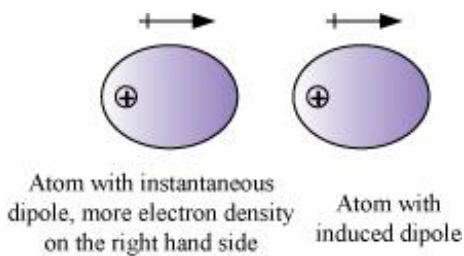


States of Matter

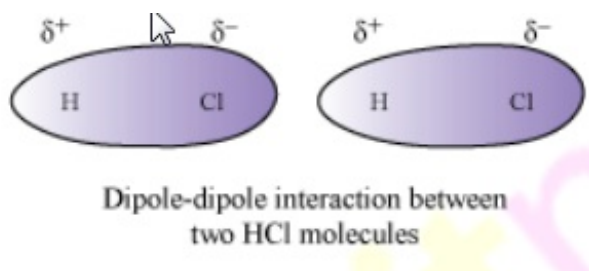
Intermolecular forces:

(Also known as **van der Waals forces**)

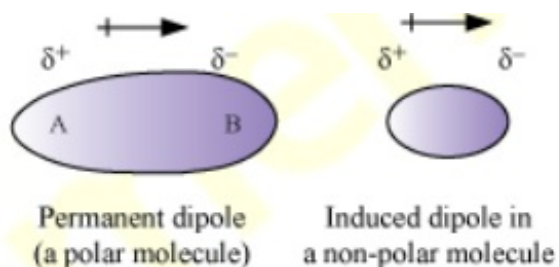
- Dispersion forces or London forces (Forces between atoms or non-polar molecules)



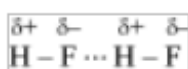
- Dipole-dipole forces (Forces between molecules possessing permanent dipole)



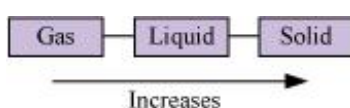
- Dipole-induced forces (Forces between a molecule having permanent dipole and a molecule lacking permanent dipole)



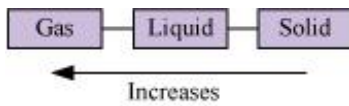
- Ion-dipole interaction (Interaction of an ion with the oppositely charged site of a polar molecule)
- Hydrogen bond (Force between hydrogen attached to an electronegative atom of one molecule and an electronegative atom of different molecule)



Intermolecular interactions:



Thermal energy: Energy of a body due to motion of its atoms or molecules

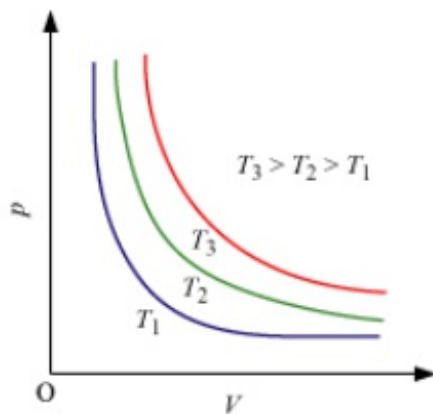


Gas Laws:

- **Boyle's Law:**

At constant temperature, the pressure of a fixed amount (number of moles) of gas is inversely proportional to its volume.

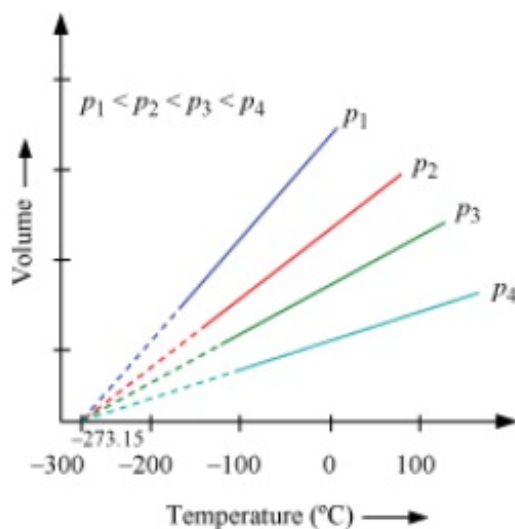
$$p \propto \frac{1}{V}; \text{ (at constant } T \text{ and } n\text{)}$$



p vs V graph

- **Charles' Law:**

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



Kinetic molecular theory of gases:

- Gases are made of large number of identical particles called molecules, which are very small. The actual volume of the molecules is negligible in comparison to the total volume of the molecules. They are considered as the point masses.
- The forces of attraction and repulsion between the particles are supposed to be negligible at ordinary temperature and pressure.
- Particles of a gas are always in constant and random motion.
- Molecules are supposed to be perfectly hard spheres and the collisions between them are perfectly elastic.
- The pressure exerted on the walls of the containing vessel is due to the collision of the molecules on the walls of the container.
- The average kinetic energy of the particles of a gas is directly proportional to the absolute temperature of the gas.

Deviation from ideal gas behaviour (real gas):

van der Waals equation

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

$$\text{Compressibility factor (Z)} = \frac{pV}{nRT}$$

$Z=1$ (for ideal gas, at all temperatures and pressures)

At very low pressures, $Z \approx 1$

At high pressures, $Z > 1$

At intermediate pressures, $Z < 1$

The temperature at which a real gas obeys ideal gas law over an appreciable range of pressure is called **Boyle temperature** or **Boyle point**.

Liquefaction of gases:

A gas can be liquefied by cooling or applying pressure or by the combined effect of both.

Critical temperature (T_c), critical volume (V_c), and critical pressure (p_c):

The temperature at which a gas liquefies is called its critical temperature.

The volume of one mole of a gas at critical temperature is called its critical volume.

The pressure of a gas at its critical temperature is called its critical pressure.

Methods of Liquefaction of Air

- (1) Lind process
- (2) Linde-Claude process

Liquid State

Vapour Pressure: Vapour pressure in the state of equilibrium between liquid phase and vapour phase.

Boiling point: The temperature at which the vapour pressure of a liquid is equal to the external pressure.

Surface tension (γ):

- Force acting per unit length perpendicular to the line drawn on the surface of liquid
- Its SI unit is Nm^{-1} .

Viscosity:

It is the measure of the resistance to flow, which arises due to the internal friction between the layers of fluid as they slip past one another while the liquid flows.

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

Where, F = Force

A = Area of contact

$\frac{du}{dz}$ = Velocity gradient

η = Proportionality constant known as coefficient of viscosity

Coefficient of viscosity (η) is the measure of viscosity. Its SI unit is N s m^{-2} .

Greater the viscosity, more slowly does the liquid flow.