

States of Matter Class 11 Notes Chemistry

Chapter 5

• Intermolecular Forces

Intermolecular forces are the forces of attraction and repulsion between interacting particles

have permanent dipole moments. This interaction is stronger than the London forces but is weaker than ion-ion interaction because only partial charges are involved.

The attractive forces decrease with the increase of distance between dipoles. The interaction energy is proportional to $1/r^6$ where r is the distance between polar molecules.

Ion-Dipole Interaction: This is the force of attraction which exists between the ions (cations or anions) and polar molecules. The ion is attracted towards the oppositely charged end of dipolar molecules.

The strength of attraction depends upon the charge and size of the ion and the dipole moment and the size of the polar molecule.

For example: Solubility of common salt (NaCl) in water.

• Ion-induced Dipolar Interactions

In this type of interaction permanent dipole of the polar molecule induces dipole on the electrically neutral molecule by deforming its electronic cloud. Interaction energy is proportional to $1/r^6$ where r is the distance between two molecules.

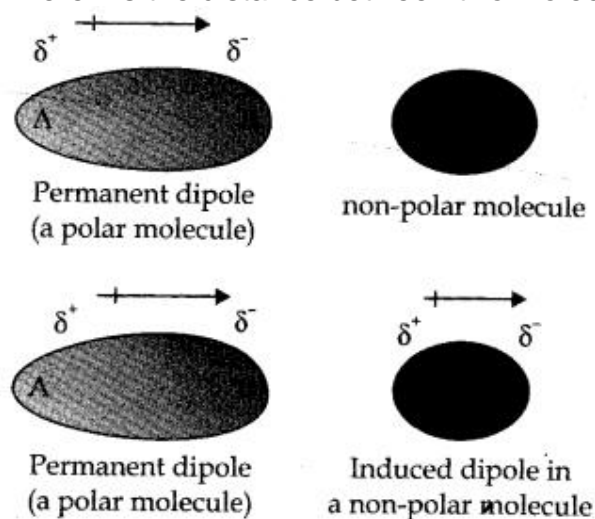


Fig. 5.1 Dipole-induced dipole interaction between permanent dipole and induced dipole.

• London Forces or Dispersion Forces

As we know that in non-polar molecules, there is no dipole moment because their electronic charge cloud is symmetrically distributed. But, it is believed that at any instant of time, the electron cloud of the molecule may be distorted so that an instantaneous dipole or momentary dipole is produced in which one part of the molecule is slightly more negative than the other part. This momentary dipole induces dipoles in the neighbouring molecules. Thus, the force of attraction exists between them and are exactly same as between permanent dipoles. This force of attraction is known as London forces or Dispersion forces. These forces are always attractive and the interaction energy is inversely proportional to the sixth power of the distance between two interacting particles, (i.e. $1/r^6$ where r is the distance between

two particles).

This can be shown by fig. given below.

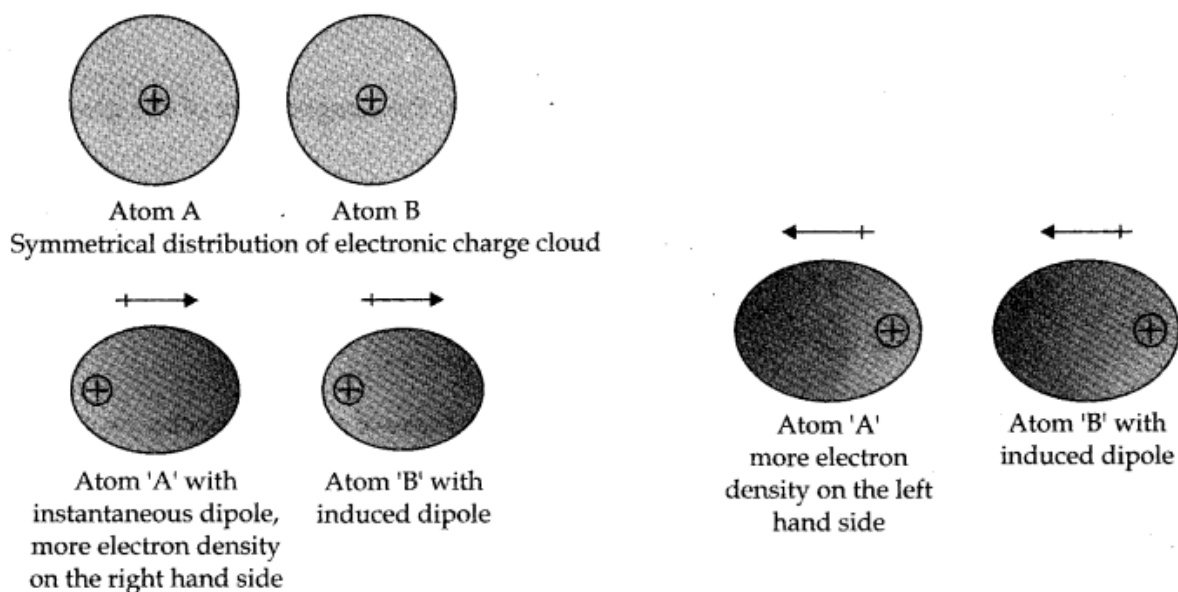
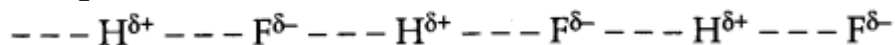


