CHAPTER IV - STATES OF MATTER (GASES AND LIQUIDS)

UNIT : 1

1 Graham's law of diffusion, Dalton's Law of partial pressures and Avagadro's law.

2 Ideal behaviour, empirical derivation of gas equation, ideal gas equation.

Very Short Answer Questions :

1. What is the value of the gas constant if pressure is expressed in N/m^2 and volume is expressed in dm^3 ?

A. $8.314\ joule\, mole^{-1}k^{-1}$

$1.987\,\text{calories}\, mole^{-1}k^{-1}$

2. What is the standard temperature and pressure conditions ? What is the volume of one mole of an ideal gas under these conditions ?

A. 1) Standard temperature is $0°C = 273K$

2) Standard pressure is 1 atmosphere = 76cm = 760 mm of Hg

3) Volume of 1 mole of gas STP = 22.4 lit or 22,400 ml

3. What are the isotherms ?

A. The curves which shows the relationship between variation of volume of a given mass of gas and pressure at constant temperature.

4. If the value of gas constant is 82.1. Give its units.

A. $ml - atm\, mole^{-1}k^{-1}$

5. State Graham's law of diffusion.

A. Graham's law of diffusion : At a given temperature and pressure the rate of diffusion of a gas is inversely proportional to the square root of its density, vapour density (or) molecular weight

$$r \propto \frac{1}{\sqrt{M}}$$

6. When do real gas show ideal behaviour ?

A. real gases behave as ideal gas at low pressure and high temperature .

7. Find the volume of 3 grams of H_2 at STP.

A. MW of H_2 is 2.

2 grams of H_2 at STP occupies 22.4 litres

3 grams of H_2 at STP occupies

$$= \frac{22.4 \times 3}{2} = 33.6 \text{ litres}$$

8. State Charle's law and Boyle's law.

A. 1) Charle's law : At constant pressure the volume of a given mass of gas is directly proportional to absolute temperature

$V \propto T$ (m, p constant)

2) Boyle's law : At constant temperature, the volume of a given mass of gas is inversely proportional to its pressure

9. How does volume of a given mass of gas change when its pressure is reduced to half and absolute temperature is increased by four times ?

A. frist gas second gas

Let $P_1 = P$ $P_2 = \frac{P}{2}$

$V_1 = V$ $V_2 = ?$

$T_1 = T$ $T_2 = 4T$

$\therefore \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} = \frac{P_1V_1T_2}{P_2T_1} = \frac{P \times V \times 4T}{P/2 \times T} = 8V$

$\therefore$ The volume increases by 8 times

10. Why is gas constant 'R' called universal gas constant ? What is boltzmann constant ?

A. (i) Since the value of gas constant 'R' is same for all gases. It is called universal gas constant.

(ii) Boltzmann constant is the gas constant per molecule.(K)

$= \frac{R}{N} = 1.38 \times 10^{-16}\, erg/k$ - molecule

$= 1.38 \times 10^{-23}\, joule/k$ - molecule

11. What is an ideal gas ?

A. The gas which obeys all gas laws at all temperatures and pressures is called an ideal gas.

12. Which gas diffuses faster among N_2, O_2 and CH_4 ? Why ?

A. CH_4 (16) diffuses faster than the other two gases since it has less molecular weight than N_2 (28) and oxygen (32) .

13. How many times methane diffuses faster than sulphur dioxide ?

A. $\frac{r_{CH_4}}{r_{SO_2}} = \sqrt{\frac{M_{SO_2}}{M_{CH_4}}} = \sqrt{\frac{64}{16}} = 2$

methane diffuses two times faster than SO_2

14. State Dalton's law of partial pressures.

A. The total pressure exerted by a mixture of chemically non reacting gases at a given temperature and volume, is equal to the sum of the partial pressures of the component gases.

