Kinetic Theory of Gases & Radiation

- **Kinetic Theory of Matter:-**
  - **(a) Solids:-** It is the type of matter which has got fixed shape and volume. The force of attraction between any two molecules of a solid is very large.
  - **(b) Liquids:-** It is the type of matter which has got fixed volume but no fixed shape. Force of attraction between any two molecules is not that large as in case of solids.
  - **(c) Gases:-** It is the type of matter does not have any fixed shape or any fixed volume.
- **Ideal Gas:-** A ideal gas is one which has a zero size of molecule and zero force of interaction between its molecules.
- **Ideal Gas Equation:-** A relation between the pressure, volume and temperature of an ideal gas is called ideal gas equation.

  PV/T = Constant or $PV = nRT$

  Here, n is the number of moles and R is the universal gas constant.
- **Gas Constant:-**
  - **(a) Universal gas constant (R):-**

    $R = P_0 V_0/T_0$

    $= 8.311\ J\ mol^{-1}K^{-1}$
  - **(b) Specific gas constant (r):-**

    $PV = (R/M)\ T = rT$,

    Here, $r = R/M$
- **Real Gas:**-The gases which show deviation from the ideal gas behavior are called real gas.
- **Vander wall's equation of state for a real gas:-**

  $[P + (na/V)^2?][V - nb] = nRT$

  Here n is the number of moles of gas.
- **Avogadro's number (N):-** Avogadro's number (N), is the number of carbon atoms contained in 12 gram of carbon-12.

  $N = 6.023\times10^{23}$
  - **(a) To calculate the mass of an atom/molecule:-**

    Mass of one atom = atomic weight (in gram)/N

    Mass of one molecule = molecular weight (in gram)/N
  - **(b) To calculate the number of atoms/molecules in a certain amount of substance:-**

    Number of atoms in m *gram* = (N/atomic weight) × m

    Number of molecules in m *gram* = (N/molecular weight) × m

(c) Size of an atom:-

Volume of the atom, $V = (4/3)\pi r^3$

Mass of the atom, $m = A/N$

Here, A is the atomic weight and N is the Avogadro's number.

Radius, $r = [3A/4\pi N\rho]^{1/3\backslash}$

Here ρ is the density.

- **Gas laws:-**

(a) Boyle's law:- It states that the volume of a given amount of gas varies inversely as its pressure, provided its temperature is kept constant.

PV = Constant

(b) Charlers law or Gey Lussac's law:- It states that volume of a given mass of a gas varies directly as its absolute temperature, provided its pressure is kept constant.

V/T= Constant

$V - V_0/V_0 t = 1/273 = \gamma_p$

Here γ_p (=1/273) is called volume coefficient of gas at constant pressure.

Volume coefficient of a gas, at constant pressure, is defined as the change in volume per unit volume per degree centigrade rise of temperature.

(c) Gay Lussac's law of pressure:- It states that pressure of a given mass of a gas varies directly as its absolute temperature provided the volume of the gas is kept constant.

$P/T = P_0/T_0$ or $P - P_0/P_0 t = 1/273 = \gamma_p$

Here γ_p (=1/273) is called pressure coefficient of the gas at constant volume.

Pressure coefficient of a gas, at constant volume, is defined as the change in pressure per unit pressure per degree centigrade rise of temperature.

(d) Dalton's law of partial pressures:-

Partial pressure of a gas or of saturated vapors is the pressure which it would exert if contained alone in the entire confined given space.

$P = p_1 + p_2 + p_3 + \ldots\ldots$

$nRT/V = p_1 + p_2 + p_3 + \ldots\ldots$

(e) Grahm's law of diffusion:- Grahm's law of diffusion states that the rate of diffusion of gases varies inversely as the square root of the density of gases.

$R \propto 1/\sqrt{\rho}$ or $R_1/R_2 = \sqrt{\rho_2/\rho_1}$

So, a lighter gas gets diffused quickly.

(f) Avogadro's law:- It states that under similar conditions of pressure and temperature, equal volume of all gases contain equal number of molecules.

For m gram of gas, $PV/T = nR = (m/M)\, R$

- **Pressure of a gas (P):-** $P = 1/3\,(M/V)\,C^2 = 1/3\,(\rho)\,C^2$

- **Root mean square (r.m.s) velocity of the gas:-** Root mean square velocity of a gas is the square root of the mean of the squares of the velocities of individual molecules.

$C = \sqrt{[c_1^2 + c_2^2 + c_3^2 + \ldots + c_n^2]/n} = \sqrt{3P/\rho}$

- **Pressure in terms of kinetic energy per unit volume:-** The pressure of a gas is equal to two-third of kinetic energy per unit volume of the gas.

  P= 2/3 *E*

- **Kinetic interpretation of temperature:-** Root mean square velocity of the molecules of a gas is proportional to the square root of its absolute temperature.

