14 The Behavior of Gases

INSIDE:

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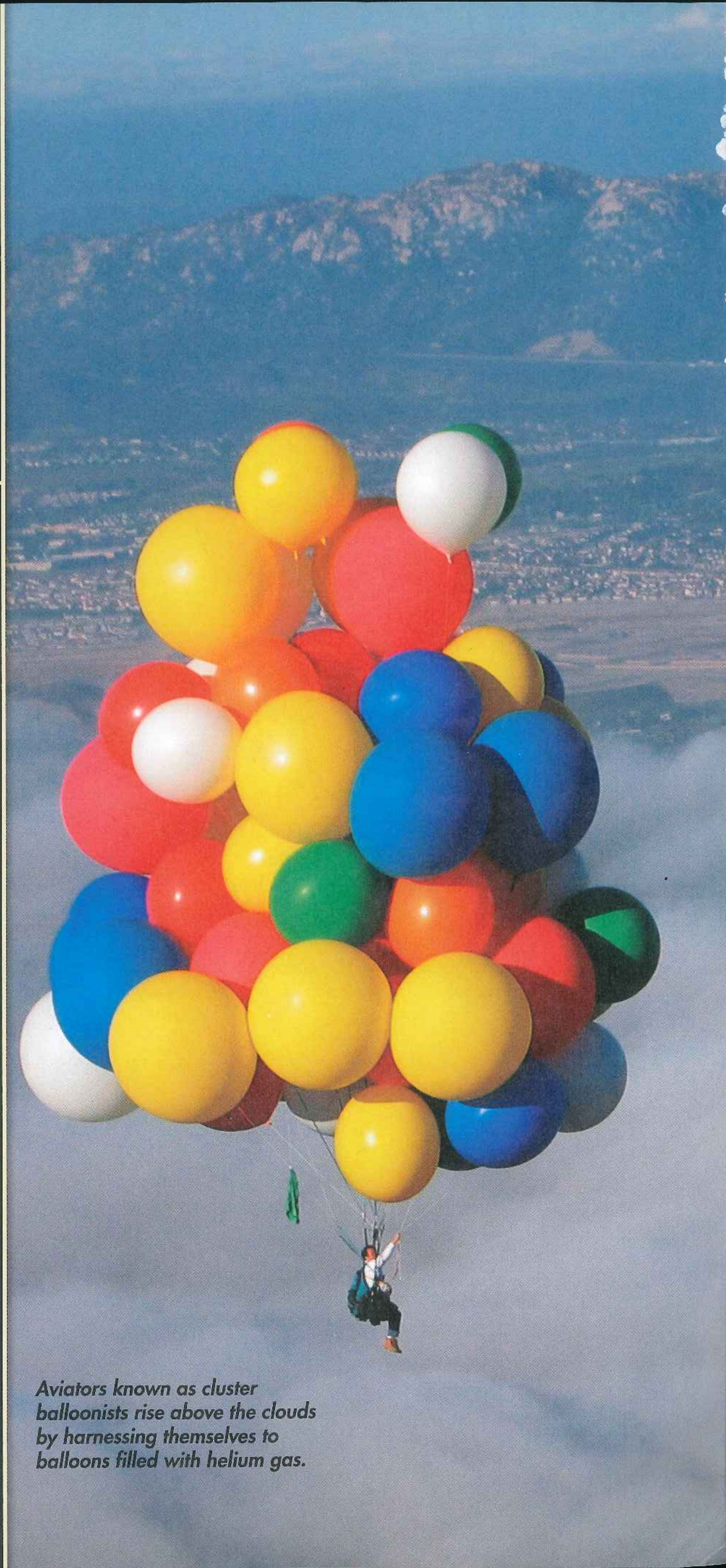
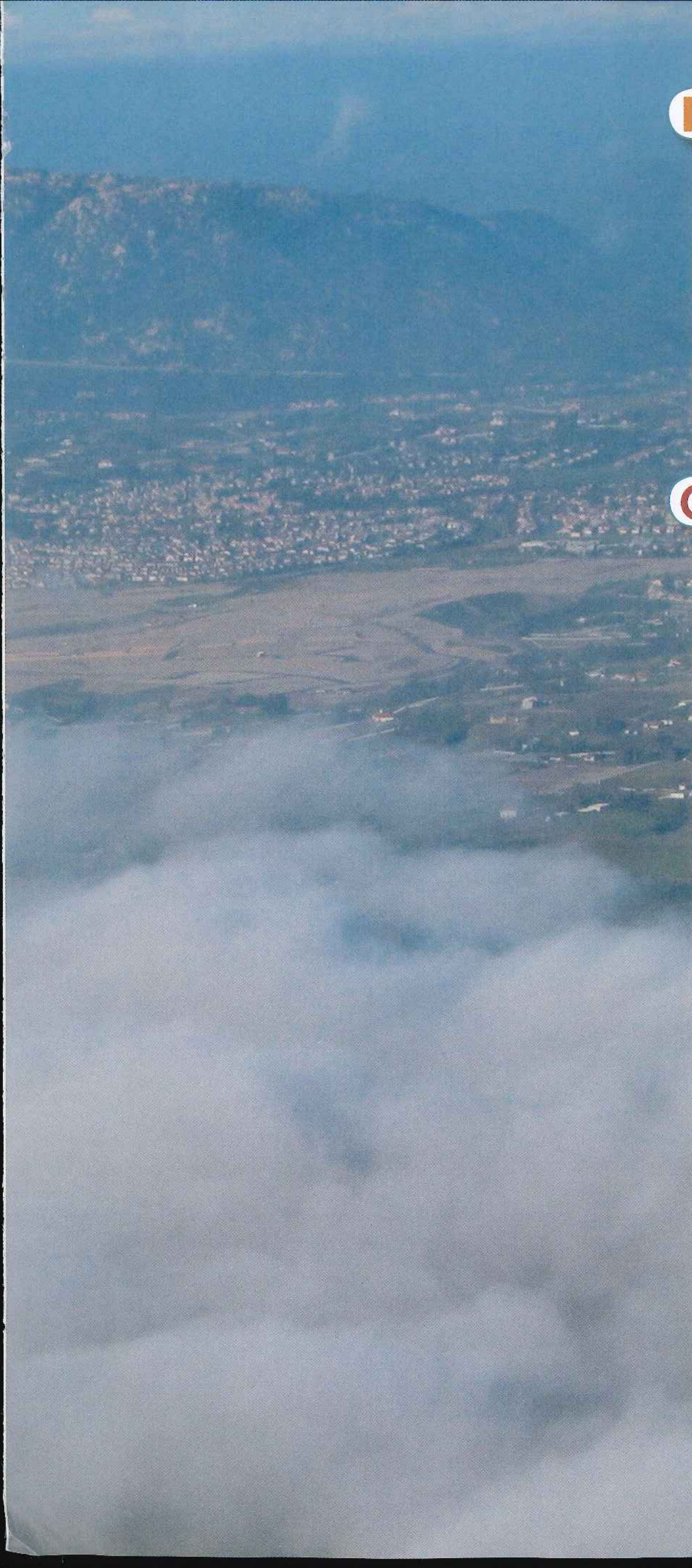


Aviators known as cluster balloonists rise above the clouds by harnessing themselves to balloons filled with helium gas.



BIGIDEA

KINETIC THEORY

Essential Questions:

1. *How do gases respond to changes in pressure, volume, and temperature?*
2. *Why is the ideal gas law useful even though ideal gases do not exist?*

CHEMYSTERY

Under Pressure

Just after 2 P.M., Becki completes her eighth scuba dive over a four-day period off the coast of Belize. After the dive, she feels fine.

A few hours later at dinner, Becki feels tired. She thinks that her fatigue is probably due to the many hours she had spent swimming during her vacation. But she also begins to feel itchy and notices a blotchy rash on her skin. Did she get stung by a sea creature during her last dive? Becki decides to go back to her hotel room to get some rest. As she is walking, she begins to feel severe pains in the joints of her arms and legs and feels achy all over her body. Becki feels like she is coming down with the flu, but she realizes that her symptoms are related to her dives. What is wrong with Becki?

▶ Connect to the **BIG**IDEA As you read about the behavior of gases, think about what may have caused Becki's symptoms.

NATIONAL SCIENCE EDUCATION STANDARDS:

A-1, A-2, B-2, B-5, E-2, F-1

14.1 Properties of Gases

CHEMISTRY & YOU

Q: *Why is there a recommended pressure range for the air inside a soccer ball?* In organized soccer, there are rules about the equipment used in a game. For example, in international competitions, the ball's mass must not be more than 450 grams and not less than 410 grams. The pressure of the air inside the ball must be no lower than 0.6 atmospheres and no higher than 1.1 atmospheres at sea level.

In this lesson, you will study variables that affect the pressure of a gas. As you will discover, gas pressure is useful in a number of different objects, including auto air bags, inflatable rafts, aerosol sprays, and, yes, soccer balls.

Key Questions

Why are gases easier to compress than solids or liquids?

What are the three factors that affect gas pressure?

Vocabulary

- compressibility

Compressibility

Why are gases easier to compress than solids or liquids?

Recall from Chapter 13 that a gas can expand to fill its container, unlike a solid or liquid. The reverse is also true. Gases are easily compressed, or squeezed into a smaller volume. **Compressibility** is a measure of how much the volume of matter decreases under pressure.

The compressibility of a gas plays an important role in auto safety. When a car comes to a sudden stop, the people in the car will continue to move forward unless they are restrained. The driver and any passengers are more likely to survive a collision if they are wearing seat belts to restrict their forward movement. Cars also contain air bags as a second line of defense. A sudden reduction in speed triggers a chemical reaction inside an air bag. One product of the reaction is nitrogen gas, which causes the bag to inflate. An inflated air bag keeps the driver from colliding with the steering wheel. On the front passenger side of the car, an inflated air bag keeps a passenger from colliding with the dashboard or windshield.

Why does a collision with an inflated air bag cause much less damage than a collision with a steering wheel or dashboard? When a person collides with an inflated air bag, as shown in Figure 14.1, the impact forces the molecules of nitrogen gas in the bag closer together. The compression of the gas absorbs the energy of the impact.

Figure 14.1 Compression of a Gas
Because gases can be compressed, the air bag absorbs some of the energy from the impact of a collision. Air bags work best when combined with seat belts.
Describe ***What happens to the gas molecules inside an air bag when a driver collides with the bag?***

Kinetic theory can explain why gases are compressed more easily than liquids or solids. **Gases are easily compressed because of the space between the particles in a gas.** Remember that the volume of the particles in a gas is small compared to the overall volume of the gas. So the distance between particles in a gas is much greater than the distance between particles in a liquid or solid. Under increased pressure, the particles in a gas are forced closer together, or compressed.

Figure 14.2 is a model of identical air samples in two different containers. Only oxygen and nitrogen—the two main gases in air—are represented. Each container has 8 nitrogen molecules and 2 oxygen molecules. In the larger container, the molecules are farther apart. In the smaller container, the air sample is compressed, and the molecules are closer together. Note that at STP, the distance between particles in an enclosed gas is about 10 times the diameter of a particle. However, it isn't practical to represent the actual distances between particles in all the molecular drawings of gases in this book. In order for the drawings to fit easily on a page, the particles are drawn closer together.

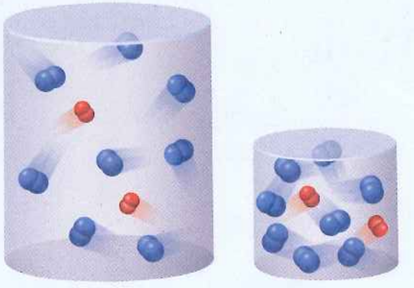


Figure 14.2 Modeling Air at Two Different Pressures
Air is primarily a mixture of two gases, nitrogen (N_2) and oxygen (O_2). A sample of air contains about 4 nitrogen molecules for every oxygen molecule.

Factors Affecting Gas Pressure

What are the three factors that affect gas pressure?

Kinetic theory can help explain other properties of gases, such as their ability to expand and take the shape and volume of their containers. Recall these assumptions about the particles in a gas. The particles move along straight-line paths until they collide with other particles or the walls of their container. The motion of the particles is constant and random. Because kinetic theory assumes there are no significant forces of attraction or repulsion among particles in a gas, particles in a gas can move freely.

Four variables are generally used to describe a gas. The variables and their common units are pressure (P) in kilopascals, volume (V) in liters, temperature (T) in kelvins, and the number of moles (n). **The amount of gas, the volume, and the temperature are factors that affect gas pressure.**

Learn more about gas properties *online* CONCEPTS IN ACTION

Figure 14.3 Gas in a Flexible Container
The volume of an air-filled raft is much larger than its volume before it was inflated. Using a pump to force air into a raft increases the pressure of the air inside the raft.
Compare ***How does an underinflated raft compare with a fully inflated raft? Why do you think an underinflated raft might be dangerous to ride in?***

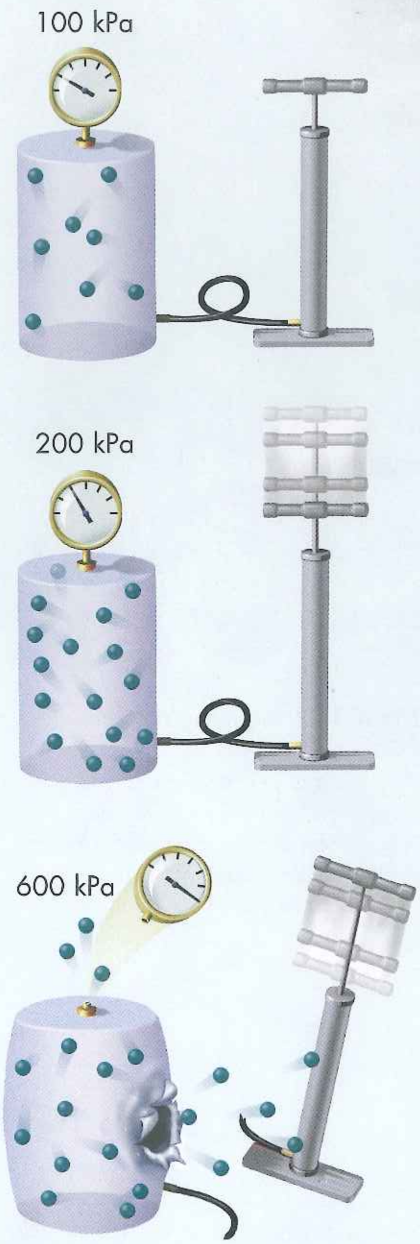


Figure 14.4 Gas in a Rigid Container When a gas is pumped into a closed rigid container, the pressure increases as more particles are added. If the number of particles is doubled, the pressure will double.