15. Give the relationship between partial pressure and mole fraction of a gas.

A. Partial pressure = Mole fraction $\times$ total pressure

16. Find the kinetic energy of 2 moles of an ideal gas inc calories at $27°C$?

A. Kinetic energy = $\frac{3}{2}nRT$

n = 2 moles, $T = 27°C = 27 + 273 = 300k$

$= \frac{3}{2} \times 2 \times 2 = 300 \left(R = 2 calories/mol \right)$ - Kelvin

Kinetic energy of 2 moles gas = 1800 cal

17. What is Boltzmann constant ? Give its value ?

A. The gas constant per one molecule is called Boltzmann constant $K = \frac{R}{N}$

Its value is $1.38 \times 10^{-23} joule\, k^{-1} mol^{-1}$ (or) $1.38 \times 10^{-16} erg\, k^{-1} mol^{-1}$

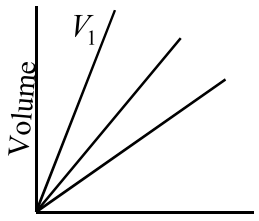
Short Answer Questions :

1. Write a brief note on 'Absolute temperature'.

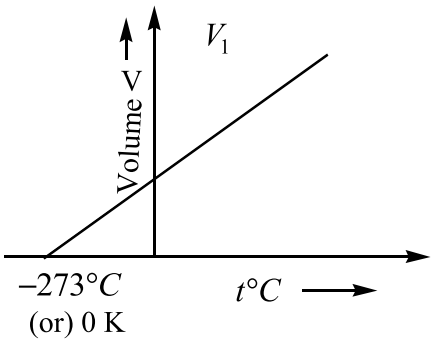
A. "The volume of a given mass of a gas is directly proportional to absolute temperature at constant pressure".

i.e, $\frac{V}{T}$ = constant

The curves plotted between volume and absolute temperature at constant pressure are called "isobars". These curves are straight lines passing through the origin.



When the line is extended to the lower temperature, it touches temperature axis at $-273.15°C$. It represents zero volume.



"The temperature at which the volume of the gas is zero temperature"

It was suggested by Lord Kelvin hence it is known as "Kelvin scale" or "Absolute scale". 0K or $0°A$ $0°A = -273.15°C$

2. Derive ideal gas equation.

A. By combining Boyle's law, Charles' law and Avogadro's law, we get an equation which relates to volume, pressure, absolute temperature and number of moles. This equation is known as ideal gas equation.

$V \propto \frac{1}{P} \longrightarrow$ Boyle's law

$V \propto T \longrightarrow$ Charles law

$V \propto n \longrightarrow$ Avogadro's law

Combining the above three laws, we can write

$V \propto \frac{1}{P} \times T \times n$ (or) $V = R \times \frac{1}{P} \times T \times n$ (or) PV = nRT

Where V = volume of the gas, P = Pressure of the gas, n=Number of moles of gas, T= Absolute temperature and R = universal gas constant.

3. State and explain Graham's law of diffusion.

A. **Graham's law of diffusion :** It states that "At constant temperature and pressure, the rate of diffusion of a gas is inversely proportional to the square root of its density"

Rate of diffusion $(r) \propto \frac{1}{\sqrt{d}}$

1) At constant temperature and pressure, if r_1 and r_2 are rates of diffusion of two gases and densities are d_1 and d_2, according to Graham's law

$$r_1 \propto \frac{1}{\sqrt{d_1}} \text{ and } r_2 \propto \frac{1}{\sqrt{d_2}}$$

$$\therefore \quad \frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} \quad \text{..... (1)}$$

2) At constant temperature and pressure, the molar volumes of two gases are same

$d_1 = \frac{M_1}{V}$ and $d_2 = \frac{M_2}{V}$ where M_1, M_2 are molecular weights of gases

$$\frac{d_1}{d_2} = \frac{M_1}{M_2}$$

From equation (1)

$$\therefore \frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{V.D_2}{V.D_1}} \quad \text{.....(2)}$$

$\therefore$ The rate of diffusion of a gas is inversely proportional to square root of its molecular weight, i.e, also $\therefore \quad r \propto \frac{1}{\sqrt{M}}$

3) The relation between molecular weight and vapour density is

Molecular weight = 2 $\times$ vapour density

$\therefore \quad \frac{M_1}{M_2} = \frac{V.D_1}{V.D_2}$ where $V.D_1$ and $V.D_2$ are vapour densities of two gases

From equation (2)

$$\therefore \frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}} = \sqrt{\frac{V.D_2}{V.D_1}}$$

$\therefore$ The rate of diffusion is also inversely proportional to square root of vapour densities of gases.

$$\therefore \quad r \propto \frac{1}{\sqrt{V.D}}$$

4. State and explain Dalton's law of partial pressures.

A. "The total pressure exerted by a mixture of chemically non – reacting gases at given temperature and volume, is equal to the sum of the partial pressures of the component gases."

