Chemistry

Corbin

Week 3 & 4

SECTION 11.1

Gases and Pressure

Name: ______________________

Period: ______________________

Teacher: ______________________

In the chapter "States of Matter," you read about the kinetic-molecular theory of matter. You were also introduced to how this theory explains some of the properties of ideal gases. In this chapter, you will study the predictions of kinetic-molecular theory for gases in more detail. This includes the relationship among the temperature, pressure, volume, and amount of gas in a sample.

KEY TERMS

- pressure
- newton
- barometer
- millimeters of mercury
- atmosphere of pressure
- pascal
- partial pressure
- Dalton's law of partial pressures

Collisions of air molecules generate pressure.

You may be familiar with the concept of tire pressure. When you pump more air into a tire, the number of molecules of air inside the tire increases. This causes an increase in the number of collisions between air molecules and the sides of the tire. An increase in pressure is the result of the increase in collisions.

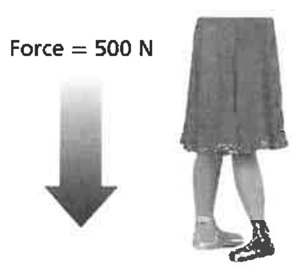
Pressure depends on force and area.

Pressure is the force per unit area on a surface. The SI unit for force is the newton, N. The pressure exerted by a particular force is given by this equation.

$$\text{pressure} = \frac{\text{force}}{\text{area}}$$

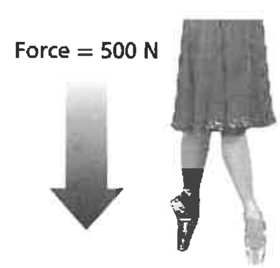
READING CHECK

1. Compute the pressure exerted by the dancer for photos (b) and (c) below.



a. Area of contact = 325 cm^2

$$\text{Pressure} = \frac{\text{force}}{\text{area}} = \frac{500\ \text{N}}{325\ \text{cm}^2} = 1.5\ \text{N/cm}^2$$

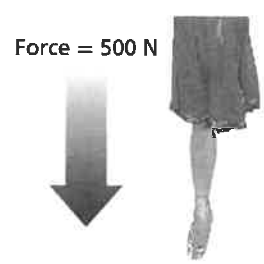


b. Area of contact = 13 cm^2

$$\text{Pressure} = \frac{\text{force}}{\text{area}}$$

=

=



c. Area of contact = 6.5 cm^2

$$\text{Pressure} = \frac{\text{force}}{\text{area}}$$

=

=

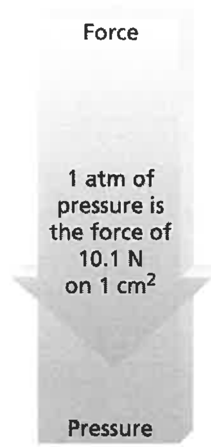
The pressure a dancer exerts against the floor depends on the area of contact. The smaller the area, the greater the pressure.

Pressure One way to think of force is to consider it the result of a mass times an acceleration. A **newton** is the force that will increase the speed of a one-kilogram mass by one meter per second each second that the force is applied.

Consider the ballet dancer on the previous page. Earth exerts a gravitational force on all objects on its surface that accelerates them toward Earth at 9.8 m/s^2. The ballet dancer has a mass of 51 kg. Therefore, the force that Earth exerts on her is given by the following equation.

$$51 \text{ kg} \times 9.8 \text{ m/s}^2 = 500 \text{ N}$$

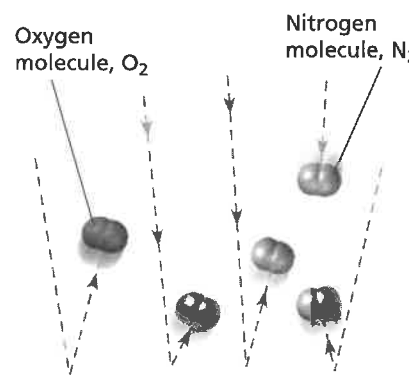
As a result of gravity, the dancer is pulled against the floor with a force of 500 N. The pressure that the dancer exerts on the floor depends on the surface area of the dancer touching the floor. If the dancer is standing on one toe, she exerts more pressure on that part of the floor than if she is standing on two flat feet. The same force applied to a smaller area results in greater pressure.



Atmospheric Pressure The atmosphere, which is the air that surrounds Earth, exerts pressure on Earth's surface. This pressure is equivalent to a 1.03 kg mass sitting on every square centimeter of Earth. The resulting pressure is 10.1 N/cm^2. This quantity is also called 1 atmosphere, or 1 atm.

Atmospheric pressure is the result of all of the different gases in the air striking Earth's surface. The atmosphere contains 78% nitrogen gas, 21% oxygen gas, and 1% other gases by volume.

Atmospheric pressure decreases higher up in the atmosphere. This is because there are fewer air particles farther away from Earth's surface. When you fly in an airplane, the decrease in pressure outside your ears relative to the pressure inside your ears makes your ears "pop."



