

**Points to Remember****Subject: Chemistry****Class: XI****Chapter: States of matter****Top concepts**

1. Intermolecular forces are the forces of attraction and repulsion between interacting particles (atoms and molecules).
2. Attractive intermolecular forces are known as van der Waals forces which include dispersion forces or London forces, dipole – dipole forces, and dipole – induced dipole forces.
3. Dispersion forces or London forces exist between neutral atoms like that of noble gases or non – polar molecules like O<sub>2</sub>, N<sub>2</sub>, H<sub>2</sub> etc. Magnitude of dispersion forces depends on the polarisability of the neutral atom or molecule.
4. Dipole – Dipole forces occur between the molecules have permanent dipole such as HCl, NH<sub>3</sub>, H<sub>2</sub>O etc.
5. Dipole – Induced Dipole forces operate between the polar molecule having permanent dipole and the molecule having no permanent dipole.
6. Hydrogen bond is a special case of dipole-dipole interaction. When hydrogen atom is bonded to atoms of highly electronegative elements such as fluorine, oxygen or nitrogen, hydrogen atom forms a weak bond with the electronegative atom of the other molecule. This weak bond is called hydrogen bond.
7. Thermal energy is the energy of a body arising from motion of its atoms or molecules.
8. Three states of matter are the result of balance between intermolecular forces and the thermal energy of the molecules.

9. Predominance of molecular interactions result into change of gases to liquid to solid state while predominance of thermal energy results into change of solid to liquid to gas.

#### 10. Physical properties of gas

- Gases are highly compressible
- Gases exert pressure equally in all directions
- Gases have density much lower than the solids and liquids
- Gases do not have definite volume and the shape. They assume volume and shape of the container
- Gases mix evenly and completely in all proportions without any mechanical aid.

11. According to Boyle's law at constant temperature, the pressure of a fixed amount (i.e., number of moles  $n$ ) of gas varies inversely with its volume. Mathematically, it is written as

$$p \propto \frac{1}{V} \quad (n, T \text{ are constant})$$

$$\text{or } pV = k_1 \text{ (constant)}$$

Value of  $k_1$  depends on amount of gas and temperature and units in which  $p$  and  $V$  are expressed

12. Boyle's law expresses in a quantitative manner that gases are compressible because when a given mass of a gas is compressed, the same number of molecules occupy a smaller space.

13. According to Charles' Law at constant pressure, the volume of fixed mass of a gas is directly proportional to absolute temperature. Mathematically, it is written as

$$V \propto T \quad (n, P \text{ are constant})$$

$$\text{or } \frac{V}{T} = k_2 \text{ (constant)}$$

Value of  $k_2$  depends on amount of gas and pressure and units in which volume  $V$  is expressed

14. The lowest hypothetical or imaginary temperature at which gases are supposed to occupy zero volume is called Absolute zero.

15. According to Gay Lussac's Law at constant volume, pressure of a fixed amount of a gas varies directly with the temperature. Mathematically, it is written as

$$p \propto T \quad (n, V \text{ are constant})$$

$$\text{or } \frac{p}{T} = k_3 \text{ (constant)}$$

16. Standard temperature and pressure means 273.15 K (0°C) temperature and 1 bar (i.e., exactly  $10^5$  Pascal) pressure.

17. According to Avogadro Law equal volumes of all gases under the same conditions of temperature and pressure contain equal number of molecules. Mathematically, it is written as

$$V \propto n \quad (p, T \text{ are constant})$$

$$\text{or } \frac{V}{n} = k_4 \text{ (constant)}$$

18. A gas that follows Boyle's Law, Charles' law and Avogadro law strictly is called an ideal gas.

19. Ideal Gas Equation :  $pV = nRT$

$R$  is called gas constant. It is same for all gases. Therefore it is also called Universal Gas Constant.

