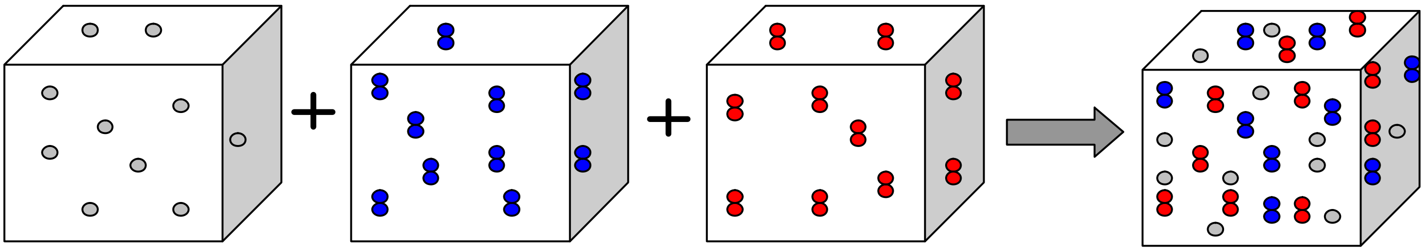
GAS MIXTURES

- Consider mixing equal volumes of 3 different gases, all at the same temperature and pressure in a container of the same size.



1 L He	1 L N_2	1 L O_2	1 L mixture
$t = 0^0C$	$t = 0^0C$	$t = 0^0C$	$t = 0^0C$
P(He) = 2 atm	$P(N_2)$ = 2 atm	$P(O_2)$ = 2 atm	P(mixture) = ?

P(mixture) = 6 atm

NOTE:

1. Each gas in a mixture of gases acts as though it were the only gas in the mixture

2. P(mixture) = P(He) + $P(N_2)$ + $P(O_2)$

PARTIAL PRESSURES

DALTON'S LAW OF PARTIAL PRESSURES

- **The Total Pressure (P_T) of a mixture of gases is equal to the sum of the partial pressures of the gases in the mixture.**

$$P_T = P_A + P_B + P_C + \ldots\ldots$$

<u>MOLE FRACTION OF A GAS IN A MIXTURE OF GASES</u>

Gas A + Gas B + Gas C = Mixture of gases
Total Number of Moles of gas in the Mixture = N

N_A = number of moles of gas A
N_B = number of moles of gas B
N_C = number of moles of gas C

$$N = N_A + N_B + N_C$$

$$\frac{N_A}{N} = \text{Mole Fraction of Gas A} = X_A$$

$$\frac{N_B}{N} = \text{Mole Fraction of Gas B} = X_B$$

$$\frac{N_C}{N} = \text{Mole Fraction of Gas C} = X_C$$

<u>Mole Percent</u> = <u>Mole fraction X 100</u>

Example: Mole % of A = $\frac{N_A}{N} \times 100$

DALTON'S LAW OF PARTIAL PRESSURES

- The **Pressure** of a sample gas is **directly proportional** to the number of molecules (expressed as **number of moles**) of gas.

$$\frac{P_A}{P} = \frac{N_A}{N} \qquad \frac{P_B}{P} = \frac{N_B}{N} \qquad \frac{P_C}{P} = \frac{N_C}{N}$$

Where:

- N_A , N_B , N_C = number of moles of gas A, B and C respectively
- N = total number of moles of gas mixture
- P_A , P_B , P_C = partial pressures of gas A, B and C respectively
- P = total pressure of gas mixture

Alternately,

$$P_A = X_A\, P \qquad P_B = X_B\, P \qquad P_C = X_C\, P$$

Examples:

1. Calculate the total pressure (in atm) and partial pressure of each gas in a mixture composed of 0.0200 moles helium gas and 0.0100 moles hydrogen gas in a 5.00 L flask at 10^0 C.