Air molecules, mostly nitrogen and oxygen, collide with Earth's surface, exerting a pressure of 10.1 N/cm^2.

READING CHECK

2. What element is responsible for most of the pressure exerted by the atmosphere on the surface of Earth?

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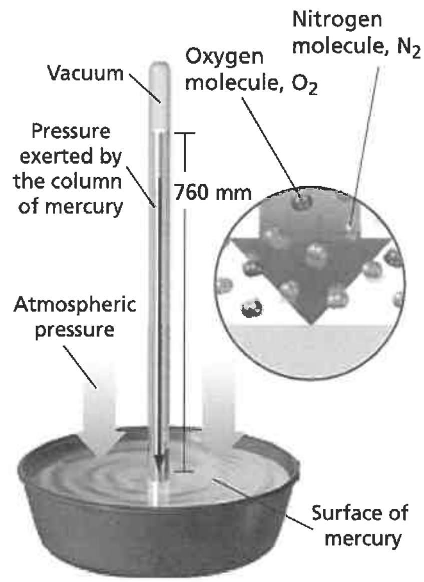
Measuring Pressure

A **barometer** is a device used to measure atmospheric pressure. The first type of barometer was introduced by Evangelista Torricelli in the early 1600s. Torricelli wondered why water pumps could only raise water to a height of 34 ft. He hypothesized that the height of the water depends on the weight of the water compared with the weight of air. He predicted that mercury, which is 14 times as dense as water, would only rise 1/14 as high as water.

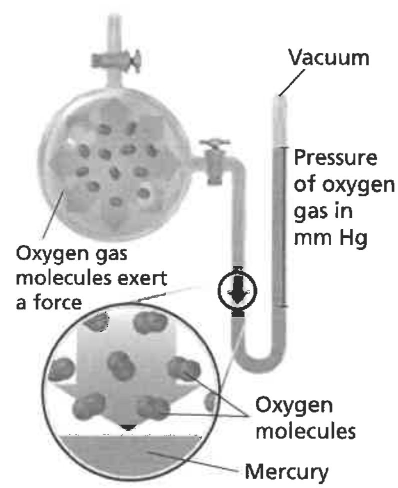
To find out if he was right, Torricelli held a tube upside down over a dish of liquid mercury. The top of the tube was a vacuum, so only the weight of the liquid mercury pressed down on the mercury in the dish. This column of fluid was prevented from falling because air pushed down on the liquid in the dish, preventing it from rising due to fluid leaving the tube. Torricelli found that the pressure of the air was equivalent to the pressure of about a 760 mm column of mercury. And 760 mm was about 1/14 the height of 34 feet (10.4 m).



This barometer measures the pressure exerted by the gas in the atmosphere in millimeters of mercury. Atmospheric pressure can support the weight of a column of mercury that is about 760 mm high.

The atmospheric pressure at any place on Earth depends on elevation and weather conditions. If the pressure rises, then the column of mercury in Torricelli's barometer becomes taller. If the pressure drops, then the column becomes shorter. So, air pressure can be measured by the height of a column of mercury in a barometer.

A manometer uses a similar method to measure the pressure of a specific gas sample. The gas is enclosed but is able to exert pressure on a column of mercury in a U-shaped tube. The difference in height of the mercury columns in the two arms of the "U" is a measure of the pressure exerted by the gas.



This manometer is measuring the pressure in a sample of oxygen gas. The pressure is indicated by the difference in height of the mercury in the two arms of the U-shaped tube.

Critical Thinking

3. **Apply** Suppose the air is completely removed from inside an empty plastic water bottle. What happens to the water bottle? Explain your answer.

Units of Pressure

There are several units for measuring pressure. The usefulness of the mercury barometer has led to the use of the height of a column of mercury as a measure of pressure. The common unit of pressure is **millimeters of mercury,** or mm Hg. A pressure of 1 mm Hg is also called 1 torr in honor of Torricelli's invention of the barometer.

Another unit of pressure is the atmosphere. One **atmosphere of pressure,** atm, is the equivalent of 760 mm Hg. The average atmospheric pressure at sea level at 0°C is 1 atm.

The SI unit of pressure is the pascal. A pascal is a derived unit named after the French mathematician and philosopher Blaise Pascal. One **pascal,** Pa, is equal to the pressure exerted by a force of 1 N acting on an area of 1 m^2. In many situations, it is more convenient to use the unit kilopascal, kPa. For example, one atmosphere of pressure, 1 atm, is equal to 101.325 kPa. The units of pressure used in this book are summarized in the table.

Remember

A derived unit is a unit that is a combination of two or more base units in the SI system. The pascal is a derived unit equal to one kilogram per meter per second squared.

Standard Temperature and Pressure

When comparing the volumes of two different gases, the temperature and pressure at which the volumes were measured must be specified. For purposes of comparison, scientists have agreed on standard conditions for comparing gases. The term *standard temperature and pressure* refers to a pressure of 1 atm and a temperature of 0°C. These conditions are also referred to as STP.

