal gas that the ideal gas law gives useful results. However, as the volume of a real gas decreases, the particles attract one another more strongly, and the volume is less than the ideal gas law would predict. At extremely high pressures, the volume of the particles themselves is close to the total volume, so the actual volume will be higher than calculated. This relationship is shown in **Figure 16.**

Figure 16
For an ideal gas, the ratio of PV/nRT is 1, which is represented by the dashed line. Real gases deviate somewhat from the ideal gas law and more at very high pressures.



SAMPLE PROBLEM E

Using the Ideal Gas Law

How many moles of gas are contained in 22.41 liters at 101.325 kPa and 0°C?

1 Gather information.

Three variables for a gas are given, so you can use the ideal gas law to compute the value of the fourth.

$V = 22.41$ L $\qquad T = 0°C$

$P = 101.325$ kPa $\qquad n = ?$

2 Plan your work.

Convert the temperature to kelvins.

$$0°C + 273 = 273 \text{ K}$$

Place the known quantities into the correct places in the equation relating the four gas variables, $PV = nRT$.

$$(101.325 \text{ kPa})(22.41 \text{ L}) = n\left(\frac{8.314 \text{ L}\bullet\text{kPa}}{\text{mol}\bullet\text{K}}\right)(273 \text{ K})$$

3 Calculate.

Solve for the unknown, n.

$$n = \frac{(101.325 \text{ kPa})(22.41 \text{ L})}{\left(\frac{8.314 \text{ L}\bullet\text{kPa}}{\text{mol}\bullet\text{K}}\right)(273 \text{ K})} = 1.00 \text{ mol}$$

4 Verify your results.

The product of pressure and volume is a little over 2000, and the product of R and the temperature is a little over 2000, so one divided by the other should be about 1. Also, recall that at STP the volume of 1 mol of gas is 22.41 L. The calculated result agrees with this value.

PRACTICE HINT

When using the ideal gas law, be sure that the units of P and V match the units of the value of R that is used.

PRACTICE

PROBLEM SOLVING SKILL

1. How many moles of air molecules are contained in a 2.00 L flask at 98.8 kPa and 25.0°C?
2. How many moles of gases are contained in a can with a volume of 555 mL and a pressure of 600.0 kPa at 20°C?
3. Calculate the pressure exerted by 43 mol of nitrogen in a 65 L cylinder at 5°C.
4. What will be the volume of 111 mol of nitrogen in the stratosphere, where the temperature is –57°C and the pressure is 7.30 kPa?

Gas Behavior and Chemical Formulas

So far we have dealt only with the general behavior of gases. No calculation has depended on knowing which gas was involved. For example, in **Sample Problem E,** we determined the number of moles of a gas. The calculation would have been the same no matter which gas or mixture of gases was involved. Now we will consider situations in which it is necessary to know more about a gas, such as its molar mass.

Diffusion

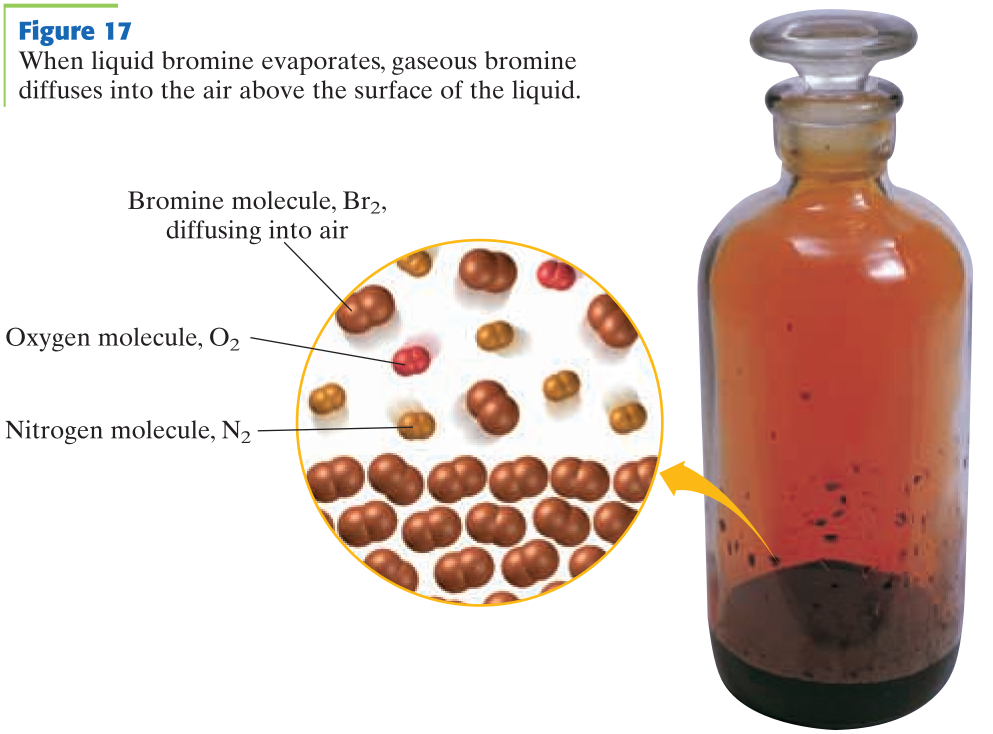
diffusion

the movement of particles from regions of higher density to regions of lower density

Household ammonia is a solution of ammonia gas, NH_3, in water. When you open a bottle of household ammonia, the odor of ammonia gas doesn't take long to fill the room. Gaseous molecules, including molecules of the compounds responsible for the smell, travel at high speeds in all directions and mix quickly with molecules of gases in the air in a process called **diffusion.** Gases diffuse through each other at a very rapid rate. Even if the air in the room was very still, it would only be a matter of minutes before ammonia molecules were evenly distributed throughout the room. During diffusion, a substance moves from a region of high concentration to a region of lower concentration. Eventually, the mixture becomes homogeneous, as seen in the closed bottle of bromine gas in **Figure 17.** Particles of low mass diffuse faster than particles of higher mass.