N_T = 0.0200 mol + 0.0100 mol = 0.0300 mol

n_T = 0.0300 mol
V = 5. 00 L
T = 283 K
P = ?

$$P\,V = n_T\,R\,T \qquad P = \frac{n_T\,R\,T}{V}$$

$$P = \frac{(0.300\ \cancel{mol})(0.0821\ \frac{\cancel{L}\ atm}{\cancel{mol}\ \cancel{K}})(283\ \cancel{K})}{5.00\ \cancel{L}} = 0.139\ atm$$

$P_{He} = X_{He}\, P =$

$P_{H2} = X_{H2}\, P =$

2. The atmosphere in a sealed diving bell contains oxygen and helium. If the gas mixture has 0.200 atmospheres of oxygen and a total pressure of 3.00 atmospheres, calculate the mass of helium in 1.00 L of gas mixture at 20^0C.

PART A : Calculate the pressure of helium

$P_T = P_{Oxygen} + P_{Helium}$ $\qquad$ $P_{Helium} = P_T - P_{Oxygen}$

$P_{Helium} = 3.00 \text{ atm} - 0.200 \text{ atm} = 2.80 \text{ atm}$

PART B: Calculate the number of moles of Helium

$P_{He} = 2.80$ atm
$V_{He} = 1.00$ L
$T_{He} = 293$ K
$n_{He} = ?$

$PV = \mathbf{n} RT$ $\qquad$ $\mathbf{n} = \frac{PV}{RT}$

$$n = \frac{(2.80 \cancel{\text{atm}})(1.00 \cancel{\text{L}})}{(0.0821 \frac{\cancel{\text{L}}\ \cancel{\text{atm}}}{\text{mol}\ \cancel{\text{K}}})(293 \cancel{\text{K}})} = 0.11\underline{6}5 \text{ mol}$$

PART C: Calculate the mass of Helium

$$? \text{ g He} = 0.11\underline{6}5 \text{ moles He} \times \frac{4.00 \text{ g He}}{1 \text{ mole He}} = 0.466 \text{ g He}$$

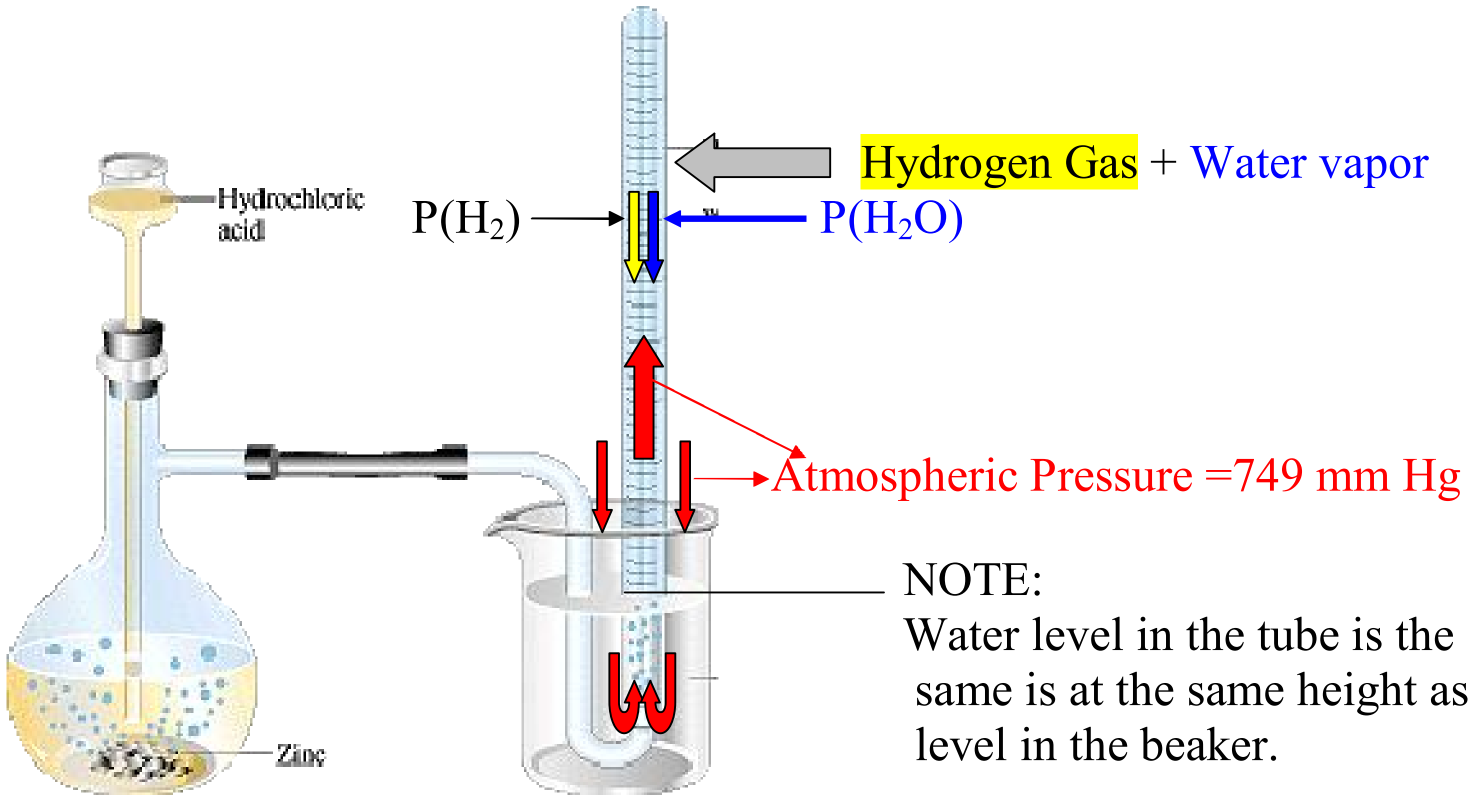
3. A mixture of 0.25 mol H_2 and 0.45 mol of N_2 are placed in a 1.50-L flask at 27°C. Calculate the partial pressures of each gas and the total pressure.