READING CHECK

4. What does each letter in the acronym STP stand for?

S________________________

T________________________

P________________________

Units of Pressure

Unit	Symbol	Definition/relationship
pascal	Pa	SI pressure unit, $1\ Pa = \frac{1\ N}{m^2}$
millimeter of mercury	mm Hg	pressure that supports a 1 mm mercury column in a barometer
torr	torr	1 torr = 1 mm Hg
atmosphere	atm	average atmospheric pressure at sea level and 0°C 1 atm = 760 mm Hg = 760 torr = $1.013\ 25 \times 10^5$ Pa = 101.325 kPa
pounds per square inch	psi	1 psi = $6.892\ 86 \times 10^3$ Pa 1 atm = 14.700 psi

The total pressure of a gas mixture is the sum of the pressures of the gases in it.

The pressure exerted by each gas in a mixture is called the **partial pressure** of that gas. John Dalton, the English chemist who proposed the atomic theory, also studied gas mixtures. He proposed that the pressure exerted by each gas in an unreactive mixture is independent of the pressures exerted by the other gases in the mixture. In other words, if a sample of oxygen gas exerts a pressure of 0.28 atm when it is isolated, then the molecules of that sample will exert a partial pressure of 0.28 atm when mixed with one or more gases.

Dalton's law of partial pressures states that the total pressure exerted by a gas mixture is the sum of the partial pressures of the component gases. Dalton's law may be expressed as an equation

$$P_T = P_1 + P_2 + P_3 + \ldots$$

where P_T is the total pressure of the mixture, P_1 is the partial pressure of the first gas, P_2 is the partial pressure of the second gas, and so on.

The kinetic-molecular theory of matter can explain Dalton's law. Each particle in a mixture of gases has an equal chance of colliding with the walls of a container. Therefore, each collection of gas molecules exerts a pressure independent of that exerted by the other gases. The total pressure is the result of the total number of collisions on a unit of wall area in a given time.

LOOKING CLOSER

5. Define each of these terms separately in your own words.

partial: ______________________

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pressure: ______________________

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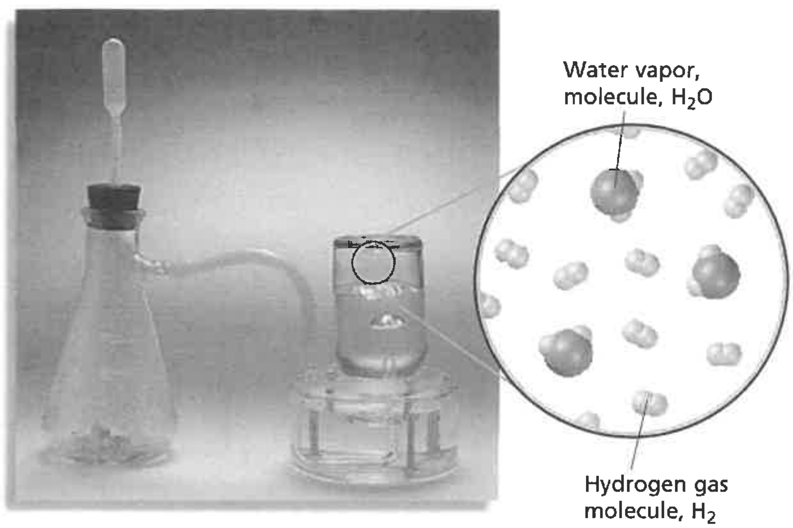
PRACTICE

A. Convert a pressure of 1.75 atm to kPa and mm Hg.

B. What is the partial pressure of O_2 in a mixture of CO_2, N_2, and O_2 at 1 atm if $P_{CO_2} = 0.285$ torr and $P_{N_2} = 593.525$ torr?



Hydrogen is collected by water displacement during a reaction of zinc with sulfuric acid

Gases Collected by Water Displacement

Often it is difficult to determine the amount of gas given off by a chemical reaction. The low density of the gas makes it hard to measure the mass that a gas adds to a container. Scientists have developed a method of collecting gases by water displacement to address this problem.

In this method, a collection bottle is partially filled with water and placed upside down over a reservoir of water. The gas released by a reaction passes into the collection bottle through a tube. When the gas enters the bottle, it displaces water from the bottle. The pressure of the gas forces the water down toward the reservoir.

After the gas is collected, the collection bottle does not contain a pure sample of the gas from the chemical reaction. The gas in the bottle is a mixture of the gas from the reaction and water vapor that is in a state of equilibrium with the water in the bottle. During the collection, the water level adjusts so that the total pressure on the water surface inside the water bottle equals the atmospheric pressure, P_{atm}. Therefore, the partial pressures of the two gases, the gas from the reaction and water vapor, must satisfy this equation:

$$P_{atm} = P_{gas} + P_{H_2O}$$

The value of P_{atm} can be read from a barometer in the laboratory. The value of P_{H_2O} can be found in the table on the next page, using the recorded temperature of the experiment. The value of P_{gas} can be determined using these two values.

