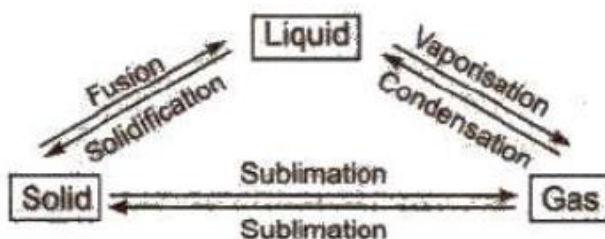


Class 11, Chapter- 5: States of Matter

1. Interconversion of States of Matter:



2. Factors Deciding Physical State of a Substance:

For gaseous state, Forces of attraction \ll thermal energy

For liquid state, Forces of attraction $>$ Thermal energy

For solid state, Forces of attraction \gg Thermal energy

3. Intermolecular Forces (Vanderwaals forces): The forces of attraction existing among the molecules of a substance (gaseous, liquid or solid) are called intermolecular forces [Greater the intermolecular forces, higher is the melting and boiling point.

Attractive intermolecular forces are known as van der Waals' forces.]

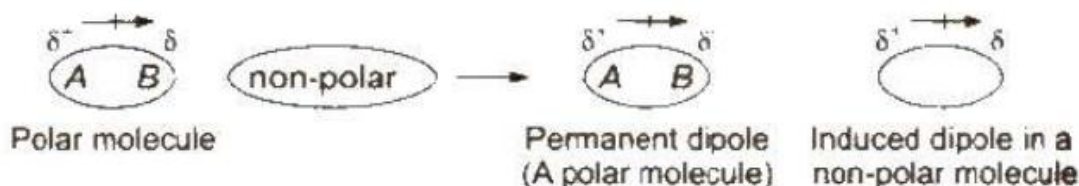
The different types of intermolecular forces are briefly explained below

(i) **Dipole-dipole interactions** Dipole-dipole forces act between the molecules possessing permanent dipoles.



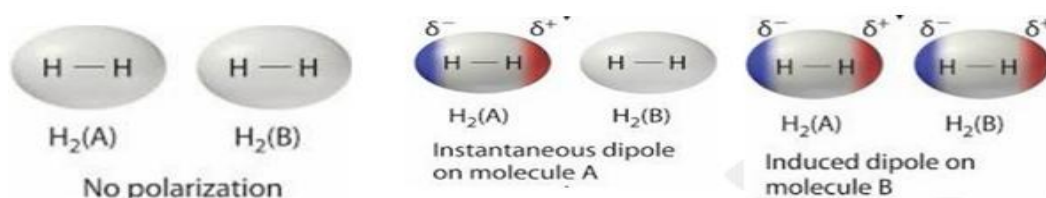
The interaction is stronger than London forces and weaker than ion-ion interaction.

(ii) **Dipole-induced dipole forces** Dipole-induced dipole forces act between the polar molecules having permanent dipole and the molecules lacking permanent dipole



(iii) **Dispersion forces or London forces** Dispersion forces or London forces are present among non-polar atoms and molecules, e.g., among the atoms or chlorine molecules. These are the weakest intermolecular forces. These forces increases with

- (i) Increase in number of electrons in molecules,
 (ii) Increase in molecular size



4. The Gaseous State: This is the simplest state of matter. Characteristics of this state of matter are as following

- In gases, the intermolecular forces are weakest.
- Gases are highly compressible.
- Gases exert pressure equally in all directions.
- Gases have much lower density than the solids and liquids.
- The volume and the shape of gases are not fixed.
- Gases mix evenly and completely in all proportions without any mechanical aid.

5. Diffusion: Without pressure, mixing of unreacted gas.

6. Effusion: High pressure to low pressure, mixing of unreacted gas. Example is removed of air from ballon.

7. Rate of Diffusion (Effusion):

Rate (r) = volume of diffused gas (v) / time (t) or

Rate (r) = mole of diffused gas (n) / time (t) or

Rate (r) = distance of diffused gas (l) / time (t)

At constant temp and pressure,

$r \propto 1 / d$ (Here d = vapor density or density of gas)

$r \propto 1 / M$ (Here M = atomic or molar mass of gas)

At constant temp ,

$r \propto P / d$ (Here P = pressure of gas)

