

## States of Matter

### **Inter molecular Interactions, Properties of gases, Gas Laws .Ideal gas equation**

1. In a substance there exists inter molecular attractions between different molecules. The physical state of a substance changes with the change of these attractive forces. These intermolecular forces are usually called van der waals forces. As the intermolecular attractions decreases the physical state of substance change from solid to liquid and then to gas. The intermolecular forces are of different types.

i. Ion-Dipole interactions

ii. Dipole-Dipole interactions

iii. London dispersion forces (Induced dipole - induced dipole)

iv. Dipole - Induced dipole interactions.

ii, iii and iv are together termed as van der Waals forces.

2. All intermolecular forces are electrical. They result from mutual attractions of unlike charges (or) repulsions of like charges.

3. If the species are ions, full charges are present and the attractions between them are very strong (500 – 1000 kJ/mol) called as ionic bonds.

i) Ion-dipole interactions

These interactions are significant when ionic compounds are dissolved in polar solvents like water.

**E.g.** When NaCl is added to water, it ionises to give  $\text{Na}^+$  and  $\text{Cl}^-$  ions. The water molecules orient around the ions in such a way that the negatively charged oxygen atoms are towards  $\text{Na}^+$  ions and positively charged hydrogen atoms are towards  $\text{Cl}^-$  ions.

The magnitude of ion-dipole interactions depend upon charge present on the ion ( $z$ ), dipole moment of the polar solvent or dipole ( $\mu$ ) and on the square of the distance 'r' between the ion and dipole.

$$\text{The interaction energy (E)} = \frac{Z\mu}{r^2}$$

ii. Dipole - Dipole interactions

These are interactive forces that exist between the dipoles. These are the electrostatic forces. The attractive forces exist between unlike poles and repulsive forces exist between like poles. These are weak forces and operate when the molecules are in close contact. These interactive energies are of the order 3–4 kJ/mole. When the polarity of dipoles is more, the interactive energies are more.

The dipole - dipole interaction energy between rotating molecules is proportional to  $\frac{1}{r^6}$  and between stationary solid polar molecules is proportional to  $\frac{1}{r^3}$ .

r = distance between the polar molecules.

iii. London dispersion forces

These are the interactive forces that exist between the atoms of non-polar molecules. **E.g.** Benzene is not a gas at room temperature. It is a liquid. It has a boiling point 80.1<sup>0</sup> C. The forces that exist between non-polar benzene molecules are known as London dispersion forces. These forces arise due the motion of electrons around the nucleus.

Interaction energy is proportional to  $\frac{1}{r^6}$  .

- i. The larger molecules have larger London dispersion forces due to their more polarisability. The smaller molecules have smaller dispersion forces due to their lesser polarisability.

**Ex.** In halogens from  $F_2$  to  $I_2$  the polarisability gradually increases. So London dispersion forces increases. Hence  $F_2$ ,  $Cl_2$  are gases.  $Br_2$  is a liquid and  $I_2$  is a solid.

- ii. The molecules which are more spread have more dispersion forces than compact molecules.

**E.g.:** n-butane has more boiling point than isobutane. n-pentane has more boiling point (309.4K) than 2, 2-dimethyl propane(282.7K)

- iv. Dipole - Induced dipole interactions

These forces exist between the polar and non-polar molecules. The permanent dipole of the polar molecule induces the dipole in the non-polar molecule. The attractive forces between these permanent dipoles and induced dipoles are called dipole-induced dipole interactions. The interaction energy is proportional to

$$\frac{1}{r^6}$$

Where  $r$  = distance between the molecules

The interactive forces also depend on the dipole moment of dipole and polarisability of non-polar molecule. Here a cumulative effect of dispersion force and dipole - induced dipole interactions exist.

2. Thermal energy is due to the motion of atoms or molecules of the substance. This energy is directly proportional to the average kinetic energy of the molecules of the substance. The movement of particles is called thermal motion.
3. Intermolecular forces tend to keep the molecules together but thermal energy tends to keep them apart.
4. If thermal energy predominates over intermolecular forces the substances would change from solid liquid gas. If the intermolecular forces predominate then substances change from gas to liquid to solid.
5. **Hydrogen Bonding**:-It is a weak electrostatic attraction that occurs between covalently bonded hydrogen and high electro negative atom of same molecule or

different molecule. High electro negative elements with small size (F, O, N) form hydrogen bonding.