Amount of Gas An air-filled raft blasts through a narrow opening between rocks and plummets over a short waterfall into churning white water below. The raft bends and twists, absorbing some of the pounding energy of the river. The strength and flexibility of the raft rely on the pressure of the gas inside the raft. The raft must be made of a material that is strong enough to withstand the pressure of the air inside the raft. The material must also keep air from leaking out of the raft. The volume of the inflated raft in Figure 14.3 is dramatically larger than the volume of the raft before it is inflated. As air is added, the raft expands to its intended volume. The pressure of the air inside the raft keeps the raft inflated.

You can use kinetic theory to predict and explain how gases will respond to a change of conditions. If you inflate an air raft, for example, the pressure inside the raft will increase. Collisions of gas particles with the inside walls of the raft result in the pressure that is exerted by the enclosed gas. By adding gas, you increase the number of particles. Increasing the number of particles increases the number of collisions, which explains why the gas pressure increases.

Figure 14.4 shows what happens when gas is added to an enclosed, rigid container. Because the container is rigid, the volume of the gas is constant. Assume also that the temperature of the gas does not change. Under these conditions, doubling the number of particles of gas doubles the pressure. Tripling the number of particles triples the pressure, and so on. With a powerful pump and a strong container, you can generate very high pressures by adding more and more gas. However, once the pressure exceeds the strength of the container, the container will burst. Removing gas from a rigid container has the opposite effect. As the amount of gas is reduced, the pressure inside the container is reduced. If the number of particles in the container were cut in half, the pressure would drop by half.

If the pressure of the gas in a sealed container is lower than the outside air pressure, air will rush into the container when the container is opened. This movement causes the whoosh you hear when you open a vacuum-packed container. When the pressure of a gas in a sealed container is higher than the outside air pressure, the gas will flow out of the container when the container is unsealed.

The operation of an aerosol can depends on the movement of a gas from a region of high pressure to a region of lower pressure. Aerosol cans may contain whipped cream, hair mousse, or spray paint. Figure 14.5 shows how a can of spray paint works. The can contains a gas stored at high pressure. The air outside the can is at a lower pressure. Pushing the spray button creates an opening between the inside of the can and the air outside. The gas flows through the opening to the lower pressure region outside. The movement of the gas propels, or forces, the paint out of the can. As the gas is depleted, the pressure inside the can decreases until the gas can no longer propel paint from the can.

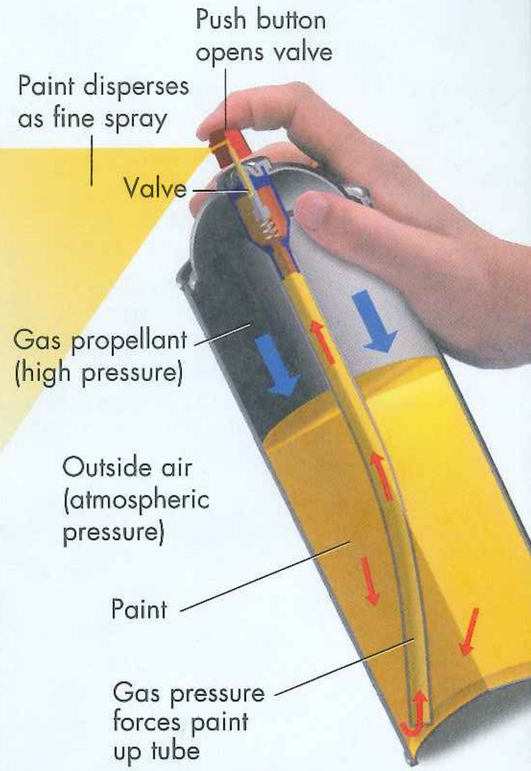


Figure 14.5 Aerosol Can
The pressure of the gas inside a new can of spray paint is greater than the air pressure outside the can. When gas rushes though an opening in the top of the can, it propels, or forces, paint out of the can. As the can is used, the pressure of the propellant decreases.
Relate Cause and Effect ***What happens when the pressure of the propellant equals the air pressure outside the can?***

Volume You can raise the pressure exerted by a contained gas by reducing its volume. The more the gas is compressed, the more pressure the gas exerts inside the container. When gas is in a cylinder, as in an automobile engine, a piston can be used to reduce its volume. The snug-fitting piston keeps gas from escaping as the cylinder moves down and up.

Figure 14.6 shows a cylinder of gas under two different conditions. When the cylinder has a volume of 1 L, the gas exerts a pressure of 100 kPa. When the volume is halved to 0.5 L, the pressure is doubled to 200 kPa. Increasing the volume of the contained gas has the opposite effect. If the volume is doubled, the particles can expand into a volume that is twice the original volume. With the same number of particles in twice the volume, the pressure of the gas is cut in half.

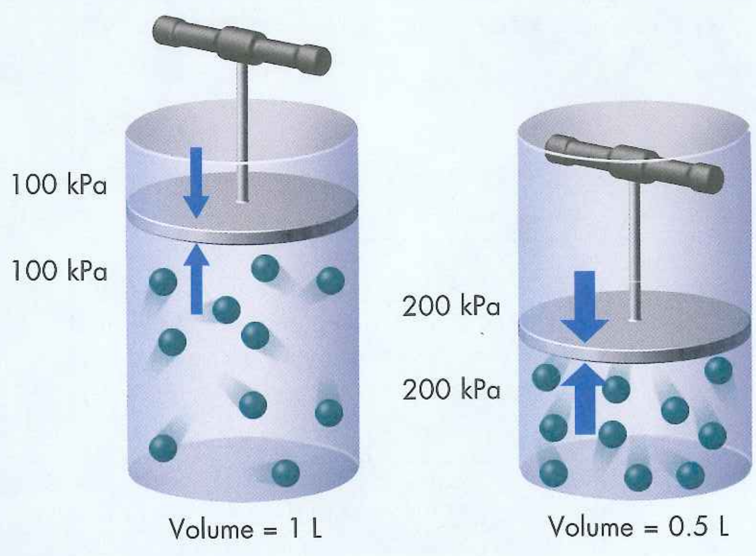


Figure 14.6 Pressure and Volume
A piston can be used to force a gas in a cylinder into a smaller volume. When the volume is decreased, the pressure the gas exerts is increased.
Interpret Diagrams ***What happens to the gas pressure when the volume is reduced from 1 L to 0.5 L?***

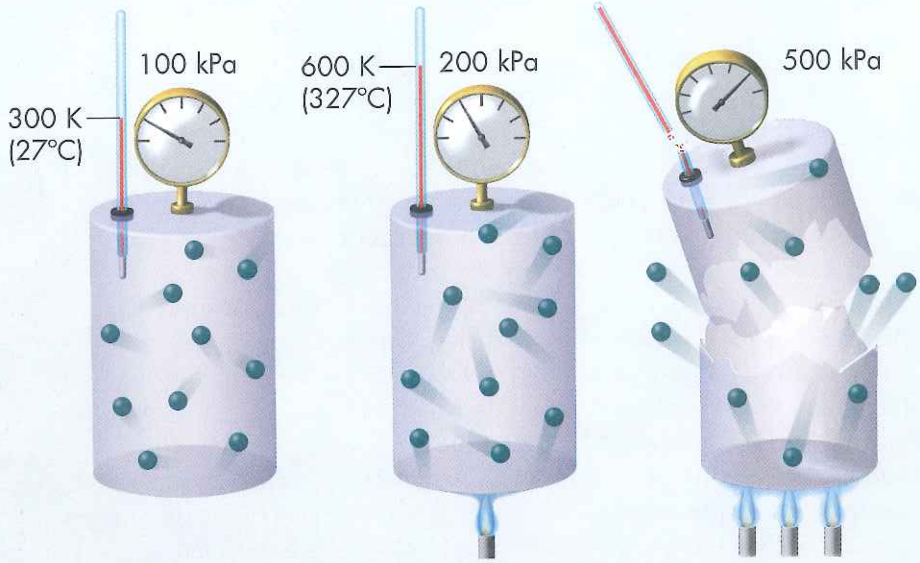


Figure 14.7 Temperature and Pressure
An increase in temperature causes an increase in the pressure of an enclosed gas. The container can explode if there is too great an increase in the gas pressure.

Temperature A sealed bag of potato chips bulges at the seams when placed in a sunny location. The bag bulges because an increase in the temperature of an enclosed gas causes an increase in its pressure. You can use kinetic theory to explain what happens. As a gas is heated, the temperature increases and the average kinetic energy of the particles in the gas increases. Faster-moving particles strike the walls of their container with more energy.

Look at Figure 14.7. The volume of the container and the amount of gas is constant. When the Kelvin temperature of the enclosed gas doubles from 300 K to 600 K, the pressure of the enclosed gas doubles from 100 kPa to 200 kPa. A gas in a sealed container may generate enormous pressure when heated. For that reason, an aerosol can, even an "empty" one, may explode if thrown onto a fire.

By contrast, as the temperature of an enclosed gas decreases, the pressure decreases. The particles, on average, move more slowly and have less kinetic energy. They strike the container walls with less force. Halving the Kelvin temperature of a gas in a rigid container decreases the gas pressure by half.

CHEMISTRY & YOU

Q: *Which do you think would travel farther if kicked with the same amount of force: a properly inflated soccer ball or an under-inflated soccer ball? What might happen to an overinflated soccer ball if you kicked it too hard?*

ONLINE PROBLEMS 14.1 LessonCheck

1. **Review** Why is a gas easy to compress?
2. **Identify** List three factors that can affect gas pressure.
3. **Compare and Contrast** Why does a collision with an air bag cause less damage than a collision with a steering wheel?
4. **Explain** How does a decrease in temperature affect the pressure of a contained gas?
5. **Apply Concepts** If the temperature is constant, what change in volume would cause the pressure of an enclosed gas to be reduced to one quarter of its original value?
6. **Apply Concepts** Assuming the gas in a container remains at a constant temperature, how could you increase the gas pressure in the container a hundredfold?
7. **Summarize** Write a paragraph explaining how a pressurized garden sprayer works. Make sure to describe what happens to the air pressure inside the sprayer as it is pumped by hand.

BIGIDEA KINETIC THEORY

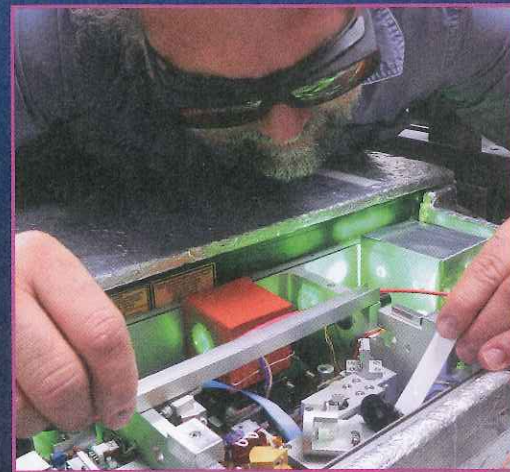
8. Use the kinetic theory of gases to explain why a gas can be easily squeezed into a smaller volume.

Atmospheric Chemist

Earth's atmosphere is a mixture of many gases, including oxygen, nitrogen, water vapor, carbon dioxide, methane, and ozone. Each one has an impact on life on Earth. The study of the chemical composition of the atmosphere is called atmospheric chemistry. Atmospheric chemists analyze the concentrations of atmospheric gases and determine how these gases chemically interact.

An important part of atmospheric research involves developing models that can predict the effects of fossil fuel emissions and other pollutants on air quality, climate, and the biosphere. Some atmospheric chemists study volcanic plumes, which are mixtures of hot gases and dust given off by volcanoes. Atmospheric chemistry is not limited to Earth's atmosphere. With the aid of telescopes, atmospheric chemists can study the composition of atmospheres of distant planets.

Atmospheric research is often a collaboration among scientists from different disciplines, including chemistry, physics, climatology, and oceanography. Atmospheric chemists typically have a bachelor's degree in chemistry or atmospheric science. Many also have a graduate degree in a specific field of research.



TOOLS AND TECHNOLOGY An atmospheric chemist adjusts a device used to analyze the motion and composition of air in the atmosphere.

AIR QUALITY Smog is a form of air pollution caused by tailpipe and smokestack emissions. The work of atmospheric chemists can help communities better understand how human activity impacts local air quality.

Take It Further

1. Infer What kinds of data do you think atmospheric chemists collect to study gases in the atmosphere?

2. Research a Problem Ozone (O_3) is one of many gases that atmospheric chemists study. Research the ozone layer and describe how atmospheric ozone levels have changed over time.

14.2 The Gas Laws

CHEMISTRY & YOU

Q: *How do you fill up a hot air balloon?* A hot air balloon works on the principle that warm air is less dense than cooler air. To make a hot air balloon rise, the pilot heats the air inside the balloon. To make the balloon descend, the pilot releases hot air through a vent in the top of the balloon. In this section, you'll study the laws that allow you to predict gas behavior.

Key Question

How are the pressure, volume, and temperature of a gas related?

Vocabulary

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

Boyle's Law

How are the pressure and volume of a gas related?

Kinetic theory tells you that there is empty space between the particles in a gas. Imagine how an increase in pressure would affect the volume of a contained gas. **If the temperature is constant, as the pressure of a gas increases, the volume decreases.** In turn, as the pressure decreases, the volume increases. Robert Boyle was the first person to study this pressure-volume relationship in a systematic way. In 1662, Boyle proposed a law to describe the relationship. **Boyle's law** states that for a given mass of gas at constant temperature, the volume of the gas varies inversely with pressure.

Look at Figure 14.8. A gas with a volume of 1.0 L (V_1) is at a pressure of 100 kPa (P_1). As the volume increases to 2.0 L (V_2), the pressure decreases to 50 kPa (P_2). The product $P_1 \times V_1$ (100 kPa $\times$ 1.0 L = 100 kPa·L) is the same as the product $P_2 \times V_2$ (50 kPa $\times$ 2.0 L = 100 kPa·L). As the volume decreases to 0.5 L (V_3), the pressure increases to 200 kPa (P_3). Again, the product of the pressure and the volume equals 100 kPa·L.

Interpret Graphs

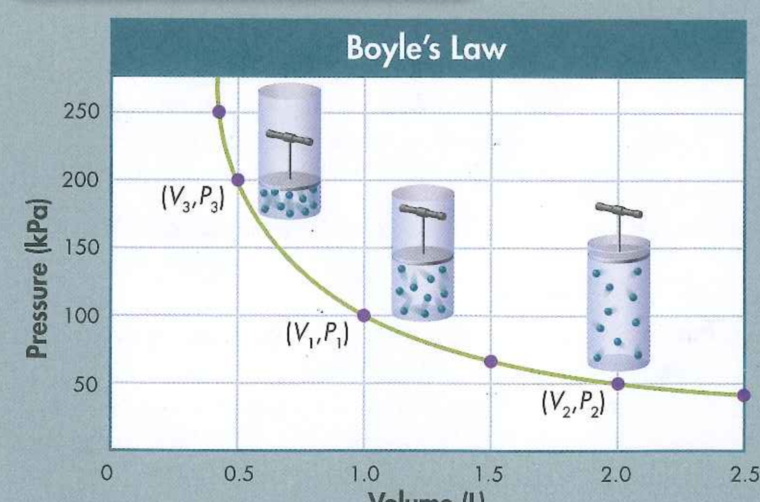


Figure 14.8 The pressure of a gas changes as the volume changes.

a. Read Graphs When the volume is 2.0 L, what is the pressure?

b. Predict What would the pressure be if the volume were increased to 3.0 L?

c. Draw Conclusions Based on the shape of the graph, describe the general pressure-volume relationship.