20. Different values of Universal Gas Constant

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

21. Combined Gas Law :  $\frac{p_1 V_1}{T_1} = \frac{p_2 V_2}{T_2}$

22. According to Dalton's Law of Partial Pressures the total pressure exerted by the mixture of non - reactive gases is equal to the sum of the partial pressures of individual gases. Mathematically, it is written as

$$p_{\text{total}} = p_1 + p_2 + p_3 + \dots \text{ (T, V are constant)}$$

$p_{\text{total}}$  is the total pressure exerted by the mixture of gases

$p_1, p_2, p_3$  etc. are partial pressures of gases

23. Partial pressure in terms of mole fraction is expressed as

$$p_i = x_i \times p_{\text{total}}$$

$p_i$  is partial pressure of  $i$ th gas

$x_i$  is mole fraction of  $i$ th gas

24. Assumptions or postulates of Kinetic Molecular Theory of Gases:

- Gases consist of large number of minute identical particles (atoms or molecules) which are very small.
- Gas molecules are so far apart from each other that the actual volume of the molecules is negligible as compared to the total volume of gas. They are thus considered as point masses.
- There is no force of attraction between the particles of a gas at ordinary temperature and pressure.
- Particles of a gas are always in constant and random motion.
- Particle of a gas move in all possible directions in straight lines. During their random motion, they collide with each other and with the walls of the container. Pressure is exerted by the gas as a result of collision of the particles with the walls of the container.
- Collisions of gas molecules are perfectly elastic. This means that total energy of molecules before and after the collision remains same.
- At any particular time, different particles in the gas have different speeds and hence different kinetic energies.

- In kinetic theory it is assumed that average kinetic energy of the gas molecules is directly proportional to the absolute temperature.
25. Gases show deviation from ideal behaviour because of two faulty assumptions
- There is no force of attraction between the molecules of a gas.
  - Volume of the molecules of a gas is negligibly small in comparison to the space occupied by the gas.

26. At low temperature and high pressure gases deviate from ideal behaviour i.e. gases behaves as real gases

27. At low pressure and high temperature gases show ideal behaviour i.e. gases behaves as ideal gases

28. van der Waals equation of state is given below

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

a, b are van der Waals constant

n is the number of moles of gas

29. The deviation from ideal behaviour can be measured in terms of compressibility factor Z, which is the ratio of product pV and nRT.

$$\text{Mathematically: } Z = \frac{pV}{nRT} = \frac{V_{\text{real}}}{V_{\text{ideal}}}$$

For ideal gas  $Z = 1$

For real gas  $Z > 1$  or  $Z < 1$

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

31. Liquefaction of Gases: Gases can be liquefied by lowering temperature and increasing pressure.

32. Critical temperature ( $T_c$ ) is the temperature above which a gas cannot be liquefied however high pressure may be.

