

PROPERTIES OF GASES

- *Gases* are the **least dense** and **most mobile** of the three phases of matter.
- *Particles of matter* in the gas phase are **spaced far apart** from one another and **move rapidly** and **collide** with each other often.
- Gases occupy much greater space than the same amount of liquid or solid. This is because the gas particles are spaced apart from one another and are therefore compressible. Solid or liquid particles are spaced much closer and cannot be compressed further.
- Gases are characterized by four properties. These are:
 1. Pressure (P).
 2. Volume (V)
 3. Temperature (T)
 4. Amount (n)

Kinetic-Molecular Theory of Gases

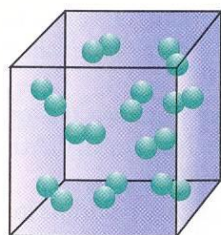
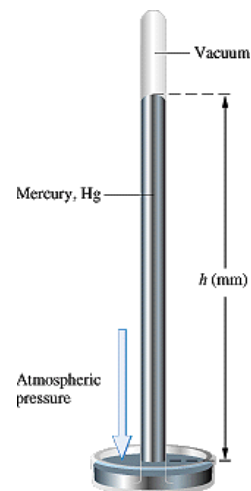
- Scientists use the **kinetic-molecular theory (KMT)** to describe the **behavior of gases**. The KMT consists of several postulates:
 1. Gases consist of **small particles** (atoms or molecules) that move **randomly** with **rapid velocities**.
 2. Gas particles have **little attraction** for one another. Therefore, **attractive forces** between gas molecules can be **ignored**.
 3. The **distance** between the particles is **large** compared to their size. Therefore the **volume** occupied by gas molecules is **small** compared to the volume of the gas.
 4. Gas particles move in **straight lines** and **collide** with each other and the container **frequently**. The force of **collisions** of the gas particles with the walls of the container causes pressure.
 5. The **average kinetic energy** of gas molecules is **directly proportional** to the **absolute temperature (Kelvin)**.

PRESSURE AND ITS MEASUREMENT

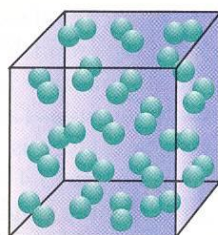
- **Pressure** is the result of *collision* of gas particles with the sides of the container. Pressure is defined as the *force per unit area*.
- **Pressure** is measured in *units of atmosphere (atm)* or *mmHg*. The SI unit of pressure is pascal (Pa) or kilopascal (kPa).

$$\begin{aligned} 1 \text{ atm} &= 760 \text{ mmHg} \\ 1 \text{ atm} &= 101.325 \text{ kPa} \end{aligned}$$

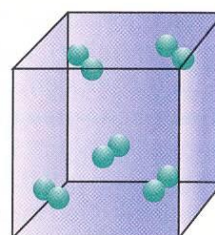
- Atmospheric **pressure** can be *measured* with the use of a **barometer**. Mercury is used in a barometer due to its high density. At sea level, the mercury stands at 760 mm above its base.
- The **pressure** of a gas is *directly proportional* to the *number of particles (moles)* present.



1 mol H₂
P = 1 atm



2 mol H₂
P = 2 atm



0.5 mol H₂
P = 0.5 atm

Examples:

1. The atmospheric pressure at Walnut, CA is 740. mmHg. Calculate this pressure in atm.

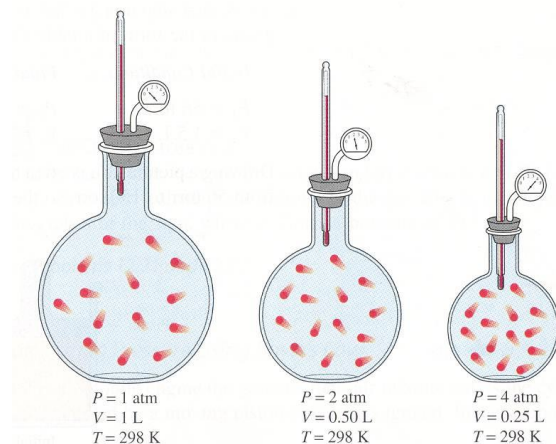
$$740. \text{ mmHg} \times \text{—————} = \text{atm}$$

2. The barometer at a location reads 1.12 atm. Calculate the pressure in mmHg.

RELATIONSHIP BETWEEN PRESSURE & VOLUME
BOYLE'S LAW

- At *constant temperature*, the *volume of a fixed amount* of gas is *inversely proportional* to its *pressure*.

$$P_1V_1 = P_2V_2$$


Examples:

- A sample of H_2 gas has a volume of 5.0 L and a pressure of 1.0 atm. What is the new pressure if the volume is decreased to 2.0 L at constant temperature?

