

Chapter 5: Gases and the Kinetic-Molecular Theory

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

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

Rmk: Atmospheric pressure $\downarrow$ with altitude

Units: $1 \text{ atm} = 760 \text{ mm Hg} = 760 \text{ torr} = 1.01325 \times 10^5 \text{ Pa} = 1.01325 \text{ bar}$
 Exact Quantities

- Gas laws

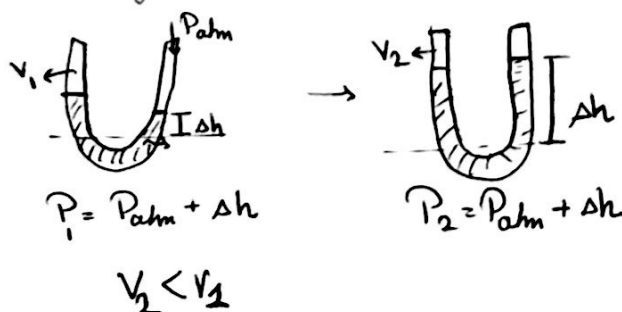
Rmk: ideal Gas = exhibits linear Relationships among these variables (n, T, V, P)

I Boyle's law

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

$$\begin{array}{cc} P \uparrow & V \downarrow \\ P \downarrow & V \uparrow \end{array}$$

At fixed T and n



II Charles's law

$$V \propto T$$

$\uparrow$
Kelvin

At fixed P and n

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

$$PV = nRT$$

Pressure $\rightarrow$ P , Volume $\rightarrow$ V , moles $\rightarrow$ n , temperature $\rightarrow$ T , constant $\rightarrow$ R

$R = 0.0821 \text{ atm} \cdot \text{L} / \text{mol} \cdot \text{K}$

Relation Between ideal Gas law and other laws:-

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

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

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

Rank 1:-

- Standard temperature and pressure (STP)

$$P = 1 \text{ atm}$$

$$T = 0^\circ \text{C} = 273.15 \text{ K}$$

- Standard molar Volume

Volume of 1 mole at STP

$$SMV = 22.4 \text{ L}$$

$$\text{Number of particles} = 6.022 \times 10^{23}$$

• The ideal Gas law and Gas Density

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

• 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:-

$$P_A = X_A \times P_{\text{total}}$$

↖
mole fraction: $X_A = \frac{n_A}{n_{\text{total}}}$

• A Model for Gas Behavior

- Gas Particles are tiny with large spaces between them
- ~ ~ are in • Constant except when they collide
 - random
 - Straight line motion

- Collisions are elastic:- Total Kinetic Energy is Constant

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

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

↑
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 inversely 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 1: 5.5

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