The process of diffusion involves an increase in entropy. Entropy can be thought of as a measure of randomness. One way of thinking of randomness is to consider the probability of finding a particular particle at a particular location. This probability is much lower in a mixture of gases than in a pure gas. Diffusion of gases is a way in which entropy is increased.

Figure 17
When liquid bromine evaporates, gaseous bromine diffuses into the air above the surface of the liquid.



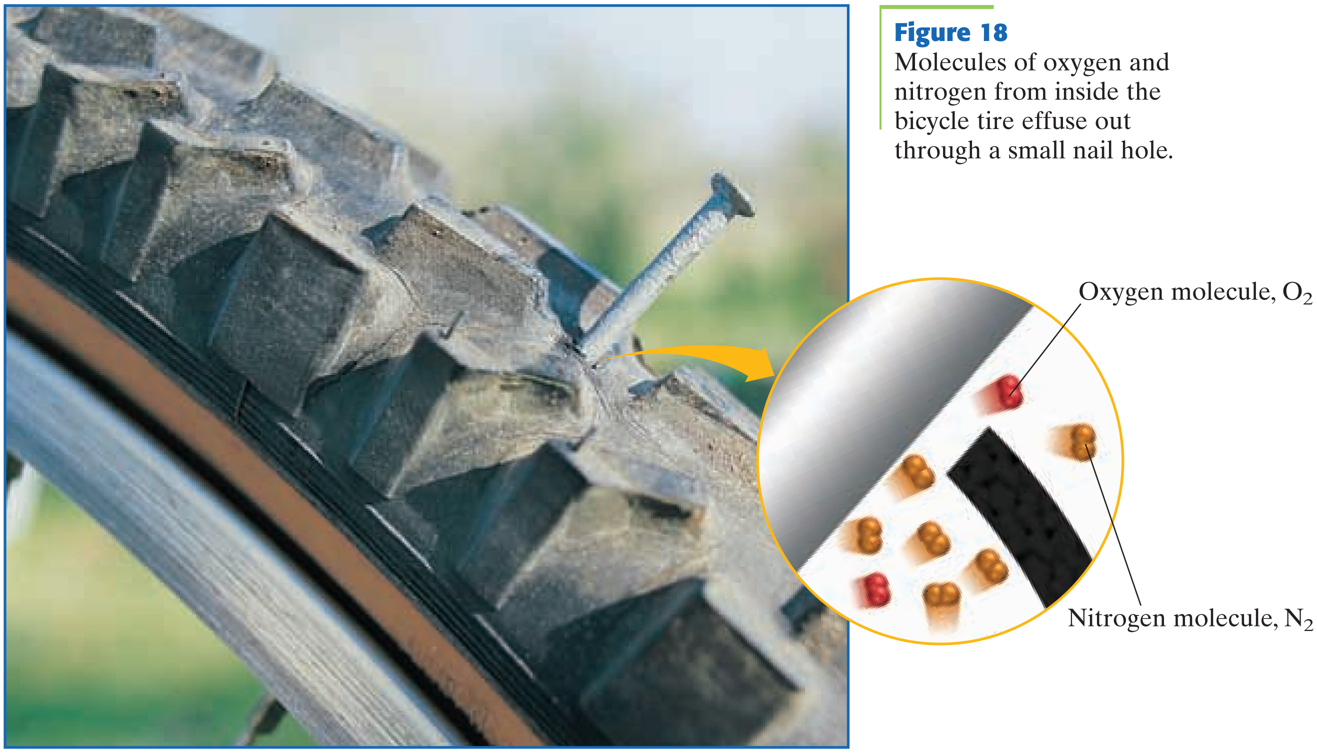


Figure 18
Molecules of oxygen and nitrogen from inside the bicycle tire effuse out through a small nail hole.

Effusion

The passage of gas particles through a small opening is called **effusion.** An example of effusion is shown in **Figure 18.** In a tire with a very small leak, the air in the tire effuses through the hole. The Scottish scientist Thomas Graham studied effusion in detail. He found that at constant temperature and pressure, the rate of effusion of a gas is inversely proportional to the square root of the gas's molar mass, *M*. This law also holds when one compares rates of diffusion, and molecular speeds in general. The molecular speeds, v_A and v_B, of gases A and B can be compared according to **Graham's law of diffusion,** shown below.

effusion
the passage of a gas under pressure through a tiny opening

Graham's law of diffusion
the law that states that the rate of diffusion of a gas is inversely proportional to the square root of the gas's density

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

For example, compare the speed of effusion of H_2 with that of O_2.

$$\frac{v_{H_2}}{v_{O_2}} = \sqrt{\frac{M_{O_2}}{M_{H_2}}} = \sqrt{\frac{32 \text{ g/mol}}{2 \text{ g/mol}}} = \sqrt{16} = 4$$

Hydrogen gas effuses four times as fast as oxygen. Particles of low molar mass travel faster than heavier particles.

According to the kinetic-molecular theory, gas particles that are at the same temperature have the same average kinetic energy. Therefore, the kinetic energy of two gases that are at the same temperature is

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

Solving this equation for the ratio of speeds between v_A and v_B gives Graham's law of diffusion.

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Topic: Effusion/Diffusion
SciLinks code: HW4041
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SAMPLE PROBLEM F

Comparing Molecular Speeds

Oxygen molecules have an average speed of about 480 m/s at room temperature. At the same temperature, what is the average speed of molecules of sulfur hexafluoride, SF_6?

1 Gather information.

$$v_{O_2} = 480 \text{ m/s}$$

$$v_{SF_6} = ? \text{ m/s}$$

2 Plan your work.

This problem can be solved using Graham's law of diffusion, which compares molecular speeds.

$$\frac{v_{SF_6}}{v_{O_2}} = \sqrt{\frac{M_{O_2}}{M_{SF_6}}}$$

You need molar masses of O_2 and SF_6.
Molar mass of O_2 = 2(16.00 g/mol) = 32.00 g/mol
Molar mass of SF_6 = 32.07 g/mol + 6(19.00 g/mol) = 146 g/mol

3 Calculate.

Place the known quantities into the correct places in Graham's law of diffusion.

$$\frac{v_{SF_6}}{480 \text{ m/s}} = \sqrt{\frac{32 \text{ g/mol}}{146 \text{ g/mol}}}$$

Solve for the unknown, v_{SF_6}.

$$v_{SF_6} = (480 \text{ m/s})\sqrt{\frac{32 \text{ g/mol}}{146 \text{ g/mol}}} = 480 \text{ m/s} \times 0.47 = 220 \text{ m/s}$$

4 Verify your results.

SF_6 has a mass about 4 times that of O_2. The square root of 4 is 2, and the inverse of 2 is $\frac{1}{2}$. SF_6 should travel about half as fast as O_2. The answer is therefore reasonable.