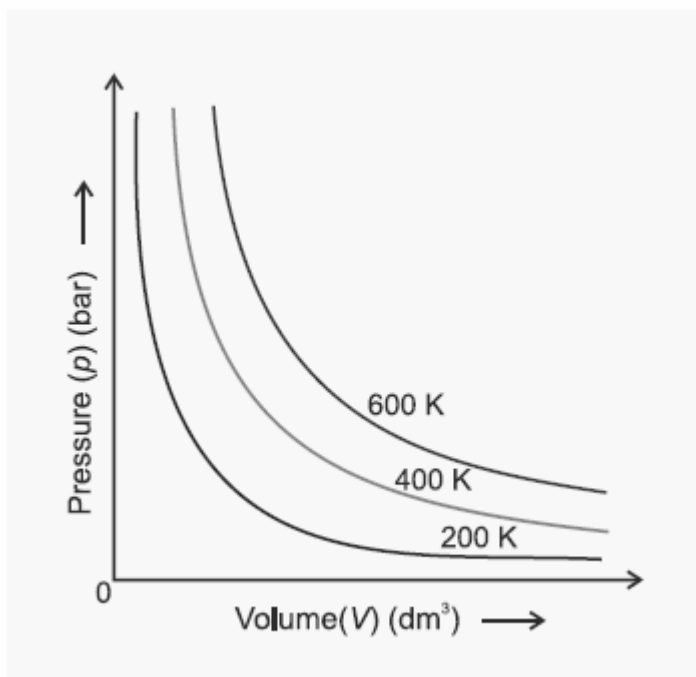
33. Critical pressure ( $p_c$ ) is minimum pressure required to liquefy gas at the critical temperature.
34. Critical volume ( $V_c$ ) is the volume occupied by one mole of gas at critical temperature and critical pressure.
35. The temperature at which vapour pressure of liquid is equal to the external pressure is called boiling temperature at that pressure. At 1 atm pressure boiling temperature is called normal boiling point. If pressure is 1 bar then the boiling point is called standard boiling point of the liquid.
36. Standard boiling point of the liquid is slightly lower than the normal boiling point because 1 bar pressure is slightly less than 1 atm pressure.
37. Surface tension ( $\gamma$ ) is defined as the force acting per unit length perpendicular to the line drawn on the surface of liquid. SI unit is expressed as  $\text{N m}^{-1}$ .
38. Surface tension decreases with rise in temperature.
39. Coefficient of viscosity ( $\eta$ ) is the force when velocity gradient is unity and the area of contact is unit area. SI unit of coefficient of viscosity is  $\text{N s m}^{-2}$  or  $\text{kg m}^{-1} \text{s}^{-1}$ . In cgs system the unit of coefficient of viscosity is poise
40. 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 between the layers.
41. In Boyle's law pressure versus volume graph is plotted at constant temperature. Such a graph is called as an isotherm because temperature remains constant
42. In Charles law volume versus temperature graph is plotted at constant pressure. Such a graph is called as isobar because pressure remains constant.

43. In Gay Lussac's law pressure versus temperature graph is plotted at constant volume. Such a graph is called as an isochore because volume remains constant.

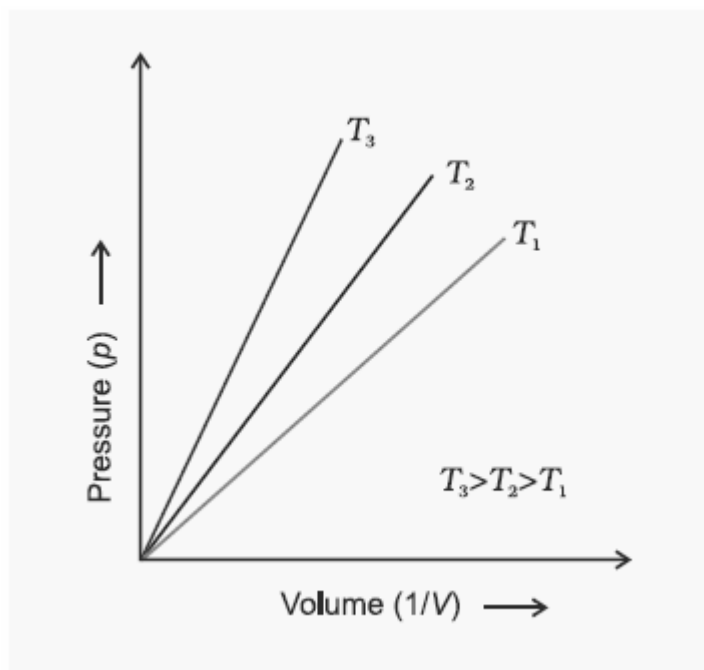
### **Top Diagrams**

1. Representation of Boyle's law graphically:

Graph 1: Plot of pressure ( $p$ ) versus volume of gas ( $V$ ) at different temperatures