In an inverse relationship, the product of the two variable quantities is constant. So the product of pressure and volume at any two sets of pressure and volume conditions is always constant at a given temperature. The mathematical expression of Boyle's law is as follows.

$$P_1 \times V_1 = P_2 \times V_2$$

The graph of an inverse relationship is always a curve, as in Figure 14.8.

SampleProblem 14.1

Using Boyle's Law

A balloon contains 30.0 L of helium gas at 103 kPa. What is the volume of the helium when the balloon rises to an altitude where the pressure is only 25.0 kPa? (Assume that the temperature remains constant.)

KNOWNS	UNKNOWN
$P_1 = 103$ kPa	$V_2 = ?$ L
$V_1 = 30.0$ L	
$P_2 = 25.0$ kPa	

1 Analyze **List the knowns and the unknown.** Use Boyle's law ($P_1 \times V_1 = P_2 \times V_2$) to calculate the unknown volume (V_2).

2 Calculate **Solve for the unknown.**

Start with Boyle's law.

$$P_1 \times V_1 = P_2 \times V_2$$

Isolate V_2 by dividing both sides by P_2:

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

Rearrange the equation to isolate V_2.

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

Substitute the known values for P_1, V_1, and P_2 into the equation and solve.

$$V_2 = \frac{103\ \cancel{\text{kPa}} \times 30.0\ \text{L}}{25.0\ \cancel{\text{kPa}}}$$

$$= 1.24 \times 10^2\ \text{L}$$

3 Evaluate **Does the result make sense?** A decrease in pressure at constant temperature must correspond to a proportional increase in volume. The calculated result agrees with both kinetic theory and the pressure-volume relationship. The units have canceled correctly.

9. Nitrous oxide (N_2O) is used as an anesthetic. The pressure on 2.50 L of N_2O changes from 105 kPa to 40.5 kPa. If the temperature does not change, what will the new volume be?

10. A gas with a volume of 4.00 L at a pressure of 205 kPa is allowed to expand to a volume of 12.0 L. What is the pressure in the container if the temperature remains constant?

Solve Problem 10 by rearranging Boyle's law to isolate P_2.

Figure 14.9 Cooling Balloons in Liquid Nitrogen
When the gas in a balloon is cooled at constant pressure, the volume of the gas decreases.
Predict ***What would happen if you removed the balloons from the beaker and allowed them to warm back up to room temperature?***

Charles's Law

How are the temperature and volume of a gas related?

Figure 14.9 shows inflated balloons being dipped into a beaker of liquid nitrogen. For each balloon, the amount of air and the pressure are constant. As the air inside rapidly cools, the balloon shrinks. In fact, the gas volume decreases so much that all the cooled balloons can easily fit inside the beaker.

In 1787, the French physicist Jacques Charles studied the effect of temperature on the volume of a gas at constant pressure. When he graphed his data, Charles observed that a graph of gas volume versus temperature (in °C) is a straight line for any gas. **As the temperature of an enclosed gas increases, the volume increases if the pressure is constant.** When Charles extrapolated, or extended, the line to zero volume ($V = 0$), the line always intersected the temperature axis at −273.15°C. This value is equal to 0 on the Kelvin temperature scale. The observations that Charles made are summarized in Charles's law. **Charles's law** states that the volume of a fixed mass of gas is directly proportional to its Kelvin temperature if the pressure is kept constant. Look at the graph in Figure 14.10. When the temperature is 300 K, the volume is 1.0 L. When the temperature is 900 K, the volume is 3.0 L. In both cases, the ratio of V to T is 0.0033.

Interpret Graphs

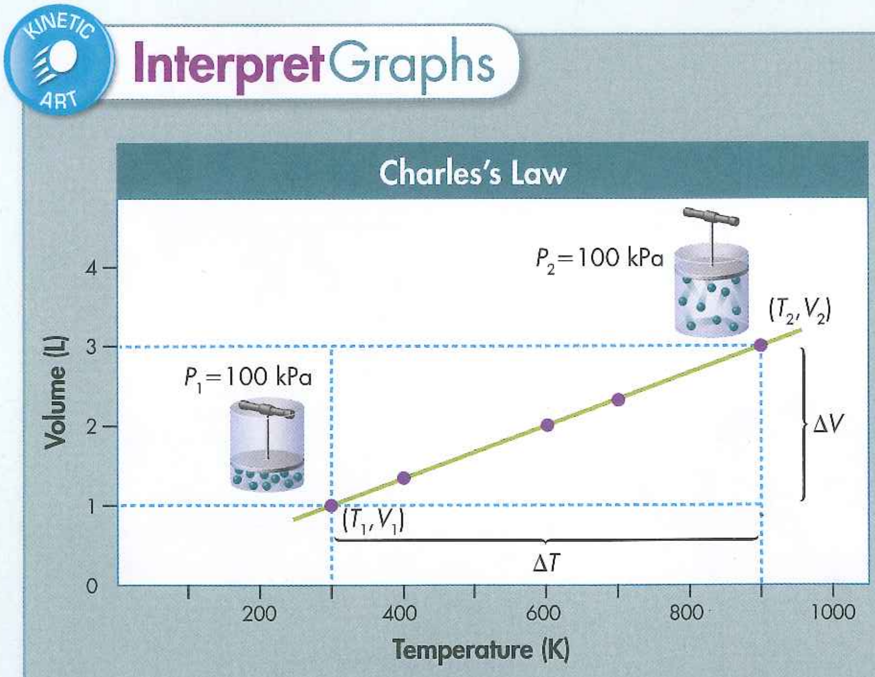


Figure 14.10 The graph shows how the volume changes as the temperature of a gas changes. View the Kinetic Art to see an online simulation of Charles's law.
a. Read Graphs In what unit is the temperature data expressed?
b. Draw Conclusions What happens to the volume as the temperature rises?
c. Predict If the temperature of a gas were 0 K, what would the volume of the gas be?

Hint: ΔV is the change in gas volume resulting from temperature change ΔT.

The ratio V_1/T_1 is equal to the ratio V_2/T_2. Because this ratio is constant at all conditions of temperature and volume, when the pressure is constant, you can write Charles's law as follows.

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

CHEMISTRY & YOU

Q: *A hot air balloon contains a propane burner onboard to heat the air inside the balloon. What happens to the volume of the balloon as the air is heated?*

The ratio of the variables is always a constant in a direct relationship, and the graph is always a straight line. It is not a direct relationship if the temperatures are expressed in degrees Celsius. So when you solve gas law problems, the temperature must always be expressed in kelvins.

CHEM TUTOR

Sample Problem 14.2

Using Charles's Law

A balloon inflated in a room at 24°C has a volume of 4.00 L. The balloon is then heated to a temperature of 58°C. What is the new volume if the pressure remains constant?

1 Analyze List the knowns and the unknown. Use Charles's law ($V_1/T_1 = V_2/T_2$) to calculate the unknown volume (V_2).

KNOWNS	UNKNOWN
$V_1 = 4.00$ L	$V_2 = ?$ L
$T_1 = 24°C$	
$T_2 = 58°C$	

2 Calculate Solve for the unknown.

Because you will use a gas law, start by expressing the temperatures in kelvins.

$T_1 = 24°C + 273 = 297$ K

$T_2 = 58°C + 273 = 331$ K

Write the equation for Charles's law.

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

Isolate V_2 by multiplying both sides by T_2:

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

Rearrange the equation to isolate V_2.

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

Substitute the known values for T_1, V_1, and T_2 into the equation and solve.

$$V_2 = \frac{4.00\text{ L} \times 331\cancel{\text{K}}}{297\cancel{\text{K}}} = \boxed{4.46\text{ L}}$$

3 Evaluate Does the result make sense? The volume increases as the temperature increases. This result agrees with both the kinetic theory and Charles's law.

11. If a sample of gas occupies 6.80 L at 325°C, what will its volume be at 25°C if the pressure does not change?

12. Exactly 5.00 L of air at −50.0°C is warmed to 100.0°C. What is the new volume if the pressure remains constant?

Gay-Lussac's Law

How are the pressure and temperature of a gas related?

When tires are not inflated to the recommended pressure, fuel efficiency and traction decrease. Treads can wear down faster. Most importantly, improper inflation can lead to tire failure. A driver should not check tire pressure after driving a long distance because the air in a tire heats up during a drive. **As the temperature of an enclosed gas increases, the pressure increases if the volume is constant.**

Joseph Gay-Lussac (1778–1850), a French chemist, discovered the relationship between the pressure and temperature of a gas in 1802. The gas law that describes the relationship bears his name. **Gay-Lussac's law** states that the pressure of a gas is directly proportional to the Kelvin temperature if the volume remains constant. Look at Figure 14.11. When the temperature is 300 K, the pressure is 100 kPa. When the temperature is doubled to 600 K, the pressure doubles to 200 kPa. Because Gay-Lussac's law involves direct proportions, the ratios P_1/T_1 and P_2/T_2 are equal at constant volume. You can write Gay-Lussac's law as follows:

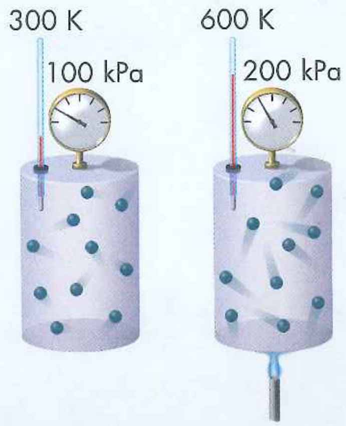


Figure 14.11 Gay-Lussac's Law
When a gas is heated at constant volume, the pressure increases.
Interpret Diagrams ***How can you tell from the drawings that there is a fixed amount of gas in the cylinders?***

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

Gay-Lussac's law can be applied to reduce the time it takes to cook food. One cooking method involves placing food above a layer of water and heating the water. The water vapor, or steam, that is produced cooks the food. Steam that escapes from the pot is at a temperature of about 100°C when the pressure is near one atmosphere. In a pressure cooker, like the one shown in Figure 14.12, steam is trapped inside the cooker. The temperature of the steam reaches about 120°C. The food cooks faster at this higher temperature, but the pressure rises, which increases the risk of an explosion. A pressure cooker has a valve that allows some vapor to escape when the pressure exceeds the set value.

Figure 14.12 Pressure Cooker
A pressure cooker is a gas-tight container in which pressurized steam is used to cook food. With the lid locked, the volume of steam and the number of water molecules are constant. So any increase in temperature causes an increase in pressure.

Sample Problem 14.3

Using Gay-Lussac's Law

Aerosol cans carry labels warning not to incinerate (burn) the cans or store them above a certain temperature. This problem will show why it is dangerous to dispose of aerosol cans in a fire. The gas in a used aerosol can is at a pressure of 103 kPa at 25°C. If the can is thrown onto a fire, what will the pressure be when the temperature reaches 928°C?

1 Analyze **List the knowns and the unknown.** Use Gay-Lussac's law ($P_1/T_1 = P_2/T_2$) to calculate the unknown pressure (P_2). Remember, because this problem involves temperatures and a gas law, the temperatures must be expressed in kelvins.

KNOWNS	UNKNOWN
$P_1 = 103\ \text{kPa}$	$P_2 = ?\ \text{kPa}$
$T_1 = 25°\text{C}$	
$T_2 = 928°\text{C}$	

2 Calculate **Solve for the unknown.**

Start by converting the two known temperatures from degrees Celsius to kelvins.

$$T_1 = 25°\text{C} + 273 = 298\ \text{K}$$
$$T_2 = 928°\text{C} + 273 = 1201\ \text{K}$$

Write the equation for Gay-Lussac's law.

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

Isolate P_2 by multiplying both sides by T_2:

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

Rearrange the equation to isolate P_2.

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

Substitute the known values for P_1, T_2, and T_1 into the equation and solve.

$$P_2 = \frac{103\ \text{kPa} \times 1201\ \cancel{\text{K}}}{298\ \cancel{\text{K}}}$$
$$= 415\ \text{kPa}$$
$$= \boxed{4.15 \times 10^2\ \text{kPa}}$$

3 Evaluate **Does the result make sense?** From the kinetic theory, one would expect the increase in temperature of a gas to produce an increase in pressure if the volume remains constant. The calculated value does show such an increase.

13. The pressure in a sealed plastic container is 108 kPa at 41°C. What is the pressure when the temperature drops to 22°C? Assume that the volume has not changed.

14. The pressure in a car tire is 198 kPa at 27°C. After a long drive, the pressure is 225 kPa. What is the temperature of the air in the tire? Assume that the volume is constant.

To solve Problem 14, rearrange Gay-Lussac's law to isolate T_2.

The Combined Gas Law

How are the pressure, volume, and temperature of a gas related?

There is a single expression, called the **combined gas law,** that combines Boyle's law, Charles's law, and Gay-Lussac's law.

$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

When only the amount of gas is constant, the combined gas law describes the relationship among pressure, volume, and temperature.

Sample Problem 14.4

CHEM TUTOR

Using the Combined Gas Law

The volume of a gas-filled balloon is 30.0 L at 313 K and 153 kPa pressure. What would the volume be at standard temperature and pressure (STP)?

KNOWNS	UNKNOWN
$V_1 = 30.0$ L	$V_2 = ?$ L
$T_1 = 313$ K	
$P_1 = 153$ kPa	
$T_2 = 273$ K (standard temperature)	
$P_2 = 101.3$ kPa (standard pressure)	

1 Analyze **List the knowns and the unknown.** Use the combined gas law ($P_1V_1/T_1 = P_2V_2/T_2$) to calculate the unknown volume (V_2).

2 Calculate **Solve for the unknown.**

State the combined gas law.