PRACTICE HINT

If you get confused, remember that molecular speeds and molar masses are inversely related, so the more massive particle will move at a slower speed.

PRACTICE

PROBLEM SOLVING SKILL

1. At the same temperature, which molecule travels faster, O_2 or N_2? How much faster?
2. At room temperature, Xe atoms have an average speed of 240 m/s. At the same temperature, what is the speed of H_2 molecules?
3. What is the molar mass of a gas if it diffuses 0.907 times the speed of argon gas?
4. Uranium isotopes are separated by effusion. What is the relative rate of effusion for $^{235}UF_6$ (M = 349.03 g/mol) and $^{238}UF_6$ (M = 352.04 g/mol)?

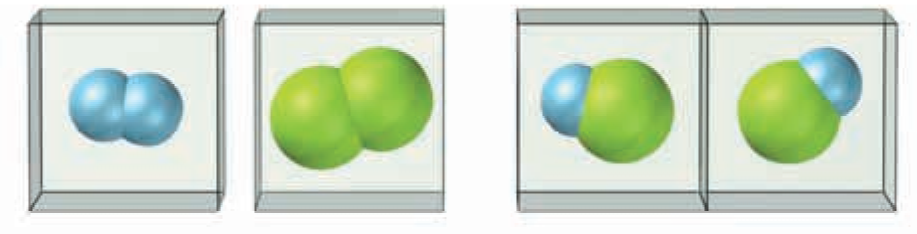


Figure 19
Hydrogen molecules combine with chlorine molecules in a 1:1 volume ratio to produce twice the volume of hydrogen chloride.

Gas Reactions Allow Chemical Formulas to Be Deduced

In 1808, Joseph Gay-Lussac made an important discovery: if the pressure and temperature are kept constant, gases react in volume proportions that are whole numbers. This is called **Gay-Lussac's law of combining volumes.** Consider the formation of gaseous hydrogen chloride from the reaction of hydrogen gas and chlorine gas. Gay-Lussac showed in an experiment that one volume of hydrogen gas reacts with one volume of chlorine gas to form two volumes of hydrogen chloride gas. **Figure 19** illustrates the volume ratios in the reaction in the form of a model. Let us use Avogadro's law and assume that two molecules of hydrogen chloride are formed, one in each of the two boxes on the right in **Figure 19.** Therefore, we must start with two atoms of hydrogen and two atoms of chlorine. Each box must contain one molecule of hydrogen and one molecule of chlorine, so there must be two atoms in each of these molecules.

Gay-Lussac's law of combining volumes
the law that states that the volumes of gases involved in a chemical change can be represented by the ratio of small whole numbers

Using several other reactions, such as the reaction of gaseous hydrogen and gaseous oxygen to form water vapor, the Italian chemist Stanislao Cannizzaro was able to deduce that oxygen is also diatomic and that the formula for water is H_2O. Dalton had guessed that the formula for water was HO, because this seemed the most likely combination of atoms for such a common compound. Before knowing atomic masses, chemists had only a set of relative weights. For example, it was known that 1 g of hydrogen can react with 8 g of oxygen to form water, so it was assumed that an oxygen atom was eight times as heavy as a hydrogen atom. It was not until the mid 1800s, just before the Civil War, that chemists knew the correct formula for water.

Dalton's Law of Partial Pressure

In 1805, John Dalton showed that in a mixture of gases, each gas exerts a certain pressure as if it were alone with no other gases mixed with it. The pressure of each gas in a mixture is called the **partial pressure.** The total pressure of a mixture of gases is the sum of the partial pressures of the gases. This principle is known as **Dalton's law of partial pressure.**

partial pressure
the pressure of each gas in a mixture

Dalton's law of partial pressure
the law that states that the total pressure of a mixture of gases is equal to the sum of the partial pressures of the component gases

$$P_{total} = P_A + P_B + P_C$$

P_{total} is the total pressure, and P_A, P_B, and P_C are the partial pressures of each gas.

How is Dalton's law of partial pressure explained by the kinetic-molecular theory? All the gas molecules are moving randomly, and each has an equal chance to collide with the container wall. Each gas exerts a pressure proportional to its number of molecules in the container. The presence of other gas molecules does not change this fact.

Gas Stoichiometry

The ideal gas law relates amount of gaseous substance in moles, *n,* with the other gas variables: pressure, volume, and temperature. Now that you have learned how to use the ideal gas law, an equation that relates the number of moles of gas to its volume, you can use it in calculations involving gases that react.

Gas Volumes Correspond to Mole Ratios

Topic Link

Refer to the "Stoichiometry" chapter for a discussion of stoichiometry.

Ratios of gas volumes will be the same as mole ratios of gases in balanced equations. Avogadro's law shows that the mole ratio of two gases at the same temperature and pressure is the same as the volume ratio of the two gases. This greatly simplifies the calculation of the volume of products or reactants in a chemical reaction involving gases. For example, consider the following equation for the production of ammonia.

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

Consequently, 3 L of H_2 react with 1 L of N_2 to form 2 L of NH_3, and no H_2 or N_2 is left over (assuming an ideal situation of 100% yield).

Consider the electrolysis of water, a reaction expressed by the following chemical equation.

$$2H_2O(l) + \text{electricity} \longrightarrow 2H_2(g) + O_2(g)$$

The volume of hydrogen gas produced will be twice the volume of oxygen gas, because there are twice as many moles of hydrogen as there are moles of oxygen. As you can see in **Figure 20,** the volume of hydrogen gas produced is, in fact, twice as large as the volume of oxygen produced.

