Preview

- Lesson Starter
- Objectives
- Pressure and Force
- Dalton’s Law of Partial Pressures

Lesson Starter

- Make a list of gases you already know about.
- Separate your list into elements, compounds, and mixtures.
- Share your list with the class by writing it on the board.

Objectives

- **Define** *pressure,* give units of pressure, and describe how pressure is measured.
- **State** the standard conditions of temperature and pressure and convert units of pressure.
- **Use** Dalton's law of partial pressures to calculate partial pressures and total pressures.

Pressure and Force

- **Pressure** (P) is defined as the force per unit area on a surface.

- Gas pressure is caused by collisions of the gas molecules with each other and with surfaces with which they come into contact.

- The pressure exerted by a gas depends on volume, temperature, and the number of molecules present.
    - The greater the number of collisions of gas molecules, the higher the pressure will be.

Pressure

Click below to watch the Visual Concept.

Equation for Pressure

Click below to watch the Visual Concept.

Visual Concept

Pressure and Force

- The SI unit for force is the **newton,** (N), the force that will increase the speed of a one-kilogram mass by one meter per second each second that the force is applied.
    - example: consider a person with a mass of 51 kg. At Earth's surface, gravity has an acceleration of 9.8 m/s^2.
    - The force the person exerts on the ground is therefore 51 kg × 9.8 m/s^2 = 500 kg • m/s^2 = 500 N

End Of Slide

Pressure and Force

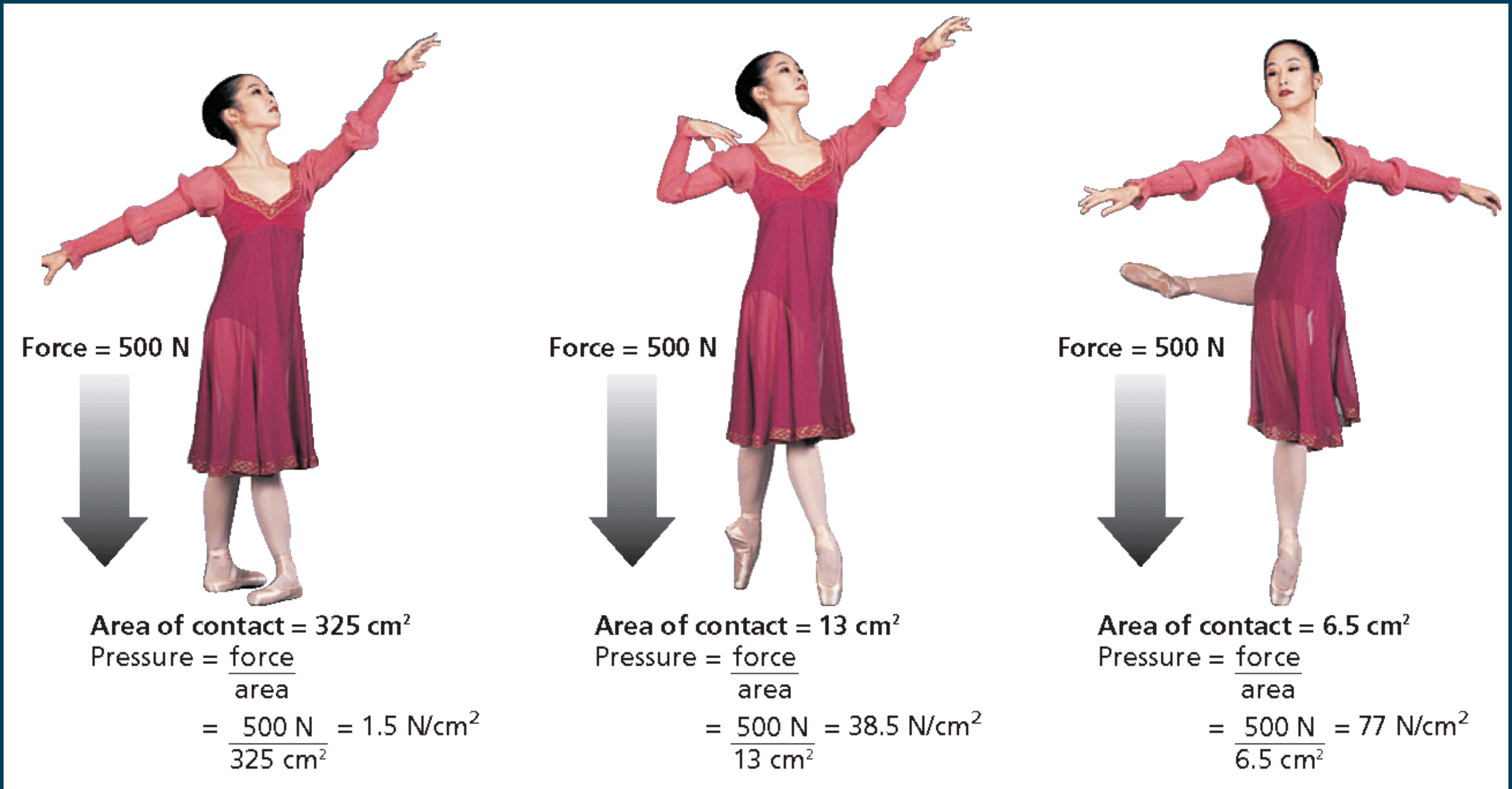
- Pressure is force per unit area, so the pressure of a 500 N person on an area of the floor that is 325 cm^2 is:

$$500 \text{ N} \div 325 \text{ cm}^2 = 1.5 \text{ N/cm}^2$$

- The greater the force on a given area, the greater the pressure.

- The smaller the area is on which a given force acts, the greater the pressure.

Relationship Between Pressure, Force, and Area



Pressure and Force, *continued*

Measuring Pressure

- A **barometer** is a device used to measure atmospheric pressure. The first barometer was introduced by Evangelista Torricelli in the early 1600s.
- Torricelli noticed that water pumps could raise water only to a maximum height of about 34 feet.
- He wondered why this was so, and thought the height must depend somehow on the weight of water compared with the weight of air.

End Of Slide

Pressure and Force, *continued*

Measuring Pressure, *continued*

- Torricelli reasoned that if the maximum height of a water column depended on its weight, then mercury, which is about 14 times as dense as water, could be raised only about 1/14 as high as water.

- He tested this idea by sealing a long glass tube at one end and filling it with mercury. Inverting the tube into a dish of mercury, the mercury rose to a height of about 30 in. (760 mm), which is about 1/14 of 34 feet.

End Of Slide

Barometer

Click below to watch the Visual Concept.

Visual Concept

Pressure and Force, *continued*

Measuring Pressure, *continued*

- The common unit of pressure is **millimeters of mercury,** symbolized mm Hg.
  - A pressure of 1 mm Hg is also called 1 torr in honor of Torricelli for his invention of the barometer.
- Pressures can also be measured in units of atmospheres. Because the average atmospheric pressure at sea level at 0°C is 760 mm Hg, one **atmosphere of pressure** (atm) is defined as being exactly equivalent to 760 mm Hg.