Explanation : Consider a gaseous mixture of the three gases which do not chemically react with one another present in a vessel with volume 'V' at temperature T. n_1, n_2, n_3 are the number of moles of the three gases respectively.

$$P_1 = \frac{n_1 RT}{V} \text{ …….. (1)}$$

$$P_2 = \frac{n_2 RT}{V} \text{ …….. (2)}$$

$$P_3 = \frac{n_3 RT}{V} \text{ …….. (3)}$$

According to Dalton's law the pressure exerted by the gaseous mixture is $p = p_1 + p_2 + p_3$

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

But $n_1 + n_2 + n_3 = n$ ($n \Rightarrow$ total number of moles present in the mixture)

$$P = \frac{nRT}{V} \text{ ……. (4)}$$

Equation (1) divided by (4)

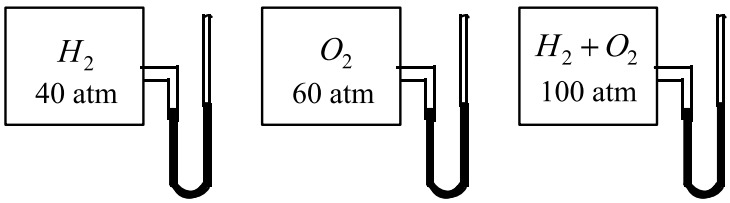
$$\frac{P_1}{P} = \frac{n_1 RT}{V} \times \frac{V}{nRT} = \frac{n_1}{n} = x_1 \text{ (or) } P_1 = P \times x_1 \text{ where } \left(x_1 = \frac{n_1}{n} = \text{mole fraction of first gas} \right)$$

Similarly

$$P_2 = P \times x_2$$

$$P_3 = P \times x_3$$

Partial pressure = Mole fraction × Total pressure



5. How is the pressure of gas calculated when it is collected over water, using Dalton's law of partial pressures ?

A. Dalton's law of partial pressures is useful in calculating the pressure of the gases collected over water in the laboratory. The mixture of gas and water vapour has a total pressure equal to atmospheric pressure when the water level is same inside and outside the vessel.

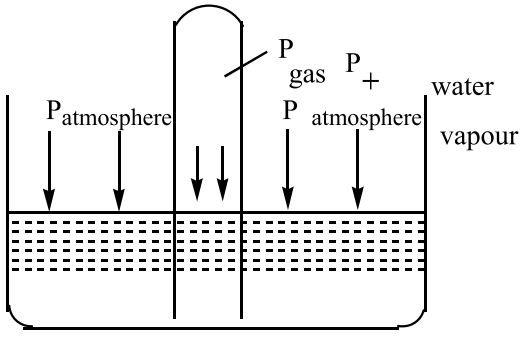
The pressure of saturated water vapour at a given temperature is called aqueous tension.

The pressure due to the gas alone can be obtained

$$P_{gas} = P_{atmosphere} - \text{aqueous solution}$$

i.e, Actual pressure of the gas = pressure of the gas noted at laboratory conditions – aqueous solution

Aqueous tension varies with temperature. It increases with an increase in temperature and is equal to the atmospheric pressure (760mm of Hg) at $100°C$



Long Answer Questions :

1. State and explain Boyle's, Charles and Avogadro's laws. Derive ideal gas equation.

A. **AVOGADRO'S LAW :**

Equal volumes of all gases under similar conditions of temperature and pressure contain equal number of molecules or moles.

i.e, $V \propto n$ (at constant P and T)

$$V = Kn = K\frac{m}{M} \quad \left(n = \frac{m}{M}\right)$$

$$M = k\frac{m}{V} = kd \quad \frac{m}{V} = d$$

BOYLE'S LAW :

This law gives relation between volume and pressure of gas at constant temperature

The volume (V) of a given mass of a gas at constant temperature is inversely proportional to the pressure (P)

$V \propto \frac{1}{P}$ (at constant T and m) or $V = \frac{K}{P}$ or PV = K where 'K' is constant

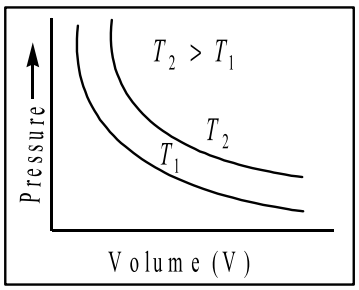
At constant temperature, the product of volume and pressure of a mass of gas is constant

i.e, P_1V_1 = constant and P_2V_2 = constant

$\therefore \quad P_1V_1 = P_2V_2$

P_1 = Initial pressure $\qquad$ P_2 = Final pressure

V_1 = initial volume $\qquad$ V_2 = Final volume



The curves plotted between volume and pressure at constant temperature are called "isotherms". The shape of isotherm is "rectangular hyberbola".

CHARLE'S LAW: At constant pressure, the volume of given mass of gas at $0°C$, increases or decreases by $1/273$ times of its original volume for every degree raise or fall in temperature respectively.

At $0°C$ the volume of gas = V_o

At $1°C$ the increase in volume = $\frac{V_o}{273}$

$\therefore$ At $1°C$ the total volume = $V_o + \frac{V_o}{273} = V_o\left(1 + \frac{1}{273}\right)$

At $t°C$ volume of gas, $V = V_o\left(1 + \frac{t}{273}\right) = V_o\left(\frac{273 + t}{273}\right)$

At $t_1°C$ the volume, $V_1 = V_o\left(\frac{273 + t_1}{273}\right)$

At $t_2°C$, the volume, $V_2 = V_o\left(\frac{273 + t_2}{273}\right)$

Then $\frac{V_1}{V_2} = \frac{(273 + t_1)}{(273 + t_2)}$

$\therefore \quad \frac{V_1}{V_2} = \frac{T_1}{T_2}$ i,e, absolute temperature, $T = t°C + 273$ or $V \propto T$ (at constant P and m)

V_1 = Initial volume V_2 = Final volume

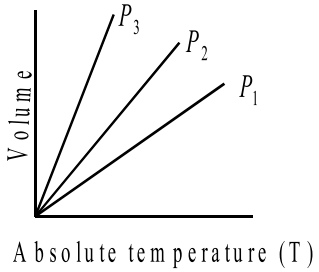
T_1 = Initial temperature T_2 = Final temperature

$\therefore$ Charles law can be defined as,

The volume of a given mass of a gas is directly proportional to absolute temperature at constant pressure.

i.e, $\frac{V}{T}$ = constant

The curves plotted between volume and temperature at constant pressure are called as "isobars". These curves are straight lines passing through the origin.



Ideal gas :

A gas which follows Boyle's law, Gay – Lussac's law and Avogadro's law under all conditions of temperature and pressure is called an ideal gas.

IDEAL GAS EQUATION OR PERFECT GAS EQUATION :

The ideal gas equation is obtained by combining Boyle's law, Charle's law and Avogadro's law.

If V is the volume, P is the pressure. T is the absolute temperature and 'n' the number of moles of the ideal gas.

$V \propto \frac{1}{P}$ (T, n are constant) Boyle's law

$V \propto T$ (P, n are constant) Gay – Lussac's law (for Charles law)

$V \propto n$ (P, T are constant) Avogadro's law

$\therefore V \propto \frac{nT}{P} \Rightarrow V = R\frac{nT}{P}$ or $PV = nRT$ or $R = \frac{PV}{nT}$

This is called **"ideal gas equation"** or **"equation of state"** where 'R' is called gas constant and independent of the amount of gas.

Value of gas constant – R :

The value of gas constant 'R' is same for all gases. So, it is called as **"Universal gas constant"** or **"Molar gas constant"**

The value of gas constant depends on units of measurement i.e, units of pressure and volume.

CHAPTER IV - STATES OF MATTER (GASES AND LIQUIDS)

UNIT: 2

1 Kinetic molecular theory of gases

2 Kinetic gas equation- Deduction of gas laws from kinetic gas equation

3 Distribution of molecular velocities and types of molecular velocities

Very Short Answer Questions:

1. What is most probable velocity? Write any one formula used in calculating it?

A. At a given temp the velocity possessed by the maximum number of molecules is called most probable velocity.

$$C_p = \sqrt{\frac{2RT}{M}}$$

R = Gas constant, T = Absolute temperature, M = Molecular weight

2. Give the formulae to calculate average velocity of gas molecules.

A. Average velocity $\left(\overline{C}\right) = \frac{C_1 + C_2 + C_3 ..... + C_n}{n}$

$$= \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8PV}{\pi M}} = \sqrt{\frac{8P}{\pi d}} = 0.9123 \times \text{RMS velocity}$$

R = Gas constant, P = pressure, T = Absolute temperature, d = density, M = Molecular weight

3. With what factor RMS velocity is multiplied to give average velocity and most probable velocities of gas molecules.

A. Average velocity = $0.9123 \times \text{RMS velocity}$

Most probable velocity = $0.8166 \times \text{RMS velocity}$

4. How is pressure of gas created?

A. The pressure exerted by a gas is due to its collisions made by gas molecules on the walls of the container.

5. Give the ratio of molecular velocities.

A. $C_p : \overline{C} : C = 1 : 1.128 : 1.224$

6. Why does the velocity of gas molecules increase with an increase in temperature?

A. The kinetic energy of gas molecules increases with the increase in temperature. Therefore the velocity of gas molecules increases in their movement,. it is because average $KE \propto T$

7. Give the mathematical relation of RMS velocities of two different gases at a given temperature.

A. $C_1 = \sqrt{\frac{3RT}{M_1}}$; $C_2 = \sqrt{\frac{3RT}{M_2}}$

R and T are constant hence $\frac{C_1}{C_2} = \sqrt{\frac{M_2}{M_1}}$

8. Why RMS velocity is used in the derivation of kinetic gas equation.

A. Velocity is a vector quantity. The molecules of a gas will move in all possible directions. In one direction if the velocity is taken as positive, in the opposite direction it is negative. The average may sometimes becomes zero. If we square the values then it is positive. Hence, square root of average of the squares of the velocities is taken.

Short Answer Questions:

1. Write five important postulates of 'Kinetic molecular theory' of gases.

A. **Postulates:**

1) Gases are composed of minute particles called molecules. All the molecules of a gas are identical.
2) Gaseous molecules are always, at a random movement. The molecules are moving in all possible directions in straight lines with very high velocities.

They keep on colliding against each other and against the walls of the vessel at very small intervals of time.

3) The actual volume occupied by the molecules is negligible when compared to the total volume occupied by the gas.
4) There is no appreciable attraction or repulsion between the molecules
5) There is no loss of kinetic energy when the molecules collide with each other or with the wall of vessel. This is because the molecules are spherical and perfectly elastic in nature.
6) The pressure exerted by the gas is due to the bombardment of the molecules of the gas on the walls of the vessel.
7) The average kinetic energy of the molecules of the gas is directly proportional to the absolute temperature, Average $K.E \propto T$
8) The force of gravity has no effect on the speed of gas molecules.

2. Deduce Boyle's law from kinetic gas equation.

A. $PV = \frac{1}{3}mnC^2$

P = Pressure of the gas

V = Volume of the gas

m = Mass of gas molecule

n = number of molecules

C = r.m.s velocity

Deduction of Boyle's law from K.G.E:

According to kinetic gas equation, $PV = \frac{1}{3}mnC^2$ (or) $PV = \frac{2}{3} \times \frac{1}{2}mnC^2$

or $PV = \frac{2}{3}K.E$ …….. (1)

$\frac{1}{2}mnC^2$ is the kinetic energy of 'n' molecules

But the average kinetic energy of a gas molecules is directly proportional to the absolute temperature

i.e, K.E $\propto$ T (or) K.E. = K.T …… (K constant) (2)

From equations (1) and (2)

$$PV = \frac{2}{3}KT$$

At constant temperature, PV = constant. This is Boyle's law.

3. Deduce Charle's law from kinetic gas equation.

A. According to kinetic gas equation,

$$PV = \frac{1}{3}mnC^2 \text{ (or) } PV = \frac{2}{3}\times\frac{1}{2}mnC^2$$

(or) $PV = \frac{2}{3}K.E$ …….. (1)

$\frac{1}{2}mnC^2$ is the kinetic energy of 'n' molecules.

But the average kinetic energy of a gas molecules is directly proportional to the absolute temperature.

i.e, i.e, K.E $\propto$ T (or) K.E. = K.T …… (2)

From equations (1) and (2)

$$P = \frac{2}{3}\frac{KT}{P} \quad \text{(or)} \quad L = \frac{2}{3}\frac{< T}{P}$$

At constant pressure, V = constant $\times$ T or $V \propto T$. This is Charle's law.

4. Deduce an expression for kinetic energy from kinetic gas equation.

A. In kinetic gas equation, $PV = \frac{1}{3}mnC^2$ if the number of molecules 'n' is equal to Avogadro's number, 'mn' becomes equal to gram molecular mass or molar mass 'M' of the gas

$$PV = \frac{1}{3}MC^2 = \frac{2}{3}\times\frac{1}{2}MC^2$$

$= \frac{2}{3}\times K.E$ (or) $K.E = \frac{3}{2}PV$

K.E. is the kinetic energy of one mole of a gas. For one mole of gas, ideal gas equation is

PV = RT …… (2)

From equations (1) and (2), $K.E = \frac{3}{2}RT$ …….. (3)

The value of kinetic energy of one mole of gas depends on the units of gas constant 'R' and directly proportional to the absolute temperature

$K.E \propto T$

For one mole of a gas, K.E. is the independent of the nature of gas, i.e, one mole of any gas has the same kinetic energy at a given temperature.

For one molecule, $K.E = \frac{3}{2}\frac{RT}{N}$

or $K.E = \frac{3}{2}KT$ …….. (4)

where k = Boltzmann's constant = $\frac{R}{N}$

N = Avogadro number

Boltzmann's constant is the gas constant per molecule and its value is

$1.38 \times 10^{-16} erg.K^{-1}.molecule^{-1}$

or $1.38 \times 10^{-23} joule.K^{-1}.molecule^{-1}$

For 'n' moles of gas $K.E = \frac{3}{2}nRT$ ……. (5)

5. Write different molecular velocities of gases and give their inter-relationship.

A. Molecular velocities are of three types. They are

1) RMS velocity

2) Average velocity

3) Most probable velocity

1. RMS velocity (C):

It is defined as, "the square root of the average of the squares of velocities of all gas molecules present in the given sample of gas"

$$C = \sqrt{\frac{C_1^2 + C_2^2 + C_3^2 + ... + C_n^2}{n}}$$

C_1, C_2, C_3 are the individual velocities, 'n' is the total number of molecules.

2. Average velocity:

The mean of velocities of all the gas molecules in the given same of the gas is called average velocity.

$$\left(\overline{C}\right) = \frac{C_1 + C_2 + C_3 + ... + C_n}{n}$$

C_1, C_2, C_3 are the individual velocities, 'n' is the total number of molecules.

3. Most probable velocity:

The velocity possessed by the most of the gas molecules in the given sample of gas is called most probable velocity.

Most probable velocity $\left(C_p\right) = 0.8166 \times \text{RMS velocity}$

Most probable velocity $\left(C_p\right) = 0.8865 \times \text{Average velocity}$

Ratio of molecular velocity:

$$C_p : \overline{C} : C = \sqrt{\frac{2RT}{M}} : \sqrt{\frac{8RT}{M}} : \sqrt{\frac{3RT}{M}}$$

$$= \sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3}$$

= 1 : 1.128 : 1.224

Long Answer Questions:

1. Give the important assumptions of kinetic molecular theory of gases and explain them.

A. **Assumptions:**

1) Gases are composed of minute particles called molecules. All the molecules of a gas are identical.
2) Gaseous molecules are always, at a random movement. The molecules are moving in all possible directions in straight lines with very high velocities. They keep on colliding against each other and against the walls of the vessel at very small intervals of time.
3) The actual volume occupied by the molecules is negligible when compared to the total volume occupied by the gas.
4) There is no appreciable attraction or repulsion between the molecules
5) There is no loss of kinetic energy when the molecules collide with each other or with the wall of vessel. This is because the molecules are spherical and perfectly elastic in nature.
6) The pressure exerted by the gas is due to the bombardment of the molecules of the gas on the walls of the vessel.
7) The average kinetic energy of the molecules of the gas is directly proportional to the absolute temperature, Average $K.E \propto T$
8) The force of gravity has no effect on the speed of gas molecules.

Boyle's law: According to kinetic theory of gases, the pressure of the gas is due to collisions of gas molecules on the walls of the vessel. At a particular temperature the molecules make definite number of collisions with the walls of the vessel. When the volume of the vessel is reduced the molecules have to travel lesser distance only before making collisions on the walls. As a result the number of collisions per unit increases. The pressure then increases i.e, the pressure increases when the volume is reduced at constant temperature. This explains Boyle's law.

Charle's law: According to kinetic theory of gases, the average kinetic energy of the molecules is directly proportional to the absolute temperature of the gas.

$$K.E \propto T$$

But $K.E = \frac{1}{2}mc^2$

As temperature increases, the velocity of the molecules also increases. As a result the molecules makes more number of collisions against the walls of the vessel. This results in an increase of pressure if the volume is kept constant. If the volume is allowed to increase the number of collisions decrease due to the increased distance between the molecules and the walls of the vessel. The pressure then decreases. In other words, with rise of temperature, the volume should increase in order to keep the pressure constant.

$V \propto T$ at constant pressure

This is Charle's law.

2. Deduce (a) Boyle's law (b) Charles law and (c) Graham's law from kinetic gas equation.

A. **KINETIC GAS EQUATION:**

It is a mathematical expression for kinetic molecular theory of gases.

(or) $PV = \frac{1}{3}mnC^2$

P = Pressure of the gas

V = Volume of the gas

m = Mass of gas molecule

n = number of molecules

C = r.m.s velocity

Deduction of Boyle's law from K.G.E :

According to kinetic gas equation, $PV = \frac{1}{3}mnC^2$ (or) $PV = \frac{2}{3} \times \frac{1}{2}mnC^2$

or $PV = \frac{2}{3}K.E$ …….. (1)

$\frac{1}{2}mnC^2$ is the kinetic energy of 'n' molecules

But the average kinetic energy of a gas molecules is directly proportional to the absolute temperature

i.e, K.E $\propto$ T (or) K.E. = K.T …… (K constant) (2)

From equations (1) and (2)

$$PV = \frac{2}{3}KT$$

At constant temperature, PV = constant. This is Boyle's law.

Deduction of Charle's law from KGE:

According to kinetic gas equation,

$$PV = \frac{1}{3}mnC^2 \text{ (or) } PV = \frac{2}{3} \times \frac{1}{2}mnC^2$$

(or) $PV = \frac{2}{3}K.E$ …….. (1)

$\frac{1}{2}mnC^2$ is the kinetic energy of 'n' molecules.

But the average kinetic energy of a gas molecules is directly proportional to the absolute temperature.

i.e, i.e, K.E $\propto$ T (or) K.E. = K.T …… (2)

From equations (1) and (2)

$$P = \frac{2}{3}\frac{KT}{P} \quad \text{(or)} \quad L = \frac{2}{3}\frac{< T}{P}$$

At constant pressure, V = constant $\times$ T or $V \propto T$. This is Charle's law.

Deduction of Graham's law of diffusion from KGE:

According to kinetic gas equation, $PV = \frac{1}{3}mnC^2$

If the number of molecules 'n' is equal to Avogadro's number, 'mn' becomes equal to gram molecular weight 'M' of gas.

$\therefore\ PV = \frac{1}{3}MC^2$ (or) $\therefore\ C^2 = \frac{3PV}{M}$

(or) $\therefore\ C = \sqrt{\frac{3PV}{M}}$

(or) $C = \sqrt{\frac{3P}{d}}\ \left(\because \frac{V}{M} = \frac{1}{d}\right)$

At constant pressure, $C = \frac{\text{constant}}{\sqrt{d}}$ (or) $C \propto \frac{1}{\sqrt{d}}$

i.e, RMS velocity or rate of diffusion of gas inversely proportional to the square root of its density. This is "Graham's law of diffusion".

3. Deduce Dalton's law of partial pressures and Graham's law of diffusion from kinetic gas equation.

A. **Deduction of Graham's law of diffusion from KGE:**

According to kinetic gas equation, $PV = \frac{1}{3}mnC^2$

If the number of molecules 'n' is equal to Avogadro's number, 'mn' becomes equal to gram molecular weight 'M' of gas.

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i.e, RMS velocity or rate of diffusion of gas inversely proportional to the square root of its density. This is "Graham's law of diffusion".