READING CHECK

6. What law states that the sum of the partial pressures of two gases in a mixture is the total pressure of the mixture?

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Water-Vapor Pressure

Temperature (°C)	Pressure (mm Hg)	Pressure (kPa)	Temperature (°C)	Pressure (mm Hg)	Pressure (kPa)
0.0	4.6	0.61	23.0	21.1	2.81
10.0	9.2	1.23	23.5	21.7	2.90
15.0	12.8	1.71	24.0	22.4	2.98
16.0	13.6	1.82	25.0	23.8	3.17
16.5	14.1	1.88	26.0	25.2	3.36
17.0	14.5	1.94	27.0	26.7	3.57
17.5	15.0	2.00	28.0	28.3	3.78
18.0	15.5	2.06	29.0	30.0	4.01
18.5	16.0	2.13	30.0	31.8	4.25
19.0	16.5	2.19	35.0	42.2	5.63
19.5	17.0	2.27	40.0	55.3	7.38
20.0	17.5	2.34	50.0	92.5	12.34
20.5	18.1	2.41	60.0	149.4	19.93
21.0	18.6	2.49	70.0	233.7	31.18
21.5	19.2	2.57	80.0	355.1	47.37
22.0	19.8	2.64	90.0	525.8	70.12
22.5	20.4	2.72	100.0	760.0	101.32

For example, suppose oxygen gas is collected by water displacement from a decomposition reaction of potassium chlorate, $KClO_3$. If the temperature during the experiment is 20.0°C, then the partial pressure of water vapor atmospheric pressure is 17.5 torr. If the atmospheric pressure during the experiment is 731.0 torr, then the partial pressure of the oxygen gas can be determined as follows.

$$P_{O_2} = P_{atm} - P_{H_2O} = 731.0 \text{ torr} - 17.5 \text{ torr} = 713.5 \text{ torr}$$

PRACTICE

c. Some hydrogen gas is collected over water at 20.0°C. The levels of water inside and outside the gas-collection bottle are the same. The partial pressure of the hydrogen is 742.5 torr. What is the barometric pressure at the time the gas is collected?

P_{atm} = __________ + P_{H_2O} = __________ torr + __________ torr = __________ torr

SECTION 11.1 REVIEW

VOCABULARY

1. Define *pressure*.

REVIEW

2. Name at least four different units that are used to express measurements of pressure.

3. Convert the following pressures to pressures in standard atmospheres.

   a. 151.98 kPa

   b. 456 torr

4. A sample of nitrogen gas is collected over water at a temperature of 23.0°C. What is the pressure of the nitrogen gas if atmospheric pressure is 785 mm Hg?

Critical Thinking

5. **EVALUATING METHODS** Clean rooms used for sterile biological research are sealed and operate at slightly above atmospheric pressure. Explain why.

6. **INFERRING RELATIONSHIPS** Explain why helium-filled balloons deflate over time faster than air-filled balloons do.

SECTION 11.2

The Gas Laws

Name: ____________________________________

Period: ________

Teacher: ________________________

The *gas laws* are simple mathematical relationships among the volume, temperature, pressure, and amount of a gas. They are the result of hundreds of years of study on the physical properties of gases. The scientists who discovered them made careful observations of the relationship between two or more variables in samples of various gases. The gas laws can all be explained using the kinetic-molecular theory of matter.

KEY TERMS

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

In this section, four gas laws will be covered. These gas laws are listed in the table at the right. These four laws are true for any choice of units for pressure and volume. The units of the constant k depend on the units of pressure and volume. In any comparison between a sample of a gas at two different times, the same units of pressure and volume should be used.

Gas law	Equation	Variables held constant
Boyle's law	$PV = k$	temperature, amount of gas
Charles's law	$\frac{V}{k} = k$	pressure, amount of gas
Gay-Lussac's law	$\frac{P}{T} = k$	volume, amount of gas
combined gas law	$\frac{PV}{T} = k$	amount of gas

The gas laws are only true if temperature is measured in kelvins. The Kelvin temperature scale starts at absolute zero and expresses temperatures in kelvins, K. **Absolute zero** is the temperature –273.15°C. The relationship between the Celsius scale and the Kelvin scale is given by

$$K = 273.15 + °C$$

So, the temperature –273.15°C is equivalent to 0 K.

TIP For calculations in this book, the value of absolute zero is rounded off to –273°C. The temperature 0°C is then considered equivalent to 273 K.

PRACTICE

Convert the following temperatures to either the Celsius temperature scale or the Kelvin temperature scale.

A. 20°C ________________ **C.** –10°C ________________ **E.** 273°C ________________

B. 300 K ________________ **D.** 100 K ________________ **F.** 297 K ________________

Gas volume and pressure are indirectly proportional.

The first relationship that was discovered between quantities in a sample of gas was between the pressure and the volume. In 1662, Robert Boyle discovered that doubling the pressure on a gas at constant temperature reduced its volume by one-half. Reducing the pressure on a gas by one-half allowed the volume of the gas to double.

