Kinetic Theory of Gases

Kinetic Theory of Gases: The kinetic theory of gases (also known as kinetic-molecular theory) explains the behavior of a hypothetical ideal gas. According to this theory, gases are made up of tiny particles in random, straight line motion. They move rapidly and continuously and make collisions with each other and the walls. This was the first theory to describe gas pressure in terms of collisions with the walls of the container, rather than from static forces that push the molecules apart. Kinetic theory also explains how the different sizes of the particles in a gas can give them different, individual speeds.

Postulates of Kinetic Theory of Gases:

1 The molecules in a gas are small and very far apart. Most of the volume which a gas occupies is empty space.

2 Gas molecules are in constant random motion. Just as many molecules are moving in one direction as in any other.

3 Molecules can collide with each other and with the walls of the container. Collisions with the walls account for the pressure of the gas.

4 When collisions occur, the molecules lose no kinetic energy; that is, the collisions are said to be perfectly elastic. The total kinetic energy of all the molecules remains constant unless there is some outside interference with the molecules.

5 The molecules exert no attractive or repulsive forces on one another except during the process of collision. Between collisions, they move in straight lines.

Ideal Gas Equation:

At low densities there is a predictable relationship between a gas pressure (P), volume (v) and temperature (T) as follows:

For a given mass of gas the volume is inversely proportional to the pressure provided the temperature remains constant. That is,

$$P \propto \frac{1}{V} \text{ ; where, T is constant}$$

$$\text{Or, } PV = \text{Constant}$$

$$\text{Thus, } P_1V_1 = P_2V_2 \qquad (1)$$

This is called Boyles law.

Again, for a given mass the volume is directly proportional to the absolute temperature provided the pressure remains constant. That is,

$$V \propto T \quad \text{; where, P is constant}$$

$$\text{or } \frac{V}{T} = \text{Constant}$$

$$\text{Thus, } \frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad (2)$$

This is called Charles law.

Now combining Boyles and Charles law we get,

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

$$\text{or } V = K\frac{T}{P}$$

$$\text{or } PV=KT \qquad (3)$$

Here K is a constant.

Now for $T_1, T_2, T_3 \ldots\ldots T_n$ absolute temperature, $P_1, P_2, P_3 \ldots\ldots P_n$ pressure and $V_1, V_2, V_3 \ldots\ldots V_n$ volume of the gas then according to equation (3), we can write,

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = \frac{P_3V_3}{T_3} = \ldots\ldots\ldots\ldots\ldots\ldots\ldots\ldots = \frac{P_nV_n}{T_n} = \text{Constant}$$

For 1 mole gas we can write equation (3) as

$$PV=RT \qquad (4)$$

Where, R is the molar gas constant.

Now for a given mass of m, volume V and atomic mass of the gas M, one gram of molecule gas volume is $\frac{M}{m}$ V. So that we can replace by V by $\frac{M}{m}$ V in equation 4, we get,

$$P \times \frac{M}{m}V=RT$$

$$\text{or, } PV=\frac{m}{M}RT$$

$$\text{or, } PV= nRT \quad [\because \frac{m}{M} = n, \text{ number of molecule}]$$

This is the ideal gas equation.

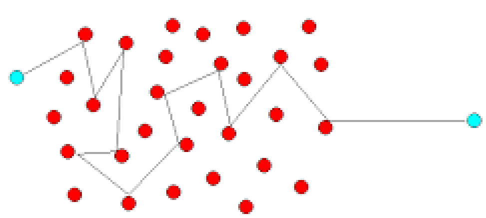
Example: If R=8.31JK^{-1}mol^{-1}, Mercury pressure is 72cm and temperature is 27°C then determine the volume of 20g of Oxygen.

Solution:

We know,	Given that,
$PV = \frac{m}{M}RT$	m=20g=20$\times 10^{-3}$ kg
=>$V = \frac{mRT}{PM}$	M=32$\times 10^{-3}$ kgmol^{-1}
=>$V = \frac{20\times10^{-3}\times8.31\times300}{72\times10^{-2}\times13.6\times10^{3}\times9.8\times32\times10^{-3}}$	R=8.31 $JK^{-1}mol^{-1}$
=.0162369 m^3	T=(27+273)=300K
=16.24$\times 10^{-3}$ m^3	h=72cm=72$\times 10^{-2}$ m
	p=hρg=$72 \times 10^{-2} \times 13.6 \times 10^{3} \times 9.8 \times 32 \times 10^{-3}$

Mean free Path: According to kinetic theory, the molecules are continuously colliding with each other. The direction of the molecule is changes after every collision, but between any two consecutive collisions, the molecules move with constant speed along a straight line. This distance between any two consecutive collision is known as free path. After a number of collisions, the total path appears to be zig-zag and the free path is not constant between any two consecutive collisions. The average distance travelled by a molecule between two collisions is called the mean free path ($\boldsymbol{\lambda}$). If S is the total distance travelled during N collisions, then

$$\boldsymbol{\lambda = \frac{S}{N}}$$



Free Path of Gas

Reversible Process: A process which can be retraced in the opposite direction so that the working substance passes through exactly the same states in all respects as in the direct process is called a reversible process.

Examples of reversible reactions include dissolving, evaporation, melting and freezing. Freezing is a reversible change. For example you can freeze juice to make ice lollies. The lollies can be changed back into juice by heating.

Conditions of Reversibility:

i) The changes in pressure and volume of the working substance must take place at an extremely slow rate, so that the substance undergoing a reversible change is, at all instants, in thermodynamic equilibrium with its surroundings.

ii) There should be no friction and no loss of heat by conduction, convection or radiation.

Irreversible Process: A process which cannot be retraced in the opposite direction is called irreversible process. In an irreversible process the system with the surroundings can never be completely restored to its initial condition by reversing the controlling factors. All changes which occur suddenly like explosions etc. may be considered as irreversible. Heat produced by friction, heat generated when a current flows through an electrical resistance, sudden unbalanced expansion are some examples of irreversible process. It is to be noted, however, that all natural processes are irreversible.