$$P_1 = 1.0 \text{ atm} \quad P_2 = ??? \quad P_2 = \frac{P_1V_1}{V_2} = \frac{(1.0 \text{ atm})(5.0 \text{ L})}{2.0 \text{ L}} = 2.5 \text{ atm}$$

$$V_1 = 5.0 \text{ L} \quad V_2 = 2.0 \text{ L}$$

- A sample of gas has a volume of 12-L and a pressure of 4500 mmHg. What is the volume of the gas when the pressure is reduced to 750 mmHg?

$$P_1 = \quad P_2 =$$

$$V_1 = \quad V_2 = ???$$

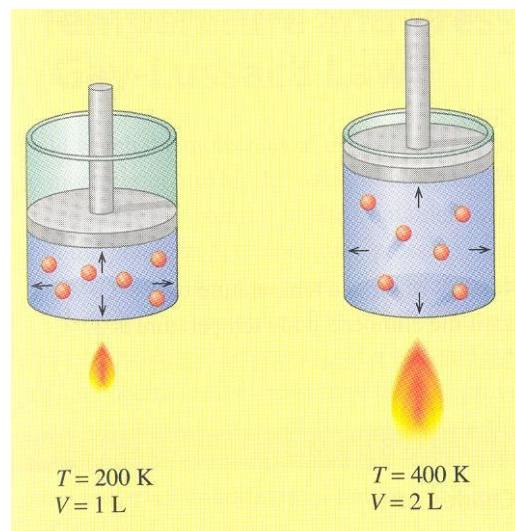
- A sample of hydrogen gas occupies 4.0 L at 650 mmHg. What volume would it occupy at 2.0 atm?

RELATIONSHIP BETWEEN TEMP. & VOLUME
CHARLES'S LAW

- At *constant pressure*, the *volume of a fixed amount* of gas is *directly proportional* to its *absolute temperature*.

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

Note: T must be in unit of K



Examples:

- A 2.0-L sample of a gas is cooled from 298K to 278 K, at constant pressure. What is the new volume of the gas?

$$V_1 = 2.0 \text{ L} \quad V_2 = ??? \quad V_2 = V_1 \times \frac{T_2}{T_1} = 2.0 \text{ L} \times \frac{278 \text{ K}}{298 \text{ K}} = 1.9 \text{ L}$$

$$T_1 = 298 \text{ K} \quad T_2 = 278 \text{ K}$$

- A sample of gas has a volume of 5.0-L and a temperature of 20°C. What is the volume of the gas when the temperature is increased to 50°C, at constant pressure?

$$V_1 = \quad V_2 =$$

$$T_1 = \quad T_2 =$$

- If 20.0 L of oxygen gas is cooled from 100 °C to 0 °C, what is the new volume?

RELATIONSHIP BETWEEN PRESSURE, VOL. & TEMP.
COMBINED GAS LAW

- All *pressure-volume-temperature* relationships can be combined into a single relationship called the *combined gas law*. This law is useful for studying the effect of *changes in two* variables.

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

- The individual gas laws studied previously are embodied in the combined gas law.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \text{Boyle's Law}$$

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \text{Charles's Law}$$

Examples:

- A 25.0-mL sample of gas has a pressure of 4.00 atm at a temperature of 10°C. What is the volume of the gas at a pressure of 1.00 atm and a temperature of 18°C?

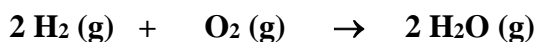
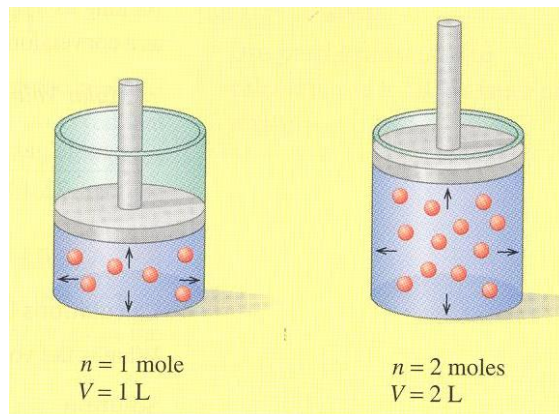
$$\begin{array}{lll}
 P_1 = & P_2 = & V_2 = V_1 \times \frac{P_1}{P_2} \times \frac{T_2}{T_1} \\
 V_1 = & V_2 = & \\
 T_1 = & T_2 = & V_2 = 25.0 \text{ mL} \times \frac{4.00 \text{ atm}}{1.00 \text{ atm}} \times \frac{291 \text{ K}}{283 \text{ K}} = 103 \text{ mL}
 \end{array}$$

- A sample of ammonia has a volume of 20.0 mL at 5 °C and 700 mmHg. What is the volume of the gas at 50°C and 850 mmHg?