Furthermore, if we know the number of moles of a gaseous substance, we can use the ideal gas law to calculate the volume of that gas. **Skills Toolkit 1** on the following page shows how to find the volume of product given the mass of one of the reactants.

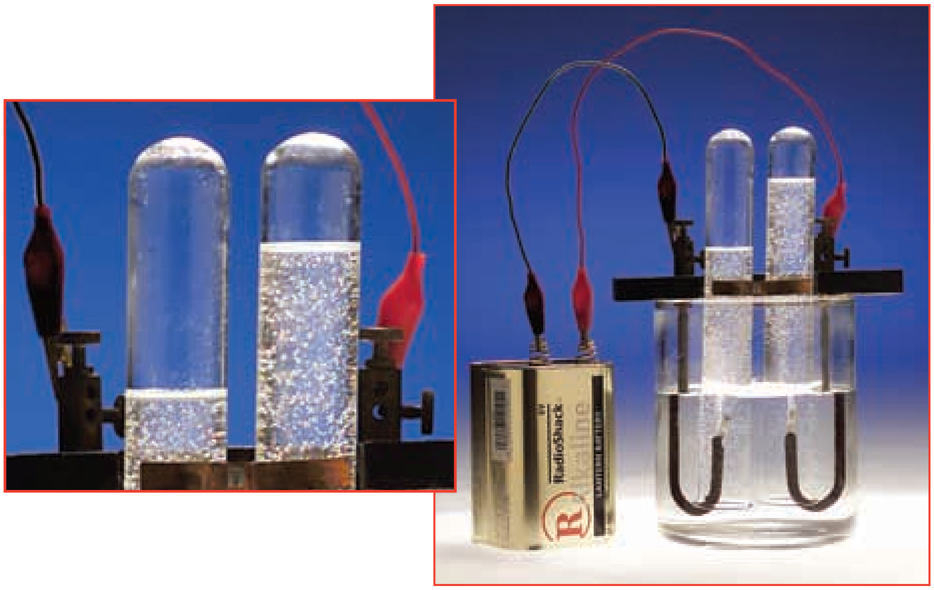
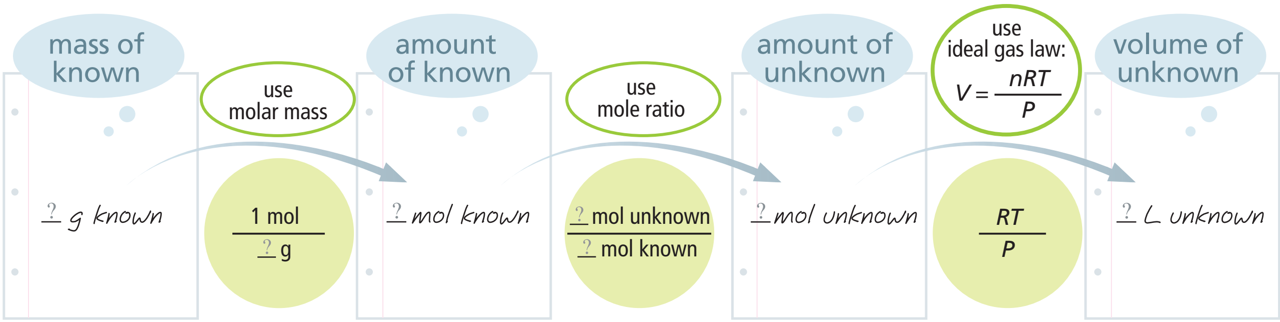


Figure 20
In the electrolysis of water, the volume of hydrogen is twice the volume of oxygen, because the mole ratio between the two gases is two to one.

SKILLS Toolkit 1

Finding Volume of Unknown



SAMPLE PROBLEM G

Using the Ideal Gas Law to Solve Stoichiometry Problems

How many liters of hydrogen gas will be produced at 280.0 K and 96.0 kPa if 1.74 mol of sodium react with excess water according to the following equation?

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

1 Gather information.

$T = 280.0$ K $\quad P = 96.0$ kPa $\quad R = 8.314$ L•kPa/mol•K

$n = ?$ mol $\quad V = ?$ L

2 Plan your work.

Use the mole ratio from the equation to compute moles of H_2.

$$1.74\ \text{mol Na} \times \frac{1\ \text{mol H}_2}{2\ \text{mol Na}} = 0.870\ \text{mol H}_2$$

Rearrange the ideal gas law to solve for the volume of hydrogen gas.

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

3 Calculate.

Substitute the three known values into the rearranged equation.

$$V = \frac{(0.870\ \text{mol H}_2)\left(\frac{8.314\ \text{L•kPa}}{\text{mol•K}}\right)(280.0\ \text{K})}{(96.0\ \text{kPa})} = 21.1\ \text{L H}_2$$

4 Verify your results.

Recall that 1 mol of gas at 0°C and 1 atm occupies 22.4 L. The conditions in this problem are close to this, so the volume should be near 22.4 L.

Practice problems on next page

PRACTICE HINT

Whenever you can relate the given information to moles, you can solve stoichiometry problems. In this case, the ideal gas law is the bridge that you need to get from moles to the answer.

PRACTICE

PROBLEM SOLVING SKILL

1. What volume of oxygen, collected at 25°C and 101 kPa, can be prepared by decomposition of 37.9 g of potassium chlorate?

$$2KClO_3(s) \longrightarrow 2KCl(s) + 3O_2(g)$$

2. Liquid hydrogen and oxygen are burned in a rocket. What volume of water vapor, at 555°C and 76.4 kPa, can be produced from 4.67 kg of H_2?

$$2H_2(l) + O_2(l) \longrightarrow 2H_2O(g)$$

3. How many grams of sodium are needed to produce 2.24 L of hydrogen, collected at 23°C and 92.5 kPa?

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

3 Section Review

UNDERSTANDING KEY IDEAS

1. What gas laws are combined in the ideal gas law?
2. State the ideal gas law.
3. What is the relationship between a gas's molecular weight and speed of effusion?
4. How did Gay-Lussac's experiment allow the chemical formulas of hydrogen gas and chlorine gas to be deduced?
5. How does the total pressure of a mixture of gases relate to the partial pressures of the individual gases in a mixture?
6. In gas stoichiometry problems, what is the "bridge" between amount in moles and volume?