Fig. 5.2 Dispersion forces or London forces between atoms.

Hydrogen bonding: When hydrogen atom is attached to highly electronegative element by covalent bond, electrons are shifted towards the more electronegative atom. Thus a partial positive charge develops on the hydrogen atom. Now, the positively charged hydrogen atom of one molecule may attract the negatively charged atom of some other molecule and the two molecules can be linked together through a weak force of attraction.



Thermal Energy: The energy arising due to molecular motion of the body is known as thermal energy. Since motion of the molecules is directly related to kinetic energy and kinetic energy is directly proportional to the temperature.

• The Gaseous State

Physical Properties of Gaseous State

- (i) gases have no definite volume and they do not have specific shape,
- (ii) Gases mix evenly and completely in all proportions without any mechanical aid.
- (iii) Their density is much lower than solids and liquids. :
- (iv) They are highly compressible and exert pressure equally in all directions.

• Boyle's Law (Pressure-Volume Relationship)

At constant temperature, the volume of a given mass of gas is inversely proportional to its pressure.

$$V \propto \frac{1}{P} \quad (\text{At constant temp.})$$

$$V = \frac{K}{P}$$

or $PV = K$ (constant)

Here, V is the volume, P is the pressure.

$K =$ constant of proportionality.

The value of constant K , depends upon the amount of gas, the temperature of the gas and the units in which P and V are expressed.

Let V_1 be the volume of the gas at a pressure P_1 . By keeping the temperature constant, if the pressure is increased to P_2 then the volume will decrease to V_2 . According to the law,

$$P_1 V_1 = P_2 V_2 = \text{constant.}$$

or
$$\frac{P_1}{P_2} = \frac{V_2}{V_1}$$

Charles' law: At constant pressure, the volume of a given mass of a gas is directly proportional to its absolute temperature.

$$V \propto T$$

or
$$\frac{V}{T} = \text{constant.}$$

Let V_1 be the volume of a gas at temperature T_1 . Pressure remains constant.

If the temperature of the gas increased to T_2 , then the volume will also increase to V_2 . Therefore, according to law,

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

• **Gay Lussac's Law (Pressure-Temperature Relationship)**

At constant volume, pressure of a given mass of a gas is directly proportional to the temperature.

Mathematically,

$$P \propto T$$

$$\Rightarrow \frac{P}{T} = \text{constant. (At constant volume, and mass)}$$

Pressure Vs. temperature (Kelvin) graph at constant molar volume is shown in figure.

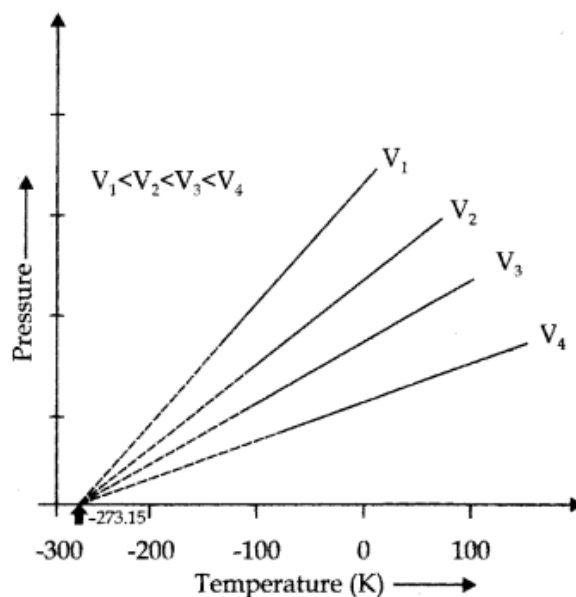


Fig. 5.3 Pressure vs temperature (K) graph (Isochores) of a gas.

Isochore: Each line of graph is called isochore.

• **Avogadro Law (Volume-Amount Relationship)**

Avogadro's law states that equal volumes of all gases under the same conditions of temperature and pressure contain equal number of molecules.

$$V \propto n$$

Where n is the number of moles of the gas.

Avogadro constant: The number of molecules in one mole of a gas

$$= 6.022 \times 10^{23}$$

Ideal Gas: A gas that follows Boyle's law, Charles' law and Avogadro law strictly, is called an ideal gas.

Real gases follow these laws only under certain specific conditions. When forces of interaction are practically negligible.

• **Ideal Gas Equation**

This is the combined gas equation of three laws and is known as ideal gas equation.

At constant T and n : $V \propto \frac{1}{P}$ Boyle's law

At constant P and n : $V \propto T$ Charles' law

At constant P and T : $V \propto n$ Avogadro law.

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

$$V = \frac{RnT}{P}$$

or

$$PV = nRT$$

R is called gas constant. It is also called Universal Gas Constant.

$$\boxed{R = \frac{PV}{nT}} \quad R = \frac{1 \text{ atm} \times 22400 \text{ cm}^3}{1 \text{ mol} \times 273 \text{ K}}$$

$$= 82.1 \text{ cm}^3 \text{ atm K}^{-1} \text{ mol}^{-1}$$

• Dalton's Law of Partial Pressure

When two or more non-reactive gases are enclosed in a vessel, the total pressure exerted by the gaseous mixture is equal to the sum of the partial pressure of individual gases.

Let P_1 , P_2 , and P_3 be the pressure of three non reactive gases A, B, and C. When enclosed separately in the same volume and under same condition.

$$P_{\text{Total}} = P_1 + P_2 + P_3$$

Where, $P_{\text{Total}} = P$ is the total pressure exerted by the mixture of gases.

• Aqueous Tension

Pressure of non reacting gases are generally collected over water and therefore are moist. Pressure of dry gas can be calculated by subtracting vapour pressure of water from total pressure of moist gas.

$$P_{\text{Dry gas}} = P_{\text{Total}} - \text{Aqueous Tension}$$

• Partial Pressure in terms of Mole Fraction

Let at the temperature T , three gases enclosed in the volume V , exert partial pressure P_1 , P_2 and P_3 respectively, then

$$P_1 = \frac{n_1 RT}{V}$$

$$P_2 = \frac{n_2 RT}{V}$$

$$P_3 = n_3 \frac{RT}{V}$$

$$P_{\text{total}} = P_1 + P_2 + P_3$$

$$= (n_1 + n_2 + n_3) \frac{RT}{V}$$

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

$$= \frac{n_1}{n} = x_1$$

Where

$$(n = n_1 + n_2 + n_3)$$

∴

$$P_1 = n_1 P_{\text{total}}$$

• Kinetic Molecular Theory of Gases

- (i) Gases consist of large number of very small identical particles (atoms or molecules),
- (ii) Actual volume occupied by the gas molecule is negligible in comparison to empty space between them.
- (iii) Gases can occupy all the space available to them. This means they do not have

any force of attraction between their particles.

(iv) Particles of a gas are always in constant random motion.

(v) When the particles of a gas are in random motion, pressure is exerted by the gas due to collision of the particles with the walls of the container.

(vi) Collision of the gas molecules are perfectly elastic. This means there is no loss of energy after collision. There may be only exchange of energy between colliding molecules.

(vii) At a particular temperature distribution of speed between gaseous particles remains constant.

(viii) Average kinetic energy of the gaseous molecule is directly proportional to the absolute temperature.

• Deviation From Ideal Gas Behaviour

Real Gas: A gas which does not follow ideal gas behaviour under all conditions of temperature and pressure, is called real gas.