$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

Isolate V_2 by multiplying both sides by T_2 and dividing both sides by P_2:

$$\frac{T_2}{P_2} \times \frac{P_1 \times V_1}{T_1} = \frac{\cancel{P_2} \times V_2}{\cancel{T_2}} \times \frac{\cancel{T_2}}{\cancel{P_2}}$$

Rearrange the equation to isolate V_2.

$$V_2 = \frac{P_1 \times V_1 \times T_2}{P_2 \times T_1}$$

Substitute the known quantities into the equation and solve.

$$V_2 = \frac{153\ \cancel{\text{kPa}} \times 30.0\ \text{L} \times 273\ \cancel{\text{K}}}{101.3\ \cancel{\text{kPa}} \times 313\ \cancel{\text{K}}} = 39.5\ \text{L}$$

3 Evaluate **Does the result make sense?** A decrease in temperature and a decrease in pressure have opposite effects on the volume. To evaluate the increase in volume, multiply V_1 (30.0 L) by the ratio of P_1 to P_2 (1.51) and the ratio of T_2 to T_1 (0.872). The result is 39.5 L.

15. A gas at 155 kPa and 25°C has an initial volume of 1.00 L. The pressure of the gas increases to 605 kPa as the temperature is raised to 125°C. What is the new volume?

16. A 5.00-L air sample has a pressure of 107 kPa at a temperature of −50.0°C. If the temperature is raised to 102°C and the volume expands to 7.00 L, what will the new pressure be?

Weather balloons, like the one in Figure 14.13, carry a package of data-gathering instruments up into the atmosphere. At an altitude of about 27,000 meters, the balloon bursts. The combined gas law can help to explain this situation. Both outside temperature and pressure drop as the balloon rises. These changes have opposite effects on the volume of the weather balloon. A drop in temperature causes the volume of an enclosed gas to decrease. A drop in outside pressure causes the volume to increase. Given that the balloon bursts, the drop in pressure must affect the volume more than the drop in temperature does.

The combined gas law can also help you solve gas problems when only two variables are changing. It may seem challenging to remember four different expressions for the gas laws. But you actually only need to remember one expression—the combined gas law. You can derive the other laws from the combined gas law by holding one variable constant.

To illustrate, suppose you hold the temperature constant ($T_1 = T_2$). Rearrange the combined gas law so that the two temperature terms are on the same side of the equation. Because $T_1 = T_2$, the ratio of T_1 to T_2 is equal to one. Multiplying by 1 does not change a value in an equation. So when the temperature is constant, you can delete the temperature ratio from the rearranged combined gas law. What you are left with is the equation for Boyle's law.

$$P_1 \times V_1 = P_2 \times V_2 \times \frac{\cancel{T_1}}{\cancel{T_2}}$$

$$P_1 \times V_1 = P_2 \times V_2$$

A similar process yields Charles's law when pressure remains constant and Gay-Lussac's law when volume remains constant.

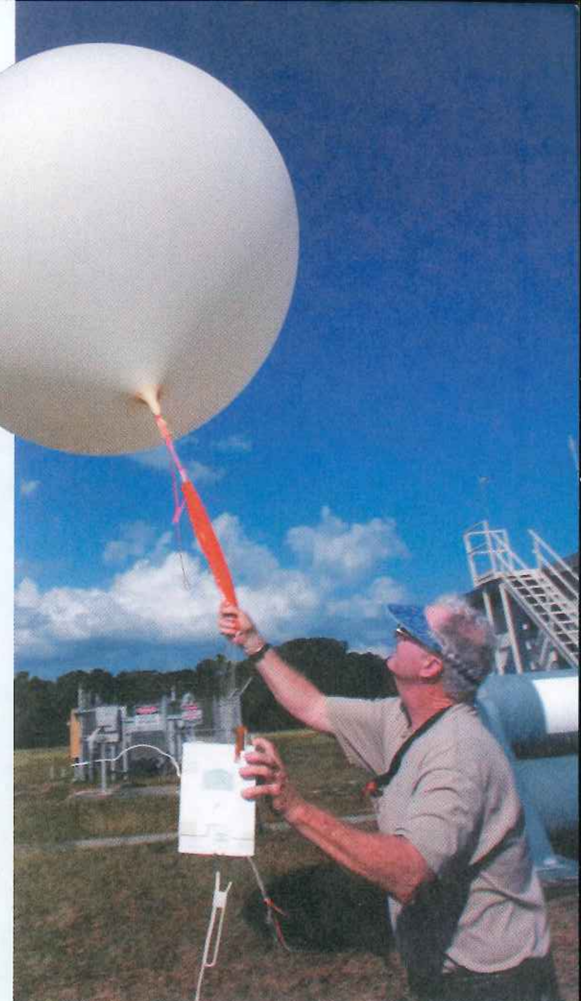


Figure 14.13 Weather Balloon
Meteorologists use weather balloons to gather data about Earth's atmosphere.
Infer ***Why is helium more likely to be used in weather balloons than air?***

ONLINE PROBLEMS 14.2 LessonCheck

17. **Review** How are the pressure and volume of a gas related at constant temperature?

18. **Review** If pressure is constant, how does a change in temperature affect the volume of a gas?

19. **Review** What is the relationship between the temperature and pressure of a contained gas at constant volume?

20. **Describe** In what situations is the combined gas law useful?

21. **Define** Write the mathematical equation for Boyle's law and explain the symbols.

22. **Calculate** A given mass of air has a volume of 6.00 L at 101 kPa. What volume will it occupy at 25.0 kPa if the temperature does not change?

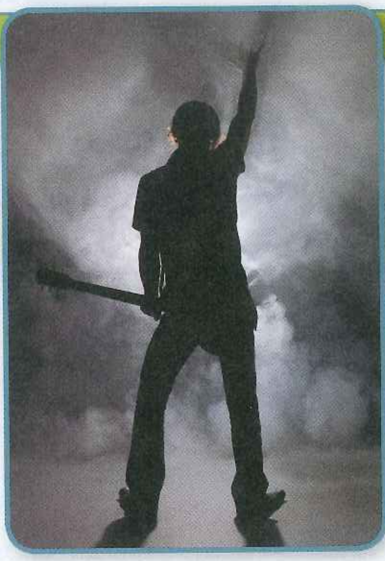
23. **Explain** How can Charles's law be derived from the combined gas law?

24. **Apply Concepts** The volume of a weather balloon increases as the balloon rises in the atmosphere. Why doesn't the drop in temperature at higher altitudes cause the volume to decrease?

BIGIDEA KINETIC THEORY

25. Why do you think scientists cannot collect temperature and volume data for an enclosed gas at temperatures near absolute zero?

14.3 Ideal Gases



CHEMISTRY & YOU

Q: *How can you make fog indoors?* Carbon dioxide freezes at −78.5°C, which is much colder than the ice in your freezer. Solid carbon dioxide, or dry ice, can be used to make stage fog. Dry ice doesn't melt—it sublimes. As solid carbon dioxide changes to gas, water vapor in the air condenses and forms a white fog. Dry ice can exist because gases don't obey the assumptions of kinetic theory at all conditions. In this section, you will learn how real gases differ from the ideal gases on which the gas laws are based.

Key Questions

How can you calculate the amount of a contained gas when the pressure, volume, and temperature are specified?

Under what conditions are real gases most likely to differ from ideal gases?

Vocabulary

- ideal gas constant
- ideal gas law

Ideal Gas Law

How can you calculate the amount of a contained gas when the pressure, volume, and temperature are specified?

Up to this point, you have worked with three variables that describe a gas: pressure, volume, and temperature. There is a fourth variable still to be considered: the amount of gas in the system, expressed in terms of the number of moles.

Suppose you want to calculate the number of moles (n) of a gas in a fixed volume at a known temperature and pressure. By modifying the combined gas law, you can solve for n. First, you must recognize that the volume occupied by a gas at a specified temperature and pressure depends on the number of particles. The number of moles of gas is directly proportional to the number of particles. So moles must be directly proportional to volume as well. You can now introduce moles into the combined gas law by dividing each side of the equation by n.

$$\frac{P_1 \times V_1}{T_1 \times n_1} = \frac{P_2 \times V_2}{T_2 \times n_2}$$

This equation shows that $(P \times V)/(T \times n)$ is a constant. This constant holds for what are called ideal gases—gases that conform to the gas laws.

If you know the values for P, V, T, and n for one set of conditions, you can calculate a value for the constant. Recall that 1 mol of every gas occupies 22.4 L at STP (101.3 kPa and 273 K). You can use these values to find the value of the constant, which has the symbol R and is called the ideal gas constant. Insert the values of P, V, T, and n into $(P \times V)/(T \times n)$.

$$\mathrm{R} = \frac{P \times V}{T \times n} = \frac{101.3\ \text{kPa} \times 22.4\ \text{L}}{273\ \text{K} \times 1\ \text{mol}} = 8.31\ (\text{L}\cdot\text{kPa})/(\text{K}\cdot\text{mol})$$

The **ideal gas constant** (R) has the value 8.31 $(L \cdot kPa)/(K \cdot mol)$. The gas law that includes all four variables—P, V, T, and n—is called the **ideal gas law.** It is usually written as follows.

$$P \times V = n \times R \times T \text{ or } PV = nRT$$

When the pressure, volume, and temperature of a contained gas are known, you can use the ideal gas law to calculate the number of moles of the gas. The amount of helium in a balloon, the amount of air in a scuba tank or a bicycle tire—each of these quantities can be calculated using the ideal gas law as long as you know the values for P, V, and T in each case.

CHEM TUTOR

Sample Problem 14.5

Using the Ideal Gas Law

At 34°C, the pressure inside a nitrogen-filled tennis ball with a volume of 0.148 L is 212 kPa. How many moles of nitrogen gas are in the tennis ball?

1 Analyze **List the knowns and the unknown.** Use the ideal gas law ($P \times V = n \times R \times T$) to calculate the number of moles (n).

KNOWNS	UNKNOWN
$P = 212$ kPa	$n = ?$ mol N_2
$V = 0.148$ L	
$T = 34°C$	
$R = 8.31\ (L \cdot kPa)/(K \cdot mol)$	

2 Calculate **Solve for the unknown.**

Convert degrees Celsius to kelvins.

$$T = 34°C + 273 = 307\text{ K}$$

State the ideal gas law.

$$P \times V = n \times R \times T$$

Isolate n by dividing both sides by $(R \times T)$:

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

Rearrange the equation to isolate n.

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

Substitute the known values for P, V, R, and T into the equation and solve.

$$n = \frac{P \times V}{R \times T} = \frac{212\ \cancel{kPa} \times 0.148\ \cancel{L}}{8.31\ (\cancel{L} \cdot \cancel{kPa})/(\cancel{K} \cdot mol) \times 307\ \cancel{K}} = 0.0123 \text{ mol } N_2$$

$$= \boxed{1.23 \times 10^{-2} \text{ mol } N_2}$$

3 Evaluate **Does the result make sense?** A tennis ball has a small volume and is not under great pressure. It is reasonable that the ball contains a small amount of nitrogen.

26. When the temperature of a rigid hollow sphere containing 685 L of helium gas is held to 621 K, the pressure of the gas is 1.89×10^3 kPa. How many moles of helium does the sphere contain?

27. What pressure will be exerted by 0.450 mol of a gas at 25°C if it is contained in a 0.650-L vessel?

Solve Problem 27 by rearranging the ideal gas law to isolate P.

Sample Problem 14.6

Using the Ideal Gas Law

A deep underground cavern contains 2.24×10^6 L of methane gas (CH_4) at a pressure of 1.50×10^3 kPa and a temperature of 315 K. How many kilograms of CH_4 does the cavern contain?

1 Analyze List the knowns and the unknown. Calculate the number of moles (n) using the ideal gas law. Use the molar mass of methane to convert moles to grams. Then convert grams to kilograms.

KNOWNS	UNKNOWN
$P = 1.50 \times 10^3$ kPa	$m = ?$ kg CH_4
$V = 2.24 \times 10^6$ L	
$T = 315$ K	
$R = 8.31$ (L·kPa) / (K·mol)	
molar mass$_{CH_4}$ = 16.0g	

2 Calculate Solve for the unknown.

State the ideal gas law.

$$P \times V = n \times R \times T$$

Rearrange the equation to isolate *n*.

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

Substitute the known quantities into the equation to find the number of moles of methane.

$$n = \frac{(1.50 \times 10^3\ \cancel{\text{kPa}}) \times (2.24 \times 10^6\ \cancel{\text{L}})}{8.31\ \frac{\cancel{\text{L}} \cdot \cancel{\text{kPa}}}{\cancel{\text{K}} \cdot \text{mol}} \times 315\ \cancel{\text{K}}} = 1.28 \times 10^6\ \text{mol}\ CH_4$$

Do a mole-mass conversion.

$$1.28 \times 10^6\ \cancel{\text{mol}\ CH_4} \times \frac{16.0\ \text{g}\ CH_4}{1\ \cancel{\text{mol}\ CH_4}} = 20.5 \times 10^6\ \text{g}\ CH_4$$

$$= 2.05 \times 10^7\ \text{g}\ CH_4$$

Convert from grams to kilograms.

$$2.05 \times 10^7\ \cancel{\text{g}}\ CH_4 \times \frac{1\ \text{kg}}{10^3\ \cancel{\text{g}}} = \boxed{2.05 \times 10^4\ \text{kg}\ CH_4}$$

3 Evaluate Does the result make sense? Although the methane is compressed, its volume is still very large. So it is reasonable that the cavern contains a large mass of methane.

28. A child's lungs can hold 2.20 L. How many grams of air do her lungs hold at a pressure of 102 kPa and a body temperature of 37°C? Use a molar mass of 29 g for air, which is about 20% O_2 (32 g/mol) and 80% N_2 (28 g/mol).

29. What volume will 12.0 g of oxygen gas (O_2) occupy at 25°C and a pressure of 52.7 kPa?

In Problems 28 and 29, make sure to express the temperature in kelvins before substituting for *T* in the ideal gas law equation.

Ideal Gases and Real Gases

Under what conditions are real gases most likely to differ from ideal gases?

An ideal gas is one that follows the gas laws at all conditions of pressure and temperature. Such a gas would have to conform precisely to the assumptions of kinetic theory. Its particles could have no volume, and there could be no attraction between particles in the gas. As you probably suspect, there is no gas for which these assumptions are true. So an ideal gas does not exist. Nevertheless, at many conditions of temperature and pressure, a real gas behaves very much like an ideal gas.