**Example HF,**

The strength of Hydrogen bond is 10 to 100 KJ/mole.

It is strongest of all inter molecular forces including the Dipole-Dipole attraction and London forces.

**II-Measurable Properties of Gases**

A gas sample can be described in terms of four measurable properties

- 1) Amount
- 2) Volume
- 3) Pressure (p)
- 4) Temperature

1) **Amount (n):** The number of moles, n, of a sample of gas in a container can be found by dividing the mass, m, of the sample by the molar mass.

2) **Volume, V:** The volume of the container is the volume of the gas sample, it is usually given in litre (L) or millilitres (ml)

$$1\text{L} = 1\text{ dm}^3 = 10^{-3}\text{ m}^3 = 10^3\text{ cm}^3$$

3) **Pressure, P:** The force exerted by the gas per unit area on the walls of the container is equal to its pressure.

The S.I unit is pascal

$$1\text{ atm} = 760\text{ mm Hg} = 760\text{ torr} = 76\text{ cm Hg}$$

$$1\text{atm} = 1.01325\text{ bar} = 1.01325 \times 10^5\text{ Pa (or)} 1\text{ bar} = 0.987\text{ atm}$$

i. Atmospheric pressure is measured by using barometer. Atmospheric pressure  $p = h \rho g$

h = Height of the mercury column,  $\rho$  = Density of mercury,  $g$  = Acceleration to gravity

ii. The pressure of a gas in a closed vessel is measured by using manometer

4) **Temperature:** The temperature of a may be measured in centigrade degrees ( $^{\circ}\text{C}$ ) (or) celsius degrees. The S.I unit of temperature is Kelvin

$$\text{K} = ^{\circ}\text{C} + 273$$

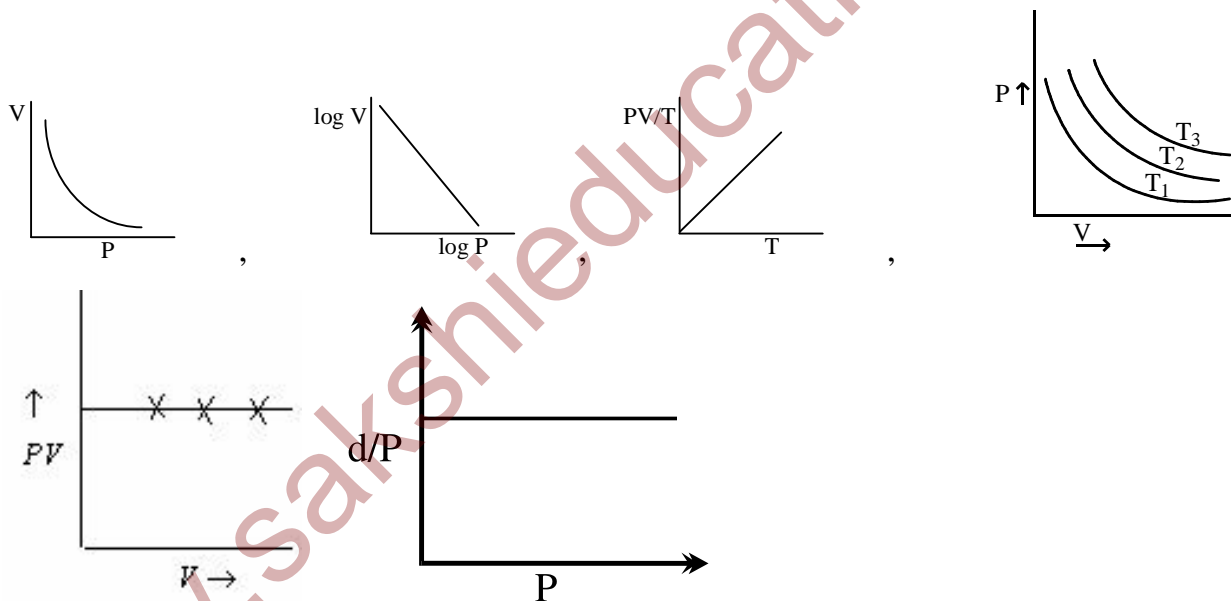
On Kelvin scale water freezes at 273.15 K & boils at 373.15K at the atmospheric pressure.