PRACTICE PROBLEMS

7. How many moles of argon are there in 20.0 L, at 25°C and 96.8 kPa?
8. A sample of carbon dioxide has a mass of 35.0 g and occupies 2.5 L at 400.0 K. What pressure does the gas exert?
9. How many moles of SO_2 gas are contained in a 4.0 L container at 450 K and 5.0 kPa?
10. Two gases effuse through a hole. Gas A has nine times the molecular mass of gas B. What is the ratio of the two molecular speeds?
11. What volume of ammonia can be produced from the reaction of 22.5 L of hydrogen with nitrogen?

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

12. What will be the volume, at 115 kPa and 355 K, of the nitrogen from the decomposition of 35.8 g of sodium azide, NaN_3?

$$2NaN_3(s) \longrightarrow 2Na(s) + 3N_2(g)$$

CRITICAL THINKING

13. Explain why helium-filled balloons deflate over time faster than air-filled balloons do.
14. Nitrous oxide is sometimes used as a source of oxygen gas:

$$2N_2O(g) \longrightarrow 2N_2(g) + O_2(g)$$

What volume of each product will be formed from 2.22 L N_2O? At STP, what is the density of the product gases when mixed? (Hint: Keep in mind Avogadro's law.)

Where Is N?

Earth's crust: <0.01% by mass

Air: 75% by mass

Sea water: 0.0016% by mass

Element Spotlight

7
N
Nitrogen
14.0067
$[He]2s^22p^3$

Nitrogen

Nitrogen gas is the most abundant gas in the atmosphere. Nitrogen is important for making the proteins, nucleic acids, vitamins, enzymes, and hormones needed by plants and animals to live. However, nitrogen gas is too unreactive for plants and animals to use directly. Nitrogen-fixing bacteria convert atmospheric nitrogen into substances that green plants absorb from the soil. Animals then eat these plants or eat other animals that feed on these plants. When the animals and plants die and decay, the nitrogen in the decomposed organic matter returns as nitrogen gas to the atmosphere and as compounds to the soil. The nitrogen cycle then starts all over again.

Legumes, such as this soybean plant, have nitrogen-fixing bacteria on their roots. These bacteria can convert atmospheric nitrogen into compounds that can be used by plants in the formation of proteins.

Industrial Uses

- Nitrogen is used in the synthesis of ammonia.
- Ammonia, NH_3, is used to produce fertilizer, explosives, nitric acid, urea, hydrazine, and amines.
- Liquid nitrogen is used in superconductor research and as a cryogenic supercoolant for storing biological tissues.
- Nitrogen gas is used as an inert atmosphere for storing and processing reactive substances.
- Nitrogen gas, usually mixed with argon, is used for filling incandescent light bulbs.

Real-World Connection The atmosphere contains almost 4×10^{18} kg of N_2. That's more than the mass of 10 000 typical mountains!

A Brief History

1600 — 1700 — 1800 — 1900

1772: Nitrogen is discovered by Daniel Rutherford in Scotland, Joseph Priestley and Henry Cavendish in England, and Carl Scheele in Sweden.

1774–1777: Antoine Lavoisier determines that nitrogen is an element.

1909: Fritz Haber, a German chemist, discovers a method for synthesizing ammonia from hydrogen gas and nitrogen gas. The method is still used today and is called the Haber-Bosch process.

Questions

1. Name three nitrogen-containing products you come into contact with regularly.
2. Research the historical impacts of the invention of the Haber-Bosch process, which is mentioned in the timeline above.

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Topic: Nitrogen Cycle
SciLinks code: HW4082

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12 CHAPTER HIGHLIGHTS

KEY IDEAS

SECTION ONE Characteristics of Gases

- Gases are fluids, have low density, and are compressible, because of the relatively large intermolecular distances between gas particles.
- Gases expand to fill their entire container.
- Gases exert pressure in all directions.
- The kinetic-molecular theory states that gas particles are in constant random motion, are relatively far apart, and have volumes that are negligible when compared with the total volume of a gas.
- The average kinetic energy of a gas is proportional to its absolute temperature.

KEY TERMS

pressure
newton
pascal
standard temperature and pressure
kinetic-molecular theory

SECTION TWO The Gas Laws

- Pressure and volume of a gas at constant temperature are inversely proportional.
- The volume of a gas at constant pressure is proportional to the absolute temperature.
- The pressure of a gas at constant volume is proportional to the absolute temperature.
- Equal volumes of gas under the same conditions contain an equal number of moles of gas.

KEY TERMS

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

SECTION THREE Molecular Composition of Gases

- The ideal gas law is the complete statement of the relations between P, V, T, and n in a quantity of gas.
- Gases diffuse rapidly into each other.
- The rate of diffusion of a gas is inversely proportional to the square root of its molar mass.
- Each gas in a mixture produces a pressure as if it were alone in a container.
- Gay-Lussac's law of combining volumes can be used to deduce the chemical formula of a gas through observation of volume changes in a chemical reaction.

KEY TERMS

ideal gas
ideal gas law
diffusion
effusion
Graham's law of diffusion
Gay-Lussac's law of combining volumes
partial pressure
Dalton's law of partial pressure

KEY SKILLS

Converting Pressure Units
Sample Problem A p. 420

Solving Pressure-Volume Problems
Sample Problem B p. 425

Solving Volume-Temperature Problems
Sample Problem C p. 428

Solving Pressure-Temperature Problems
Sample Problem D p. 430

Using the Ideal Gas Law
Sample Problem E p. 435

Comparing Molecular Speeds
Sample Problem F p. 438

Using the Ideal Gas Law to Solve Stoichiometry Problems
Skills Toolkit 1 p. 441
Sample Problem G p. 441

CHAPTER REVIEW 12

USING KEY TERMS

1. What is the definition of *pressure?*
2. What is a newton?
3. Write a paragraph that describes how the kinetic-molecular theory explains the following properties of a gas: fluidity, compressibility, and pressure. WRITING SKILLS
4. How do pascals relate to newtons?
5. What relationship does Boyle's law express?
6. What relationship does Charles's law express?
7. What relationship does Gay-Lussac's law express?
8. What gas law combines the basic gas laws?
9. What are two characteristics of an ideal gas?
10. Describe in your own words the process of diffusion.