The kinetic-molecular theory explains why this pressure-volume relationship exists. Pressure is caused by gas particles hitting the walls of a container. If the volume of the container is decreased, the particles will strike the container walls more frequently. If the volume of the container is increased, the particles will spread out and fewer collisions will occur.

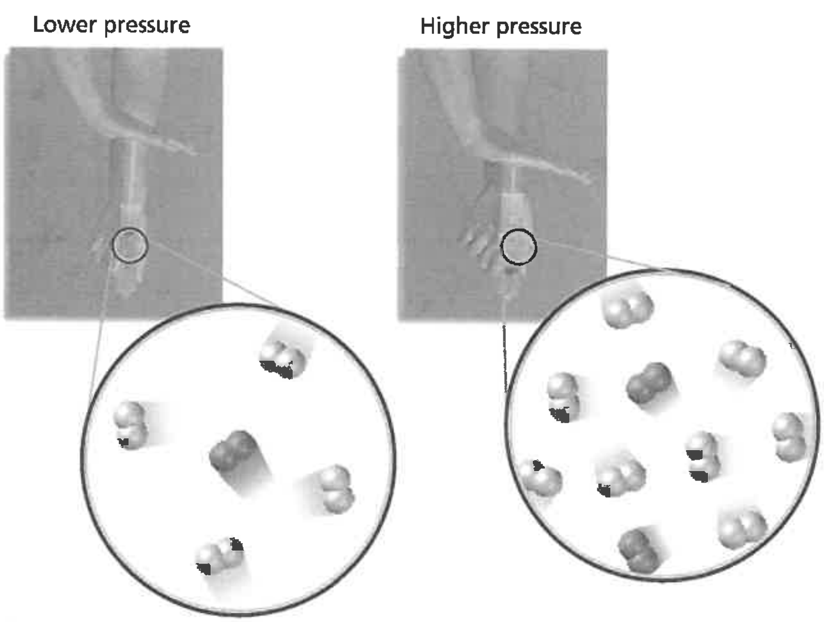
Boyle's law states that the volume of a fixed mass of gas varies inversely with the pressure at constant temperature. Mathematically, Boyle's law is expressed as

$$PV = k$$

where P is pressure, V is volume, and k is a constant.

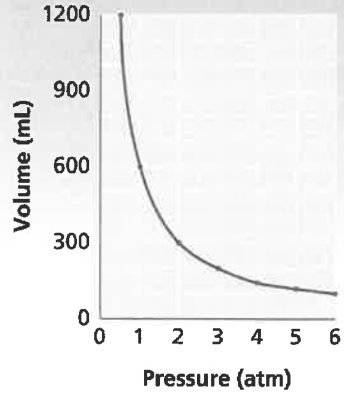
Another way to express Boyle's law is to consider two sets of measurements at different times on the same sample of a gas. If P_1 and V_1 are the first set of measurements of pressure and volume, and P_2 and V_2 are the second set, then

$$P_1V_1 = P_2V_2$$



When the plunger is pressed in, the molecules collide with the walls of the plunger more frequently, raising the pressure.

Volume vs. Pressure for a Gas at Constant Temperature



This graph shows the inverse relationship between volume and pressure for a gas sample at a constant temperature.

READING CHECK

1. If a gas is pumped from a smaller container to a container that is twice the size, what happens to the pressure of the gas?

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____

SAMPLE PROBLEM

A sample of oxygen gas has a volume of 150.0 mL when its pressure is 0.947 atm. What will the volume of the gas be at a pressure of 0.987 atm if the temperature remains constant?

SOLUTION

1 ANALYZE *Determine what information is given and unknown.*

Given: $V_1 = 150.0$ mL
$P_1 = 0.947$ atm
$P_2 = 0.987$ atm

Unknown: V_2

2 PLAN *Write the equation that can be used to find the unknown.*

Because the given data are pressure and volume, and the unknown is volume, use Boyle's law.

$$P_1V_1 = P_2V_2$$

Rearranging the equation to isolate the unknown value:

$$V_2 = \frac{P_1V_1}{P_2}$$

3 SOLVE *Substitute the given information and find the unknown value.*

$$V_2 = \frac{(0.947\,\cancel{\text{atm}})(150.0\text{ mL }O_2)}{0.987\,\cancel{\text{atm}}} = 144\text{ mL }O_2$$

4 CHECK YOUR WORK *Check to see if the answer makes sense.*

The pressure increased slightly at constant temperature. Therefore the volume should decrease slightly.

PRACTICE

G. A balloon filled with helium gas has a volume of 500 mL at a pressure of 1 atm. The balloon is released and reaches an altitude of 6.5 km, where the pressure is 0.5 atm. If the temperature has remained the same, what volume does the gas occupy at 6.5 km?

$$V_2 = \frac{P_1V_1}{P_2} = \rule{6cm}{0.4pt}$$

$$= \rule{5cm}{0.4pt}$$

Gas volume and temperature are directly related.