4. The partial pressure of CH_4 gas is 0.175 atm and that of O_2 gas is 0.250 atm in a mixture of the two gases. What is the mole fraction of each gas in the mixture?

COLLECTING GASES OVER WATER

- Many gases are commonly collected over water.
- Consider the collection of Hydrogen gas from the reaction of zinc with hydrochloric acid:

$$Zn(s) + HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$



Gas Pressure in the tube = $P_T = P(H_2) + P(H_2O)$

$P(atm) = P_T = P(H_2) + P(H_2O)$

$$P(atm) = P(H_2) + P(H_2O)$$

$P(H_2O)$ = is temperature dependent

= is given in tables (Table 5.6, page 206)

Example:

1. An aqueous solution of ammonium nitrite decomposes when heated to give off nitrogen gas and water vapor:

$$NH_4NO_2\,(aq) \longrightarrow 2\,H_2O(g) + N_2(g)$$

How many grams of NH_4NO_2 must have reacted if 4.16 L of N_2 gas was collected over water at 19 °C and 0.965 atmospheres of pressure ?

$$\underset{?\,g}{NH_4NO_2\,(aq)} \longrightarrow 2\,H_2O(g) + \underset{4.16\,L}{N_2(g)}$$

<u>PART A</u>: <u>Calculate the pressure of dry Nitrogen</u>

$P_T = P(N_2) + \text{Pressure of Water Vapor}$

$\mathbf{P(N_2) = P_T - \text{Pressure of Water Vapor}}$

$P_T = 0.965$ atm

Pressure of Water Vapor = ?

Table 5.6 at 19 °C ⇨ $P(H_2O) = 16.5$mmHg

$$16.5\text{ mm Hg} \times \frac{1\text{ atm}}{760.0\text{ mm Hg}} = 0.0217\text{ atm}$$

$P(N_2)$ = 0.965 atm – 0.0217 atm = **0.94$\underline{3}$3 atm**

<u>PART B</u>: <u>Calculate the number of moles of nitrogen</u>

P = 0.94$\underline{3}$3 atm
V = 4.16 L
T = 292 K
n = ?

$$P\,V = n\,R\,T \qquad n = \frac{P\,V}{R\,T}$$

$$n = \frac{(0.94\underline{3}3\ \cancel{atm})(4.16\ \cancel{L})}{(0.0821\frac{\cancel{L}\ \cancel{atm}}{mol\ \cancel{K}})(292\ \cancel{K})} = 0.16\underline{3}7\text{ mol}$$