Deviation with respect to pressure can be studied by plotting pressure V_s volume curve at a given temperature. (Boyle's law)

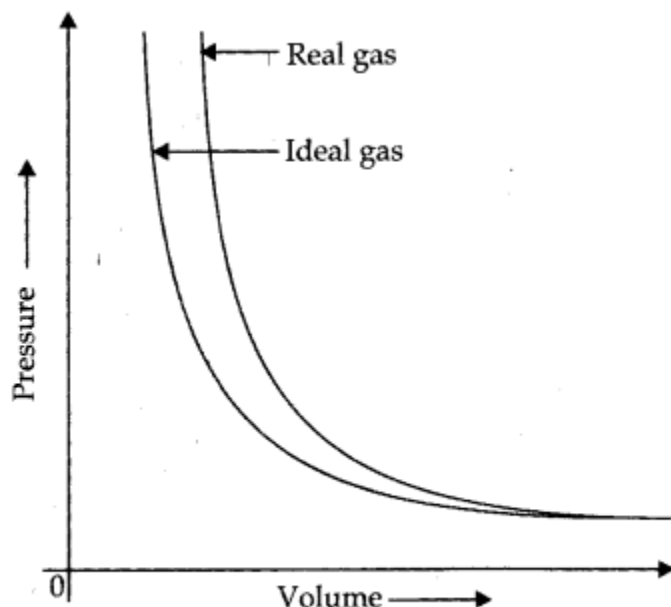


Fig. 5.4 Plot of pressure vs volume for real gas and ideal gas.

Compressibility factor (Z): Deviation from ideal behaviour can be measured in terms of compressibility factor, Z.

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

For ideal gas $Z = 1$

For real gas $Z \neq 1$

When

$Z > 1$ Real gases show positive deviation from ideality

$Z < 1$ The gas show more compressibility and deviation is regarded as negative deviation.

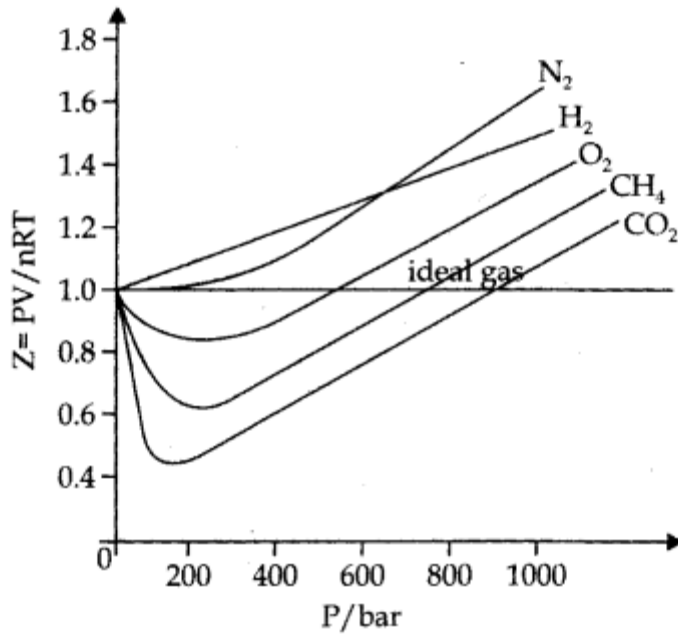


Fig. 5.5 Variation of compressibility factor for some gases.

• van der Waals Equation

For one mole of a gas $\left(P + \frac{a}{V^2}\right) (V - b) = RT$

Where a is a constant for molecular attraction while b is a constant for molecular volume.

(a) There is no force of attraction between the molecules of a gas.

(b) Volume occupied by the gas molecule is negligible in comparison to the total volume of the gas.

Above two assumptions of the kinetic theory of gas was found to be wrong at very high pressure and low temperature.

• **Liquifaction of Gases**

Liquifaction of gases can be achieved either by lowering the temperature or increasing the pressure of the gas simultaneously.

Thomas Andrews plotted isotherms of CO_2 at various temperatures shown in figure.