The particles in a real gas do have volume, and there are attractions between the particles. Because of these attractions, a gas can condense, or even solidify, when it is compressed or cooled. For example, if water vapor is cooled below 100°C at standard atmospheric pressure, it condenses to a liquid. The behavior of other real gases is similar, although lower temperatures and greater pressures may be required. Such conditions are required to produce the liquid nitrogen in Figure 14.14. **Real gases differ most from an ideal gas at low temperatures and high pressures.**

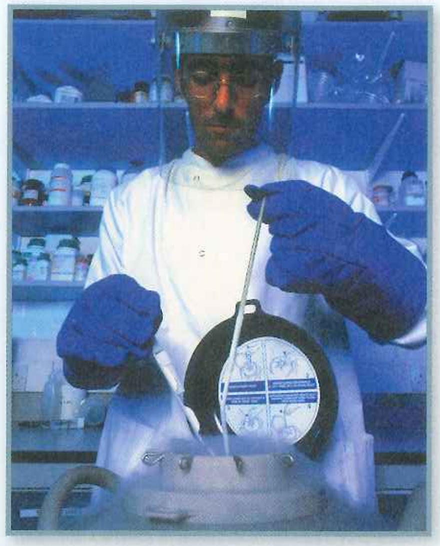


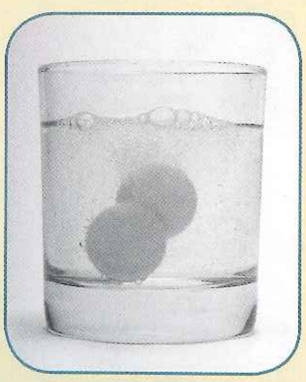
Figure 14.14 Liquid Nitrogen
A lab technician places a cell sample into an insulated tank containing liquid nitrogen. Nitrogen boils at −196°C.

Quick Lab

Purpose To measure the amount of carbon dioxide gas given off when antacid tablets dissolve in water

Materials

- 6 effervescent antacid tablets
- 3 rubber balloons (spherical)
- plastic medicine dropper
- water
- clock or watch
- metric tape measure
- graph paper
- water
- pressure sensor (optional)



Carbon Dioxide From Antacid Tablets

Procedure

1. Break six antacid tablets into small pieces. Keep the pieces from each tablet in a separate pile. Put the pieces from one tablet into the first balloon. Put the pieces from two tablets into a second balloon. Put the pieces from three tablets into a third balloon. CAUTION *If you are allergic to latex, do not handle the balloons.*

2. After you use the medicine dropper to squirt about 5 mL of cold water into each balloon, immediately tie off each balloon.

3. Shake the balloons to mix the contents. Allow the contents to warm to room temperature.

4. Measure and record the circumference of each balloon several times during the next 20 minutes.

5. Use the maximum circumference of each balloon to calculate its volume. (*Hint:* For the volume of a sphere, use $V = \frac{4}{3}\pi r^3$ and $r = \text{circumference}/2\pi$.)

Analyze and Conclude

1. Graph Make a graph of volume versus number of tablets. Use your graph to describe the relationship between the number of tablets used and the volume of the balloon.

2. Calculate Assume that the balloon is filled with carbon dioxide gas at 20°C and standard pressure. Calculate the mass and the number of moles of CO_2 in each balloon at maximum inflation.

3. Analyze Data If a typical antacid tablet contains 2.0 g of sodium hydrogen carbonate, how many moles of CO_2 should one tablet yield? Compare this theoretical value with your results.

Interpret Graphs

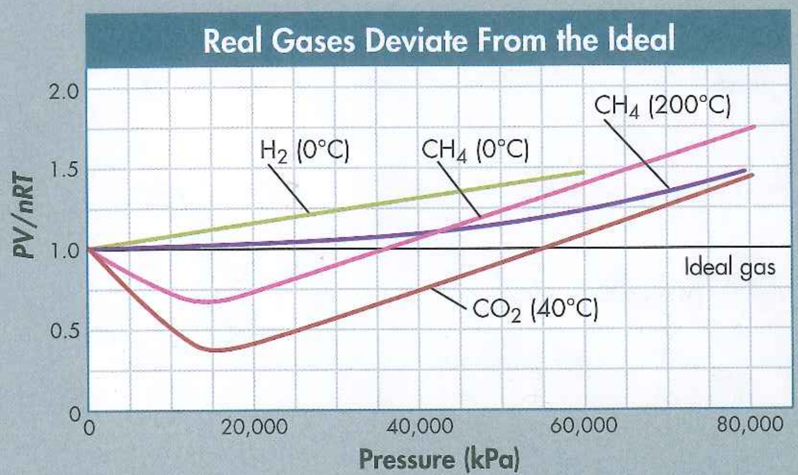


Figure 14.15 This graph shows how real gases deviate from the ideal gas law at high pressures.

a. Read Graphs What are the values of (*PV/nRT*) for an ideal gas at 20,000 and 60,000 kPa?

b. Identify What variable is responsible for the differences between the two methane (CH_4) curves?

c. Make Generalizations How does an increase in pressure affect the value of (*PV/nRT*) for real gases?

CHEMISTRY & YOU

Q: *Certain types of fog machines use dry ice and water to create stage fog. Chunks of dry ice are added to hot water, which causes the dry ice to sublime. The cold carbon dioxide gas causes water vapor in the surrounding air to condense into small droplets, resulting in a smoky fog. What phase changes occur when stage fog is made?*

Figure 14.15 shows how the value of the ratio (*PV/nRT*) changes as pressure increases. For an ideal gas, the result is a horizontal line because the ratio is always equal to 1. For real gases at high pressure, the ratio may deviate, or depart, from the ideal. When the ratio is greater than 1, the curve rises above the ideal gas line. When the ratio is less than 1, the curve drops below the line. The deviations can be explained by two factors. As attractive forces reduce the distance between particles, a gas occupies less volume than expected, causing the ratio to be less than 1. But the actual volume of the molecules causes the ratio to be greater than 1.

In portions of the curves below the line, intermolecular attractions dominate. In portions of the curves above the line, molecular volume dominates. Look at the curves for methane (CH_4) at 0°C and at 200°C. At 200°C, the molecules have more kinetic energy to overcome intermolecular attractions. So the curve for CH_4 at 200°C never drops below the line.

14.3 LessonCheck

30. Review How can you determine the number of moles of a contained gas when the pressure, volume, and temperature are known values?

31. Identify Under what conditions do real gases deviate most from ideal behavior?

32. Calculate Determine the volume occupied by 0.582 mol of a gas at 15°C if the pressure is 81.8 kPa.

33. Calculate You fill a rigid steel cylinder that has a volume of 20.0 L with nitrogen gas to a final pressure of 2.00×10^4 kPa at 28°C. How many kilograms of N_2 does the cylinder contain?

34. Compare What is the difference between a real gas and an ideal gas?

35. Analyze Data At standard pressure, ammonia condenses at −33.3°C but nitrogen does not condense until −195.79°C. Use what you know about bond polarity to explain this difference.

BIG IDEA KINETIC THEORY

36. Use the kinetic theory of gases to explain this statement: No gas exhibits ideal behavior at all temperatures and pressures.

14.4 Gases: Mixtures and Movements

CHEMISTRY & YOU

Q: *Why do balloons filled with helium deflate faster than balloons filled with air?* You have probably seen party balloons inflated with air or helium. The surface of a latex balloon has tiny pores through which gas particles can pass, causing the balloon to deflate over time. The rate at which the balloon deflates depends on the gas it contains.

Key Questions

How is the total pressure of a gas mixture related to the partial pressures of the component gases?

How does the molar mass of a gas affect the rate at which the gas diffuses or effuses?

Vocabulary

- partial pressure
- Dalton's law of partial pressures
- diffusion
- effusion
- Graham's law of effusion

Dalton's Law

How is the total pressure of a gas mixture related to the partial pressures of the component gases?

Gas pressure results from collisions of particles in a gas with an object. If the number of particles increases in a given volume, more collisions occur. If the average kinetic energy of the particles increases, more collisions occur. In both cases, the pressure increases. Gas pressure depends only on the number of particles in a given volume and on their average kinetic energy. Particles in a mixture of gases at the same temperature have the same average kinetic energy. So the kind of particle is not important.

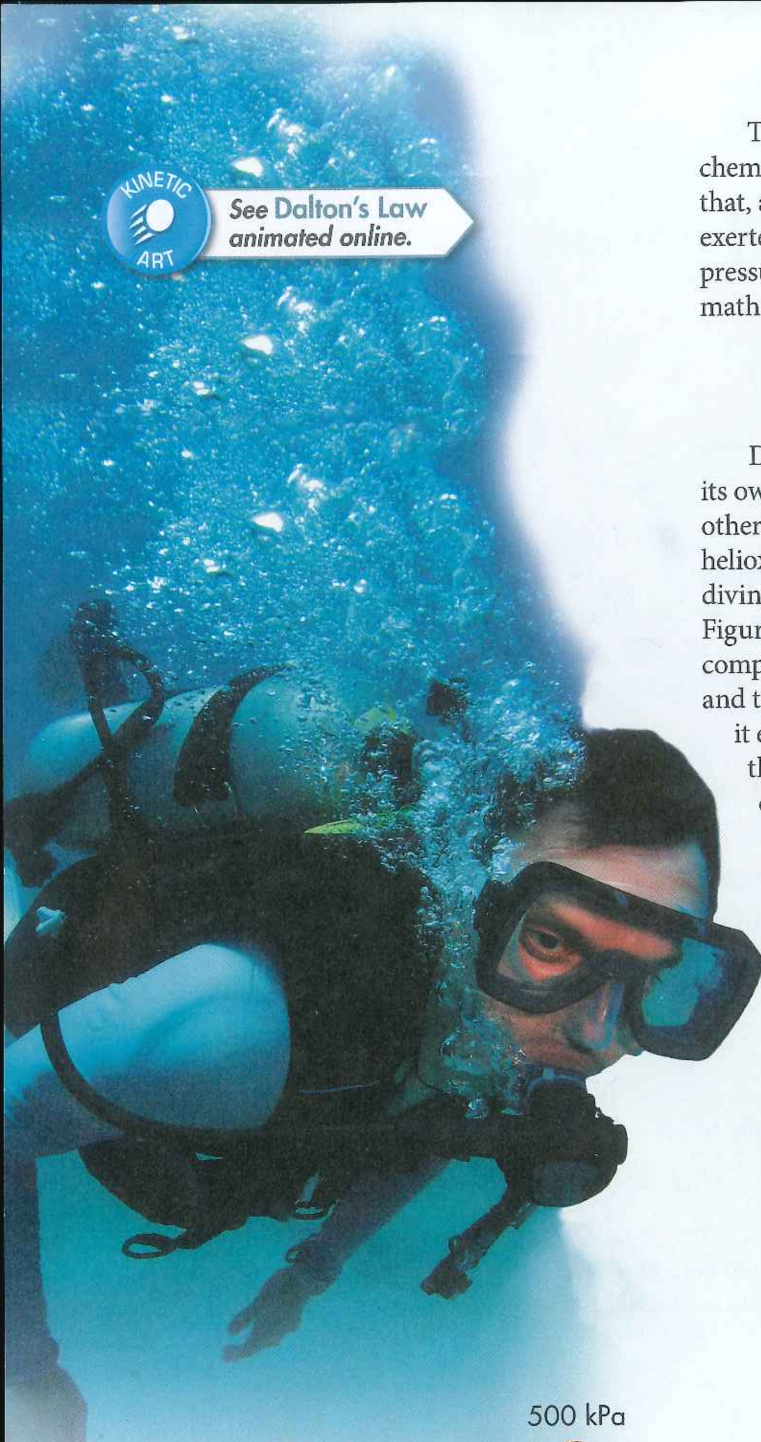
Table 14.1 shows the composition of dry air, or air that does not contain any water vapor. The contribution each gas in a mixture makes to the total pressure is called the **partial pressure** exerted by that gas. In dry air, the partial pressure of nitrogen is 79.11 kPa. **In a mixture of gases, the total pressure is the sum of the partial pressures of the gases.**

Interpret Data

Composition of Dry Air

Component	Volume (%)	Partial pressure (kPa)
Nitrogen	78.08	79.11
Oxygen	20.95	21.22
Carbon dioxide	0.04	0.04
Argon and others	0.93	0.95
Total	**100.00**	**101.32**

Table 14.1 The total pressure of dry air is the sum of the partial pressures of its component gases.
a. Read Tables What is the partial pressure of oxygen in dry air?
b. Predict As altitude increases, atmospheric pressure decreases. What do you think happens to the partial pressure of oxygen in the air as altitude increases?



KINETIC ART See Dalton's Law animated online.

The data in Table 14.1 illustrate a law proposed by the chemist John Dalton. **Dalton's law of partial pressures** states that, at constant volume and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the partial pressures of the component gases. You can express Dalton's law mathematically as follows:

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

Dalton's law holds true because each component gas exerts its own pressure independent of the pressure exerted by the other gases. Look at Figure 14.16a. The container is filled with heliox, a helium-oxygen gas mixture used in deep-sea scuba diving. The helium component of this mixture is shown in Figure 14.16b at the same volume and temperature. The oxygen component is shown in Figure 14.16c, also at the same volume and temperature. Each gas in the mixture exerts the pressure it exerted before the gases were mixed to make heliox. So the pressure in the container of heliox (500 kPa) is the sum of the pressures in the containers of helium and oxygen (400 kPa + 100 kPa).

If the percent composition of a mixture of gases does not change, the fraction of the pressure exerted by a gas does not change as the total pressure changes. This fact is important for people who must operate at high altitudes. For example, at the top of Mount Everest, the total atmospheric pressure is 33.73 kPa. This pressure is about one third of its value at sea level. The partial pressure of oxygen is also reduced by one third, to 7.06 kPa. But in order to support respiration in humans, the partial pressure of oxygen must be 10.67 kPa or higher. So climbers of Mount Everest need an oxygen mask and a cylinder of compressed oxygen to survive the ascent.

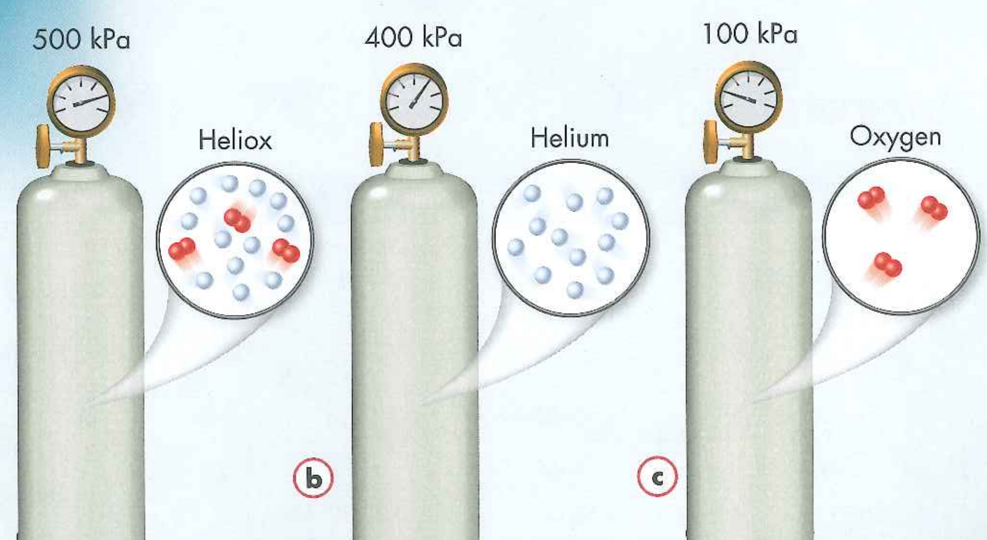


Figure 14.16 Dalton's Law
Heliox is a mixture of helium and oxygen gas. The pressure exerted by the helium in the mixture is independent of the pressure exerted by the oxygen. **Compare** ***How does the ratio of helium atoms to oxygen molecules in heliox compare to the ratio of the partial pressures?***

Sample Problem 14.7

Using Dalton's Law of Partial Pressures

Air contains oxygen, nitrogen, carbon dioxide, and trace amounts of other gases. What is the partial pressure of oxygen (P_{O_2}) at 101.30 kPa of total pressure if the partial pressures of nitrogen, carbon dioxide, and other gases are 79.10 kPa, 0.040 kPa, and 0.94 kPa, respectively?

