Chapter 5: Gases and the Kinetic-Molecular Theory

- Properties:
  - pressure changes it's volume
  - Temprature changes it's volume (expand when → heated, shrink when cooled)
  - have low densities
  - form a sol in any proportions

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

Rmk: Atmospheric pressure ↓ with altitude

Units: 1 atm = 760 mmHg = 760 torr = 1.01325×10^5 Pa = 1.01325 bar

Exact Quantities

Rmk: ideal Gas = exhibits linear Relationships among these variables (n, T, V, P)

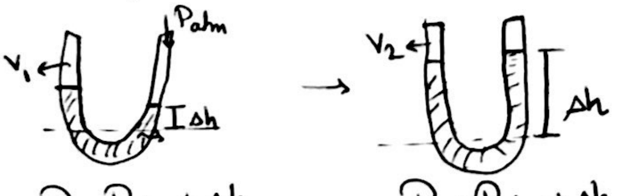
- Gas laws

$\boxed{1}$ Boyles law

$V \propto \frac{1}{P}$ At fixed T and n

P ↑ V ↓

P ↓ V ↑



$P_1 = P_{atm} + \Delta h$ $\quad$ $P_2 = P_{atm} + \Delta h$

$V_2 < V_1$

$\boxed{2}$ Charle's law

$V \propto T$ At fixed P and n

↑ Kelvin

Rmk: • At 0 Kelvin an ideal gas would have zero volume

[3] Avogadro's law

- Equal volumes of any ideal gas contain equal number of particles (or moles) at fixed temperature and pressure

[4] The ideal Gas law

$$\boxed{\underset{\text{pressure}}{P}\,\underset{\text{volume}}{V} = \overset{\text{mole}}{n}\,\underset{\text{constant}}{R}\,\overset{\text{temperature}}{T}}$$

$R = 0.0821\ atm.L/mol.K$

Rmk :-

- Standard temperature and pressure (STP)

$P = 1\ atm$

$T = 0\,°C = 273.15\ K$

- Standard molar volume

Volume of 1 mole at STP

SMV = 22.4 L

Number of particles = 6.022×10^{23}

Relation Between ideal Gas law and other laws :-

[1] n, T fixed ~ $PV = \underset{\text{constant}}{R}$

[2] n, P fixed ~ $V = \underset{\text{constant}}{R}T$

[3] P, T fixed ~ $V = n\underset{\text{constant}}{R}$

- **The ideal Gas law and Gas Density**

$$\text{density} = \frac{\overset{\text{Mass}}{m}}{V} = \frac{\underset{\text{molecular Mass}}{M} \times \overset{\text{pressure}}{P}}{R \times \underset{\text{Temp}}{T}}$$

- **Molar Mass from the ideal Gas law**

$$M = \frac{mRT}{PV}$$

5 Dalton's law

Pressure in a mixture is the sum of partial pressures of the component gases

For gas A:-

$P_A = X_A \times P_{total}$

mole fraction: $X_A = \frac{nA}{n_{total}}$

• A Model for Gas Behavior

- Gas Particles are tiny with large spaces between them
- " " are in
  - Constant
  - random
  - Straight line motion

  except when they colide
- Collisions are elastic:- Total Kinetic Energy is Constant

$\boxed{\text{Temp} \uparrow \rightarrow \text{Speed} \uparrow \rightarrow \text{Kinetic Energy} \uparrow}$

Kinetic Energy = $\frac{1}{2} m V^2$

Speed

- At the same T heavier gas particle moves slowly than a lighter one

- Mol. Mass ↑ molecular speed ↑

Graham's law of Effusion

- rate of Effusion is enversly proportional to the Square root of It's Molar Mass

$$\text{Rate of Effusion} \propto \frac{1}{\sqrt{M}}$$ at the same T

Real Gases : Deviations from ideal Behavior

↳ at High pressure and low temperature

- Back to Fig 5.21, 22, 23

The van der Waals Equation

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

- b : particle Volume
- a : factors that influence the attraction between particles

look at Question :- 5.5

$$\frac{h_{H_2O}}{h_{Hg}} = \frac{d_{Hg}}{d_{H_2O}}$$