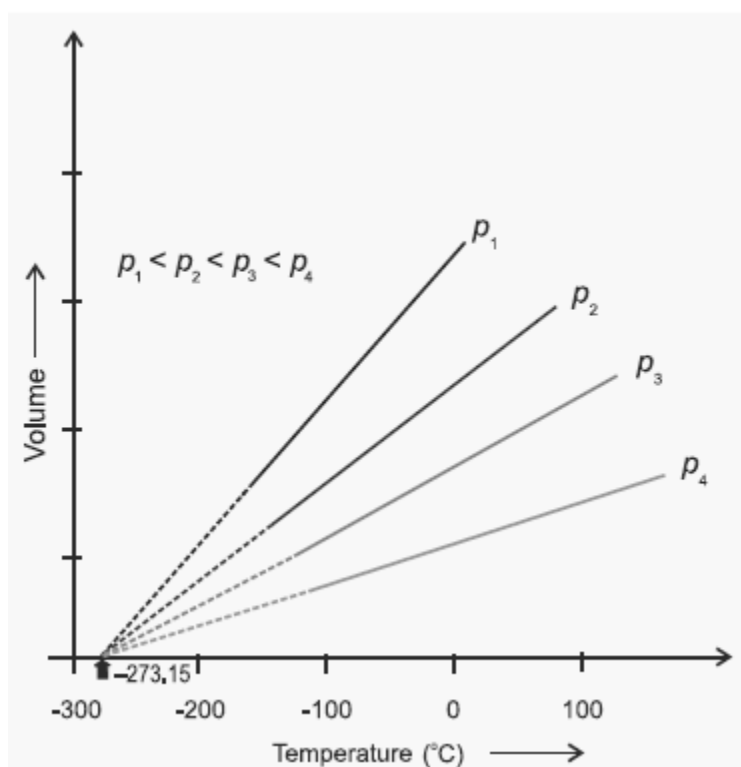


Graph 2: Plot of pressure ( $p$ ) versus ( $1/V$ ) at different temperatures



## 2. Representation of Charles law graphically

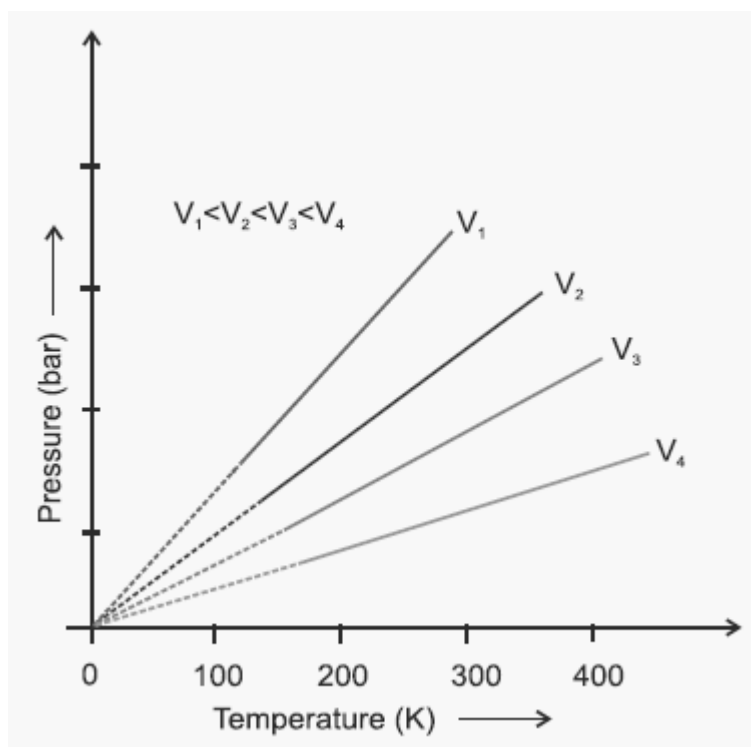
Plot of volume versus temperature at different pressures