## Gas Laws

**1. Boyle's Law:** 'At constant temperature, the volume of given mass of gas is inversely proportional to its pressure.'

$$V \propto \frac{1}{P} (T, n \text{ const}) \quad (\text{Or}) \quad PV = K (T, n \text{ Constant})$$

The graph between P & V at constant temperature is called isotherm. The following are Isotherms.



If  $P_1, V_1$  are the initial & volume of a given sample of gas &  $P_2, V_2$  the changed pressure & volume, we can write

$$P_1 V_1 = P_2 V_2$$

**2. Charles's Law:** 'The volume of given mass of a gas is directly proportional to absolute temperature at constant pressure'

i.e  $V \propto T$  ( $n, p = \text{constant}$ ) or  $\frac{V_1}{T_1} = \frac{V_2}{T_2}$

**(OR)**

At constant pressure, the volume of a given mass of a gas at 0°C, increases or decreases by 1/273 times of its volume at 0°C for every degree rise or fall in temperature respectively.

The law may be mathematically expressed as  $V_t = V_0 (1 + \alpha t)$  (pressure, mass or number of moles constant)

Here ‘ $\alpha$ ’ is a constant known as coefficient of thermal expansion,  $t$  is the temperature (°C) and  $V_0$  is the volume of the gas at 0°C. From the experimental observations is found to be equal to 1/273.

Let the volume of a given mass of a gas at 0°C be  $V_0$ . Then at 1°C, the volume of a gas increases by  $V_0 / 273$ .

Volume of the gas at 1°C =  $\left( V_0 + \frac{V_0}{273} \right) = V_0 \left( 1 + \frac{1}{273} \right)$ .....(1)

In general, at  $t$ °C, volume of the gas,  $V_t = V_0 \left( 1 + \frac{t}{273} \right)$ .....(2)

Or  $V_1 = V_0 \left( \frac{273+t_1}{273} \right); V_2 = V_0 \left( \frac{273+t_2}{273} \right)$  i.e  $\therefore \frac{V_2}{V_1} = \frac{(273+t_2)}{(273+t_1)} = \frac{T_2}{T_1}$ .....(3)

$V \propto T$ , at constant pressure, where  $V$  is the volume and  $T$  is the absolute temperature.

Introducing proportionality constant  $K$ ,

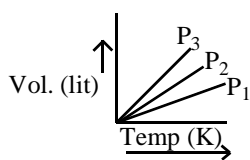
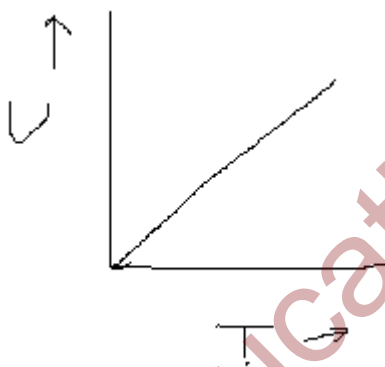
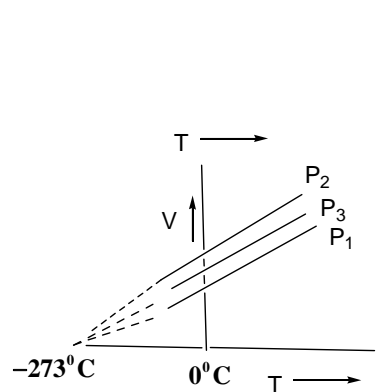
$V = KT$  or  $V/ T = K = \text{constant}$

At constant pressure, the ratio of volumes of a gas at two different conditions is nothing but the ratio of temperatures under those conditions. If  $V_1$  and  $V_2$  are initial

and final volumes,  $T_1$  and  $T_2$  are initial and final temperatures in absolute scale, at constant pressure.

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

\*The graph between  $T$  &  $V$  at constant pressure is called isobar. The following are Isobars.