Pressure and Force, *continued*

Measuring Pressure, *continued*

- In SI, pressure is expressed in pascals. One **pascal** (Pa) is defined as the pressure exerted by a force of one newton (1 N) acting on an area of one square meter.
    - The unit is named for Blaise Pascal, a French mathematician and philosopher who studied pressure during the seventeenth century.
- One pascal is a very small unit of pressure, so in many cases, it is more convenient to express pressure in kilopascals (kPa). 1 atm is equal to 101.325 kPa.

Units of Pressure

Unit	Symbol	Definition/relationship
Pascal	Pa	SI pressure unit $1\ \text{Pa} = \frac{1\ \text{N}}{\text{m}^2}$
Millimeter of mercury	mm Hg	pressure that supports a 1 mm mercury column in a barometer
Atmosphere	atm	average atmospheric pressure at sea level and 0°C $1\ \text{atm} = 760\ \text{mm Hg}$ $= 1.013\ 25 \times 10^5\ \text{Pa}$ $= 101.325\ \text{kPa}$
Torr	torr	$1\ \text{torr} = 1\ \text{mm Hg}$

Pressure and Force, *continued*

Sample Problem A

The average atmospheric pressure in Denver, Colorado is 0.830 atm. Express this pressure in

a. millimeters of mercury (mm Hg) and

b. kilopascals (kPa)

Pressure and Force, *continued*

Sample Problem A Solution

Given: atmospheric pressure = 0.830 atm

Unknown: **a.** pressure in mm Hg

b. pressure in kPa

Solution:

conversion factor

a. $\text{atm} \rightarrow \text{mm Hg}; \quad \text{atm} \times \frac{760 \text{ mm Hg}}{\text{atm}} = \text{mm Hg}$

b. $\text{atm} \rightarrow \text{kPa}; \quad \text{atm} \times \frac{101.325 \text{ kPa}}{\text{atm}} = \text{kPa}$

End Of Slide

Pressure and Force, *continued*

Sample Problem A Solution, *continued*

conversion factor

a. $0.830 \text{ atm} \times \dfrac{760 \text{ mm Hg}}{\text{atm}} = \boxed{631 \text{ mm Hg}}$

b. $0.830 \text{ atm} \times \dfrac{101.325 \text{ kPa}}{\text{atm}} = \boxed{84.1 \text{ kPa}}$

End Of Slide

Dalton's Law of Partial Pressures

- The pressure of each gas in a mixture is called the **partial pressure** of that gas.
- John Dalton, the English chemist who proposed the atomic theory, discovered that the pressure exerted by each gas in a mixture is independent of that exerted by other gases present.
- **Dalton's law of partial pressures** states that the total pressure of a gas mixture is the sum of the partial pressures of the component gases.

Dalton's Law of Partial Pressures

Click below to watch the Visual Concept.

Visual Concept

Equation for Dalton’s Law of Partial Pressures

Click below to watch the Visual Concept.

Visual Concept

Dalton's Law of Partial Pressures, *continued*

Gases Collected by Water Displacement

- Gases produced in the laboratory are often collected over water. The gas produced by the reaction displaces the water in the reaction bottle.
- Dalton's law of partial pressures can be applied to calculate the pressures of gases collected in this way.
- Water molecules at the liquid surface evaporate and mix with the gas molecules. Water vapor, like other gases, exerts a pressure known as *vapor pressure.*

End of Slide

Dalton's Law of Partial Pressures, *continued*

Gases Collected by Water Displacement, *continued*

- To determine the pressure of a gas inside a collection bottle, you would use the following equation, which is an instance of Dalton's law of partial pressures.

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

- If you raise the bottle until the water levels inside and outside the bottle are the same, the total pressure outside and inside the bottle will be the same.
- Reading the atmospheric pressure from a barometer and looking up the value of P_{H_2O} at the temperature of the experiment in a table, you can calculate P_{gas}.

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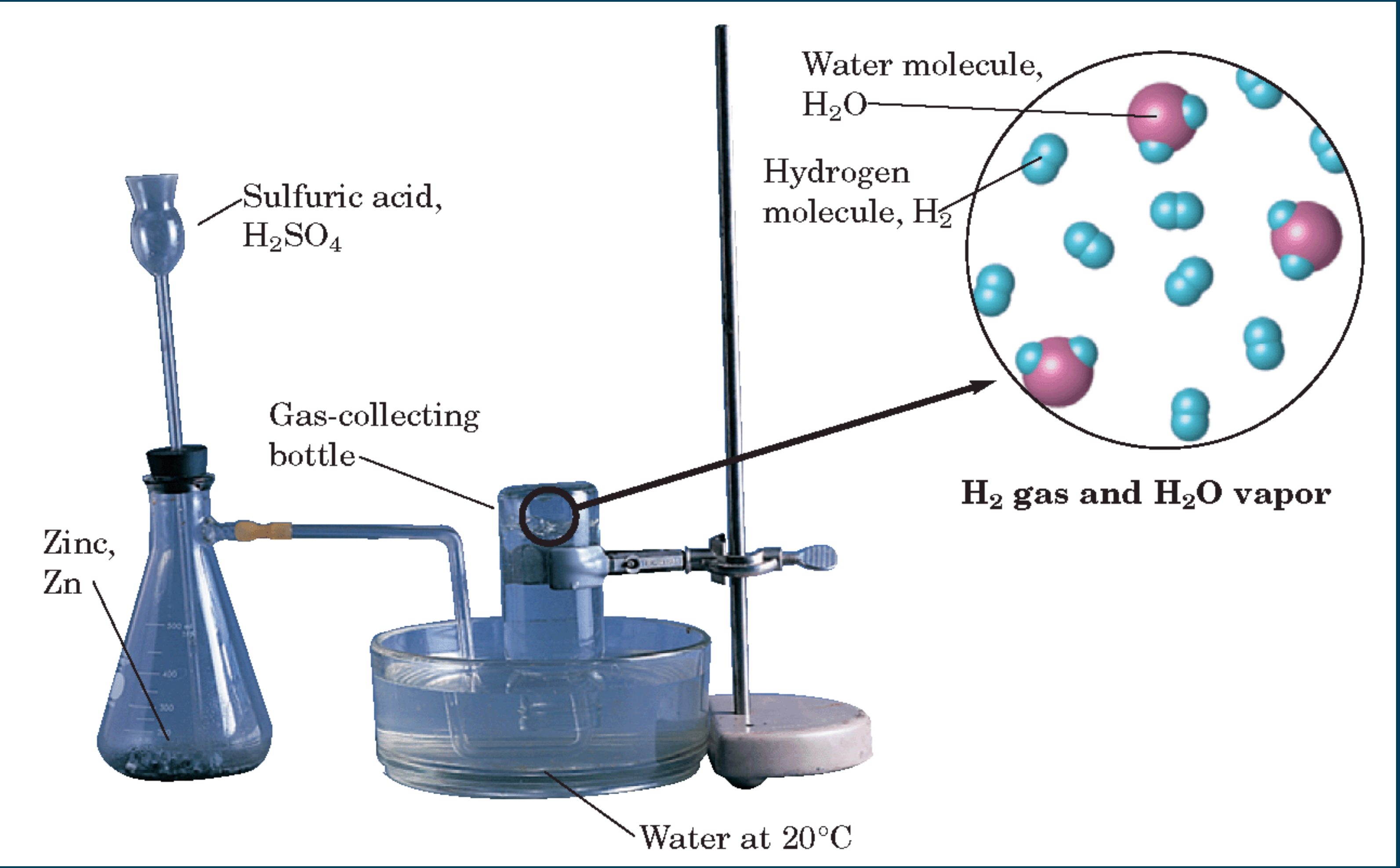
Particle Model for a Gas Collected Over Water



Dalton's Law of Partial Pressures, *continued*

Sample Problem B

Oxygen gas from the decomposition of potassium chlorate, $KClO_3$, was collected by water displacement. The barometric pressure and the temperature during the experiment were 731.0 torr and 20.0°C. respectively. What was the partial pressure of the oxygen collected?

End of Slide

Dalton's Law of Partial Pressures, *continued*

Sample Problem B Solution

Given: $P_T = P_{atm} = 731.0$ torr

$P_{H_2O} = 17.5$ torr (vapor pressure of water at 20.0°C, from table A-8 in your book)

$P_{atm} = P_{O_2} + P_{H_2O}$

Unknown: P_{O_2} in torr

Solution:

start with the equation: $P_{atm} = P_{O_2} + P_{H_2O}$

rearrange algebraically to: $P_{O_2} = P_{atm} - P_{H_2O}$

Dalton's Law of Partial Pressures, *continued*

Sample Problem B Solution, *continued*

substitute the given values of P_{atm} *and* P_{H_2O} *into the equation:*

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

Preview

Objectives

- **Use** the kinetic-molecular theory to explain the relationships between gas volume, temperature and pressure.

- **Use** Boyle's law to calculate volume-pressure changes at constant temperature.

- **Use** Charles's law to calculate volume-temperature changes at constant pressure.

End Of Slide

Objectives, *continued*

- **Use** Gay-Lussac's law to calculate pressure-temperature changes at constant volume.
- **Use** the combined gas law to calculate volume-temperature-pressure changes.

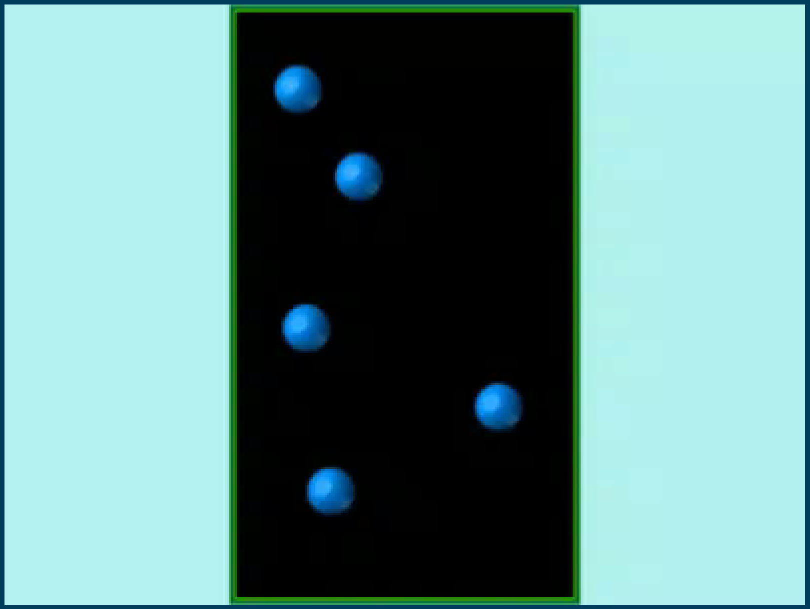
Boyle's Law: Pressure-Volume Relationship

- Robert Boyle discovered that doubling the pressure on a sample of gas at constant temperature reduces its volume by one-half.
- This is explained by the kinetic-molecular theory:
  - The pressure of a gas is caused by moving molecules hitting the container walls.
  - If the volume of a gas is decreased, more collisions will occur, and the pressure will therefore increase.
  - Likewise, if the volume of a gas is increased, less collisions will occur, and the pressure will decrease.

End Of Slide

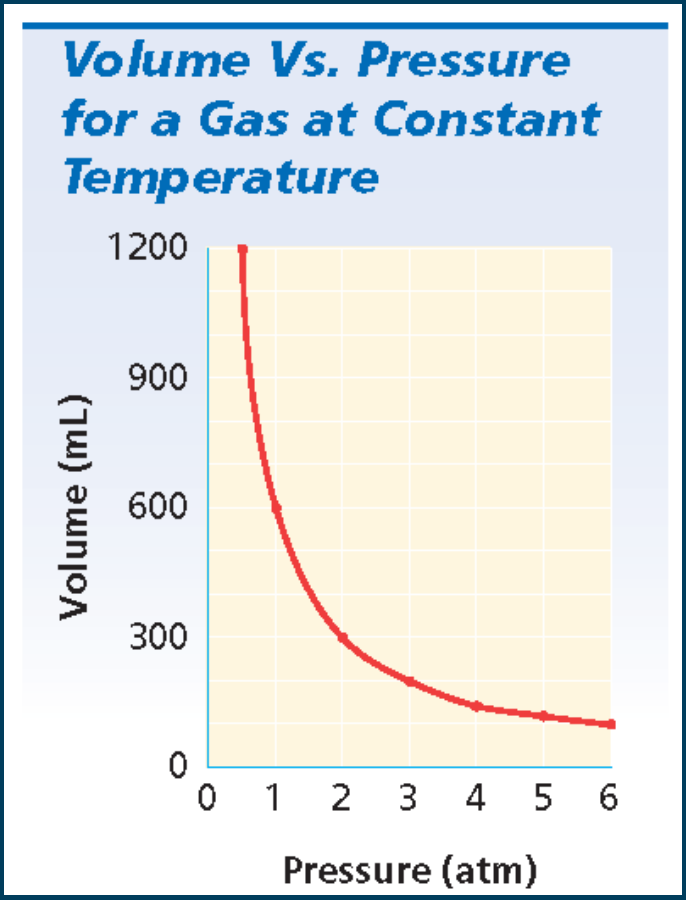
Boyle’s Law

Click below to watch the Visual Concept.



Boyle's Law: Pressure-Volume Relationship

- **Boyle's Law** states that the volume of a fixed mass of gas varies inversely with the pressure at constant temperature.
- Plotting the values of volume versus pressure for a gas at constant temperature gives a curve like that shown at right.



Boyle's Law: Pressure-Volume Relationship

- Mathematically, Boyle's law can be expressed as:

$$PV = k$$

- P is the pressure, V is the volume, and k is a constant. Since P and V vary inversely, their product is a constant.

Boyle's Law: Pressure-Volume Relationship, *continued*

- Because two quantities that are equal to the same thing are equal to each other, Boyle's law can also be expressed as:

$$P_1V_1 = P_2V_2$$

- P_1 and V_1 represent initial conditions, and P_2 and V_2 represent another set of conditions.

- Given three of the four values P_1, V_1, P_2, and V_2, you can use this equation to calculate the fourth value for a system at constant temperature.

Equation for Boyle’s Law

Click below to watch the Visual Concept.

Visual Concept

Boyle's Law: Pressure-Volume Relationship, *continued*

Sample Problem C

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?

Boyle's Law: Pressure-Volume Relationship, *continued*

Sample Problem C Solution

Given: V_1 of O_2 = 150.0 mL

P_1 of O_2 = 0.947 atm

P_2 of O_2 = 0.987 atm

Unknown: V_2 of O_2 in mL

Solution:

Rearrange the equation for Boyle's law ($P_1V_1 = P_2V_2$) to obtain V_2.

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

End of Slide

Boyle's Law: Pressure-Volume Relationship, *continued*

Sample Problem C Solution, *continued*

Substitute the given values of P_1, V_1, and P_2 into the equation to obtain the final volume, V_2:

$$V_2 = \frac{P_1V_1}{P_2} = \frac{(0.947\ \text{atm})(150.0\ \text{mL O}_2)}{0.987\ \text{atm}} = \boxed{144\ \text{mL O}_2}$$

End of Slide

Charles’s Law: Volume-Temperature Relationship, *continued*

- If pressure is constant, gases expand when heated.
  - When the temperature increases, the volume of a fixed number of gas molecules must increase if the pressure is to stay constant.
    - At the higher temperature, the gas molecules move faster. They collide with the walls of the container more frequently and with more force.
    - The volume of a flexible container must then increase in order for the pressure to remain the same.

End Of Slide

Charles's Law

Click below to watch the Visual Concept.

Charles's Law: Volume-Temperature Relationship, *continued*

- The quantitative relationship between volume and temperature was discovered by the French scientist Jacques Charles in 1787.
- Charles found that the volume changes by 1/273 of the original volume for each Celsius degree, at constant pressure and at an initial temperature of 0°C.
- The temperature –273.15°C is referred to as **absolute zero,** and is given a value of zero in the Kelvin temperature scale. The relationship between the two temperature scales is K = 273.15 + °C.

Absolute Zero

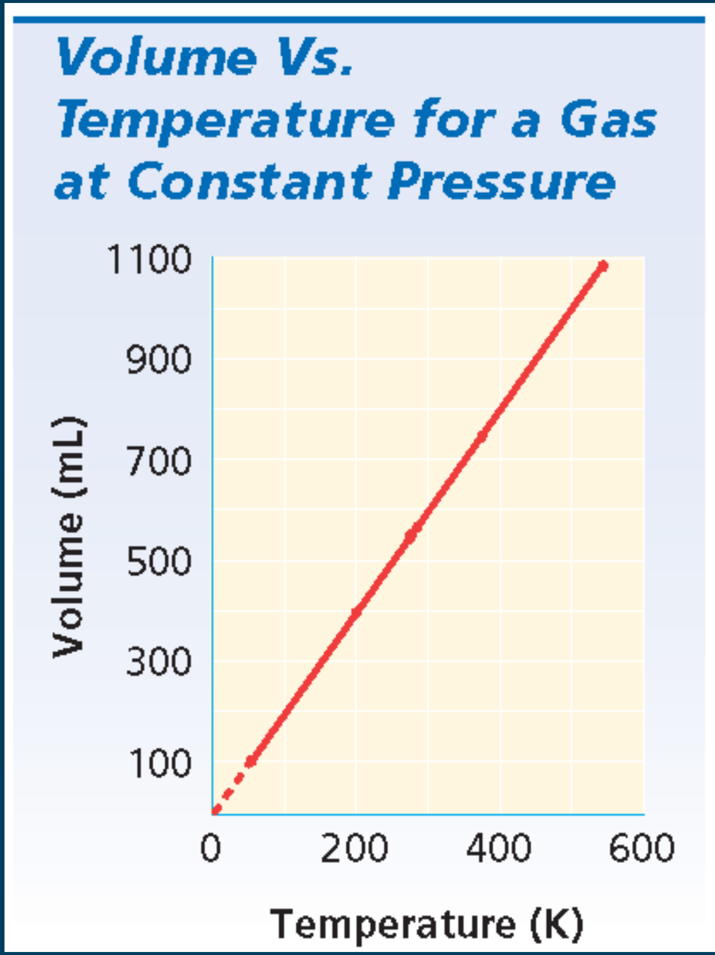
Click below to watch the Visual Concept.

Visual Concept

Charles’s Law: Volume-Temperature Relationship, *continued*

- **Charles’s law** states that the volume of a fixed mass of gas at constant pressure varies directly with the Kelvin temperature.

- Gas volume and Kelvin temperature are directly proportional to each other at constant pressure, as shown at right.



Charles's Law: Volume-Temperature Relationship, *continued*

- Mathematically, Charles's law can be expressed as:

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

- V is the volume, T is the Kelvin temperature, and k is a constant. The ratio V/T for any set of volume-temperature values always equals the same k.

- This equation reflects the fact that volume and temperature are directly proportional to each other at constant pressure.

End Of Slide

Charles's Law: Volume-Temperature Relationship, *continued*

- The form of Charles's law that can be applied directly to most volume-temperature gas problems is:

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

- V_1 and T_1 represent initial conditions, and V_2 and T_2 represent another set of conditions.

- Given three of the four values V_1, T_1, V_2, and T_2, you can use this equation to calculate the fourth value for a system at constant pressure.

Equation for Charles’s Law

Click below to watch the Visual Concept.

Visual Concept

Charles's Law: Volume-Temperature Relationship, *continued*

Sample Problem D

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?

Charles's Law: Volume-Temperature Relationship, *continued*

Sample Problem D Solution

Given: V_1 of Ne = 752 mL

T_1 of Ne = 25°C + 273 = 298 K

T_2 of Ne = 50°C + 273 = 323 K

Unknown: V_2 of Ne in mL

Solution:

Rearrange the equation for Charles's law $\left(\frac{V_1}{T_1} = \frac{V_2}{T_2}\right)$ to obtain V_2.