UNDERSTANDING KEY IDEAS

Characteristics of Gases

11. How does wind illustrate that gases are fluids?
12. Using air, water, and a syringe, how can you show the difference in compressibility of liquids and gases?
13. When you drive a car on hot roads, the pressure in the tires increases. Explain.
14. As you put more air in a car tire, the pressure increases. Explain.
15. When you expand your lungs, air flows in. Explain.
16. Even when there is air in bicycle tires, you can still push down on the handle of the pump rather easily. Explain.
17. When you put air in a completely flat bicycle tire, the entire tire expands. Explain.
18. What assumptions does the kinetic-molecular theory make about the nature of a gas?
19. How does the average kinetic energy of a gas relate to its temperature?

The Gas Laws

20. How are the volume and pressure of a gas related, if its temperature is kept constant?
21. Explain why pressure increases as a gas is compressed into a smaller volume.
22. How are the absolute temperature and volume of a gas related, at constant pressure?
23. Explain Charles's law in terms of the kinetic-molecular theory.
24. How is a gas's pressure related to its temperature, at constant volume?
25. Explain Gay-Lussac's law in terms of the kinetic-molecular theory.
26. What does Avogadro's law state about the relationship between gas volumes and amounts in moles?

Molecular Composition of Gases

27. When using the ideal gas law, what is the proportionality constant, and in what units is it usually expressed?
28. Ammonia, NH_3, and alcohol, C_2H_6O, are released together across a room. Which will you smell first?

29. How can Gay-Lussac's law of combining volumes be used to deduce chemical formulas?

30. Write the equation that expresses Dalton's law of partial pressures.

PRACTICE PROBLEMS

PROBLEM SOLVING SKILL

Sample Problem A Converting Pressure Units

31. The standard pressure at sea level is 101 325 pascals. What force is being exerted on each square meter of Earth's surface?

32. The vapor pressure of hydrogen peroxide is 100.0 torr at 97.9°C. What is this pressure in kPa?

33. The gauge pressure in a tire is 28 psi, which adds to atmospheric pressure of 14.0 psi. What is the internal tire pressure in kPa?

34. The weather bureau reports the atmospheric pressure as 925 millibars. What is this pressure in kPa?

Sample Problem B Solving Pressure-Volume Problems

35. A gas sample has a volume of 125 mL at 91.0 kPa. What will its volume be at 101 kPa?

36. A 125 mL sample of gas at 105 kPa has its volume reduced to 75.0 mL. What is the new pressure?

37. A diver at a depth of 1.0×10^2 m, where the pressure is 11.0 atm, releases a bubble with a volume of 100.0 mL. What is the volume of the bubble when it reaches the surface? Assume a pressure of 1.00 atm at the surface.

38. In a deep-sea station 2.0×10^2 m below the surface, the pressure in the module is 20.0 atm. How many liters of air at sea level are needed to fill the module with 2.00×10^7 L of air?

39. The pressure on a 240.0 mL sample of helium gas is increased from 0.428 atm to 1.55 atm. What is the new volume, assuming constant temperature?

40. A sample of air with volume 6.6×10^7 L changes pressure from 99.4 kPa to 88.8 kPa. Assuming constant temperature, what is the new volume?

Sample Problem C Solving Volume-Temperature Problems

41. Use Charles's law to solve for the missing value in the following. $V_1 = 80.0$ mL, $T_1 =$ 27°C, $T_2 = 77$°C, $V_2 = ?$

42. A balloon filled with helium has a volume of 2.30 L on a warm day at 311 K. It is brought into an air-conditioned room where the temperature is 295 K. What is its new volume?

43. The balloon in item 42 is dipped into liquid nitrogen at −196°C. What is its new volume?

44. A gas at 65°C occupies 4.22 L. At what Celsius temperature will the volume be 3.87 L, at the same pressure?

45. A person breathes 2.6 L of air at −11°C into her lungs, where it is warmed to 37°. What is its new volume?

46. A scientist warms 26 mL of gas at 0.0°C until its volume is 32 mL. What is its new temperature in degrees Celsius?

Sample Problem D Solving Pressure-Temperature Problems

47. Use Gay-Lussac's law to solve for the unknown. $P_1 = 111$ kPa, $T_1 = 273$ K, $T_2 =$ 373 K, $P_2 = ?$

48. A sample of hydrogen exerts a pressure of 0.329 atm at 47°C. What will the pressure be at 77°C, assuming constant volume?

49. A sample of helium exerts a pressure of 101 kPa at 25°C. Assuming constant volume, what will its pressure be at liquid nitrogen temperature, −196°C?

50. The pressure inside a tire is 39 psi at 20°C. What will the pressure be after the tire is driven at high speed on a hot highway, when the temperature in the tire is 48°C?

51. A tank of oxygen for welding is at 31°C and 11 atm. What is the pressure when it is taken to the South Pole, where the temperature is –41°C?

Sample Problem E Using the Ideal Gas Law

52. How many moles of argon are there in 20.0 L, at 25°C and 96.8 kPa?

53. How many moles of air are in 1.00 L at –23°C and 101 kPa?

54. A 4.44 L container holds 15.4 g of oxygen at 22.55°C. What is the pressure?

55. A polyethylene plastic weather balloon contains 65 L of helium, which is at 20.0°C and 94.0 kPa. How many moles of helium are in the balloon?

56. What will be the volume of the balloon in item 56 in the stratosphere at –61°C and 1.1×10^3 Pa?

Sample Problem F Comparing Molecular Speeds

57. An unknown gas effuses at a speed one-quarter of that of helium. What is the molar mass of the unknown gas? It is either sulfur dioxide or sulfur trioxide. Which gas is it?