$$\begin{array}{ll}
 P_1 = & P_2 = \\
 V_1 = & V_2 = \\
 T_1 = & T_2 =
 \end{array}$$

RELATIONSHIP BETWEEN VOLUME & MOLES
AVOGADRO'S LAW

- At *constant temperature and pressure*, the *volume* of a gas is *directly proportional* to the *number of moles*.
- As a result of Avogadro's Law, *equal volumes* of different gases at the *same temperature and pressure* contain *equal number of moles* (molecules).
- This relationship also allows chemists to relate volumes and moles of a gas in a chemical reaction. For example:



2 molecules	1 molecule	2 molecules
2 moles	1 mole	2 moles
2 Liters	1 Liter	2 Liters

Examples:

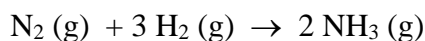
1. A sample of helium gas with a mass of 18.0 g occupies 1.6 Liters at a particular temperature and pressure. What mass of oxygen would occupy 1.6 L at the same temperature and pressure?

$$\text{mol of He} = 18.0 \text{ g} \times \frac{1 \text{ mol}}{4.00 \text{ g}} = 4.50 \text{ mol He}$$

$$\text{mol of O}_2 = \text{mol He} = 4.50 \text{ mol} \quad (\text{since at same T \& P})$$

$$\text{mass of O}_2 = 4.50 \text{ mol} \times \frac{32.0 \text{ g}}{1 \text{ mol}} = 144 \text{ g}$$

2. How many Liters of NH_3 can be produced from reaction of 1.8 L of H_2 with excess N_2 , as shown below?



$$1.8 \text{ L H}_2 \times \frac{2 \text{ L NH}_3}{3 \text{ L H}_2} = 1.2 \text{ L NH}_3$$

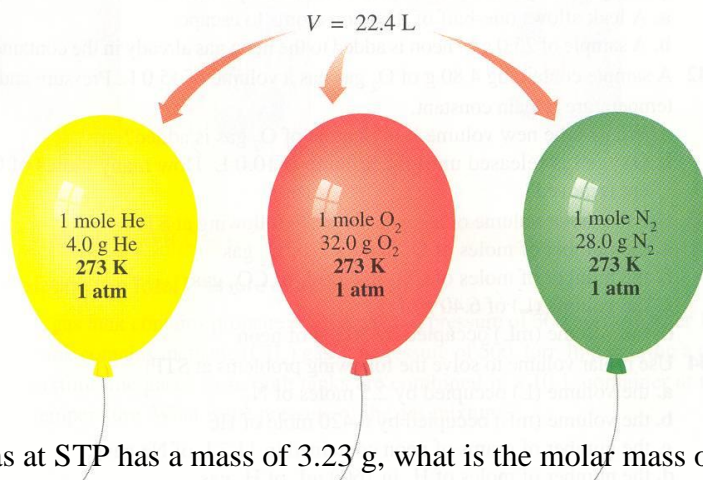
STP & MOLAR VOLUME

- To better understand the factors that affect gas behavior, a set of *standard conditions* have been chosen for use, and are referred to as *Standard Temperature and Pressure (STP)*.

STP: 0°C (273 K) and 1 atm (760 mmHg)

- At *STP* conditions, *one mole of any gas* is observed to occupy a *volume of 22.4 L*. This quantity is referred to as *Molar Volume*.

Molar Volume of a gas at STP = 22.4 L



Examples:

- If 2.00 L of a gas at STP has a mass of 3.23 g, what is the molar mass of the gas?

$$\text{mol of gas} = 2.00 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 0.0893 \text{ mol}$$

$$\text{molar mass} = \frac{\text{g}}{\text{mol}} = \frac{3.23 \text{ g}}{0.0893 \text{ mol}} = 36.2 \text{ g/mol}$$

- A sample of gas has a volume of 2.50 L at 730 mmHg and 20°C. What is the volume of this gas at STP?

$$P_1 = \quad P_2 =$$

$$V_1 = \quad V_2 =$$

$$T_1 = \quad T_2 =$$

IDEAL GAS LAW

- Combining all the laws that describe the behavior of gases, one can obtain a useful relationship that relates the volume of a gas to the temperature, pressure and number of moles.