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

End Of Slide

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Charles's Law: Volume-Temperature Relationship, *continued*

Sample Problem D Solution, *continued*

Substitute the given values of V_1, T_1, and T_2 into the equation to obtain the final volume, V_2:

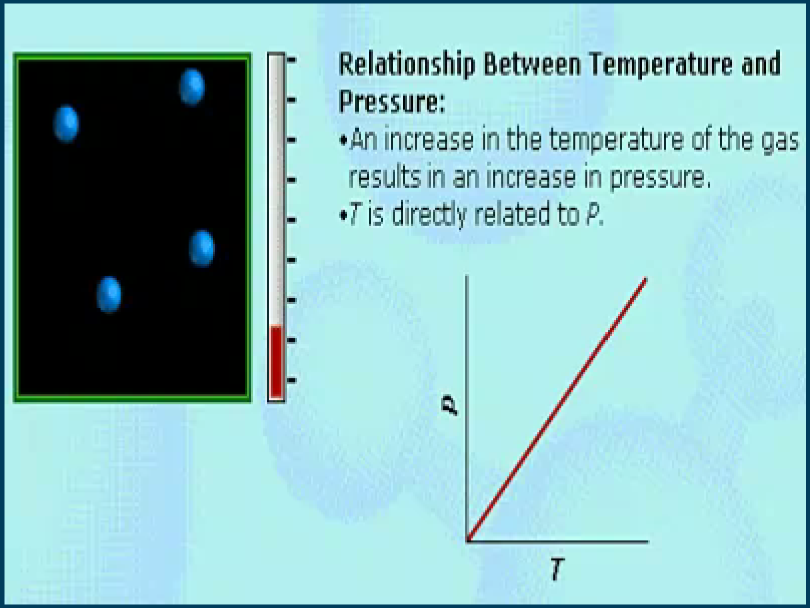
$$V_2 = \frac{V_1 T_2}{T_1} = \frac{(752 \text{ mL Ne})(323 \text{ K})}{298 \text{ K}} = \boxed{815 \text{ mL Ne}}$$

Gay-Lussac's Law: Pressure-Temperature Relationship

- At constant volume, the pressure of a gas increases with increasing temperature.
  - Gas pressure is the result of collisions of molecules with container walls.
  - The energy and frequency of collisions depend on the average kinetic energy of molecules.
  - Because the Kelvin temperature depends directly on average kinetic energy, pressure is directly proportional to Kelvin temperature.

Gay-Lussac's Law

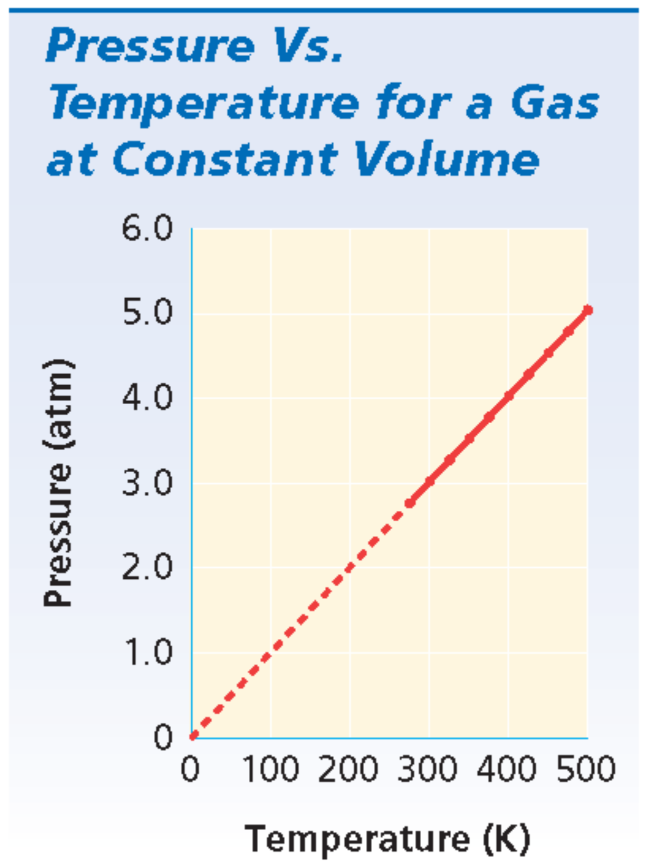
Click below to watch the Visual Concept.



Gay-Lussac's Law: Pressure-Temperature Relationship, *continued*

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

- This law is named after Joseph Gay-Lussac, who discovered it in 1802.



Gay-Lussac's Law: Pressure-Temperature Relationship, *continued*

- Mathematically, Gay-Lussac's law can be expressed as:

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

- *P* is the pressure, *T* is the Kelvin temperature, and *k* is a constant. The ratio *P*/*T* for any set of volume-temperature values always equals the same *k*.

- This equation reflects the fact that pressure and temperature are directly proportional to each other at constant volume.

End Of Slide

Gay-Lussac's Law: Pressure-Temperature Relationship, *continued*

- The form of Gay-Lussac's law that can be applied directly to most pressure-temperature gas problems is:

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

- P_1 and T_1 represent initial conditions, and P_2 and T_2 represent another set of conditions.

- Given three of the four values P_1, T_1, P_2, and T_2, you can use this equation to calculate the fourth value for a system at constant pressure.

End of Slide

Equation for Gay-Lussac’s Law

Click below to watch the Visual Concept.

Visual Concept

Gay-Lussac’s Law: Volume-Temperature Relationship, *continued*

Sample Problem E

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?

End of Slide

Gay-Lussac's Law: Volume-Temperature Relationship, *continued*

Sample Problem E Solution

Given: P_1 of gas = 3.00 atm

T_1 of gas = 25°C + 273 = 298 K

T_2 of gas = 52°C + 273 = 325 K

Unknown: P_2 of gas in atm

Solution:

Rearrange the equation for Gay-Lussac's law $\left(\frac{P_1}{T_1} = \frac{P_2}{T_2}\right)$ to obtain V_2.

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

Gay-Lussac's Law: Volume-Temperature Relationship, *continued*

Sample Problem E Solution, *continued*

Substitute the given values of P_1, T_1, and T_2 into the equation to obtain the final volume, P_2:

$$V_2 = \frac{V_1 T_2}{T_1} = \frac{(3.00\ \text{atm})(325\ \text{K})}{298\ \text{K}} = \boxed{3.27\ \text{atm}}$$

End Of Slide

Summary of the Basic Gas Laws

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

The Combined Gas Law

- Boyle's law, Charles's law, and Gay-Lussac's law can be combined into a single equation that can be used for situations in which temperature, pressure, and volume, all vary at the same time.
- The **combined gas law** expresses the relationship between pressure, volume, and temperature of a fixed amount of gas. It can be expressed as follows:

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

End Of Slide

Equation for the Combined Gas Law

Click below to watch the Visual Concept.