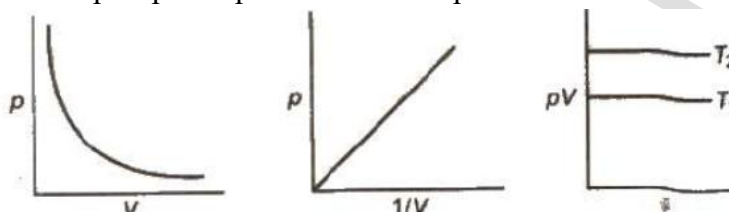
$r \propto P / M$

8. Gas Laws: The quantitative relationship b/w any two of the variables (V, n, P, T) when other two are constant.

8.1. Boyle's Law (Pressure – Volume Relationship): At constant temperature, the pressure of a fixed amount (i.e., number of moles n) of gas inversely proportional to its volume. This is known as Boyle's law.

$$p \propto \frac{1}{V} \text{ (at constant } T \text{ \& } V) \quad \text{i.e. } p = K_1 \frac{1}{V} \quad pV = K_1 \quad p_1V_1 = p_2V_2$$

Isotherms Graphs of V vs p or pV vs p at constant temperature known as Isotherms.

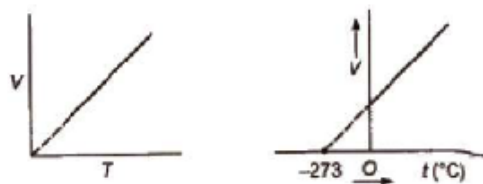


8.2. Charles Law (Temperature – Volume Relationship): At constant pressure, the volume of a fixed mass of a gas is directly proportional to its absolute temperature.

$$\text{i.e. } V \propto T \text{ (at constant } P \text{ \& } n)$$

$$V = K_2 T \quad \frac{V}{T} = K_2 \quad \boxed{\frac{V_1}{T_1} = \frac{V_2}{T_2}}$$

Isobars A graph of V vs T at constant pressure is known as isobar



8.3. Gay Lussac's Law (Pressure-Temperature Relationship): It states that at constant volume, pressure of a fixed amount of a gas is directly proportional to its temperature.

$$P \propto T \text{ (at constant } V \text{ \& } n)$$

$$p = K_3 T \quad \boxed{\frac{P}{T} = K_3}$$

8.4. Avogadro Law (Volume – Amount Relationship): It states that equal volumes of all gases under the same conditions of temperature and pressure contain equal number of molecules.

$$V \propto n \text{ i.e. } \boxed{V = K_4 n}$$

Since volume of a gas is directly proportional to the number of moles; one mole of each gas at standard temperature and pressure (STP)* will have same volume. STP means 273.15 K (0°C) temperature and 1 bar (i.e., exactly 105 pascal) pressure.

9. Ideal gas: A gas that follows strictly, Boyle law, Charle's law and Avogadro law is called Ideal gas. Nearly ideal gases are N₂ and O₂. The **Ideal Gas Equation is** $pV = nRT$
Where R is **Universal Gas Constant**, is same for all gases.

R=

$$8.314 \text{ Pa m}^3 \text{ K}^{-1} \text{ mol}^{-1} = 8.314 \times 10^{-2} \text{ bar L K}^{-1} \text{ mol}^{-1} = 8.314 \text{ J K}^{-1} \text{ mol}^{-1}$$

1 litre = 1 dm³

Equation of state:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

10. Density and Molar Mass (molecular mass) of a Gaseous Substance: $pV = nRT$

Then $pV = m/M[RT]$; $pM = m/V[RT]$; $pM = dRT$

$$\therefore M = \frac{dRT}{P}$$

11. Dalton's Law of Partial Pressures: At constant temperature and volume, the total pressure exerted by the mixture of non-reactive gases is equal to the sum of the partial pressures of individual gases

$$P_{\text{Total}} = p_1 + p_2 + p_3 + \dots \text{ (at constant } T, V)$$

Generally gases having moisture so pressure of a dry gas can be determined mathematically as

$$\text{Pressure of dry gas} = P_{\text{Total}} (\text{pressure of moist gas}) - \text{Aqueous tension.}$$

Where aqueous tension is a pressure exerted by water vapours at a particular temperature. It depends upon temperature.

12. Relation b/w partial pressure and mole fraction: According to ideal gas equation-

$PV = nRT$, Where n (no. of mole of Gas) = mass of gas / molecular mass of Gas

For mixture of three gases, partial pressure are $P_1 = n_1RT/V$, $P_2 = n_2RT/V$, $P_3 = n_3RT/V$

$$\frac{P_1}{P_{\text{total}}} = \left(\frac{n_1}{n_1 + n_2 + n_3} \right) \frac{RTV}{RTV}$$

$$= \frac{n_1}{n_1 + n_2 + n_3} = \frac{n_1}{n} = x_1$$

where $n = n_1 + n_2 + n_3$

Where x_1 = mole fraction of first gas

So partial pressure and mole fraction of first gas = $P_1 = x_1 * P_{\text{total}}$

13. KINETIC MOLECULAR THEORY OF GASES: Gases consist main assumptions of this theory are:

1. A gas consists of large number of small particles, called molecules.
2. Volume occupied by gas molecules, is negligible as compared to the total volume of the gas.
3. There is continuous rapid random motion of gas molecules so the molecules collide with each other and with the walls of container.
4. The molecules are perfect elastic bodies and there is no loss of kinetic energy during collisions.
5. There are no attractive forces between the gaseous molecules.
6. The pressure is exerted by a gas as a result, the bombardment of gas molecules against the walls of the container.
7. The different molecules possess different velocities and hence, different kinetic energies. The average KE is directly proportional to absolute temperature.

14. Real Gases: The gases which deviate from Ideal behavior at higher pressure and low temperature due to increases force of attraction b/w molecules are called Real Gases.

Cause of deviation from the ideal behaviour: It is due to two faulty assumptions of kinetic theory of gases, particularly which are not valid at high pressure and low temperature

1. Volume occupied by the gas molecules is negligible as compared to the total volume of the gas.
2. There are no attractive forces between the gas molecules.

Deviation from Ideal Gas Behaviour: Deviation in behavior of ideal gases can be expressed in terms of compressibility factor (Z) like as

$$Z = pV / nRT$$

In case of ideal gas, $pV = nRT$, $Z = 1$

In case of nearly ideal gas (N_2 and O_2), $pV = nRT$, $Z = \text{approx } 1$

In case of real gas, $pV \neq nRT$, $Z \neq 1$

Negative deviation: At intermediate pressure all gases show $Z < 1$, in that case, the gas is more compressible. Examples liquefiable gas like CO_2, NH_3

Positive deviation: At high pressure all gases show $Z > 1$, in that cases, gas are less compressible. Examples are permanent gas like H_2, He

The factors affecting the deviation in ideal gas are

- (i) Nature of Gas: The most easily liquefiable and highly soluble gases show larger deviation.
- (ii) Pressure: The deviation is more at high pressure. CO_2 and N_2 show negative deviation at low pressure and positive derivation at high pressure.
- (iii) Temperature: The deviation is more at low temperature; H_2 and He always show positive deviations at $0^\circ C$.

15. Vander Waal's Equation:

(a) Pressure correction: At high pressure, molecules do not strike the wall of container due to increasing molecular interactions so pressure exerted by the gas at high pressure is lower than the pressure exerted by the ideal gas.

$$p_{\text{ideal}} = p_{\text{real}} + a \left[\frac{n^2}{V^2} \right]$$

Where $a \left[\frac{n^2}{V^2} \right]$ is force of attraction and 'a' is a constant, which is magnitude of attractive force between gas molecules. Greater the value of 'a', the greater the strength of van der Waals' forces and greater is the ease with which a gas can be liquefied.

(b) Volume Correction: At high pressure and low temperature, volume of real gas should be calculated as $(V-nb)$ where nb is approximately the total volume occupied by the molecules themselves. Here, b is a constant.

After volume and pressure correction, van der Waals' obtained the following equation for n moles of a gas.

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

Where, (nb) = abundant volume or effective volume and b = excluded volume or co-volume = 4 * actual volume of gas molecules

Units for van der Waals' constant: The unit of van der Waals' constant (a and b) can be found out accordingly

$$p = \frac{n^2 a}{V^2} \quad \text{or} \quad a = \frac{pV^2}{n^2} = \text{atm L}^2 \text{ mol}^{-2}$$

$$V = nb \quad \text{or} \quad b = \frac{V}{n} = \text{L mol}^{-1}$$

Limitation of van der Waals' equation: There is specific range of temperature and pressure, to apply the equation. It deviates at very high pressure and very low temperature for permanent gas (H_2 and He) and liquifiable gas (like N_2 , O_2 , CO_2 etc).

Real gas equation at high pressure for permanent gas (H_2 and He):

Real gas equation at high pressure for liquifiable gas (like N_2 , O_2 , CO_2 etc):

Real gas equation at high volume for liquifiable gas (like N_2 , O_2 , CO_2 etc):

16. Liquefaction of Gases and Critical Points

The phenomenon of conversion of a gas into liquid is known as liquefaction. The liquefaction of a gas takes place when the intermolecular forces of attraction become so high that exist in the liquid. A gas can be liquefied by

- (i) Increasing pressure
- (ii) Decreasing temperature

(i) Critical temperature (T_c): The temperature above which a gas cannot be liquefied, whatever high pressure may be applied.

Critical temperature (T_c) of some gases are He (5.4), H_2 (33.2), N_2 (126.0), CO (134.4), O_2 (154.3), CO_2 (304.1), NH_3 (405.5) (Temp in *K)

Higher is the critical temperature of gas, easier its liquefaction, means intermolecular force of attraction b/w the molecules of a gas are directly proportional to its critical temperature so intermolecular forces of attraction are stronger in NH_3 than CO_2 .

$$T_c = 8a / 27Rb$$

(ii) Critical pressure (P_c): The minimum pressure required to liquify a gas at its critical temperature.

$$P_c = a / 27 b^2$$

(iii) Critical volume (V_c): The volume occupied by one mole of a gas at critical temperature and critical pressure is known as critical volume.

$$V_c = 3b$$

(iv) Boyle's temperature (T_b): Temperature at which a real gas exhibits ideal behaviour for considerable range of pressure is called Boyle's temperature.

$$T_b = a / bR$$

17. Liquid State: If a substance is having melting point below room temperature and boiling point above room temperature, the substance is known as liquid. In liquid state, matter has definite shape and molecular motion is in between solids and gases.

Properties of Liquids:

(i) Vapour pressure: The pressure exerted by the vapours above the liquid surface when these are in equilibrium with the liquid at a given temperature is known as vapour pressure of liquid.

The vapour pressure of a liquid depends on:

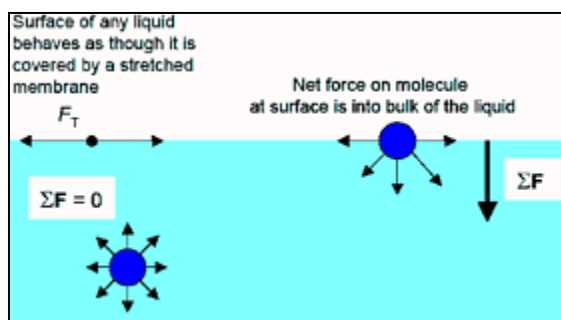
- (a) Nature of liquid
- (b) Temperature: Vapour pressure increases with increasing temperature.
- (c) **Boiling point:** the temperature, at which vapour pressure of liquids becomes equal to the atmospheric pressure, is called boiling point.

At 1 atm pressure, boiling point is known as normal boiling point.

At 1 bar pressure, boiling point is known as standard boiling point. Boiling point varies linearly with external pressure.

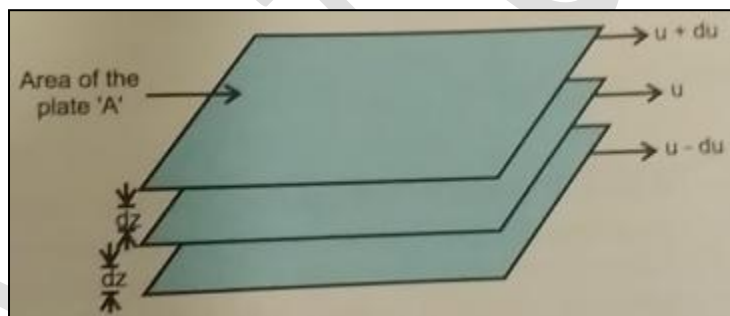
The normal boiling point of water is $100^\circ C$ and its standard boiling point is $99.6^\circ C$

(iii) **Surface tension:** It is the force acting per unit length perpendicular to the imaginary line drawn on the surface of liquid. Rise or fall of liquid in a capillary tube is due to surface tension.



- It has dimensions in SI unit is expressed as N m^{-1} or kg s^{-2} [1 Newton = $\text{Kg m} / \text{s}^2$]
- The magnitude of surface tension of a liquid depends on the attractive forces between the molecules.
- It is measured with the help of an apparatus, called stalagmometer.
- Surface tension decreases as the temperature increases.

(iv) **Viscosity:** Viscosity is a measure of resistance to flow which arises due to friction between layers of fluid.



When there is a regular gradation of velocity, in passing from one layer to the next, it is called laminar flow.

$$F = \eta A [du / dz]$$

Where, F = forces required to maintain the flow of layers, A = area of contact, du / dz = velocity gradient; (the change in velocity with distance.)

' η ' is proportionality constant and is called Viscosity coefficient.

Viscosity coefficient is the force when velocity gradient is unity and the area of contact is unit area. CGS unit of coefficient of viscosity is poise S.I. unit of coefficient of viscosity is Nsm^{-2} .

Effect of T & P on viscosity: It decreases with increase in T, and increases with increase in P.