

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.

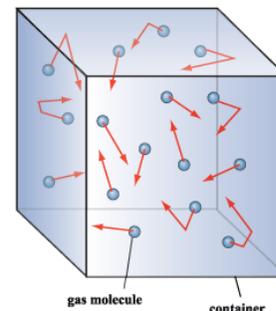
(c) Gases:- It is the type of matter does not have any fixed shape or any fixed volume.

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.

$$PV/T = \text{Constant} \text{ or } PV = nRT$$

Here, n is the number of moles and R is the universal gas constant.



Gas Constant:-

(a) Universal gas constant (R):-

$$R = P_0 V_0 / T_0$$

$$= 8.311 \text{ J mol}^{-1} \text{ K}^{-1}$$

(b) Specific gas constant (r):-

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

$$\text{Here, } r = R/M$$

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.

$$N = 6.023 \times 10^{23}$$

(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 = $(N/\text{atomic weight}) \times m$

Number of molecules in m gram = $(N/\text{molecular weight}) \times m$

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

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.

$$PV = \text{Constant}$$

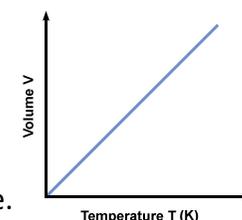
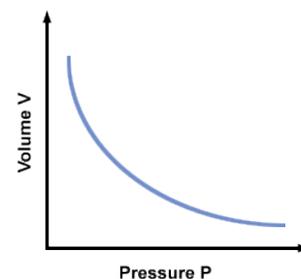
(b) Charlers law or Gey Lussac's law:- It states that volume of a given mass of a gas varies directly as its absolute temperature, provided its pressure is kept constant.

$$V/T = \text{Constant}$$

$$V - V_0/V_0t = 1/273 = \gamma_p$$

Here $\gamma_p (=1/273)$ 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.

$$P/T = P_0/T_0 \text{ or } P - P_0/P_0t = 1/273 = \gamma_p$$

Here $\gamma_p (=1/273)$ 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.

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

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

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

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

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.

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

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.

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

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.

$$\text{At, } T=0, C=0$$

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

Kinetic energy per mole of gas:-

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

Kinetic energy per gram of gas:-

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

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:-

$$\text{Kinetic energy per molecule} = \frac{1}{2} mC^2 = \frac{3}{2} kT$$

$$\text{Here, } k \text{ (Boltzmann constant)} = R/N$$

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 $v+dv$

$$n_v dv = \frac{\sqrt{2\pi nm^{3/2}} v^2}{(\pi kT)^{3/2}} e^{-mv^2/2kT} dv$$

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

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

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

$$n_p dp = \frac{\sqrt{2\pi n}}{(\pi mkT)^{3/2}} p^2 e^{-\frac{p^2}{2mkT}} dp$$

(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 $\frac{1}{2} kT$.

$$E = \frac{1}{2} kT$$

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 .

$$U = \frac{n}{2} RT$$

Here n is the degree of freedom of the gas.

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

(b) Diatomic gas:-

At low temperature ($n=3$):- $U = \frac{3}{2} RT$

At medium temperature ($n=5$):- $U = \frac{5}{2} RT$

At high temperature ($n=7$):- $U = \frac{7}{2} RT$

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 ($n=7$)): - $\gamma = [1+(2/7)] = 9/7 = 1.29$