Visual Concept

The Combined Gas Law, *continued*

- The combined gas law can also be written as follows.

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

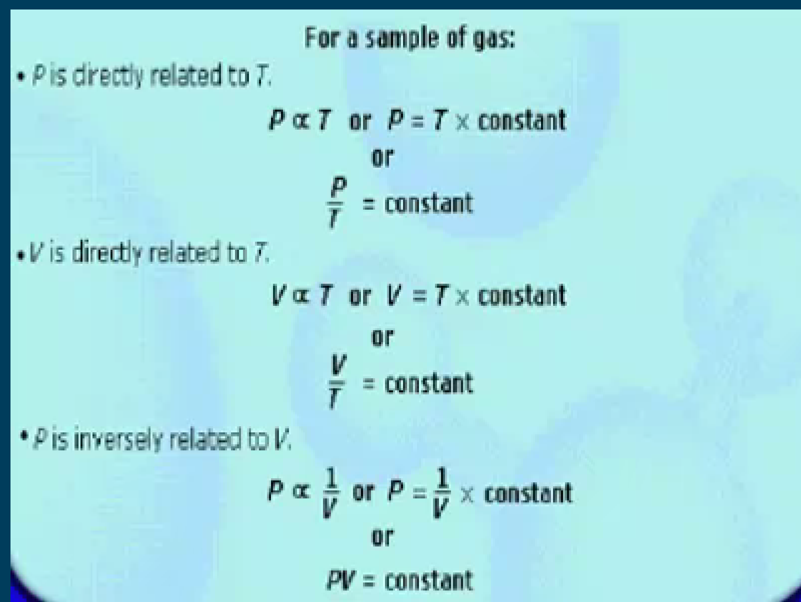
  - The subscripts 1 and 2 represent two different sets of conditions. As in Charles's law and Gay-Lussac's law, T represents Kelvin temperature.

- Each of the gas laws can be obtained from the combined gas law when the proper variable is kept constant.

End of Slide

Combined Gas Law

Click below to watch the Visual Concept.



The Combined Gas Law, *continued*

Sample Problem F

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?

End Of Slide

The Combined Gas Law, *continued*

Sample Problem F Solution

Given: V_1 of He = 50.0 L

T_1 of He = 25°C + 273 = 298 K

T_2 of He = 10°C + 273 = 283 K

P_1 of He = 1.08 atm

P_2 of He = 0.855 atm

Unknown: V_2 of He in L

The Combined Gas Law, *continued*

Sample Problem F Solution, *continued*

Solution:

Rearrange the equation for the combined gas law to obtain V_2.

$$V_2 = \frac{P_1V_1T_2}{P_2T_1} \qquad \left(\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}\right)$$

Substitute the given values of P_1, T_1, and T_2 into the equation to obtain the final volume, P_2:

$$V_2 = \frac{P_1V_1T_2}{P_2T_1} = \frac{(1.08\ \text{atm})(50.0\ \text{L He})(283\ \text{K})}{(0.855\ \text{atm})(298\ \text{K})} = \boxed{60.0\ \text{L He}}$$

Preview

Lesson Starter

- Write balanced chemical equations for the two chemical reactions indicated below.

hydrogen gas + oxygen gas → water vapor
(2 liters) (1 liter) (2 liters)

hydrogen gas + chlorine gas → hydrogen chloride
(1 liter) (1 liter) (2 liters)

- Compare the balanced equations with the expressions above. What do you notice?

End Of Slide

Objectives

- **State** the law of combining volumes.
- **State** Avogadro's law and explain its significance.
- **Define** *standard molar volume of a gas* and use it to calculate gas masses and volumes.
- **State** the ideal gas law.
- **Using** the ideal gas law, calculate pressure, volume, temperature, or amount of gas when the other three quantities are known.

Measuring and Comparing the Volumes of Reacting Gases

- In the early 1800s, French chemist Joseph Gay-Lussac observed that 2 L of hydrogen can react with 1 L of oxygen to form 2 L of water vapor.

hydrogen gas	+ oxygen gas	→ water vapor
2 L (2 volumes)	1 L (1 volume)	2 L (2 volumes)

- The reaction shows a simple 2:1:2 ratio in the volumes of reactants and products. This same ratio applies to any volume proportions: for example, 2 mL, 1 mL, and 2 mL; or 600 L, 300 L, and 600 L.

Measuring and Comparing the Volumes of Reacting Gases

- The same simple and definite volume proportions can be observed in other gas reactions.

hydrogen gas	+	chlorine gas	$\rightarrow$	hydrogen chloride gas
1 L (2 volumes)		1 L (1 volume)		2 L (2 volumes)

- **Gay-Lussac's law of combining volumes of gases** states that at constant temperature and pressure, the volumes of gaseous reactants and products can be expressed as ratios of small whole numbers.

Gay-Lussac's Law of Combining Volumes of Gases

Click below to watch the Visual Concept.

Visual Concept

Avogadro's Law

- In 1811, Amedeo Avogadro explained Gay-Lussac's law of combining volumes of gases without violating Dalton's idea of indivisible atoms.
- Avogadro reasoned that, instead of always being in monatomic form when they combine to form products, gas molecules can contain more than one atom.
- He also stated an idea known today as **Avogadro's law.** The law states that equal volumes of gases at the same temperature and pressure contain equal numbers of molecules.

Avogadro's Law

Click below to watch the Visual Concept.

Visual Concept

Avogadro's Law, *continued*

- Avogadro's law also indicates that gas volume is directly proportional to the amount of gas, at a given temperature and pressure.
- The equation for this relationship is shown below, where V is the volume, k is a constant, and n is the amount of moles of the gas.

$$V = kn$$

End Of Slide

Avogadro's Law, *continued*

- Avogadro's law applies to the combining volumes in gas reactions, and helped him to deduce chemical formulas in reactions.

- Dalton had guessed that the formula for water was HO, but Avogadro's reasoning established that water must contain twice as many H atoms as O atoms because of the volume ratios in which the gases combine:

hydrogen gas + oxygen gas $\rightarrow$ water vapor
2 L (2 volumes) 1 L (1 volume) 2 L (2 volumes)

Avogadro's Law, *continued*

- Given Avogadro's law, the simplest possible chemical formula for a water molecule indicated two hydrogen atoms and one oxygen atom.

hydrogen gas	+	oxygen gas	→	water vapor
(2 volumes)		(1 volume)		(2 volumes)

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$$

- Avogadro's idea that some gases, such as hydrogen and oxygen, must be diatomic, was thus consistent with Avogadro's law and a chemical formula for water of H_2O.

Using Gay-Lussac's Law of Combining Volumes of Gases and Avogadro's Law to Find Mole Ratios

Click below to watch the Visual Concept.

Visual Concept

Molar Volume of a Gas

- Recall that one mole of a substance contains a number of particles equal to Avogadro's constant (6.022×10^{23}).
  - example: one mole of oxygen, O_2, contains 6.022×10^{23} diatomic molecules.
- According to Avogadro's law, one mole of any gas will occupy the same volume as one mole of any other gas at the same conditions, despite mass differences.
- The volume occupied by one mole of gas at STP is known as the **standard molar volume of a gas,** which is 24.414 10 L (rounded to 22.4 L).

End of Slide

Molar Volume of a Gas, *continued*

- Knowing the volume of a gas, you can use the conversion factor 1 mol/22.4 L to find the moles (and therefore also mass) of a given volume of gas at STP.
  - example: at STP, $5.00 \text{ L of gas} \times \dfrac{1 \text{ mol}}{22.4 \text{ L}} = 0.223 \text{ mol of gas}$
- You can also use the molar volume of a gas to find the volume, at STP, of a known number of moles or a known mass of gas.
  - example: at STP, $0.768 \text{ mol of gas} \times \dfrac{22.4 \text{ L}}{1 \text{ mol}} = 17.2 \text{ L of gas}$

End of Slide

Molar Volume of a Gas, *continued*

Sample Problem G

a. What volume does 0.0685 mol of gas occupy at STP?

b. What quantity of gas, in moles, is contained in 2.21 L at STP?

Molar Volume of a Gas, *continued*

Sample Problem G Solution

a.

Given: 0.0865 mol of gas at STP

Unknown: volume of gas

Solution: Multiply the amount in moles by the conversion factor, $\frac{22.4\text{ L}}{1\text{ mol}}$.

$$0.0685\text{ mol of gas} \times \frac{22.4\text{ L}}{1\text{ mol}} = \boxed{1.53\text{ L of gas}}$$

End of Slide

Molar Volume of a Gas, *continued*

Sample Problem G Solution, *continued*

b.

Given: 2.21 L of gas at STP

Unknown: moles of gas

Solution: Multiply the volume in liters by the conversion factor, $\frac{1 \text{ mol}}{22.4 \text{ L}}$.

$$2.21 \text{ L of gas} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = \boxed{0.0987 \text{ mol of gas}}$$

End Of Slide

Gas Stoichiometry

- Gay-Lussac's law of combining volumes of gases and Avogadro's law can be applied in calculating the stoichiometry of reactions involving gases.

- The coefficients in chemical equations of gas reactions reflect not only molar ratios, but also volume ratios (assuming conditions remain the same).

    - example—reaction of carbon dioxide formation:

$$2CO(g) + O_2(g) \rightarrow 2CO_2(g)$$

2 molecules	1 molecule	2 molecules
2 mole	1 mole	2 mol
2 volumes	1 volume	2 volumes

Gas Stoichiometry, *continued*

$2CO(g)$ +	$O_2(g)$ →	$2CO_2(g)$
2 molecules	1 molecule	2 molecules
2 mole	1 mole	2 mol
2 volumes	1 volume	2 volumes

- You can use the volume ratios as conversion factors in gas stoichiometry problems as you would mole ratios:

$$\frac{2 \text{ volumes CO}}{1 \text{ volume } O_2} \text{ or } \frac{1 \text{ volume } O_2}{2 \text{ volumes CO}}$$

$$\frac{2 \text{ volumes CO}}{2 \text{ volumes } CO_2} \text{ or } \frac{2 \text{ volumes } CO_2}{2 \text{ volumes CO}}$$

etc.…

End Of Slide

Gas Stoichiometry

Click below to watch the Visual Concept.

Visual Concept

Gas Stoichiometry, *continued*

Sample Problem H

Propane, C_3H_8, is a gas that is sometimes used as a fuel for cooking and heating. The complete combustion of propane occurs according to the following balanced equation.

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

(a) What will be the volume, in liters, of oxygen required for the complete combustion of 0.350 L of propane?

(b) What will be the volume of carbon dioxide produced in the reaction? Assume that all volume measurements are made at the same temperature and pressure.

Gas Stoichiometry, *continued*

Sample Problem H Solution

a.

Given: balanced chemical equation;
V of propane = 0.350 L

Unknown: *V* of O_2

Solution: Because all volumes are to be compared at the same conditions, volume ratios can be used like mole ratios.

$$0.350\ \text{L}\ C_3H_8 \times \frac{5\ \text{L}\ O_2}{1\ \text{L}\ C_3H_8} = \boxed{1.75\ \text{L}\ O_2}$$

End of Slide

Gas Stoichiometry, *continued*

Sample Problem H Solution, *continued*

b.

Given: balanced chemical equation;
V of propane = 0.350 L

Unknown: *V* of CO_2

Solution: Because all volumes are to be compared at the same conditions, volume ratios can be used like mole ratios.

$$0.350 \text{ L } C_3H_8 \times \frac{3 \text{ L } CO_2}{1 \text{ L } C_3H_8} = \boxed{1.05 \text{ L } CO_2}$$

The Ideal Gas Law

- You have learned about equations describing the relationships between two or three of the four variables—pressure, volume, temperature, and moles—needed to describe a gas sample at a time.
- All of the gas laws you have learned thus far can be combined into a single equation, the **ideal gas law:** the mathematical relationship among pressure, volume, temperature, and number of moles of a gas.
- It is stated as shown below, where *R* is a constant:

$$PV = nRT$$

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Equation for the Ideal Gas Law

Click below to watch the Visual Concept.

Visual Concept

The Ideal Gas Law, *continued*

The Ideal Gas Constant

- In the equation representing the ideal gas law, the constant *R* is known as the **ideal gas constant.**
  - Its value depends on the units chosen for pressure, volume, and temperature in the rest of the equation.
  - Measured values of *P, V, T,* and *n* for a gas at near-ideal conditions can be used to calculate *R:*

$$R = \frac{PV}{nT} = \frac{(1\ \text{atm})(22.414\ 10\ \text{L})}{(1\ \text{mol})(273.15\ \text{K})} = 0.082\ 057\ 84\ \frac{\text{L} \bullet \text{atm}}{\text{mol} \bullet \text{K}}$$

End of Slide

The Ideal Gas Law, *continued*

The Ideal Gas Constant, *continued*

- The calculated value of *R* is usually rounded to 0.0821 (L•atm)/(mol•K).
  - Use this value in ideal gas law calculations when the volume is in liters, the pressure is in atmospheres, and the temperature is in kelvins.
- The ideal gas law can be applied to determine the existing conditions of a gas sample when three of the four values, *P, V, T,* and *n,* are known.
  - Be sure to match the units of the known quantities and the units of *R.*

Numerical Values of the Gas Constant

Units of ***R***	Numerical value of ***R***	Units of ***P***	Units of ***V***	Units of ***T***	Units of ***n***
$\frac{\text{L•mm Hg}}{\text{mol•K}}$	62.4	mm Hg	L	K	mol
$\frac{\text{L•atm}}{\text{mol•K}}$	0.0821	atm	L	K	mol
$\frac{\text{J}}{\text{mol•K}}$ *	8.314	atm	L	K	mol
$\frac{\text{L•kPa}}{\text{mol•K}}$	8.314	kPa	L	K	mol

Note: 1 L•atm = 101.325 J; 1 J = 1 Pa•m^3

* SI units

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Ideal Gas Law

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Visual Concept

The Ideal Gas Law, *continued*

Sample Problem I

What is the pressure in atmospheres exerted by a 0.500 mol sample of nitrogen gas in a 10.0 L container at 298 K?

The Ideal Gas Law, *continued*

Sample Problem I Solution

Given: V of N_2 = 10.0 L

n of N_2 = 0.500 mol

T of N_2 = 298 K

Unknown: P of N_2 in atm

Solution: Use the ideal gas law, which can be rearranged to find the pressure, as follows.

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

The Ideal Gas Law, *continued*

Sample Problem I Solution, *continued*

Substitute the given values into the equation:

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

$$P = \frac{(0.500\ \text{mol})(0.0821\ \text{L} \bullet \text{atm})(298\ \text{K})}{10.0\ \text{L}} = \boxed{1.22\ \text{atm}}$$

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Preview

Objectives

- **Describe** the process of diffusion.
- **State** Graham's law of effusion.
- **State** the relationship between the average molecular velocities of two gases and their molar masses.

Diffusion and Effusion

- The constant motion of gas molecules causes them to spread out to fill any container they are in.
- The gradual mixing of two or more gases due to their spontaneous, random motion is known as *diffusion.*
- *Effusion* is the process whereby the molecules of a gas confined in a container randomly pass through a tiny opening in the container.

Comparing Diffusion and Effusion

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Visual Concept

Graham's Law of Effusion

- Rates of effusion and diffusion depend on the relative velocities of gas molecules. The velocity of a gas varies inversely with the square root of its molar mass.
  - Recall that the average kinetic energy of the molecules in any gas depends only the temperature and equals $\frac{1}{2}mv^2$.
  - For two different gases, A and B, at the same temperature, the following relationship is true.

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

End of Slide

Graham's Law of Effusion

- From the equation relating the kinetic energy of two different gases at the same conditions, one can derive an equation relating the rates of effuses of two gases with their molecular mass:

$$\frac{\text{rate of effusion of } A}{\text{rate of effusion of } B} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$$

- This equation is known as **Graham's law of effusion,** which states that the rates of effusion of gases at the same temperature and pressure are inversely proportional to the square roots of their molar masses.

Graham's Law of Effusion

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Visual Concept

Equation for Graham's Law of Effusion

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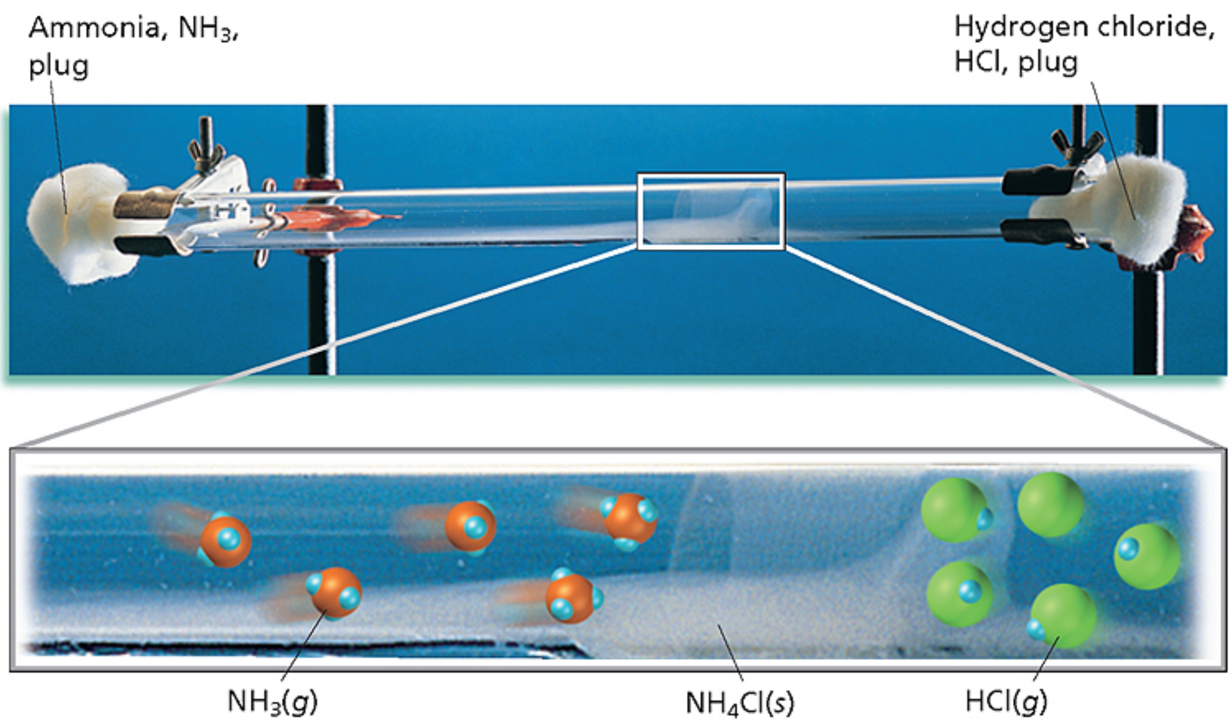
Visual Concept

Graham's Law

Cotton plugs moistened with ammonia and hydrogen chloride were placed at opposite ends of the glass tube several minutes before this photograph was taken. Why does the white ring of ammonium chloride form closer to the right end than the left end?



Graham's Law of Effusion, *continued*

Sample Problem J

Compare the rates of effusion of hydrogen and oxygen at the same temperature and pressure.

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Graham's Law of Effusion, *continued*

Sample Problem J Solution

Given: identities of two gases, H_2 and O_2

Unknown: relative rates of effusion

Solution: The ratio of the rates of effusion of two gases at the same temperature and pressure can be found from Graham's law.

$$\frac{\text{rate of effusion of } A}{\text{rate of effusion of } B} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$$

End of Slide

Graham's Law of Effusion, *continued*

Sample Problem J Solution, *continued*

Substitute the given values into the equation:

$$\frac{\text{rate of effusion of } A}{\text{rate of effusion of } B} = \frac{\sqrt{M_B}}{\sqrt{M_A}} = \frac{\sqrt{32.00 \text{ g/mol}}}{\sqrt{2.02 \text{ g/mol}}} = \sqrt{\frac{32.00 \text{ g/mol}}{2.02 \text{ g/mol}}} = 3.98$$

Hydrogen effuses 3.98 times faster than oxygen.

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