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

$$R = \text{Universal Gas Constant} = 0.0821 \frac{\text{L atm}}{\text{mol K}}$$

- This relationship is called the **Ideal Gas Law**, and commonly written as:

$$\begin{array}{ccccccc}
 \mathbf{P} & \mathbf{V} & = & \mathbf{n} & \mathbf{R} & \mathbf{T} & \\
 \uparrow & \uparrow & & \uparrow & & \uparrow & \\
 \text{atm} & \text{L} & & \text{mol} & & \text{K} &
 \end{array}$$

Examples:

- A sample of H₂ gas has a volume of 8.56 L at a temperature of 0°C and pressure of 1.5 atm. Calculate the moles of gas present.

$$P = 1.5 \text{ atm} \quad n = ??? \quad n = \frac{PV}{RT}$$

$$V = 8.56 \text{ L} \quad T = 273 \text{ K} \quad n = \frac{(1.5 \text{ atm})(8.56 \text{ L})}{(0.0821 \text{ L atm/mol K})(273 \text{ K})} = 0.57 \text{ mol}$$

- What volume does 40.0 g of N₂ gas occupy at 10°C and 750 mmHg?

$$P = \quad n =$$

$$V = ??? \quad T = \quad V = \frac{nRT}{P} =$$

- A 23.8-L cylinder contains oxygen at 20.0 °C and 732 mmHg. How many moles of oxygen does it contain?

PARTIAL PRESSURES

DALTON'S LAW

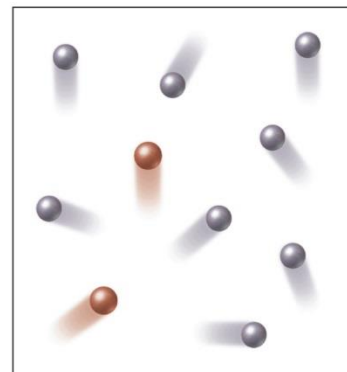
- Many gas samples are mixture of gases. For example, the air we breathe is a mixture of mostly oxygen and nitrogen gases.
- Since gas particles have no attractions towards one another, each gas in a mixture behaves as if it is present by itself, and is not affected by the other gases present in the mixture.
- In a mixture, each gas exerts a pressure as if it was the only gas present in the container. This pressure is called *partial pressure* of the gas.
- In a mixture, the sum of all the partial pressures of gases in the mixture is equal to the total pressure of the gas mixture. This is called *Dalton's law* of partial pressures.

$$\begin{aligned}
 P_{\text{total}} &= P_1 + P_2 + P_3 + \dots \\
 \text{Total pressure} &= \text{sum of the partial pressures} \\
 \text{of a gas mixture} &\quad \text{of the gases in the mixture}
 \end{aligned}$$

- The partial pressure of each gas in a mixture is proportional to the amount (mol) of gas present in the mixture. Therefore, the partial pressure of a gas is its fractional composition (X) times the total pressure of the gas. (Note: fractional composition is percent divided by 100)
- For example, in the mixture shown below, since the mixture is 80% He and 20% Ne, the partial pressure of each gas can be calculated as shown:

$$P_{\text{He}} = X_{\text{He}} P_{\text{tot}} = (0.80)(1.0 \text{ atm}) = 0.80 \text{ atm}$$

$$P_{\text{Ne}} = X_{\text{Ne}} P_{\text{tot}} = (0.20)(1.0 \text{ atm}) = 0.20 \text{ atm}$$



Gas mixture (80% He ●, 20% Ne ●)

$$\begin{aligned}
 P_{\text{tot}} &= 1.0 \text{ atm} \\
 P_{\text{He}} &= 0.80 \text{ atm} \\
 P_{\text{Ne}} &= 0.20 \text{ atm}
 \end{aligned}$$

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PARTIAL PRESSURES

Examples:

1. A mixture of AHe, Ne and Ar has a total pressure of 558 mmHg. If the partial pressure of He is 341 mmHg and the partial pressure of Ne is 112 mmHg, what is the partial pressure of Ar?

2. A scuba tank contains a mixture of oxygen and helium gases with total pressure of 7.00 atm. If the partial pressure of oxygen in the tank is 1140 mmHg, what is the partial pressure of helium in the tank?

3. Calculate the partial pressure of oxygen that a diver breathes with a mixture containing 2.0% oxygen at a depth of 100 m where the total pressure is 8.0 atm.

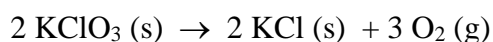
4. A diver breathing a mixture with an oxygen composition of 5.0% cannot exceed a depth where partial pressure of oxygen is greater than 0.21 atm. What must be the maximum total pressure at this depth?

GASES IN CHEMICAL REACTIONS

- Earlier we saw how Avagadro's law can be used to solve stoichiometry problems relating to reactions where all substances are in gas phase.
- For reactions where only some of the reactants or products are in gas phase, the ideal gas law can be used to find moles of that reactant or product from volume, pressure and temperature data.

Examples:

1. How many liters of oxygen gas forms, at 278 K and 755 mmHg, when 294 g of KClO_3 decomposes according to the reaction below:



2. How many grams of Ag_2O must have decomposed to produce 4.58 L of O_2 at 745 mmHg and 35°C , as shown below:

