*Chapter 5: The Gaseous State.

*5.1: Gas Pressure and its Measurement.

* Gas physical characteristics:

1) gases assume the same shape and volume of their container.
2) gases are the most compressible of the states of matter.
3) gases have much lower densities than liquid and solid.

$$\boxed{\text{Pressure} = \frac{\text{Force}}{\text{Area}}}$$

* Pressure: the force exerted per unit area of surface.

* The SI unit of pressure is $Kg/(m.s^2)$ => pascal (Pa)

* barometer: a device for measuring the pressure of the atmosphere.

* manometer: a device that measures the pressure of a gas or liquid in a vessel.

* $1 \text{ atm} = 1.01325 \times 10^5 \text{ Pa} \approx 101 \text{ kPa}$

$760 \text{ mmHg} = 1 \text{ atm}$

$1.01325 \text{ bar} = 1 \text{ atm}$

Important units of Pressure.

* note: mmHg → torr

* Example: The pressure of a gas is 797.7 mmHg, what is the pressure in Pascals and atmospheres?

$$797.7\ \cancel{\text{mmHg}} \times \frac{1.01325\times10^{5}\ \text{Pa}}{760\ \text{mmHg}} \doteq \boxed{1.064\times10^{5}\ \text{Pa}}$$

$$797.7\ \text{mmHg} \times \frac{1\ \text{atm}}{760\ \text{mmHg}} \doteq \boxed{1.05\ \text{atm}}$$

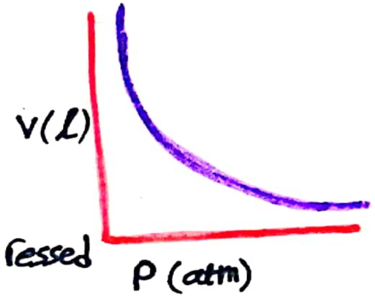
*5.2: Empirical gas laws:

1) Boyle's law: relationship between (P and V) at constant (T and n).

$$P \propto \frac{1}{V}$$

$PV = \text{constant}$

$P_1V_1 = P_2V_2$



* example: a volume of air = 12.0 dm^3 at 98.9 KPa, compressed to 119.0 KPa at constant temprature, determine the new V?

$V_1 = 12\ dm^3$

$V_2 = ??$

$P_1 = 98.9\ KPa$

$P_2 = 119\ KPa$

$$V_2 = \frac{P_1V_1}{P_2} = \frac{98.9\ \cancel{KPa} \times 12\ dm^3}{119\ \cancel{KPa}} = 9.97\ dm^3$$

2) Charles's law: relating (V and T) at constant (P and n).

$V \propto T$

$\frac{V}{T}$ = constant

$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ and T in K

* Example: a gas P = 1.0 atm T = 21 c° V = 785 ℓ

what would be V if T increases to T = 28 c°?

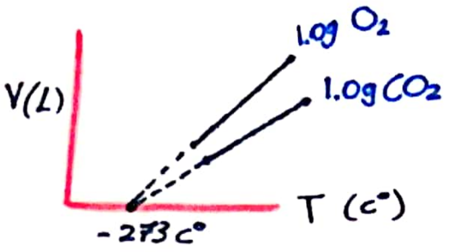
$T_1 = 21 + 273 = 294$ K

$T_2 = 28 + 273 = 301$ K

$V_1 = 785$ L

$V_2 = ??$

$V_2 = \frac{V_1 T_2}{T_1} = \frac{785\text{ L} \times 301\text{ K}}{294\text{ K}} = \boxed{804\text{ L}}$



3) Combined Gas Law: relating volume, Temperature, and Pressure

$V \propto \frac{T}{P} \Rightarrow \frac{PV}{T} = \text{constant}$

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

*Example: 39.8 mg caffine, 10.1 cm³ N_2, 23 c°, 746 mmHg, V at 0 c° and 760 mmHg ??

$T_1 = 23 + 273 = 296$ K

$T_2 = 273$ K

$V_1 = 10.1\text{ cm}^3$

$V_2 = ??$

$P_1 = 746$ mmHg

$P_2 = 760$ mmHg

$V_2 = \frac{P_1 V_1 T_2}{T_1} = 9.14\text{ cm}^3$

4) Avogdro's Law: relating Volum and Amount (n) at constant (T and P).

$$2H_2 + O_2 \longrightarrow 2H_2O$$

$$2\text{mol } H_2 + 1\text{mol } O_2 \longrightarrow 2\text{mol } H_2O$$

$$2V \qquad 1V \qquad 2V$$

*Molar gas volume (V_m): Volume of one mole of gas, at "STP" (standard temperature and pressure)

$T = 0°C$, $P = 1$ atm , $V_m = 22.4$ L/mol

$V \propto n$ — $\frac{V}{n}$ = constant $\quad \frac{V_1}{n_1} = \frac{V_2}{n_2}$ at constant T and P

*5.3: The Ideal Gas Law:

* from the combined gas law

$$\frac{PV}{T} = \text{constant} \Rightarrow V = \text{constant} \times \frac{T}{P}$$

$$V_m = R \times \frac{T}{P}$$

$$\underbrace{nV_m}_{V} = \frac{nRT}{P} \Rightarrow \boxed{PV = nRT}$$

"The Ideal gas law"

*$R = 8.3145$ J/Kmol $= \underline{0.0821}$ $\underline{\text{atm.L/mol.K}}$

*Example: How many grams of oxygen, O_2, are thee in a 50.0L gas cylinder at 21°C when the oxygen pressure is 15.7 atm?

$V = 50L$ $\quad T = 21 + 273 = 294\ K$ $\quad P = 15.7\text{atm}$

$$PV = nRT$$

$$15.7 \times 50 = n \times 0.0821 \times 294$$

$$\boxed{n = 32.5\text{mol}}$$

32.50

32.53

$$m = n \times M.W$$

$$= 32.5 \times 32 = 1.04 \times 10^3 g\ O_2$$

* Example: what is the density of O_2 in grams per liter at 25°C and 0.850 atm?

$T = 25 + 273 = 298\ K$

$V = 1\ L$

*$P = 0.85\ atm$

$$PV = \underline{\underline{n}}RT$$

$$0.85 \times 1 = n \times 0.0821 \times 298$$

$$n = \underline{0.0347\ mol} \times 32 = \boxed{1.11\ g\ O_2}$$

$$\frac{1.11\ g\ O_2}{1\ L} = \boxed{1.11\ g/L}$$

⇒ suppose you wanted the density of Cl_2 insted of O_2 at this T and P:

$$\underset{\checkmark}{P}\underset{\checkmark}{V} = n\underset{\checkmark}{R}\underset{\checkmark}{T} \qquad n_{Cl_2} = n_{O_2}$$

$$n = 0.0347\ mol \times 70.9 = 2.46\ g\ Cl_2/L$$

$$PV = nRT \rightarrow PV = \frac{m}{M.m} RT \rightarrow P\,M.m = \frac{m}{V} RT \qquad d = \frac{P\,M.m}{RT}$$

$$P\,Mm = dRT \qquad M.m = \frac{dRT}{P}$$

*** Example:** A 200.0 mL flask at 99°C and a pressure of 733 mmHg is filled with vapor of easily vaporized liquid, The mass of the substance in the flask is 0.970g. what is the molecular weight of the liquid?

- $T = 99 + 273 = 372\ K$
- $733\ mmHg = P \rightarrow P = 0.964\ atm$
- $m = 0.97\ g$
- $V = 200\ mL = 0.2\ L$

$$M.W = \frac{dRT}{P} = \frac{4.85 \times 0.0821 \times 372}{0.964} = 153.7 = 154$$

*5.4: Stoichiometry Problems Involving Gas Volumes:

Amount of A (g or V) → mol of A $\xrightarrow{}$ mol of B → Amount of B (g or V)

↓ stoichiometry

*Example: $6\,NaN_{3\,(s)} + Fe_2O_{3\,(s)} \longrightarrow 3\,Na_2O_{(s)} + 2\,Fe_{(s)} + 9\,N_{2\,(g)}$

How many of sodium azide NaN_3 would be required to provide 75.0L of nitrogen gas at 25°C and 748 mm Hg

$PV = nRT$

$P = \frac{748\,mmHg}{760} = 0.984\,atm$

$T = 25 + 273 = 298\,K$

$V = 75\,L$

$$\frac{0.984 \times 75}{0.0821 \times 298} = n$$

$n = 3.02\,mol$

$6\,NaN_3 \longrightarrow 9\,N_2$

$?\,NaN_3 \longrightarrow 3.02\,N_2$

$2.01\,mol\,NaN_3 \times 65.01 = \boxed{131\,g\,NaN_3}$

* Example: $C_6H_{12}O_6 + 6O_2 \rightarrow \underline{6CO_2} + 6H_2O$

Calculate the volume of CO_2 produced at 37°C and 1 atm when 5.6g of glucose is used up in the reaction?

$n = \text{mol } \underline{C}_6H_{12}O_6 = \frac{5.6\,g}{180\,\frac{g}{mol}} = \underline{0.031\,mol}$

$T = 37 + 273 = 310\,K$

$1\,mol\ C_6H_{12}O_6 \longrightarrow \underline{6mol}\ CO_2$

$\underline{\underline{0.031}} \longrightarrow ?\ CO_2$

$n_{CO_2} =$ ~~R~~ $0.187\,mol$

$V = \frac{nRT}{P} = \frac{0.187 \times 0.0821 \times 310}{1\,atm} = \boxed{4.75\,L}$

$V = \frac{nRT}{P} = \frac{0.031 \times 0.0821 \times 310}{1}$

$= 0.\underline{79\,L}$

$1\,L\ C_6H_{12}O_6 \longrightarrow 6\,L\ CO_2$

$0.79\,L \longrightarrow ??\ CO_2$

$\boxed{V = 4.74\,L\ CO_2}$

* 5.5: Gas Mixtures; Law of Partial Pressures.

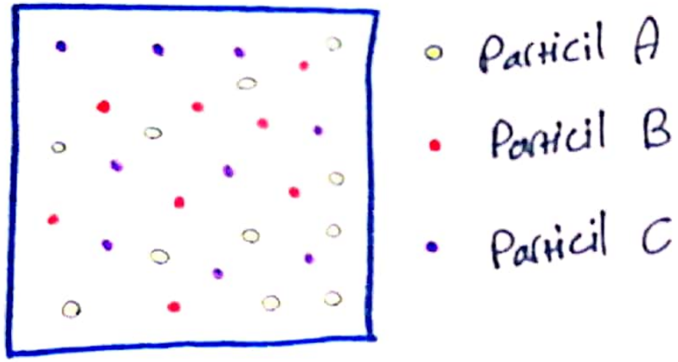
⇒ **Dalton Law:** The sum of Partial pressures of all different gases in a mixture is equal to the total pressure of the mixture.

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

* $P_A = \frac{n_A RT}{V}$, $P_B = \frac{n_B RT}{V}$, …

* mol fraction of A $= \frac{n_A}{n_T} = \frac{P_A}{P_T}$ ⇒ it is the fraction of moles of that component in the total moles of gas mixture.

* Partial pressure: The pressure exerted by a particular gas in a mixture.



* **Example:** 1.00L sample of dry air at 25°c and 786 mmHg contains 0.925g $\underline{\underline{N_2}}$, plus other gases including oxygen, argon, and carbon dioxide.

$\boxed{a}$ what is the partial pressure in mmHg of N_2 in the air sample.

$\boxed{b}$ what is the mole fraction and mole percent of N_2 in the mixture.

$T = 25 + 273 = 298$ $\quad V = 1\ L$

$P = 786\ mmHg$

$m_{N_2} = 0.925g$

$\boxed{a}$ $$\frac{0.925\ g}{28\ g/mol} = 0.0330\ mol\ N_2 = n_{N_2}$$

$$P_{N_2} = \frac{n_{N_2} RT}{V} = \frac{0.033 \times 0.0821 \times 298}{1} = 0.807\ \cancel{atm} \times \frac{760\ mmHg}{1\ \cancel{atm}} = \boxed{613.3\ mmHg}$$

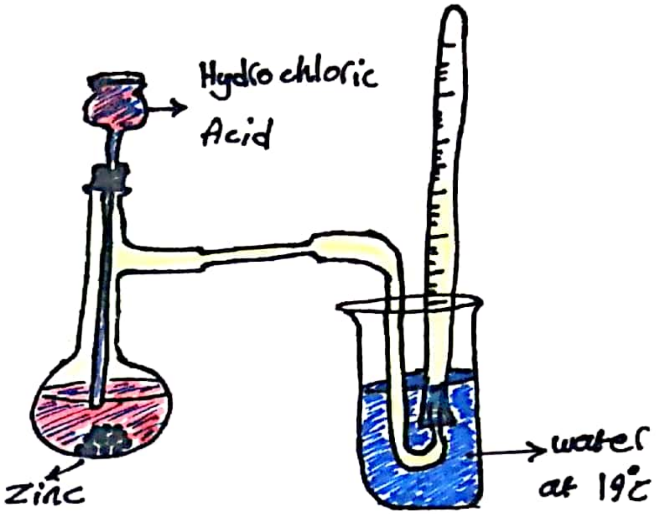
$\boxed{b}$ $$X_{N_2} = \frac{n_{N_2}^{\checkmark}}{\underset{\times}{\underline{\underline{n_{tot}}}}} = \frac{P_{N_2}^{\checkmark}}{P_{tot}^{\checkmark}} = \frac{613.3}{786} = \boxed{0.78}$$

mole percent $= 0.78 \times 100\% = 78$ mole percent
78%

*Collecting gas over water:

$$2HCl + Zn \longrightarrow ZnCl_2 + H_2\uparrow$$

$$P_T = P_{H_2} + \underbrace{P_{H_2O}}_{\text{constant at specific temperature : Table 5.6 / P.165}}$$



- 🟨 Hydrogen gas
- 🟦 water
- 🟥 Hydrochloric Acid
- ● Zinc

Example: The gas collected over water. If 156 mL of gas is collected at 19°C and 769 mmHg total pressure, what is the mass of hydrogen collected?

P_{H_2O} at 19°C = 17 mmHg

$$P_{H_2} = P_T - P_{H_2O} = 769 - 17 = \frac{752\ mmHg}{760} = 0.989\ atm$$

$T = 19 + 273 = 292\ K$

$V = 156\ mL \times 10^{-3} = 0.156\ L$

$n = ?? \longrightarrow m$

$$PV = nRT$$

$$0.989 \times 0.156 = n \times 0.0821 \times 292$$

$$n = 0.00641\ mol$$

$$m = 0.00641\ \cancel{mol} \times 2.02\ \frac{g}{\cancel{mol}} = \boxed{0.013\ g\ H_2}$$

* **Example:** A mixture of gases contain 8.24 mole CH_4, 0.421 mole of C_2H_6 and 0.116 mole of C_3H_6, if the total pressure is $P_{tot} = 1.37$ atm, calculate the partial Pressures of the gases:

$n_{CH_4} = 8.24$ mol

$n_{C_2H_6} = 0.421$ mol

$n_{C_3H_6} = 0.116$ mol

$P_{tot} = 1.37$ atm

$n_{tot} = 8.24 + 0.421 + 0.116 = 8.78$ mol

$$X_A = \frac{n_A}{n_{tot}} = \frac{P_A}{P_{tot}}$$

$$X_{CH_4} = \frac{8.24}{8.78} = 0.939$$

$P_{CH_4} = X_{CH_4} \times P_{tot}$

$= 0.939 \times 1.37 = 1.29$ atm

$$X_{C_2H_6} = \frac{0.421}{8.78} = 0.048$$

$P_{C_2H_6} = 0.048 \times 1.37$

$= 0.066$ atm

$$X_{C_3H_6} = \frac{0.116}{8.78} = 0.013$$

$P_{C_3H_6} = 0.013 \times 1.37$

$= 0.0178$

* 5.6: Kinetic Theory of an Ideal Gas:

1] Gases are composed of molecules whose size is negligible compared with the average distance between them. This means that you can usually ignore the volume occupied by the molecules, And the molecules have the same mas.

2] Molecules move randomly in straight lines in all directions and at various speeds, which mean that the properties of agas depends on the motion of molecules will be the same in all directions.