<u>PART C: Calculate the mass of NH_4NO_2 (Stoichiometry)</u>

$$?g\ NH_4NO_2 = 0.16\underline{3}7\ \cancel{moles\ N_2} \times \frac{1\ \cancel{mole\ NH_4NO_2}}{1\ \cancel{mole\ N_2}} \times \frac{64.06\text{ g } NH_4NO_2}{1\ \cancel{mole\ NH_4NO_2}} = \mathbf{10.5\ g}$$

2. Helium is collected over water at 25°C and 1.00 atm total pressure. What volume of gas must be collected to obtain 0.586 g of helium? (At 25°C the vapor pressure of water is 23.8 mmHg)

KINETIC - MOLECULAR THEORY OF IDEAL GASES

(Theory of Gas Molecules in Motion)

Ideal Gas = A gas that follows the IDEAL GAS LAW (**P V = n R T**)

BASIC IDEA:

- **A GAS CONSISTS OF MOLECULES IN CONSTANT RANDOM MOTION**

Recall: Kinetic Energy = $E_K = \frac{1}{2} mv^2$

m = mass of molecule
v = molecular speed
E_K = energy of moving molecule

POSTULATE 1:

- The size of molecules is negligible compared to the intermolecular distances

Consequence:

- **The Volume of a gas is determined by the distances between the molecules.**

POSTULATE 2:

- Molecules move randomly in straight lines in all directions and at various speeds

Consequence:

- **The Pressure of a gas is the same in all directions**

POSTULATE 3:

- The attractive or repulsive forces between molecules are very weak (negligible)

POSTULATE 4:

- When molecules collide with one another, the collisions are elastic (no energy is lost). This is similar to collisions between billiard balls.

POSTULATE 5:

- The Average Kinetic Energy is directly proportional to the Absolute Temperature

INTERPRETATION OF THE KINETIC-MOLECULAR THEORY OF GASES

The **PRESSURE** of a gas results from: -The **COLLISIONS** of the molecules with the walls of the container

The NUMBER OF COLLISIONS is determined by:
- -the number of molecules per unit volume
- -average speed of molecules

BOYLE'S LAW

Recall: The Temperature is constant; therefore: KE_{ave} = constant

An Increase In Volume → Decreases the number of molecules per unit volume → Decreases the frequency of collisions → **Decreases the Pressure**

CHARLES' LAW

Recall: The pressure is constant, therefore: Frequency and force of molecular collisions are constant.

An Increase In Temperature → Increases Average Molecular Speed → I**ncreases the Volume** (molecules spread), in order to maintain frequency of collisions unchanged

MOLECULAR SPEEDS

- In gases, molecular speeds vary over a wide range
- Meaning: Different molecules move at different speeds, that are constantly changing, as the molecules collide with one another and the walls of the container

It is more convenient to refer to :

U = AVERAGE MOLECULAR SPEED
= THE SPEED OF A MOLECULE HAVING AVERAGE KINETIC ENERGY (KE_{AVE})

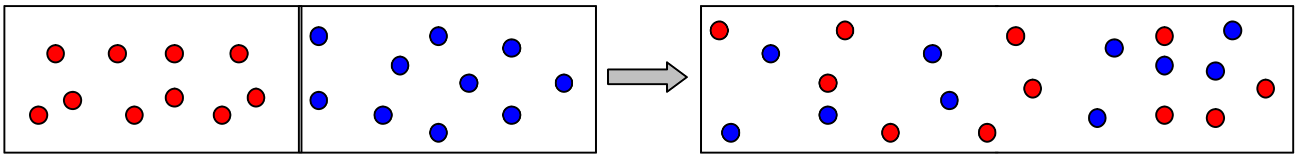
$$U = \sqrt{\frac{3\,R\,T}{M_m}}$$

R = gas constant
T = absolute temperature
M_m = molar mass

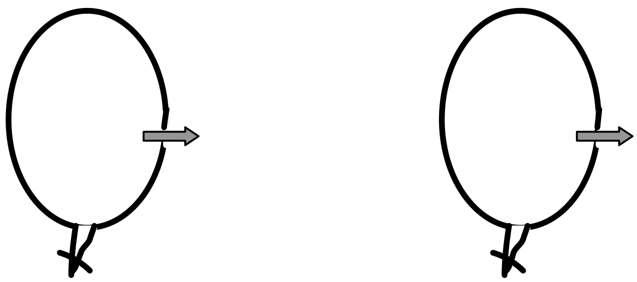
<u>**NOTE:**</u> **The higher the molar mass, the lower the average molecular speed.**
(The heavier the molecules, the slower they move)