A balloonist makes use of the expansion of a gas when it is heated at a constant pressure. When the temperature of a gas increases, the average kinetic energy of its particles increases. Because the particles are moving faster, they collide with the walls of a container more frequently. For the number of collisions, and therefore the pressure, to remain constant, the volume of the container must increase.

French scientist Jacques Charles discovered the relationship between volume and temperature in 1787. He found that each time a gas was heated from 0°C to 1°C, its volume increased by a factor of 1/273. If the temperature was increased from 0°C to 273°C, then the volume doubled.

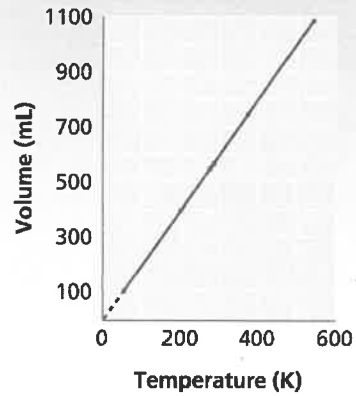
Charles's law states that the volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature. Mathematically, Charles's law is expressed as

$$V = kT \qquad \text{or} \qquad \frac{V}{T} = k$$

where V is volume, T is temperature in kelvins, and k is a constant. Just as for Boyle's law, Charles's law can be used to consider two sets of measurements at different times on the same sample of a gas. In the equation below, V_1 and T_1 are the first set of measurements, and V_2 and T_2 are the second set.

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

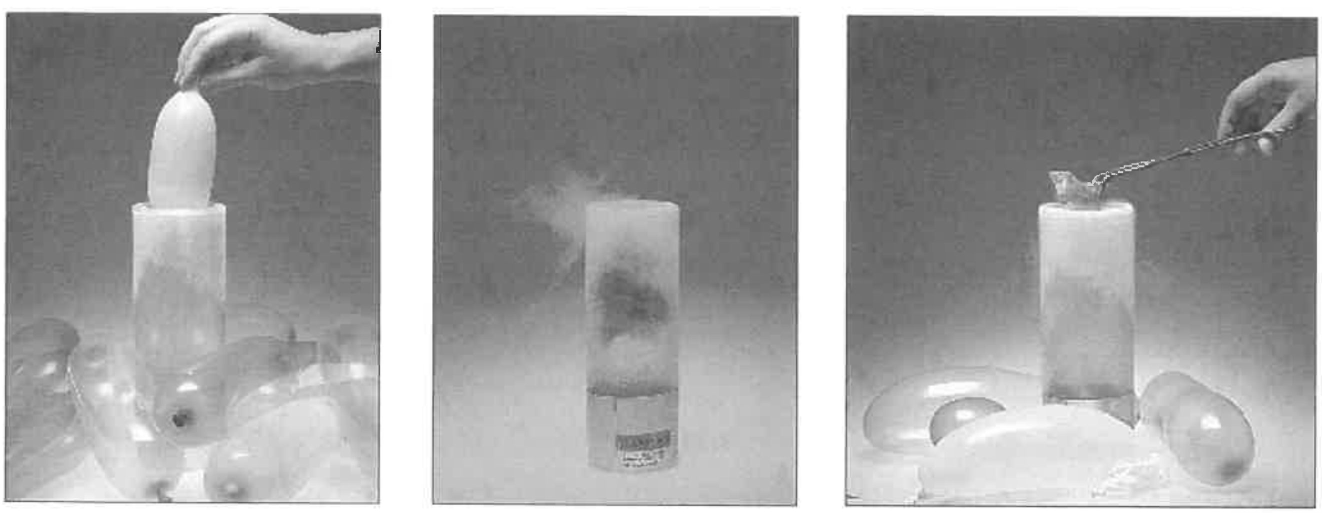
Volume Vs. Temperature for a Gas at Constant Pressure



This graph shows the linear relationship between the volume and the temperature in kelvins for a gas sample at constant pressure.

READING CHECK

2. If a gas is pumped from a smaller container to a container that is twice the size, and its pressure is kept the same, then what happens to the temperature of the gas?



As balloons filled with air are placed in a beaker of liquid nitrogen, the extremely cold temperature inside the beaker causes them to shrink. When the balloons are removed from the beaker, they expand to their original volume.

SAMPLE PROBLEM

A sample of neon gas occupies a volume of 752 mL at 25°C. What volume will the gas occupy at 50°C if the pressure remains constant?

SOLUTION

1 ANALYZE *Determine what information is given and unknown.*

Given: $V_1 = 752$ mL
$T_1 = 25°C$
$T_2 = 50°C$

Unknown: V_2

2 PLAN *Write the equation that can be used to find the unknown.*

Because the given data are volume and temperature, use Charles's law and solve for the value of V_2.

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

3 SOLVE *Substitute the given information and find the unknown value.*

First, convert the given temperatures from degrees Celsius to Kelvin: $T_1 = 25°C + 273 = 298$ K; $T_2 = 50°C + 273 = 323$ K.

$$V_2 = \frac{(323\,\cancel{K})(752 \text{ mL Ne})}{298\,\cancel{K}} = 815 \text{ mL Ne}$$

4 CHECK YOUR WORK *Check to see if the answer makes sense.*

The temperature doubled in the Celsius scale, but not in kelvins. The temperature increased slightly at constant pressure, therefore the volume should increase slightly.