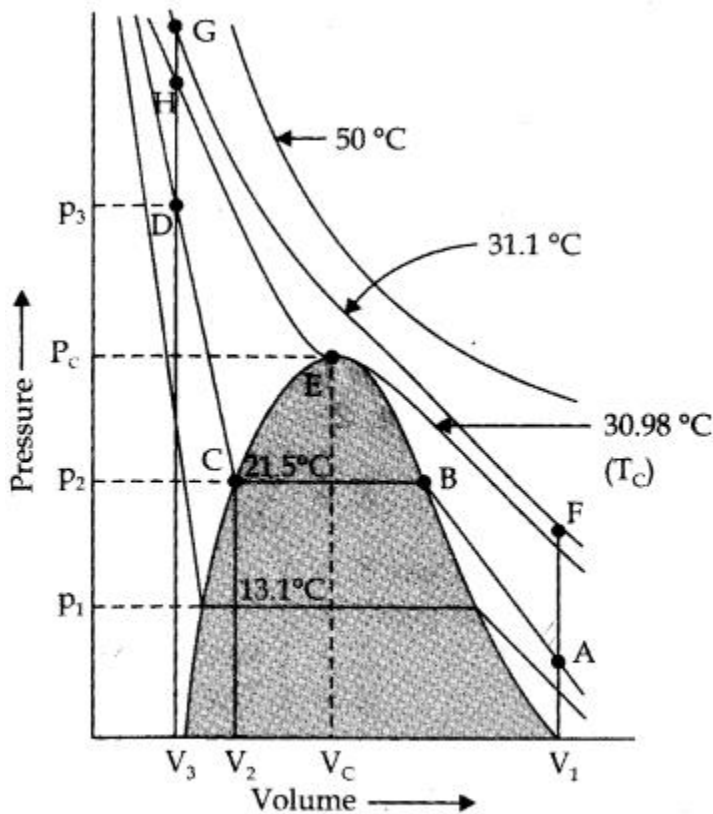


Fig. 5.6 Isotherms of carbon dioxide at various temperatures

Critical Temperature (T_c): It is defined as that temperature above which a gas cannot be liquified however high pressure may be applied on the gas.

$$T_c = 8a/27bR$$

(Where a and b are van der Waals constants)

Critical Pressure (P_c): It is the pressure required to Liquify the gas at the critical temperature.

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

The volume occupied by one mole of the gas at the critical temperature and the critical pressure is called the critical volume (V_c).

For Example. For CO_2 to Liquify.

$$T_c = 30.98^\circ C$$

$$P_c = 73,9 \text{ atm.}$$

$$V_c = 95.6 \text{ cm}^3/\text{mole}$$

All the three are collectively called critical constants.

• Liquid State

Characteristics of Liquid State

- (i) In liquid, intermolecular forces are strong in comparison to gas.
- (ii) They have definite volume but irregular shapes or we can say that they can take the shape of the container.
- (iii) Molecules of liquids are held together by attractive intermolecular forces.

Vapour Pressure: The pressure exerted by the vapour of a liquid, at a particular temperature in a state of dynamic equilibrium, is called the vapour pressure of that liquid at that temperature.

Vapour Pressure depends upon two factors:

(i) Nature of Liquid (ii) Temperature

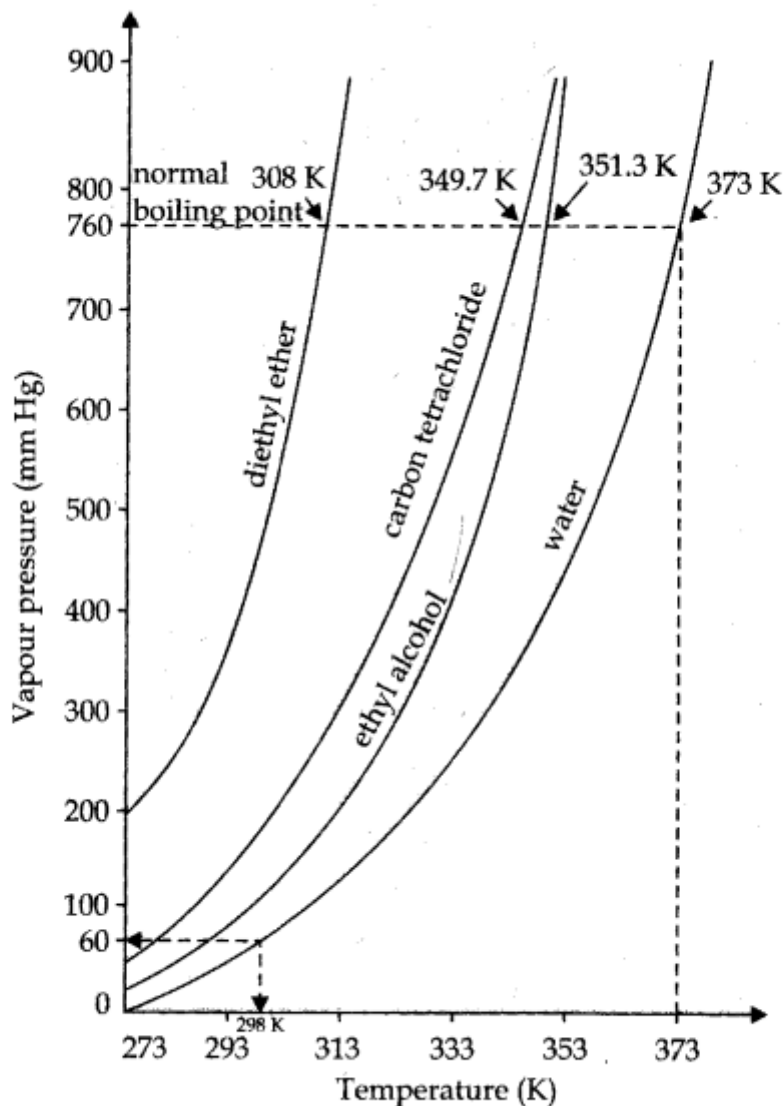


Fig. 5.7 Vapour pressure vs temperature curve of some common liquids.

• **Surface Tension**

It is defined as the force acting per unit length perpendicular to the line drawn on the surface of liquid.

S.I. unit of Surface Tension = Nm^{-1}

Surface Tension decreases with increase in temperature, because force acting per unit length decreases due to increase in kinetic energy of molecules.

• **Viscosity**

It is defined as the internal resistance to flow possessed by a liquid.

The liquids which flow slowly have very high internal resistance, which is due to strong intermolecular forces and hence are said to be more viscous.

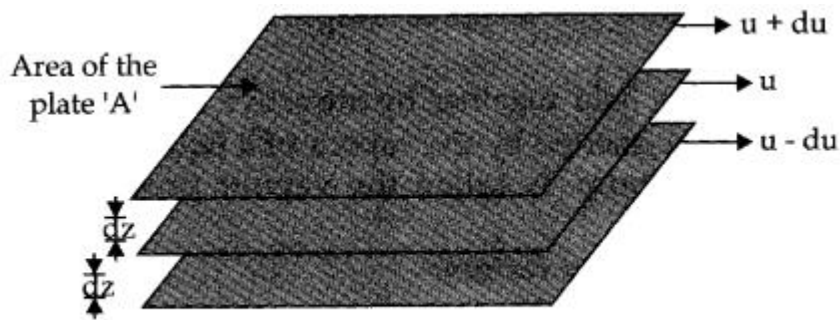


Fig. 5.8 Gradation of velocity in the laminar flow.

When liquid flows, the layer immediately below it tries to retard its flow while the one above tries to accelerate.

Thus, force is required to maintain the flow of layers.

This force

$$f \propto A \quad (A \text{ is area of contact})$$

$$f \propto \frac{du}{dz} \quad \left(\frac{du}{dz} \text{ is velocity gradient i.e., the change in velocity with distance.} \right)$$

$$f \propto A \frac{du}{dz}$$

$$f = \eta A \frac{du}{dz}$$

Where ' η ' is proportionality constant and is called coefficient of viscosity.

SI unit of viscosity coefficient = 1 newton second per square metre
= Nsm^{-2} = Pascal Second.

In C.G.S system, the unit of coefficient of viscosity is poise.

$$1 \text{ poise} = 1 \text{ g cm}^{-1} \text{ s}^{-1}.$$

Effect of Temp, on Viscosity: Viscosity of liquids decreases as the temperature rises because at high temperature, molecules have high kinetic energy and can overcome the intermolecular forces to slip past one another.

• **Boyle's Law:** It states that, under isothermal conditions pressure of a given mass of a gas is inversely proportional to its volume.

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

or $PV = \text{constant}$

or $P_1V_1 = P_2V_2$

- **Charles' Law:** It states that, pressure remaining constant, volume of a fixed amount of a gas is directly proportional to its absolute temperature.

$$V \propto T$$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ (At constant pressure)}$$

- **Avogadro law:** It states that, equal volumes of all gases under same conditions of temperature and pressure contain equal number of molecules.

- **Ideal gas equation:**

$$PV = nRT$$

Where $R = \text{Universal Constant.}$

- **Dalton's law of partial pressures:** It states that, total pressure exerted by a mixture of non-reacting gases is equal to the sum of partial pressures exerted by them.

$$P = P_1 + P_2 + P_3 \dots$$

- **van der Waals equation**

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

- **Critical Temperature:** $(T_c) = \frac{8a}{27Rb}$

Critical pressure: $(P_c) = \frac{a}{27b^2}$

Critical volume $V_c = 3b$