  $C = \sqrt{3RT/M}$

  Root mean square velocity of the molecules of a gas is proportional to the square root of its absolute temperature.

  At, T=0, C=0

  Thus, absolute zero is the temperature at which all molecular motion ceases.

- **Kinetic energy per mole of gas:-**

  K.E. per *gram mol* of gas = $\frac{1}{2} MC^2 = 3/2\ RT$

- **Kinetic energy per gram of gas:-**

  $\frac{1}{2} C^2 = 3/2\ rt$

  Here, $\frac{1}{2} C^2$ = kinetic energy per gram of the gas and r = gas constant for one gram of gas.

- **Kinetic energy per molecule of the gas:-**
  Kinetic energy per molecule = $\frac{1}{2} mC^2 = 3/2\ kT$
  Here, *k* (Boltzmann constant) = *R*/*N*
  Thus, K.E per molecule is independent of the mass of molecule. It only depends upon the absolute temperature of the gas.
- **Regnault's law:-** $P \propto T$
- **Graham's law of diffusion:-**

  $R_1/R_2 = C_1/C_2 = \sqrt{\rho_2/\rho_1}$

- **Distribution of molecular speeds:-**
  (a) Number of molecules of gas possessing velocities between *v* and *v+dv* :-

  $$n_v dv = \frac{\sqrt{2}\pi n m^{3/2} v^2}{(\pi k\mathrm{T})^{3/2}} e^{-mv^2/2kT} dv$$

  (b) Number of molecules of gas possessing energy between u and u+dv:-

  $$n_u du = \frac{2\pi n}{(\pi k\mathrm{T})^{3/2}} \sqrt{u} e^{\frac{-u}{k\mathrm{T}}} du$$

  (c) Number of molecules of gas possessing momentum between *p* and *p+dp* :-

  $$n_p dp = \frac{\sqrt{2}\pi n}{(\pi m k\mathrm{T})^{3/2}} p^2 e^{\frac{-p}{2mk\mathrm{T}}} dp$$

(d) **Most probable speed:-** It is the speed, possessed by the maximum number of molecules of a gas contained in an enclosure.

$V_m = \sqrt{[2kT/m]}$

(e) **Average speed (V_{av}):-** Average speed of the molecules of a gas is the arithmetic mean so the speeds of all the molecules.

$V_{av} = \sqrt{[8kT/\pi m]}$

(f) **Root mean square speed (V_{rms}):-** It is the square root of the mean of the squares of the individual speeds of the molecules of a gas.

$V_{rms} = \sqrt{[3kT/m]}$

- **$V_{rms} > V_{av} > V_m$**
- **Degree of Freedom (n):-** Degree of freedom, of a mechanical system, is defined as the number of possible independent ways, in which the position and configuration of the system may change.
  In general, if N is the number of particles, not connected to each other, the degrees of freedom n of such a system will be, $n = 3N$
  If K is the number of constraints (restrictions), degree of freedom n of the system will be, $n=3N–K$
- **Degree of freedom of a gas molecule:-**

  (a) **Mono-atomic gas:- Degree of freedom of monoatomic molecule, $n = 3$**

  (b) **Di-atomic gas:-**

  At very low temperature (0-250 K):- Degree of freedom, $n = 3$

  At medium temperature (250 K – 750 K):- Degree of freedom, $n = 5$ (Translational = 3, Rotational = 2)

  At high temperature (Beyond 750 K):- Degree of freedom, $n = 6$ (Translational = 3, Rotational = 2, Vibratory =1), For calculation purposes, $n = 7$
- **Law of equipartition of energy:-** In any dynamical system, in thermal equilibrium, the total energy is divided equally among all the degrees of freedom and energy per molecule per degree of freedom is ½ kT.

  $E = ½\, kT$
- **Mean Energy:-** K.E of one mole of gas is known as mean energy or internal energy of the gas and is denoted by U.

  $U = n/2\ RT$ (Here n is the degree of freedom of the gas.)

  (a) Mono-atomic gas($n = 3$):- $U = 3/2\ RT$

  (b) Diatomic gas:-

  At low temperature (n=3):- $U = 3/2\ RT$

  At medium temperature (n=5):- $U = 5/2\ RT$

  At high temperature (n=7):- $U = 7/2\ RT$
- **Relation between ratio of specific heat capacities (γ) and degree of freedom (n):-**

  $\gamma = C_p/C_v = [1+(2/n)]$

  (a) For mono-atomic gas (n=3):- $\gamma = [1+(2/n)] = 1+(2/3) = 5/3=1.67$

  (b) For diatomic gas (at medium temperatures (n=5)):- $\gamma = [1+(2/5)] = 1+(2/5) = 7/5=1.4$

  (c) For diatomic gas (at high temperatures (n=7)):- $\gamma = [1+(2/7)] = 9/7 = 1.29$