Where  $P_1 > P_2 > P_3$  ,

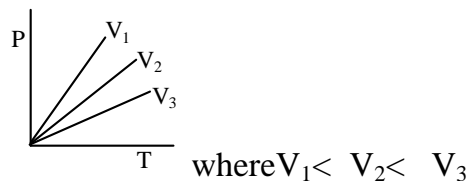
These plots when extrapolated intersect the temperature axes at  $-273$  degree centigrade i.e 'O'K, Which is called Absolute zero temperature. At absolute zero all molecular motions would stop and the volume of gas would become zero. The gas becomes a liquid and solid, this is lowest possible temperature.

**3. Pressure –Temperature Law:** The pressure of given mass of a gas is directly proportional to absolute temperature at constant volume

$$P \propto T, \text{ at constant volume}$$

$$P/T = \text{constant}$$

When pressure of gas is plotted against its absolute temperature for a fixed mass of gas we get a straight line called 'Isochore'.



#### 4. Avogadro's law

This is stated as, "equal volumes of all gases contain equal number of moles or molecules under same conditions of temperature and pressure".

The volume of gas is directly proportional to the number of molecules of the gas at constant temperature and pressure. The number of moles of given gas ( $n$ ) is obtained from the number of molecules.

$$n = \frac{\text{number of gas molecules}}{6.023 \times 10^{23}}$$

Hence the Avogadro's relationship can be represented as  $V \propto n$

If  $n_1$  and  $n_2$  are the number of moles of a gas with respective volumes  $V_1$  and  $V_2$

$$\text{Or } V_1/n_1 = V_2/n_2$$

When volume of a gas is plotted against number of moles at constant temperature and pressure, we get a straight line.

5. The latest standard temperature value is still 273.15K (0°C) but standard pressure is 1 bar i.e., exactly  $10^5$  Pascals. These values are approximate freezing temperature of water and atmospheric pressure at sea level, respectively. According to the latest STP conditions the molar volume is 22.711 lit. mol<sup>-1</sup>
6. Standard absolute temperature and pressure (SATP) conditions are also used in some specific works. SATP conditions are 298.15K temperature and 1 bar pressure (1 bar =  $10^5$  pascal). At SATP conditions the Molar volume of an ideal gas is 24.789 litre/mole.



**7. Ideal Gas Equation:** Ideal gas Equation for any mole of gas is  $PV=nRT$

Since  $n= w/M$ ;  $PV = WRT/M$  (or)  $PM=dRT$

Where  $w$  = mass of the gas and  $M$  is molecular weight of gas.

Where  $d$  = density of the gas

8. For same gases at two different conditions then

$$\frac{d_1 T_1}{P_1} = \frac{d_2 T_2}{P_2}$$

9. For a given mass of gas the gas equation can be expressed as  $PV/T = \text{constant}$  (or)

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

10. In ideal gas equation,  $R$  is called universal gas constant or molar gas constant. The value of  $R$  depends upon the units of measurement of Pressure and Volume.

**Ex.** Taking the volume in  $\text{m}^3$  and pressure in  $\text{Nm}^{-2}$ , the value of universal gas constant in SI system is given as  $8.314 \text{ J K}^{-1} \text{ mol}^{-1}$ .

11. The importance of universal gas constant is that it represents the work done by gas. It is given as,  $R=Pv/nT$

12. The different values of 'R' are

Value R	Units of R
0.0821	L -atm $\text{K}^{-1} \text{ mol}^{-1}$
82.1	cc-atm $\text{K}^{-1} \text{ mol}^{-1}$
62.4	L -torr $\text{K}^{-1} \text{ mol}^{-1}$
1.987	cal $\text{K}^{-1} \text{ mol}^{-1}$
8.314	J $\text{K}^{-1} \text{ mol}^{-1}$
$8.314 \times 10^7$	erg $\text{K}^{-1} \text{ mol}^{-1}$

13. A gas is said to be an ideal when it obeys Boyle's and Charles's law under all conditions of temperature and pressure.

Gases like  $H_2, O_2, N_2, CO_2$  etc would show ideal gas behavior at low pressure and high temperature when molecular interactions are negligible.

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