4. States of Matter

Boyle's Law at constant temperature and amount

$P_1V_1 = P_2V_2$ = Constant

Charle's Law

V = kT at constant pressure

k is the proportionality constant depends upon (i) Amount of Gas (ii) Temperature

Gay Lussac's Law :

$\frac{P_1}{T_1} = \frac{P_2}{T_2}$ at constant volume

Avogadro's Law

V ∝ n (T and P constant)

$V = K_4 n$

$V_1/n_1 = V_2/n_2$ (Constant T and P)

Ideal Gas Equation

PV = nRT

Where R is Proportionality constant is also known as Gas constant it is same for all Gases

Value of R in different units

Magnitude	Unit
0.0821	Litre-atm K^{-1} mol^{-1}
82.1	ML-atm K^{-1} mol^{-1}
62.1	Litre-mm-Hg K^{-1} mol^{-1}
0.083	Litre bar K^{-1} mol^{-1}
8.314	Pascal m^3 K^{-1} mol^{-1}
8.314×10^7	erg K^{-1} mol^{-1}
8.314	Joule K^{-1} mol^{-1}
1.987	Cal K^{-1} mol^{-1}

Density ; d = PM/RT

(d ∝ P) ; (d ∝ 1/T)

Graham's Law of Diffusion / Effusion

- Rate of diffusion $R \propto \frac{1}{\sqrt{d}}$

  where d is density of gas at constant temperature and pressure

- $\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}$

- $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$

Dalton's of Partial Pressure :

Calculate the total pressure of mixture of non-reacting gas and based on the law of conservation of amount

$P_1 = P_T \times x_1$ (where P_1 is a partial pressure, P_T is a Total pressure, x_1 is mole fraction)

Total pressure of Gaseous mixture at constant temperature : $P_T = \frac{(P_1V_1 + P_2V_2)}{(V_1 + V_2)}$

Aqueous Tension : $P_{moist} = P_{dry\ gas} + P_{water\ vapours}$

$$RH = \frac{\text{Mass of water vapour present in certain volume of air}}{\text{Maximum Mass of water vapour present in same volume of air saturated by water vapour}}$$

Molecular Speed

Most probable speed = $\sqrt{\frac{2KT}{m}} = \sqrt{\frac{2RT}{M}}$

Average speed = $\sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8KT}{m}}$

Root mean square = $\sqrt{\frac{3RT}{M}} = \sqrt{\frac{3KT}{m}}$

$V_{mp} : V_{av} : V_{rms} = \sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3}$

Kinetic Energy

Average Kinetic Energy = $\frac{3}{2}KT$

Total Kinetic Energy = $\frac{3}{2}nRT$

Compressibility Factor (Z) $\left(Z = \frac{PV}{nRT} = \frac{PV_m}{RT}\right)$

Ideal gases

- zero volume
- zero attractive force
- $PV = nRT$
- $Z = 1$

Real gases

- corrected equation $(P + an^2/V^2)(V - nb) = nRT$
- non-zero volume
- some intermolecular force

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- $Z > 1$
- Repulsive forces
- Difficult to compress
- Difficult to Liquify

(–)ve

- $Z < 1$
- Attractive forces
- Easy to compress
- Easy to Liquify

Energy-Distance for different ion-Covalent Interaction

Types of Interaction	Energy-distance Functions
Ionic bond	$1/r$
Ion-dipole	$1/r^2$
Dipole-dipole	Stationary molecules - $1/r^3$
	Rotating molecules - $1/r^6$
Ion-induced dipole	$1/r^4$
Dipole-induced dipole	$1/r^6$
London forces	$1/r^6$

Critical Constant of the Gases

T_c or critical temp : **$T_c = 8a / 27Rb$**

P_c or critical pressure : **$P_C = a/ 27b^2$**

V_C or critical volume : **$V_C = 3b$**

$Z_C = \frac{P_C V_C}{RT_C} = \frac{3}{8}$ (For all real gases)

Van der Waal's Equation Real Gas

$\left(P + \frac{an^2}{V^2}\right)(V - nb) = nRT$ where a and b are Van der Waal's constant.

Boyle's Temperature : $T_b = \frac{a}{Rb}$