PRACTICE

H. A sample of nitrogen gas is contained in a piston with a freely moving cylinder. At 0.0°C, the volume of the gas is 375 mL. To what temperature must the gas be heated to occupy a volume of 500.0 mL?

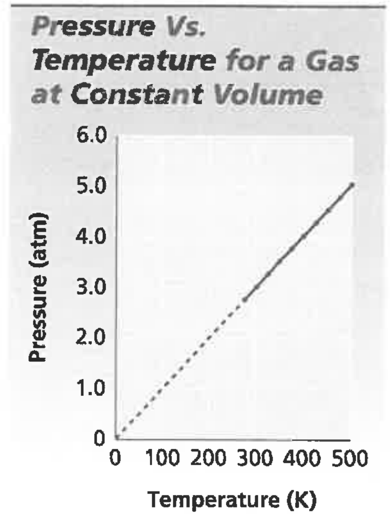
$$T_2 = \frac{(\quad)(\quad)}{(\quad)} = \underline{\qquad\qquad\qquad\qquad}$$

$$= \underline{\qquad\qquad\qquad}$$

Gas pressure and temperature are directly related.

When the temperature of a gas is increased, Charles's law explains that the gas must expand for the pressure to remain constant. However, suppose that the volume were held constant when the temperature was increased. When the temperature of a gas in a rigid container increases, the average kinetic energy of its particles increases. Since the volume of the container is fixed, the increased speed of the particles leads to more collisions with the walls of the container. As a result, the gas exerts a greater pressure on the walls of the container.

French scientist Joseph Gay-Lussac is given credit for discovering the relationship between the pressure and temperature of a gas in 1802. His results were similar to the results of Charles's experiments. Each time a gas was heated from 0°C to 1°C, the pressure of the gas increased by a factor of 1/273. And an increase from 0°C to 273°C led to a doubling of the pressure.



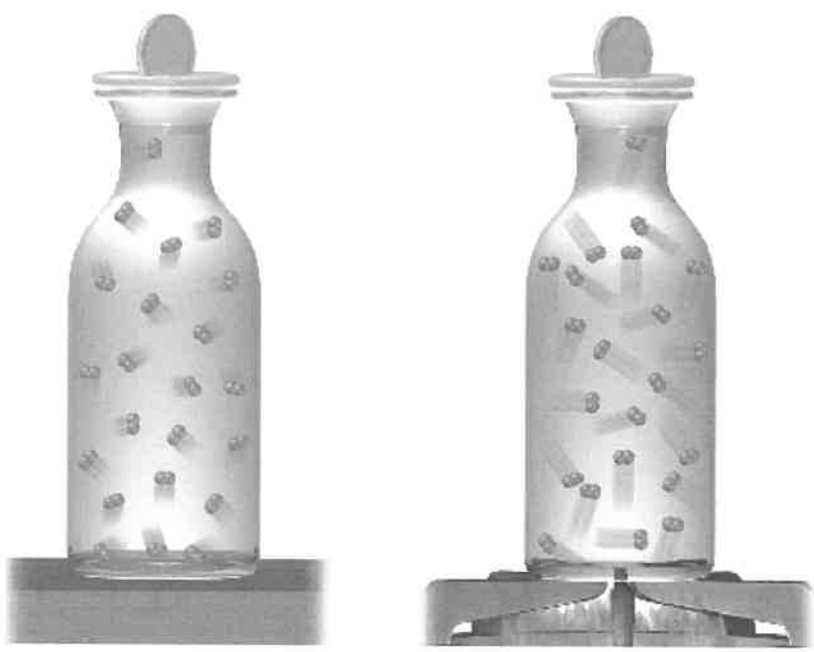
This graph shows the linear relationship between the pressure and the temperature in kelvins for a gas sample at constant volume.

Gay-Lussac's law states that the pressure of a fixed mass of gas at constant volume varies directly with the Kelvin temperature. Gay-Lussac's law can be expressed as

$$P = kT \quad \text{or} \quad \frac{P}{T} = k$$

where P is pressure, T is temperature in kelvins, and k is a constant. If P_1 and T_1 are a set of measurements on a sample of gas, and P_2 and T_2 are a second set taken at a later time, then the following equation is true.

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



As the temperature of a gas inside a sealed container increases, the pressure of the gas increases.

SAMPLE PROBLEM

The gas in a container is at a pressure of 3.00 atm at 25°C. Directions on the container warn the user not to keep it in a place where the temperature exceeds 52°C. What would the gas pressure in the container be at 52°C?

SOLUTION

1 ANALYZE

Determine what information is given and unknown.

Given: $P_1 = 3.00$ atm
$T_1 = 25°C$
$T_2 = 52°C$

Unknown: P_2

2 PLAN

Write the equation that can be used to find the unknown.