3] The forces of attraction ~~and~~ or repulsion between two molecules (intermolecular forces) in agas are very weak or negligible, except when they collide. It means that the molecule will continue moving in astraight line with undiminished speed until it collides with another gas molecule or with the walls of the container.

4] when molecules collide with one another, the collisions are elastic. in an elastic collision the total kinetic energy remains constant; it can be transfered between molecules but never is lost.

5] The average kinetic energy of a molecule is proportional to the absolute temperature. The higher the temperature, the greater the molecular kinetic energy. Any gases at the same temperature will have the same (KE).

* The Ideal Gas law:

$P \propto$ frequency of collisions × average force.

$$P \propto \left(u \times \frac{1}{V} \times N\right) \times mu$$

$$VP \propto Nmu^2$$

$PV \propto nT \longrightarrow \frac{PV}{nT}$ ~~∝~~ constant

$$\boxed{PV = nRT}$$

* You must know th Kinetic-Molecular Theory Postulates very hard.

Note

* 5.7: Molecular Speed; Diffusion and Effusion.

* root-mean-square (rms) molecular speed: 'u', is a type of average molecular speed, equal to the speed of a molecule having the average molecular kinetic energy

$$u_{rms} = \sqrt{\frac{3RT}{M_m}} = \left(\frac{3RT}{M.m}\right)^{1/2}$$

→ in this formula $R = 8.314$ Kg.m²/(s².K.mol) based on the units in the formula
$R = 8.314$ J/K.mol

→ M.m in Kg/mol not in g/mol

*** Example:** Calculate the rms speed of O_2 molecules in a cylinder at 21°C and 15.7 atm

T = 21 + 273 = 294 K

P = 15.7

m.m = 16 × 2 = 32 g/mol

= $\frac{32}{1000}$ ~~000~~ Kg/mol

$$u_{rms} = \sqrt{\frac{3 \times 8.314 \times 294}{32}}$$

$$\boxed{u_{rms} = 478.7 \text{ m/s}}$$

* Diffusion and Effusion:

* diffusion: the process whereby a gas spreads out through another gas to occupy the space uniformly.

* Effusion: is the process in which a gas flows through a small hole in a container

* Graham's law of effusion:

$$\text{Rate of effusion of molecules} \propto \frac{1}{\sqrt{M.m}}$$ (for the same container at constant T and P)

$$\to \text{Time} \propto \sqrt{M.m}$$

***Example:** Calculate th ratio of effusion rates of molecules of carbon dioxide, CO_2, and sulfur dioxide, SO_2, from the same container and at the same temperature and pressure.

$$\frac{\text{rate of effusion } CO_2}{\text{'' '' '' } SO_2} = \frac{\frac{1}{\sqrt{m.m_{CO_2}}}}{\frac{1}{\sqrt{m.m_{SO_2}}}} = \sqrt{\frac{mm_{SO_2}}{m.m_{CO_2}}} = \sqrt{\frac{64.1}{44}}$$

$$= 1.21$$

$$\boxed{CO_2} = 1.21\ SO_2$$

$$\frac{\text{time } CO_2}{\text{time } SO_2} = \frac{1}{1.21} = 0.83$$

* Example: Arrange the following gases in order of increasing average molecular speed at 25°C $[He, O_2, CO_2, N_2]$, $[m.w: He = 2, O_2 = 32, CO_2 = 44, N_2 = 28]$

Higher speed → Higher rate → lower M.W

$$\text{rate of effusion} \propto \frac{1}{\sqrt{m.w}}$$

$$He > N_2 > O_2 > CO_2$$

* Example: A sample of "O_2" was found to effuse at a rate equal to two times the rate of an unkown gas. what is the M.W of the unkown gas?

$$\text{rate } O_2 = 2 \text{ rate } A$$

$$\frac{1}{\sqrt{MW_{O_2}}} = 2 \times \frac{1}{\sqrt{m.w_A}}$$

$$\frac{1}{\sqrt{32}} = \frac{2}{\sqrt{M.W_A}}$$

$$\frac{1}{32} = \frac{4}{MW_A} \longrightarrow MW_A = 4 \times 32 = \boxed{128\ g/mol}$$