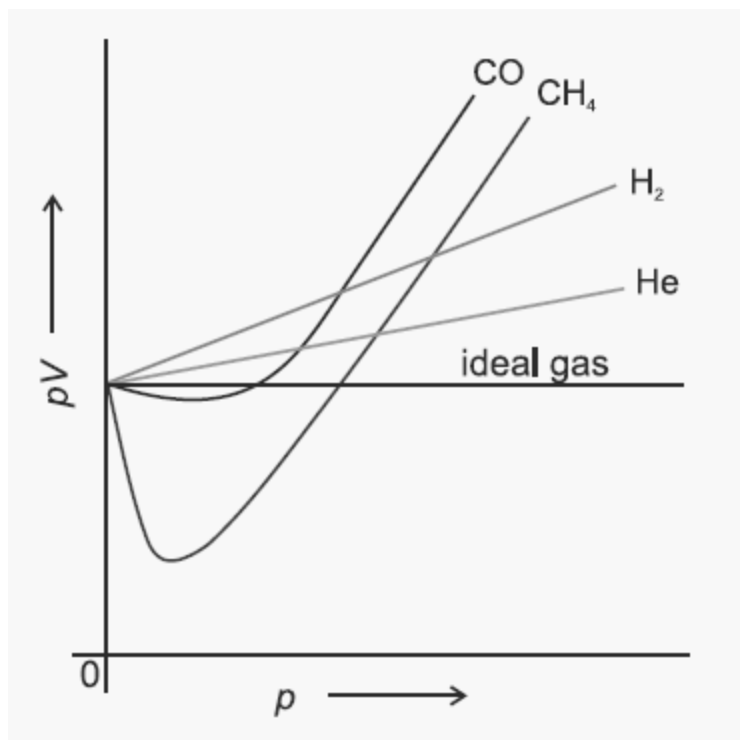
## 3. Representation of Charles law graphically

Plot of pressure versus temperature at different volume

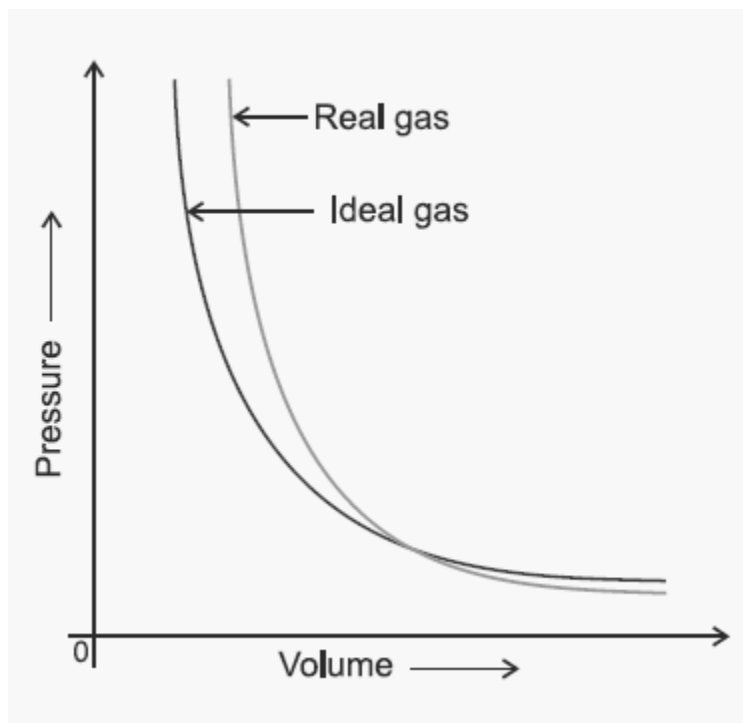


## 4. Deviation from ideal gas behaviour

Graph 1: Plot of pressure-volume ( $pV$ ) versus pressure ( $p$ ) at constant temperature for ideal and real gases



Graph 2: Plot of pressure ( $p$ ) versus volume ( $V$ ) for ideal and real gases



### 5. Plot of compressibility factor ( $Z$ ) versus pressure( $p$ ) for some gases