Because the given data are pressure and temperature, use Gay-Lussac's law and solve for the value of P_2.

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

3 SOLVE

Substitute the given information and find the unknown value.

First, convert the given temperatures from degrees Celsius to Kelvin.

$$T_1 = 25°C + 273 = 298\text{ K}$$

$$T_2 = 52°C + 273 = 325\text{ K}$$

$$P_2 = \frac{(325\,\cancel{\text{K}})(3.00\text{ atm})}{298\,\cancel{\text{K}}} = 3.27\text{ atm}$$

4 CHECK YOUR WORK

Check to see if the answer makes sense.

The temperature increased slightly at constant volume. Therefore the pressure should increase slightly.

PRACTICE

1. A sample of helium gas has a pressure of 1.20 atm at 22°C. At what Celsius temperature will the helium reach a pressure of 2.00 atm, assuming constant volume?

$$T_2 = \frac{(\quad)(\quad)}{(\quad)} = \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_$$

$$= \_\_\_\_\_\_\_\_\_\_\_\_$$

Gas pressure, temperature, and volume are interrelated.

The preceding pages of this section discussed the relationships between two quantities that describe a gas sample. However, in many cases, the pressure, volume, and temperature of a gas can all change between different sets of measurements.

The **combined gas law** expresses the relationship among pressure, volume, and temperature of a fixed amount of gas. The combined gas law can be expressed as follows.

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

In this equation, the constant k depends on the amount of gas in the sample. The combined gas law can also be written as the following equation for two sets of measurements on the same sample of gas.

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

In this equation, P_1, V_1, and T_1 are the pressure, volume and temperature of a gas at one time, and P_2, V_2, and T_2 are the pressure, volume, and temperature of the gas at a later time.

The combined gas law is a combination of Boyle's law, Charles's law, and Gay-Lussac's law. If one of the quantities (pressure, volume, or temperature) is constant, then the combined gas law simplifies to one of the original three laws.

For example, if $T_1 = T_2$, then both sides of the equation above can be multiplied by T_1 to cancel the temperature variables.

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

$$P_1V_1 = P_2V_2$$

The result is Boyle's law. Because Boyle's law describes the relationship when temperature is a constant, the derivation makes sense.

Remember

Two quantities are *inversely proportional* if their product is a constant. Two quantities are *directly proportional* if their ratio is a constant.

Critical Thinking

3. **Identify** Which of the three gas laws discussed so far describes a relationship that is directly proportional?

SAMPLE PROBLEM

A helium-filled balloon has a volume of 50.0 L at 25°C and 1.08 atm. What volume will it have at 0.855 atm and 10.0°C?

SOLUTION

1 ANALYZE *Determine what information is given and unknown.*

Given: $P_1 = 1.08$ atm, $T_1 = 25°C$, $V_1 = 50.0$ L

$P_2 = 0.855$ atm, $T_2 = 10°C$

Unknown: V_2

2 PLAN *Write the equation that can be used to find the unknown.*

Use the combined gas law because the pressure, volume, and temperature are all changing. Rearrange to solve for V_2.

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

3 SOLVE *Substitute the given information and find the unknown value.*

$$T_1 = 25°C + 273 = 298 \text{ K}$$

$$T_2 = 10°C + 273 = 283 \text{ K}$$

$$V_2 = \frac{(1.08 \cancel{\text{atm}})(50.0 \text{ L He})(283 \cancel{\text{K}})}{(0.855 \cancel{\text{atm}})(298 \cancel{\text{K}})} = 60.0 \text{ L He}$$

4 CHECK YOUR WORK *Check to see if the answer makes sense.*

The temperature decreased slightly, which should have decreased the volume slightly. The pressure decreased by a larger factor, which should have increased the volume by a larger factor. The answer reflects the net effect of a slight increase in volume.

PRACTICE

J. A 700.00 mL gas sample at STP is compressed to a volume of 200 mL, and the temperature is increased to 30.0°C. What is the new pressure of the gas in Pa?

$$P_2 = \frac{(\quad)(\quad)(\quad)}{(\quad)} = \underline{\qquad\qquad\qquad\qquad}$$

$$= \underline{\qquad\qquad\qquad}$$

SECTION 11.2 REVIEW

VOCABULARY

1. Explain Charles's law in terms of the kinetic-molecular theory.

REVIEW

2. Relate the effect of temperature and pressure on a gas to the model of a gas given by the kinetic-molecular theory.

3. A sample of helium gas has a volume of 200.0 mL at 0.960 atm. What pressure, in atmospheres, is needed to reduce the volume at constant temperature to 50.0 mL?

4. A gas occupies 2.0 m^3 at 100.0 K and exerts a pressure of 100.0 kPa. What volume will the gas occupy if the temperature is increased to 400.0 K and the pressure is increased to 200.0 kPa?

Critical Thinking

5. ANALYZING RESULTS A student has the following data: $V_1 = 822$ mL, $T_1 = 75°C$, and $T_2 = -25°C$. He calculates V_2 and gets –274 mL. Is this value correct? Explain why or why not.