1 Analyze List the knowns and the unknown. Use the equation for Dalton's law of partial pressures ($P_{total} = P_{O_2} + P_{N_2} + P_{CO_2} + P_{others}$) to calculate the unknown value (P_{O_2}).

KNOWNS	UNKNOWN
$P_{N_2} = 79.10$ kPa	$P_{O_2} = ?$ kPa
$P_{CO_2} = 0.040$ kPa	
$P_{others} = 0.94$ kPa	
$P_{total} = 101.30$ kPa	

2 Calculate Solve for the unknown.

Start with Dalton's law of partial pressures.

$$P_{total} = P_{O_2} + P_{N_2} + P_{CO_2} + P_{others}$$

Isolate P_{O_2} by subtracting the sum ($P_{N_2} + P_{CO_2} + P_{others}$) from both sides.

Rearrange Dalton's law to isolate P_{O_2}.

$$P_{O_2} = P_{total} - (P_{N_2} + P_{CO_2} + P_{others})$$

Substitute the values for P_{total} and the known partial pressures.

$$= 101.30\text{ kPa} - (79.10\text{ kPa} + 0.040\text{ kPa} + 0.94\text{ kPa})$$

$$= \boxed{21.22\text{ kPa}}$$

3 Evaluate Does this result make sense? The partial pressure of oxygen must be smaller than that of nitrogen because P_{total} is only 101.30 kPa. The other partial pressures are small, so the calculated answer of 21.22 kPa seems reasonable.

37. A gas mixture containing oxygen, nitrogen, and carbon dioxide has a total pressure of 32.9 kPa. If $P_{O_2} = 6.6$ kPa and $P_{N_2} = 23.0$ kPa, what is P_{CO_2}?

38. Determine the total pressure of a gas mixture that contains oxygen, nitrogen, and helium. The partial pressures are $P_{O_2} = 20.0$ kPa, $P_{N_2} = 46.7$ kPa, and $P_{He} = 26.7$ kPa.

In Problem 38, the unknown is P_{total}, so you can solve without having to rearrange Dalton's law.

Graham's Law

How does the molar mass of a gas affect the rate at which the gas diffuses or effuses?

READING SUPPORT

Build Vocabulary: *Prefixes* *Diffusion* and *effusion* come from the Latin *fundere* meaning "to pour." They differ only in their prefixes. The prefix *dif-* means "apart." The prefix *ex-* means "out." ***How do these prefixes help to contrast what happens to a gas during diffusion and effusion?***

Suppose you open a perfume bottle in one corner of a room. At some point, a person standing in the opposite corner will be able to smell the perfume. Molecules in the perfume evaporate and diffuse, or spread out, through the air in the room. **Diffusion** is the tendency of molecules to move toward areas of lower concentration until the concentration is uniform throughout.

The photo sequence in Figure 14.17 illustrates the diffusion process for bromine vapor. In Figure 14.17a, a glass cylinder containing air is inverted and sealed onto a cylinder containing bromine vapor. Figure 14.17b shows the bromine vapor diffusing through the air. The bromine vapor in the bottom cylinder has started to move upward into the top cylinder, where there is a lower concentration of bromine. In Figure 14.17c, the bromine has diffused to the top of the column formed by the combined cylinders. The concentration of bromine is now the same throughout the column.

Figure 14.17 Diffusion
The diffusion of one substance through another is a relatively slow process.
Describe ***How does the concentration of bromine in the bottom part of the column change during this sequence?***

(a) A cylinder of air and a cylinder of bromine vapor are sealed together.

(b) Bromine vapor diffuses upward through the air.

(c) After several hours, bromine vapors reach the top of the column.

There is another process that involves the movement of molecules in a gas. This process is called effusion. During **effusion,** a gas escapes through a tiny hole in its container. With effusion and diffusion, the type of particle is important. **Gases of lower molar mass diffuse and effuse faster than gases of higher molar mass.**

Thomas Graham's Contribution The Scottish chemist Thomas Graham studied rates of effusion during the 1840s. From his observations, he proposed a law. **Graham's law of effusion** states that the rate of effusion of a gas is inversely proportional to the square root of the gas's molar mass. This law can also be applied to the diffusion of gases.

Figure 14.18 Blimps
The cigar-shaped part of a blimp, called an envelope, is a sealed container of helium gas. **Infer** ***What properties do you think are desirable for the materials used to make the envelope of a blimp?***

CHEMISTRY & YOU

Q: *Why do balloons filled with helium deflate faster than balloons filled with air? Use Graham's law of effusion to explain your answer.*

Graham's law makes sense if you know how the mass, velocity, and kinetic energy of a moving object are related. The expression that relates the mass (m) and the velocity (v) of an object to its kinetic energy (KE) is $\frac{1}{2}mv^2$. For the kinetic energy to be constant, any increase in mass must be balanced by a decrease in velocity. For example, a ball with a mass of 2 g must travel at 5 m/s to have the same kinetic energy as a ball with a mass of 1 g traveling at 7 m/s. There is an important principle here. If two objects with different masses have the same kinetic energy, the lighter object must move faster.

Comparing Effusion Rates The blimp shown in Figure 14.18 is inflated with helium, which is less dense than air. One of the challenges in maintaining blimps is to keep the helium from seeping out. You may have noticed that party balloons filled with either helium or air gradually deflate over time. Both helium atoms and the molecules in air can pass through the tiny pores in a latex balloon. But a helium-filled balloon will deflate faster than an air-filled balloon. Kinetic theory can explain the difference.

Suppose you have two balloons, one filled with helium and the other filled with air. If the balloons are at the same temperature, the particles in each balloon have the same average kinetic energy. But helium atoms are less massive than oxygen or nitrogen molecules. So the molecules in air move more slowly than helium atoms with the same kinetic energy. Because the rate of effusion is related only to a particle's speed, Graham's law can be written as follows for two gases, A and B.

$$\frac{\text{Rate}_A}{\text{Rate}_B} = \sqrt{\frac{\text{molar mass}_B}{\text{molar mass}_A}}$$

In other words, the rates of effusion of two gases are inversely proportional to the square roots of their molar masses. Sample Problem 14.8 on the following page compares the effusion rates of helium and nitrogen.

Sample Problem 14.8

Comparing Effusion Rates

How much faster does helium (He) effuse than nitrogen (N_2) at the same temperature?

1 Analyze **List the knowns and the unknown.** Use Graham's law and the molar masses of the two gases to calculate the ratio of effusion rates.

KNOWNS

molar mass$_{He}$ = 4.0 g

molar mass$_{N_2}$ = 28.0 g

UNKNOWN

ratio of effusion rates = ?

2 Calculate **Solve for the unknown.** Helium effuses nearly three times faster than nitrogen at the same temperature.

Start with the equation for Graham's law of effusion.

$$\frac{\text{Rate}_{He}}{\text{Rate}_{N_2}} = \frac{\sqrt{\text{molar mass}_{N_2}}}{\sqrt{\text{molar mass}_{He}}}$$

Substitute the molar masses of nitrogen and helium into the equation.

$$\frac{\text{Rate}_{He}}{\text{Rate}_{N_2}} = \sqrt{\frac{28.0\,\cancel{g}}{4.0\,\cancel{g}}} = \sqrt{7.0} = \boxed{2.7}$$

3 Evaluate **Does the result make sense?** Helium atoms are less massive than nitrogen molecules, so it makes sense that helium effuses faster than nitrogen.

39. Calculate the ratio of the velocity of hydrogen molecules to the velocity of carbon dioxide molecules at the same temperature.

Use what you know about chemical formulas and the mole to write the molar mass of each gas.

14.4 LessonCheck

ONLINE PROBLEMS

40. Review In a mixture of gases, how is the total pressure determined?

41. Review What is the effect of molar mass on rates of diffusion and effusion?

42. Explain How is the partial pressure of a gas in a mixture calculated?

43. Calculate The pressure in an automobile tire filled with air is 245.0 kPa. If P_{O_2} = 51.3 kPa, P_{CO_2} = 0.10 kPa, and P_{others} = 2.3 kPa, what is P_{N_2}?

44. Compare What distinguishes effusion from diffusion? How are these processes similar?

45. Relate Cause and Effect Explain why the rates of diffusion of nitrogen gas and carbon monoxide are almost identical at the same temperature.

46. Analyze Data Both Table 14.1 on page 469 and the Elements in the Atmosphere table on page R1 list data on the composition of air. Look at the data included in each table. Identify two ways in which the tables are similar. Describe at least three differences.

Small-Scale Lab

Diffusion

(Probe or sensor version of this lab is available in the *Probeware Lab Manual.*)

Purpose

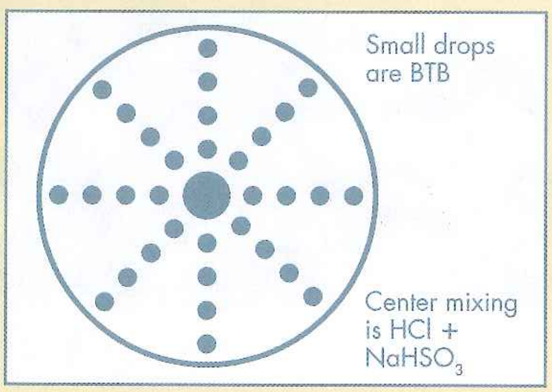
To infer diffusion of a gas by observing color changes during chemical reactions

Materials

- clear plastic cup or petri dish
- reaction surface
- dropper bottles containing bromothymol blue, hydrochloric acid, and sodium hydrogen sulfite
- ruler
- cotton swab
- NaOH, NH_4Cl (optional)

Procedure

1. Use the plastic cup or petri dish to draw the large circle shown below on a sheet of paper.



2. Place a reaction surface over the grid and add small drops of bromothymol blue (BTB) in the pattern shown by the small circles. Make sure the drops do not touch one another.

3. Mix one drop each of hydrochloric acid (HCl) and sodium hydrogen sulfite ($NaHSO_3$) in the center of the pattern.

4. Place the cup or petri dish over the grid and observe what happens.

5. If you plan to do Activity 1 in the You're the Chemist section, don't dispose of your materials yet.

Analyze

1. Observe Describe in detail the changes you observed in the drops of BTB over time. Draw pictures to illustrate the changes.

2. Describe Draw a series of pictures showing how one of the BTB drops might look over time if you could view the drop from the side.

3. Explain The BTB changed even though you added nothing to it. If the mixture in the center circle produced a gas, would this explain the change in the drops of BTB? Use kinetic theory to explain your answer.

4. Describe Translate the following word equation into a balanced chemical equation: Sodium hydrogen sulfite reacts with hydrochloric acid to produce sulfur dioxide gas, water, and sodium chloride.

You're the Chemist

The following activities allow you to develop your own procedures and analyze the results.

1. Analyze Data Carefully absorb the center mixture of the original experiment onto a cotton swab and replace it with one drop of NaOH and one drop of NH_4Cl. Describe what happens and explain in terms of kinetic theory. Ammonium chloride reacts with sodium hydroxide to produce ammonia gas, water, and sodium chloride. Write and balance a chemical equation to describe this reaction.

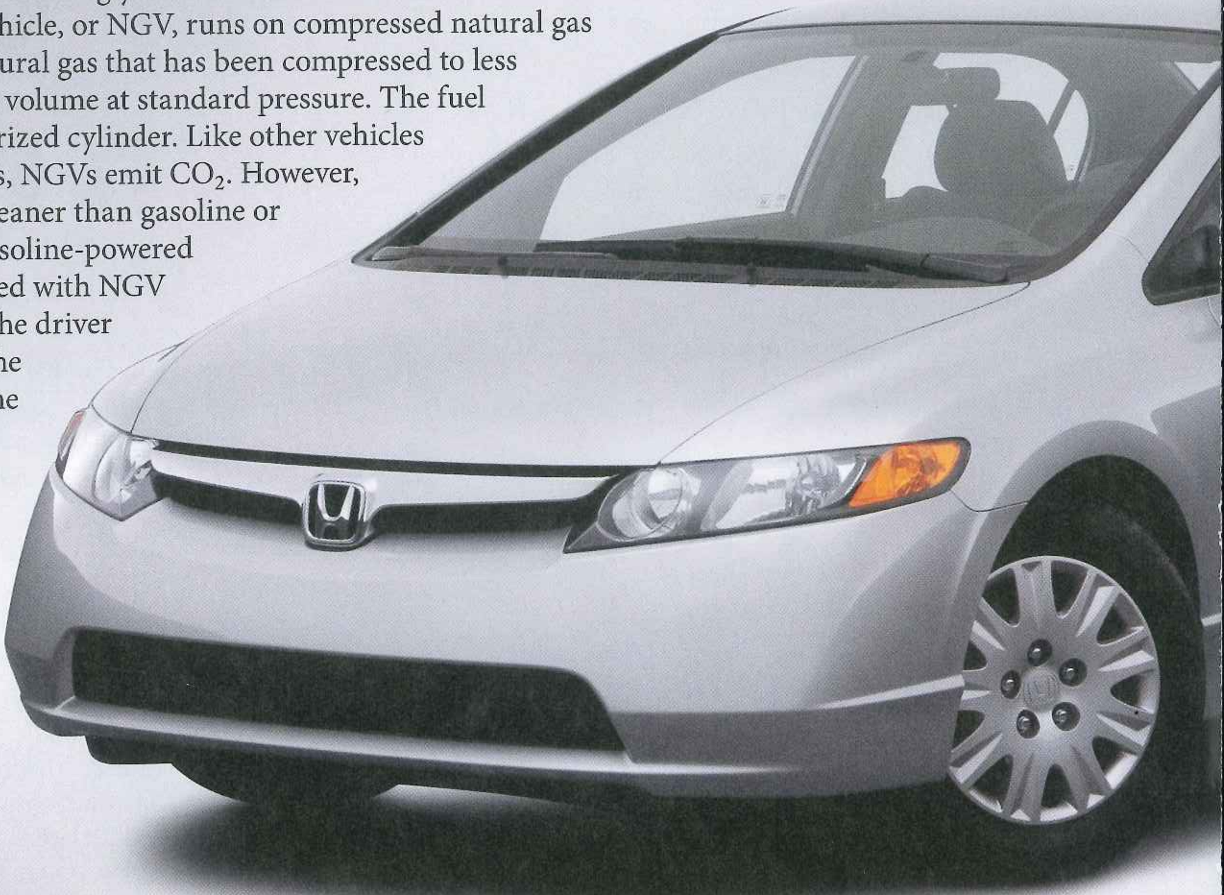
2. Design an Experiment Design an experiment to observe the effect of the size of the BTB drops on the rate at which they change. Explain your results in terms of kinetic theory.