58. An unknown gas effuses at one half the speed of oxygen. What is the molar mass of the unknown? It is either HBr or HI. Which gas is it?

59. Oxygen molecules have an average speed of 4.80×10^2 m/s at 25°C. What is the average speed of H_2 molecules at the same temperature?

Sample Problem G Using the Ideal Gas Law to Solve Stoichiometry Problems

60. How many liters of hydrogen gas can be produced at 300.0 K and 104 kPa if 20.0 g of sodium metal is reacted with water according to the following equation?

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

61. Magnesium will burn in oxygen to form magnesium oxide as represented by the following equation.

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

What mass of magnesium will react with 500.0 mL of oxygen at 150.0°C and 70.0 kPa?

62. Suppose a certain automobile engine has a cylinder with a volume of 500.0 mL that is filled with air (21% oxygen) at a temperature of 55°C and a pressure of 101.0 kPa. What mass of octane must be injected to react with all of the oxygen in the cylinder?

$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 18H_2O(g)$$

63. Methanol, CH_3OH, is made by using a catalyst to react carbon monoxide with hydrogen at high temperature and pressure. Assuming that 450.0 mL of CO and 825 mL of H_2 are allowed to react, answer the following questions. (Hint: First write the balanced chemical equation for this reaction.)

a. Which reactant is in excess?
b. How much of that reactant remains when the reaction is complete?
c. What volume of $CH_3OH(g)$ is produced?

64. What volume of oxygen, measured at 27°C and 101.325 kPa, is needed for the combustion of 1.11 kg of coal? (Assume coal is 100% carbon.)

$$C(s) + O_2(g) \longrightarrow CO_2(g)$$

MIXED REVIEW

65. How many liters of hydrogen are obtained from the reaction of 4.00 g calcium with excess water, at 37°C and 0.962 atm?

66. How many grams of carbon dioxide are contained in 1.000 L of the gas, at 25.0°C and 101.325 kPa? What is the density of the gas at these conditions? What would the density of oxygen be at these conditions?

67. Below is a diagram showing the effects of pressure on a column of mercury and on a column of water. Which system is under a higher pressure? Explain your choice.

68. Solid LiOH can be used in spacecraft to remove CO_2 from the air.

$$2LiOH(s) + CO_2(g) \longrightarrow Li_2CO_3(s) + H_2O(l)$$

What mass of LiOH must be used to absorb the carbon dioxide that exerts a partial pressure of 5.0 kPa at 15°C in a spacecraft with volume of 8.0×10^4 L?

69. How many grams of oxygen gas must be in a 10.0 L container to exert a pressure of 97.0 kPa at a temperature of 25°C?

70. Explain in terms of the kinetic-molecular theory why increasing the temperature of a gas at constant volume increases the pressure of the gas.

71. Above 100°C and at constant pressure, two volumes of hydrogen react with one volume of oxygen to form two volumes of gaseous water. Set up a diagram for this reaction that is similar to **Figure 19,** and that shows how the molecular formulas for oxygen and water can be determined.

CRITICAL THINKING

72. Clean rooms, used for sterile biological research, are sealed tightly and operate under high air pressure. Explain why.

73. The partial pressure of oxygen in the air is 0.21 atm at sea level, where the total pressure is 1.00 atm. What is the partial pressure of oxygen when this air rises to where the total pressure is 0.86 atm?

74. You have a 1 L container of a gas at 20°C and 1 atm. Without opening the container, how could you tell whether the gas is chlorine or fluorine?

75. Gas companies often store their fuel supplies in liquid form and in large storage tanks. Liquid nitrogen is used to keep the temperature low enough for the fuel to remain condensed in liquid form. Although continuous cooling is expensive, storing a condensed fuel is more economical than storing the fuel as gas. Suggest a reason that storing a liquid is more economical than storing a gas.

76. How would the shape of a curve showing the kinetic-energy distribution of gas molecules at 50°C compare with the blue and red curves in **Figure 8?**

ALTERNATIVE ASSESSMENT

77. The air pressure of car tires should be checked regularly for safety reasons and to prevent uneven tire wear. Find out the units of measurement on a typical tire gauge, and determine how gauge pressure relates to atmospheric pressure.

78. Find a local hot-air balloon group, and discuss with group members how they use the gas laws to fly their balloons. The group may be willing to give a demonstration. Report your experience to the class.

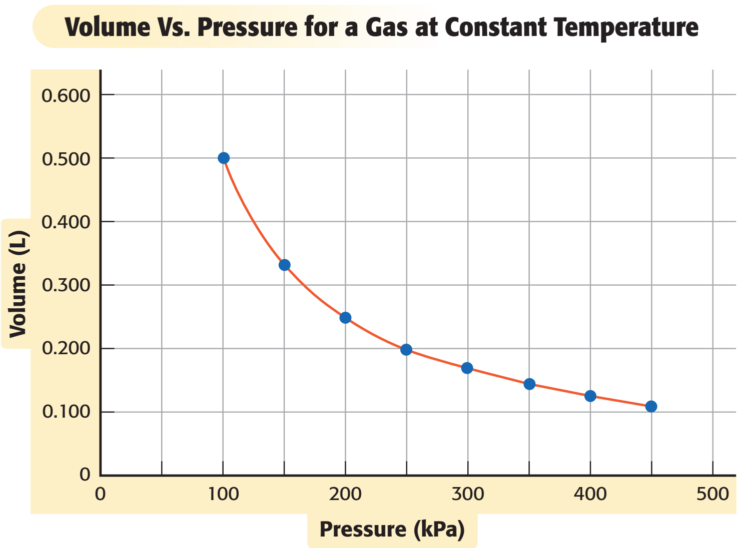
CONCEPT MAPPING

concept map

79. Use the following terms to create a concept map: *amount in moles, ideal gas law, pressure, temperature,* and *volume.*

FOCUS ON GRAPHING

Study the graph below, and answer the questions that follow.
For help in interpreting graphs, see Appendix B, "Study Skills for Chemistry."



80. What variable decreases as you go down the *y*-axis of the graph?

81. What is the volume of the gas at a pressure of 150 kPa?

82. What is the pressure of the gas at a volume of 0.200 L?

83. Describe in your own words the relationship this graph illustrates.

84. List the values of volume and pressure that correspond to any two points on the graph, and show why they demonstrate Boyle's law.