DIFFUSION & EFFUSION

- Diffusion is the process by which a gas spreads through another gas.



- Effusion is the process by which a gas flows out from a container through a small hole.



Oxygen Hydrogen

- Same Volume, Same Pressure, Same Temperature, Same Size Hole
- Which Balloon deflates faster? OR Which gas has a **higher Rate of Effusion** ?
- What is the **Ratio between the Rates of Effusion** ?

Rate of Effusion depends on:

1. The cross-sectional area of the hole → (the same)
2. The number of molecules per unit volume → (the same)
3. **The average Molecular Speed (U)** → **(different)**
   (the faster they move, the sooner they escape)

Recall: $$\mathbf{U} = \sqrt{\frac{\mathbf{3\,R\,T}}{\mathbf{M_m}}}$$

NOTE: The Rate of Effusion:

- is directly proportional to the Average Molecular Speed (U)
- is inversely proportional to the square root of the Molar Mass

GRAHAM'S LAW OF EFFUSION

- **The Rate of Effusion is inversely proportional to the square root of the Molar Mass**

$$\frac{\text{Rate of Effusion of } H_2}{\text{Rate of Effusion of } O_2} = \frac{\sqrt{\text{Molar Mass of } O_2}}{\sqrt{\text{Molar Mass of } H_2}} = \sqrt{\frac{32.0}{2.0}} = 4$$

Conclusion: Hydrogen will effuse 4 times faster than Oxygen

Examples:

1. Obtain the ratio of effusion of H_2 and H_2S under the same conditions.

$$\frac{\text{Rate } H_2}{\text{Rate } H_2S} = \frac{\sqrt{M_m\ (H_2S)}}{\sqrt{M_m (H_2)}} = \sqrt{\frac{34}{2}} = 4.1$$

2. If 0.10 mol of I_2 vapor can effuse from an opening in a heated vessel in 52 seconds, how long will it take 0.10 mol of H_2 to effuse under the same condition?

$$\frac{\text{Rate } H_2}{\text{Rate } I_2} = \frac{\sqrt{M_m\ (I_2)}}{\sqrt{M_m (H_2)}} = \sqrt{\frac{253.8}{2.02}} = 11.2$$

H_2 effuses 11.2 faster than I_2

It follows that the time t takes H_2 to effuse is 11.2 times shorter:

$$T\ (H_2) = \frac{52 \text{ seconds}}{11.2} = 4.6 \text{ seconds}$$

REAL GASES

Gases behave according to the Ideal Gas Law, provided:

1. The space occupied by molecules is truly negligible compared to the total gas volume (Postulate 1),
2. The intermolecular forces of attraction are truly negligible (Postulate 3)

THIS IS TRUE:

- at High Temperature, and
- at Low Pressure

Reasons:

- Molecules are far apart
- Molecules do not attract

IDEAL GAS

PV = nRT

THIS IS NOT QUITE TRUE:

- at Low Temperature, and
- at High Pressure

Reasons:

- Molecules are relatively closer

As such:

- **Volume occupied by molecules must be taken into account**
- **Intermolecular forces between molecules are not negligible and molecules begin to attract one another**

$$\left(P + \frac{n^2a}{V^2}\right)(V-nb) = nRT$$

Van Der Waals Equation for Real Gases

NOTE: **V** (intermolecular space) becomes **(V - nb)**
- **decreased Volume** accounts for space occupied by molecules)

P becomes $\left(P + \frac{n^2a}{V^2}\right)$ **increased Pressure** accounts for intermolecular forces of attraction

a and b- are constants
- have different values for different gases