CHEMISTRY & YOU: TECHNOLOGY

Natural Gas Vehicles

Most of the cars you see on the road run on gasoline, better known as gas—although the fuel itself is a liquid, not a gas. But in some U.S. cities, as well as parts of South America and Asia, vehicles that run on gaseous fuel are becoming increasingly common.

A natural gas vehicle, or NGV, runs on compressed natural gas (CNG), which is natural gas that has been compressed to less than 1 percent of its volume at standard pressure. The fuel is stored in a pressurized cylinder. Like other vehicles that burn fossil fuels, NGVs emit CO_2. However, natural gas burns cleaner than gasoline or diesel fuel. Many gasoline-powered cars can be retrofitted with NGV technology so that the driver can choose to run the car on either gasoline or CNG.



Pros & Cons

Advantages of NGVs

✔ **Less pollution** NGVs produce much less carbon monoxide, nitrogen oxides, and toxins than gasoline-powered vehicles.

✔ **Less maintenance** Because natural gas burns cleaner than gasoline, the engines of NGVs require less servicing than those of gasoline-powered vehicles.

✔ **Cheaper fuel** Natural gas costs less than gasoline.

✔ **Safety** The fuel tanks in NGVs are stronger and safer than gasoline storage tanks.

Disadvantages of NGVs

✘ **More expensive** NGVs tend to cost more than comparable gasoline-powered cars.

✘ **Less roomy** Due to the CNG tank, NGVs have less trunk space.

✘ **Limited travel range** On a single tank of gas, NGVs can travel only about 60 percent as far as gasoline-powered cars before needing more fuel.

✘ **Hard to refuel** CNG refueling stations are currently few and far between.

✘ **Still fossil-fueled** Like oil, natural gas is a nonrenewable resource.

ON THE ROAD Most of the NGVs found in the United States are buses. But don't be surprised if you start to see passenger cars sporting the "NGV" logo.

FILL 'ER UP Any building with a natural gas line can be equipped with a pressurized fueling device that delivers CNG. But NGV drivers must be patient—an empty CNG tank takes much longer to fill than an empty gasoline tank.

A REAL GAS IN THE GAS TANK A mid-sized NGV sedan has a fuel tank in the back with a volume of 8 gallons, or 30 L. At ambient temperatures, the pressure inside a full tank of CNG is about 3600 pounds per square inch (psi), or 25,000 kPa.

Take It Further

1. Calculate The natural gas in a 30-L NGV fuel tank has a pressure of 2.05×10^4 kPa at a temperature of 297 K. How many kilograms of fuel are in the tank? (Use a molar mass of 19 g/mol for natural gas.)

2. Calculate Natural gas is 89% methane (CH_4), 5% ethane (C_2H_6), 5% butane (C_4H_{10}), and 1% propane (C_3H_8). Use the data from Question 1 to determine the partial pressures of each component gas in the fuel tank.

14 Study Guide

BIGIDEA KINETIC THEORY

Ideal gases conform to the assumptions of kinetic theory. The behavior of ideal gases can be predicted by the gas laws. With the ideal gas law, the number of moles of a gas in a fixed volume at a known temperature and pressure can be calculated. Although an ideal gas does not exist, real gases behave ideally under a variety of temperature and pressure conditions.

14.1 Properties of Gases

Gases are easily compressed because of the space between the particles in a gas.

The amount of gas (n), volume (V), and temperature (T) are factors that affect gas pressure (P).

- compressibility (450)

14.2 The Gas Laws

If the temperature is constant, as the pressure of a gas increases, the volume decreases.

As the temperature of an enclosed gas increases, the volume increases if the pressure is constant.

As the temperature of an enclosed gas increases, the pressure increases if the volume is constant.

When only the amount of gas is constant, the combined gas law describes the relationship among pressure, volume, and temperature.

- Boyle's law (456)
- Charles's law (458)
- Gay-Lussac's law (460)
- combined gas law (462)

Key Equations

Boyle's law:
$$P_1 \times V_1 = P_2 \times V_2$$

Charles's law:
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

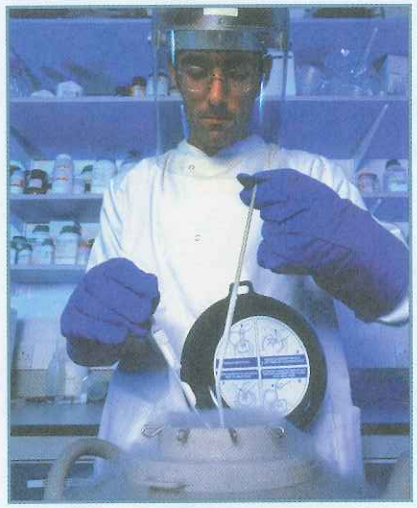
Gay-Lussac's law:
$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

combined gas law:
$$\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$$

14.3 Ideal Gases

When the pressure, volume, and temperature of a contained gas are known, you can use the ideal gas law to calculate the number of moles of the gas.

Real gases differ most from an ideal gas at low temperatures and high pressures.



- ideal gas constant (465)
- ideal gas law (465)

Key Equation

ideal gas law:
$$P \times V = n \times R \times T \text{ or } PV = nRT$$

14.4 Gases: Mixtures and Movements

In a mixture of gases, the total pressure is the sum of the partial pressures of the gases.

Gases of lower molar mass diffuse and effuse faster than gases of higher molar mass.

- partial pressure (469)
- Dalton's law of partial pressure (470)
- diffusion (472)
- effusion (472)
- Graham's law of effusion (472)

Key Equations

Dalton's law: $P_{total} = P_1 + P_2 + P_3 + ...$

Graham's law: $\frac{Rate_A}{Rate_B} = \sqrt{\frac{molar\ mass_B}{molar\ mass_A}}$

Math Tune-Up: Gas Law Problems

Problem	1 Analyze	2 Calculate	3 Evaluate
A 2.50-L sample of nitrogen gas at a temperature of 308 K has a pressure of 1.15 atm. What is the new volume of the gas if the pressure is increased to 1.80 atm and the temperature is decreased to 286 K?	Knowns: $P_1 = 1.15$ atm $V_1 = 2.50$ L $T_1 = 308$ K $P_2 = 1.80$ atm $T_2 = 286$ K Unknown: $V_2 = ?$ Use the combined gas law: $\frac{P_1 \times V_1}{T_1} = \frac{P_2 \times V_2}{T_2}$	Solve for V_2 and calculate: $V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}$ $V_2 = \frac{(1.15\ \cancel{\text{atm}})(2.50\ \text{L})(286\ \cancel{\text{K}})}{(1.80\ \cancel{\text{atm}})(308\cancel{\text{K}})}$ $V_2 = 1.48$ L	An increase in pressure causes the volume of a gas to decrease. Likewise, a decrease in temperature causes the volume of a gas to decrease. So, V_2 should be smaller than V_1. The answer makes sense. **Hint:** Review Sample Problem 14.4 if you have trouble with the combined gas law.
How many moles of helium gas fill a 6.45-L balloon at a pressure of 105 kPa and a temperature of 278 K?	Knowns: $P = 105$ kPa $V = 6.45$ L $T = 278$ K $R = 8.31$ L·kPa/K·mol Unknown: $n = ?$ Use the ideal gas law: $PV = nRT$	Solve for n and calculate: $n = \frac{PV}{RT}$ $n = \frac{(105\ \cancel{\text{kPa}})(6.45\ \cancel{\text{L}})}{\left(8.31\frac{\cancel{\text{L}}\cdot\cancel{\text{kPa}}}{\cancel{\text{K}}\cdot\text{mol}}\right)(278\cancel{\text{K}})}$ $n = 0.293$ mol	The gas is not at high pressure, nor is the volume large. So the number of moles in the balloon should be small. The answer is reasonable, and the units have canceled correctly.
A gas mixture containing argon, krypton, and helium has a total pressure of 376 kPa. If the partial pressures of argon and krypton are 92 kPa and 144 kPa, respectively, what is the partial pressure of helium?	Knowns: $P_{total} = 376$ kPa $P_{Ar} = 92$ kPa $P_{Kr} = 144$ kPa Unknown: $P_{He} = ?$ Dalton's law of partial pressures applies: $P_{total} = P_{Ar} + P_{He} + P_{Kr}$	Solve for P_{He} and calculate: $P_{He} = P_{total} - (P_{Ar} + P_{Kr})$ $P_{He} = 376\ \text{kPa} - (92\ \text{kPa} + 144\ \text{kPa})$ $P_{He} = 140$ kPa $P_{He} = 1.40 \times 10^2$ kPa	The partial pressure of helium must be less than half the total pressure. The answer is reasonable. **Dalton's law:** The total pressure exerted by a mixture of gases (P_{total}) is equal to the sum of the partial pressures of the component gases.

14 Assessment

★ Solutions appear in Appendix E

Lesson by Lesson

14.1 Properties of Gases

47. What happens to the particles in a gas when the gas is compressed?

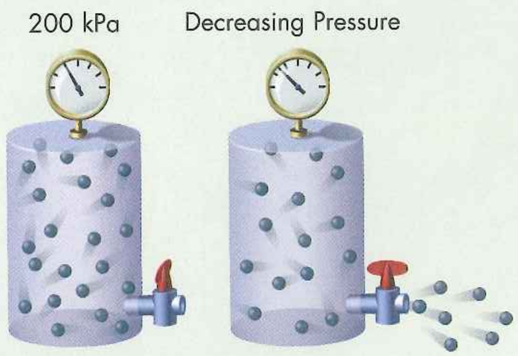
48. Explain why heating a contained gas that is held at a constant volume increases its pressure.

49. Describe what happens to the volume of a balloon when it is taken outside on a cold winter day. Explain why the observed change happens.

50. A metal cylinder contains 1 mol of nitrogen gas. What will happen to the pressure if another mole of gas is added to the cylinder, but the temperature and volume do not change?

51. If a gas is compressed from 4 L to 1 L and the temperature remains constant, what happens to the pressure?

52. Use the drawing to help explain why gas pressure decreases when gas is removed from a container with a fixed volume.



14.2 The Gas Laws

53. Write the mathematical equation for Charles's law and explain the symbols.

★**54.** The gas in a closed container has a pressure of 3.00×10^2 kPa at 30°C (303 K). What will the pressure be if the temperature is lowered to −172°C (101 K)?

55. Calculate the volume of a gas (in L) at a pressure of 1.00×10^2 kPa if its volume at 1.20×10^2 kPa is 1.50×10^3 mL.

56. A gas with a volume of 4.0 L at 90.0 kPa expands until the pressure drops to 20.0 kPa. What is its new volume if the temperature doesn't change?

★**57.** A gas with a volume of 3.00×10^2 mL at 150.0°C is heated until its volume expands to 6.00×10^2 mL. What is the new temperature of the gas if the pressure remains constant during the heating process?

★**58.** A gas with a volume of 15 L at 327°C is cooled at constant pressure until the volume reaches 5 L. What is the new temperature of the gas?

59. Write the mathematical expression for the combined gas law.

60. A sealed cylinder of gas contains nitrogen gas at 1.00×10^3 kPa pressure and a temperature of 20°C. When the cylinder is left in the sun, the temperature of the gas increases to 50°C. What is the new pressure in the cylinder?

★**61.** A sample of nitrogen gas has a pressure of 6.58 kPa at 539 K. If the volume does not change, what will the pressure be at 211 K?

62. Show how Gay-Lussac's law can be derived from the combined gas law.

14.3 Ideal Gases

63. Describe an ideal gas.

64. Explain why it is impossible for an ideal gas to exist.

★**65.** What is the volume occupied by 1.24 mol of a gas at 35°C if the pressure is 96.2 kPa?

66. What volume will 12.0 g of oxygen gas (O_2) occupy at 25°C and a pressure of 52.7 kPa?

★**67.** If 4.50 g of methane gas (CH_4) is in a 2.00-L container at 35°C, what is the pressure in the container?

68. What pressure is exerted by 0.450 mol of a gas at 25°C if the gas is in a 0.650-L container?

★**69.** A helium-filled weather balloon has a volume of 2.4×10^2 L at 99 kPa pressure and a temperature of 0°C. What is the mass of the helium in the balloon?

14.4 Gases: Mixtures and Movements

70. In your own words, state Dalton's law of partial pressure.

71. Which gas effuses faster: hydrogen or chlorine? How much faster?

72. Which gas effuses faster at the same temperature: molecular oxygen or atomic argon?

***73.** Calculate the ratio of the velocity of helium atoms to the velocity of neon atoms at the same temperature.

74. Calculate the ratio of the velocity of helium atoms to the velocity of fluorine molecules at the same temperature.

Understand Concepts

75. How does kinetic theory explain the compressibility of gases?

76. A teacher adds enough water to cover the bottom of an empty metal can with a screw cap. Using a stove, the teacher heats the can with the cap off until the water boils, and then screws on the cap tightly. When the sealed can is dunked in cold water, the sides of the can immediately collapse inward as though crushed in a trash compactor.

a. Use kinetic theory to explain why the can collapsed inward.
b. If the experiment were done with a dry can, would the results be similar? Explain.

77. Explain how the compressed gas in an aerosol can forces paint out of the can. Make sure to describe how the gas pressure inside the can changes as the paint is sprayed. (Refer to Figure 14.5 in Lesson 14.1.)

78. Why do aerosol containers display the warning, "Do not incinerate"?

79. The manufacturer of an aerosol deodorant packaged in a 150-mL container plans to produce a container of the same size that will hold twice as much gas. How will the pressure of the gas in the new product compare with that of the gas in the original container?

80. Why must Kelvin temperatures be used in calculations that involve gases?

81. Explain how using a pressure cooker reduces the time required to cook food.

82. The ratio of two variables is always a constant. What can you conclude about the relationship between the two variables?

***83.** A 3.50-L gas sample at 20°C and a pressure of 86.7 kPa expands to a volume of 8.00 L. The final pressure of the gas is 56.7 kPa. What is the final temperature of the gas in degrees Celsius?

84. Explain the reasons why real gases deviate from ideal behavior.

85. How would the number of particles of two gases compare if their partial pressures in a container were identical?

86. Why does a balloon filled with helium deflate more quickly than a balloon filled with air?

***87.** A certain gas effuses four times as fast as oxygen (O_2). What is the molar mass of the gas?

***88.** During an effusion experiment, a certain number of moles of an unknown gas passed through a tiny hole in 75 seconds. Under the same conditions, the same number of moles of oxygen gas passed through the hole in 30 seconds. What is the molar mass of the unknown gas?