85. Explain why you feel resistance if you try to compress a sample of gas inside a plugged syringe.

TECHNOLOGY AND LEARNING

86. Graphing Calculator

Calculating Pressure Using the Ideal Gas Law

The graphing calculator can run a program that calculates the pressure in atmospheres, given the number of moles of a gas (n), volume (V), and temperature (T).

Go to Appendix C. If you are using a TI-83 Plus, you can download the program IDEAL and run the application as directed. If you are using another calculator, your teacher will provide you with keystrokes and data sets to use. Use the program on your calculator to answer these questions.

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

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

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

12 STANDARDIZED TEST PREP

UNDERSTANDING CONCEPTS

Directions (1–3): **For *each* question, write on a separate sheet of paper the letter of the correct answer.**

1. A sample of oxygen gas has a volume of 150 milliliters when its pressure is 0.947 atmosphere. If the pressure is increased to 0.987 atmosphere and the temperature remains constant, what will the new volume be?
   - **A.** 140 mL
   - **B.** 160 mL
   - **C.** 200 mL
   - **D.** 240 mL

2. What does the kinetic-molecular theory state about ideal gas molecules?
   - **F.** They have weight and take up space.
   - **G.** They are in constant, rapid, random motion.
   - **H.** They have high densities compared to liquids and solids.
   - **I.** They exert forces of attraction and repulsion on one another.

3. Gases that have properties like those stated in the kinetic molecular theory are said to be ideal. Which of the following gases is **most** nearly ideal?
   - **A.** helium
   - **B.** hydrogen
   - **C.** oxygen
   - **D.** xenon

Directions (4–6): **For *each* question, write a short response.**

4. What is the final volume when the temperature of 100 mL of a gas is raised from 200 K to 400 K, while the pressure changes from 100 kPa to 200 kPa? Explain your answer.

5. Why can you quickly smell perfume if someone opens a bottle on the other side of the room, even when there is no noticeable movement of air?

6. Why does carbon dioxide gas, which has a sublimation point of –79°C, cool enough to become solid as it leaves a fire extinguisher at room temperature?

READING SKILLS

Directions (7–9): **Read the passage below. Then answer the questions.**

Hot air balloonists use large fans to force air into their balloons. When the balloon is about three quarters full, they turn off the fan and heat the air inside the balloon using a propane burner. As the air is heated, the balloon rises. Once airborne, the balloonist can turn off the burner and just use it occasionally to keep the balloon aloft. The heating of the air in the balloon is a practical application of the gas laws.

7. If the volume of the balloon is 3 000 cubic meters (m^3) at 27°C, what is its volume at 52°C, assuming a constant pressure?
   - **F.** 2 750 m^3
   - **G.** 3 250 m^3
   - **H.** 4 500 m^3
   - **I.** 6 250 m^3

8. As the balloon rises in altitude, the pressure outside the balloon decreases. How does this change in altitude affect the volume of the balloon, assuming the temperature does not change rapidly?
   - **A.** The balloon's volume becomes smaller because the pressure decreases.
   - **B.** The balloon's volume is not affected by changes in the ouside pressure.
   - **C.** The balloon's volume increases as its pressure matches the outside pressure.
   - **D.** The balloon's volume change cannot be predicted without more information.

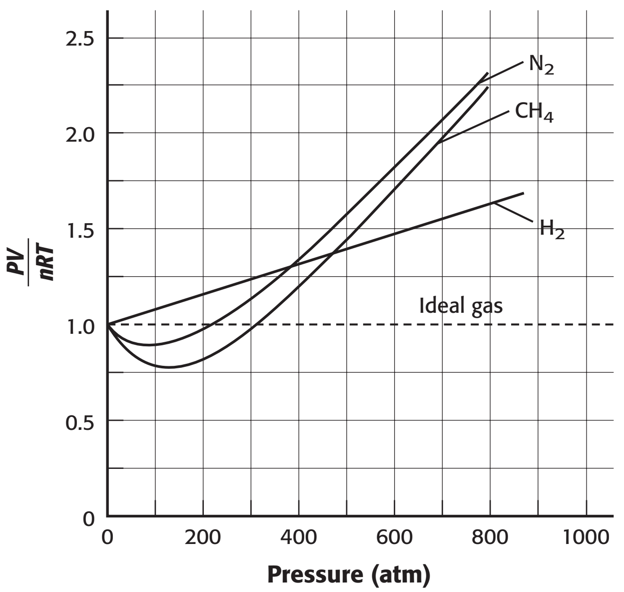
9. Why does the balloonist only fill the balloon to three quarters of its volume before heating it?

INTERPRETING GRAPHICS

***Directions (10–13):* For *each* question below, record the correct answer on a separate sheet of paper.**

As shown in the illustration below, the behavior of real gases deviates from that predicted by the ideal gas law. The amount of variation depends on the nature of the gas and the specific conditions of temperature and pressure. Use this graph to answer questions 10 through 13.

Deviation of Real Gases from Ideal Behavior



10 Under what pressure conditions is the ideal gas law useful for predicting the behavior of real gases?

F. about one atmosphere
G. about 200 atmospheres
H. equally useful at all pressures
I. pressures close to the value of R

11 Why does the value of PV/nRT for methane and nitrogen drop below the value predicted by the ideal gas law?

A. The ratio is a function of mass of the molecules.
B. Interaction between molecules reduces the volume.
C. There is an inverse relationship between pressure and PV/nRT.
D. The temperature of the gas is reduced as pressure is increased.

12 Based on information from the graph, which of these characteristics of a gas appears to have the greatest effect on the deviation at high pressures?

F. mass of the molecule
G. polarity of the molecule
H. number of atoms per molecule
I. number of molecules per mole of the gas

13 At what pressure (atm) does the behavior of methane **most** resemble that of an ideal gas?

Test TIP

When analyzing a graph, pay attention to its title. The title should tell you what is plotted on the graph and provide some context for the data.

CHAPTER 13

Solutions