# Gaseous State

### * Pressure at mercury column

(i) $P_{atm} = P_{gas}$ if both are on the same level/height

(ii) if, $P_{atm} > P_{gas}$, then

$$\boxed{P_{atm} = P_{gas} + h}$$ , h is the difference b/w heights in tube tubes

(iii) $P_{atm} < P_{gas}$ then →

$$\boxed{P_{atm} = P_{gas} - h}$$

## Units of Pressure

1 atm = 760 torr = 760 mmHg = 76 cm Hg

= $1.013 \times 10^5\ N/m^2$

= $1.013 \times 10^5$ Pa

1 bar = $10^5$ Pa = 0.9869 atm = 750.062 torr

1 atm = 1.013 bar , 1 atm = 14.7 psi

## Gas Laws

(i) Boyle's law (Relation b/w V and P)

⇒ At a constant temperature, volume $\propto \frac{1}{\text{pressure}}$

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

$\Rightarrow P = \frac{K}{V}$ $\Rightarrow PV = K$ $\boxed{\Rightarrow P_1V_1 = P_2V_2}$

Also, in terms of Density → $D_1, D_2$

$$\boxed{\frac{P_1}{P_2} = \frac{D_1}{D_2}}$$

(ii) Charles law (Relation b/w V and T)

$\boxed{V \propto T} \Rightarrow V = KT \Rightarrow \boxed{\frac{V_1}{V_2} = \frac{T_1}{T_2}}$

also in terms density, $V \propto \frac{1}{d}$

$\therefore \boxed{\frac{d_1}{d_2} = \frac{T_2}{T_1}}$

(iii) Gay-Lussac's law of Pressure :- (Relation b/w P and T)

At constant volume, $\boxed{P \propto T}$

$\boxed{\frac{P_1}{P_2} = \frac{T_1}{T_2}}$

(iv) Avogadro's law : (Relation b/w V and n)

At constant temperature and pressure,

$\boxed{V \propto n} \Rightarrow \frac{V_1}{V_2} = \frac{n_1}{n_2}$

(*) Combination of Boyle's and Charles' law :-

$\boxed{\frac{PV}{T} = K} \Rightarrow \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$

## (*) Equation of state for an ideal gas :→

Formula → $\boxed{PV = nRT}$ , $\boxed{PM = dRT}$

Dimension of R :- work (or energy) per kelvin per mol of gas.

Boltzmann Constant

$$K = 1.38 \times 10^{-23}\ J/K \ (\text{in SI})$$
$$= 1.38 \times 10^{-16}\ erg/K \ (\text{in CGS})$$

Values of K = $8.314\ J/(mol\ K)$
$= 2\ cal/(mol\ K)$
$= 0.0821\ L\ atm\ (mol\ K)$.

## (*) Graham's law of diffusion/effusion

$r \propto \frac{1}{\sqrt{d}}$   r = rate , d = density

(i) $\frac{r_A}{r_B} = \sqrt{\frac{d_B}{d_A}}$ , Also, vapour density (D) = $\frac{\text{molar Mass (M)}}{2}$

(ii) $\frac{r_A}{r_B} = \sqrt{\frac{M_B}{M_A}}$   (iii) $\frac{V_A}{V_B} = \sqrt{\frac{M_B}{M_A}}$   (iv) time, $\frac{t_A}{t_B} = \sqrt{\frac{M_A}{M_B}}$

## Kinetic Theory of Gases

Kinetic Gas equn :-

$$P = \frac{1}{3}\frac{mnC_{rms}^2}{V}$$

Average velocity and mean square velocity

(i) Average velocity :→

$$\bar{C} = \frac{n_1C_1 + n_2C_2 + \dots}{n} \quad \text{or,}\ \bar{C} = \sqrt{\frac{8RT}{\pi m}}$$

(ii) Mean square velocity :-

$$\bar{C}^2 = \frac{n_1C_1^2 + n_2C_2^2 + \dots}{n}$$

## Root Mean Square velocity

$$C_{rms} = \sqrt{\overline{c^2}} = \sqrt{\frac{n_1c_1^2 + n_2c_2^2 + \cdots}{n}}$$

$$C_{rms} = \sqrt{\frac{3RT}{M}}$$

## ⊛ Most probable velocity

$$C_m = \sqrt{\frac{2RT}{m}}$$

$$\therefore \quad C_m : \bar{c} : C_{rms}$$

$$= \sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3}$$

$$= 1 : 1.128 : 1.224$$

## ⊛ Avg K.E of gas molecule

$$= \frac{1}{2} m c_{rms}^2$$

## Compressibility factor of real gas

$$Z = \frac{PV}{nRT}$$

For ideal gas, $PV = nRT$, $Z = 1$

For real gas, $PV \neq nRT$, $Z \neq 1$.

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

for $n = 1$, $$\left(P + \frac{a}{V^2}\right)(V - b) = RT.$$

At low pressure

⊛ $(V - b) \simeq V$

At high pressure

⊛ $\left(P + \frac{a}{V^2}\right) \simeq P$

## Relation regarding critical constants :→

$$\frac{P_c V_c}{T_c} = \frac{3R}{8}$$