89. The photograph shows a tube with cotton balls at each end. The cotton ball at the left was soaked with hydrochloric acid. The cotton ball on the right was soaked with a solution of ammonia. When these compounds react, they form a white solid, ammonium chloride. Based on the location of the ammonium chloride in the tube, which gas diffuses at a faster rate, hydrogen chloride or ammonia? Explain.

Think Critically

90. Infer Figure 14.14 in Lesson 14.3 shows an insulated tank used to store liquid nitrogen. How does the vacuum between the walls of the tank prevent heat transfer?

91. Infer Gases will diffuse from a region of higher concentration to a region of lower concentration. Why don't the gases in Earth's atmosphere escape into the near-vacuum of space?

92. Apply Concepts What real gas comes closest to having the characteristics of an ideal gas? Explain your answer.

93. Predict Death Valley in California is at 86 m below sea level. Will the partial pressure of oxygen in Death Valley be the same, lower, or higher than the partial pressure of oxygen at sea level? Give a reason for your answer.

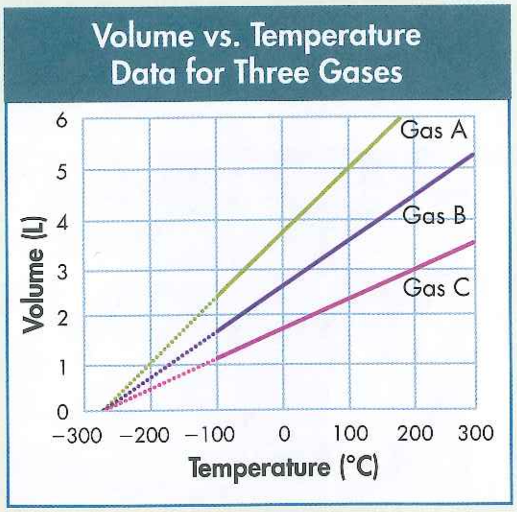
***94. Calculate** The following reaction takes place in a sealed 40.0-L container at a temperature of 120°C:

$$4NH_3(g) + 5O_2(g) \longrightarrow 4NO(g) + 6H_2O(g)$$

a. When 34.0 g of NH_3 reacts with 96.0 g of O_2, what is the partial pressure of NO in the sealed container?

b. What is the total pressure in the container?

95. Interpret Graphs The graph shows the direct relationship between volume and temperature for three different gas samples. Offer at least one explanation for why the graphs are not identical for the three samples. (*Hint:* What variables other than temperature and volume can be used to describe a gas?)



***96. Analyze Data** A student collected the following data for a fixed volume of gas.

Temperature (°C)	Pressure (mm Hg)
10	726
20	750
40	800
70	880
100	960

a. Graph the data, using pressure as the dependent variable.

b. What is the pressure of the gas at 0°C?

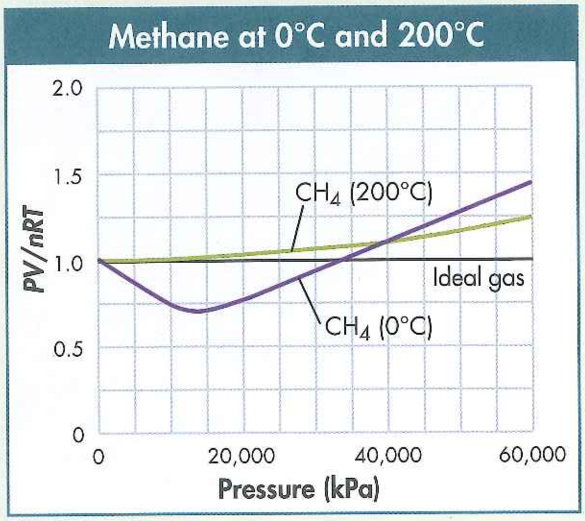
c. Is the relationship between the variables directly or inversely proportional?

d. How does the pressure of the gas change with each degree Celsius change in the temperature?

e. Write an equation relating the pressure and temperature of the gas.

f. Which gas law is illustrated by the data? Select two data points on your graph to confirm your answer.

97. Interpret Graphs The graph shows how the ratio (PV/nRT) changes with increasing pressure for methane (CH_4) at 0°C and 200°C.



a. At lower pressures, which gas behaves more like an ideal gas: methane at 0°C or methane at 200°C.

b. The curve for methane at 0°C shows that the ratio PV/nRT is less than 1 at lower pressures and greater than 1 at higher pressures. What characteristics of real gases can explain these deviations?

Enrichment

98. Analyze Data Oxygen is produced in the laboratory by heating potassium nitrate (KNO_3). The data table below gives the volume of oxygen produced at STP from different quantities of KNO_3. Use the data to determine the mole ratio by which KNO_3 and O_2 react.

Mass of KNO_3 (g)	Volume of O_2 (cL)
0.84	9.3
1.36	15.1
2.77	30.7
4.82	53.5
6.96	77.3

***99. Calculate** A mixture of ethyne gas (C_2H_2) and methane gas (CH_4) occupied a certain volume at a total pressure of 16.8 kPa. When the sample burned, the products were CO_2 gas and H_2O vapor. The CO_2 was collected and its pressure found to be 25.2 kPa in the same volume and at the same temperature as the original mixture. What percentage of the original mixture was methane?

***100. Calculate** A 0.10-L container holds 3.0×10^{20} molecules of H_2 at 100 kPa and 0°C.

a. If the volume of a hydrogen molecule is 6.7×10^{-24} mL, what percentage of the volume of the gas is occupied by its molecules?

b. If the pressure is increased to 100,000 kPa, the volume of the gas is 1×10^{-4} L. What fraction of the total volume do the hydrogen molecules now occupy?

101. Draw Conclusions Many gases that have small molecules, such as N_2 and O_2, have the expected molar volume of 22.41 L at STP. However, other gases behave in a very non-ideal manner, even if extreme pressures and temperatures are not involved. The molar volumes of CH_4, CO_2, and NH_3 at STP are 22.37 L, 22.26 L, and 22.06 L, respectively. Explain the reasons for these large departures from the ideal.

Write About Science

102. Explain Why does a tennis ball bounce higher in the summer than it does in the winter? Use what you know about gas behavior to explain your answer.

103. Research a Problem Cars that run on natural gas or hydrogen require different fuel tanks and different refueling stations than cars that run on gasoline, which is a liquid at STP. How would you design a fuel tank for storing a gas? How would you design a gas pump that pumped a gas instead of a liquid? Research a kind of vehicle that runs on gaseous fuel and explain how these questions have been addressed.

CHEMYSTERY

Under Pressure

Becki realized that she had decompression sickness, also known as the bends. Recreational divers use regulators attached on their air tank to "regulate" the air they breathe in so that it's at the same pressure as the pressure outside their bodies. Although the fractions of nitrogen and oxygen in her air supply remained constant under high pressure, the partial pressure of each component gas increased. Therefore, with each breath under water, she was receiving more nitrogen and oxygen than normal.

As Becki ascended and the pressure on her body decreased, the excess nitrogen formed bubbles in her blood and tissues, causing pain and other symptoms. Serious cases of the bends require treatment in a high-pressure chamber. The pressure is reduced gradually so that the excess nitrogen can leave the body harmlessly.

104. Infer How could Becki have prevented getting the bends?

105. Connect to the BIGIDEA What would have happened if Becki held her breath while ascending from a dive? Use the gas laws to explain.

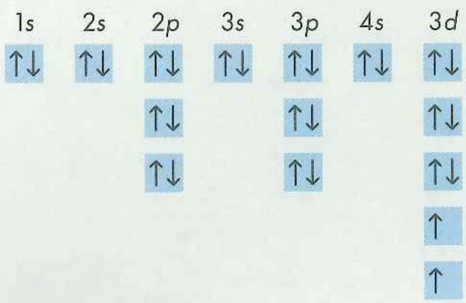
Cumulative Review

106. What is the mathematical relationship between the Kelvin and Celsius temperature scales?

*__107.__ A metal sample has a mass of 9.92 g and measures 4.50 cm × 1.30 cm × 1.60 mm. What is the density of the metal?

108. How many electrons, protons, and neutrons are there in an atom of lead-206?

109. Which element has the following electron configuration?



110. Which of these elements are metals?

a. arsenic **b.** tungsten **c.** xenon

111. Which element is most likely to form a compound with strontium?

a. neon **b.** tin **c.** selenium

112. Which compound contains at least one double bond?

a. H_2Se **b.** SO_2 **c.** PCl_3

*__113.__ Name each compound.

a. $SnBr_2$
b. $BaSO_4$
c. $Mg(OH)_2$
d. IF_5

114. An atom of lead has a mass 17.16 times greater than the mass of an atom of carbon-12. What is the molar mass of this isotope of lead?

*__115.__ Calculate the molar mass of each substance.

a. $Ca(CH_3CO_2)_2$
b. H_3PO_4
c. $C_{12}H_{22}O_{11}$
d. $Pb(NO_3)_2$

116. What is the significance of the volume 22.4 L?

*__117.__ Calculate the molecular formula of each of the following compounds.

a. The empirical formula is C_2H_4O, and the molar mass is 88 g.
b. The empirical formula is CH, and the molar mass is 104 g.
c. The molar mass is 90 g. The percent composition is 26.7% C, 71.1% O, and 2.2% H.

118. Calculate the percent composition of 2-propanol (C_3H_7OH).

119. What type of reaction is each of the following?

a. Calcium reacts with water to form calcium hydroxide and hydrogen gas.
b. Mercury and oxygen are prepared by heating mercury(II) oxide.

120. Write balanced equations for the following chemical reactions.

a. Tetraphosphorus decoxide reacts with water to form phosphoric acid.
b. Aluminum hydroxide and hydrogen sulfide form when aluminum sulfide reacts with water.

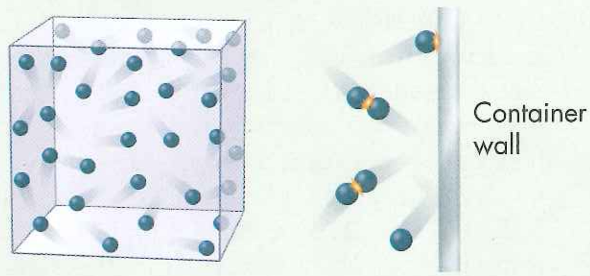
*__121.__ Aluminum oxide is formed from its elements.

$$Al(s) + O_2(g) \longrightarrow Al_2O_3(s)$$

a. Balance the equation.
b. How many grams of each reactant are needed to form 583 g $Al_2O_3(s)$?

122. Explain why a gas expands until it takes the shape and volume of its container.

123. Use the drawings to explain how gas pressure is produced.



If You Have Trouble With . . .

Question	106	107	108	109	110	111	112	113	114	115	116	117	118	119	120	121	122	123
See Chapter	3	3	4	5	6	7	8	9	10	10	10	10	10	11	11	12	13	13

Standardized Test Prep

Select the choice that best answers each question or completes each statement.

1. A gas in a balloon at constant pressure has a volume of 120.0 mL at −123°C. What is its volume at 27.0°C?
   (A) 60.0 mL
   (B) 240.0 mL
   (C) 26.5 mL
   (D) 546 mL

2. If the Kelvin temperature of a gas is tripled and the volume is doubled, the new pressure will be
   (A) 1/6 the original pressure.
   (B) 2/3 the original pressure.
   (C) 3/2 the original pressure.
   (D) 5 times the original pressure.

3. Which of these gases effuses fastest?
   (A) Cl_2
   (B) NO_2
   (C) NH_3
   (D) N_2

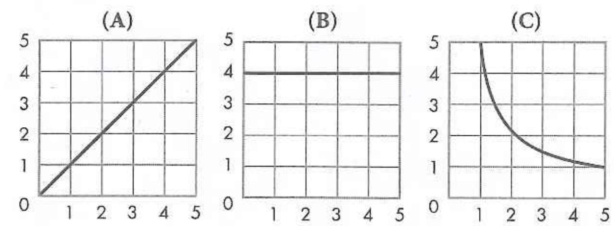
4. All the oxygen gas from a 10.0-L container at a pressure of 202 kPa is added to a 20.0-L container of hydrogen at a pressure of 505 kPa. After the transfer, what are the partial pressures of oxygen and hydrogen?
   (A) Oxygen is 101 kPa; hydrogen is 505 kPa.
   (B) Oxygen is 202 kPa; hydrogen is 505 kPa.
   (C) Oxygen is 101 kPa; hydrogen is 253 kPa.
   (D) Oxygen is 202 kPa; hydrogen is 253 kPa.

5. Which of the following changes would increase the pressure of a gas in a closed container?
   I. Part of the gas is removed.
   II. The container size is decreased.
   III. Temperature is increased.

   (A) I and II only
   (B) II and III only
   (C) I and III only
   (D) I, II, and III

6. A real gas behaves most nearly like an ideal gas
   (A) at high pressure and low temperature.
   (B) at low pressure and high temperature.
   (C) at low pressure and low temperature.
   (D) at high pressure and high temperature.

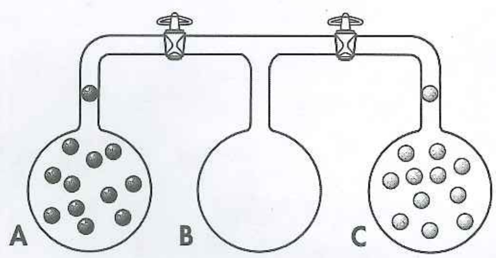
Use the graphs to answer Questions 7–10. A graph may be used once, more than once, or not at all.



Which graph shows each of the following?

7. directly proportional relationship
8. graph with slope = 0
9. inversely proportional relationship
10. graph with a constant slope

Use the drawing to answer Questions 11 and 12.



11. Bulb A and bulb C contain different gases. Bulb B contains no gas. If the valves between the bulbs are opened, how will the particles of gas be distributed when the system reaches equilibrium? Assume none of the particles are in the tubes that connect the bulbs.

Tips for Success

Constructing a Diagram If you are asked to draw a diagram, sketch lightly at first (so you can erase easily), or do a sketch on a separate piece of paper. Once you are sure of your answer, draw the final diagram.

12. Make a three-bulb drawing with 6 blue spheres in bulb A, 9 green spheres in bulb B, and 12 red spheres in bulb C. Then draw the setup to represent the distribution of gases after the valves are opened and the system reaches equilibrium.

If You Have Trouble With . . .

Question	1	2	3	4	5	6	7	8	9	10	11	12
See Lesson	14.2	14.2	14.4	14.3	14.1	14.3	14.2	14.2	14.2	14.2	14.4	14.4