Deduction of Dalton's law of partial pressures from KGE :

' n_1 ' molecules of a gas is present in a vessel of volume 'V'. If mass of each molecule is ' m_1 ' and C_1 is the RMS velocity, the pressure of the gas according to kinetic gas equation

$$P_1 = \frac{1}{3}.\frac{m_1 n_1 C_1^2}{V}$$

If another gas is present in the same vessel instead of first gas, the pressure of that gas according to kinetic gas equation.

$$P_2 = \frac{1}{3}.\frac{m_2 n_2 C_2^2}{V}$$

If the mixture of these two gases are taken in the same vessel of volume 'V', the total pressure of the mixture is

$$P = \frac{1}{3}.\frac{m_1 n_1 C_1^2}{V} + \frac{1}{3}.\frac{m_2 n_2 C_2^2}{V}$$

(or) $P_1 = P_1 + P_2$

This is Dalton's law of partial pressures

4. What are RMS velocity, most probable velocity and average velocity of gas molecules?

A. **TYPES OF MOLECULAR VELOCITIES:**

Molecular velocities are of three types. They are

1) RMS velocity

2) Average velocity

3) Most probable velocity

1. RMS velocity (C):

It is defined as, "the square root of the average of the squares of velocities of all gas molecules present in the given sample of gas"

$$C = \sqrt{\frac{C_1^2 + C_2^2 + C_3^2 + ... + C_n^2}{n}}$$

C_1, C_2, C_3 are the individual velocities, 'n' is the total number of molecules.

According to kinetic gas equation,

$$PV = \frac{1}{3}mnC^2$$

$$PV = \frac{1}{3}MC^2 \quad (\because \text{ mn = M molar mass gas})$$

$$\text{or } C^2 = \frac{3PV}{M} \quad \text{or} \quad C = \sqrt{\frac{3PV}{M}}$$

$$\text{or } C = \sqrt{\frac{3P}{d}} \left(\because \frac{V}{M} = \frac{1}{d}\right)$$

$$\text{or } C = \sqrt{\frac{3RT}{M}} \left(\because PV = RT \text{ for 1 mole of gas}\right)$$

RMS velocity = 1.224 × Most probable velocity

RMS velocity = 1.085 × Average velocity

2. Average velocity:

The mean of velocities of all the gas molecules in the given same of the gas is called average velocity.

$$\left(\overline{C}\right) = \frac{C_1 + C_2 + C_3 + ... + C_n}{n}$$

C_1, C_2, C_3 are the individual velocities, 'n' is the total number of molecules.

Average velocity

$$\left(\overline{C}\right) = \sqrt{\frac{8RT}{\pi M}} = \sqrt{\frac{8PV}{\pi M}} = \sqrt{\frac{8P}{\pi d}}$$

Average velocity $\left(\overline{C}\right) = 0.913 \times \text{RMS velocity}$

Average velocity $\left(\overline{C}\right) = 1.128 \times \text{Most probable velocity}$

3. Most probable velocity $\left(C_p\right)$:

The velocity possessed by the most of the gas molecules in the given sample of gas is called most probable velocity.

Most probable velocity $\left(C_p\right) = 0.8166 \times \text{RMS velocity}$

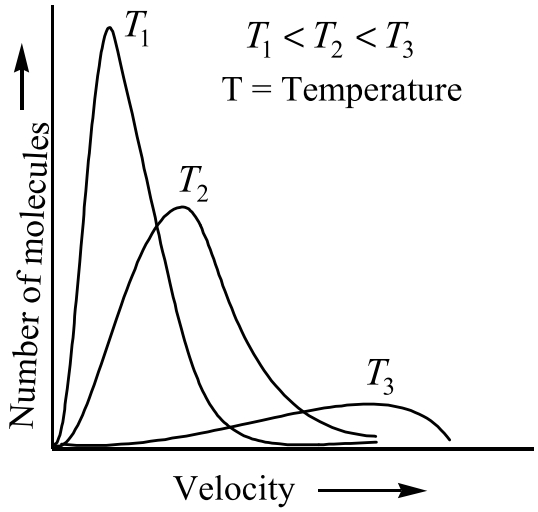
Most probable velocity $\left(C_p\right) = 0.8865 \times \text{Average velocity}$

5. Explain Maxwell-Boltzmann distribution curves of molecular velocities and give the important conclusions. Discuss the effect of temperature on the distribution of molecular velocities.

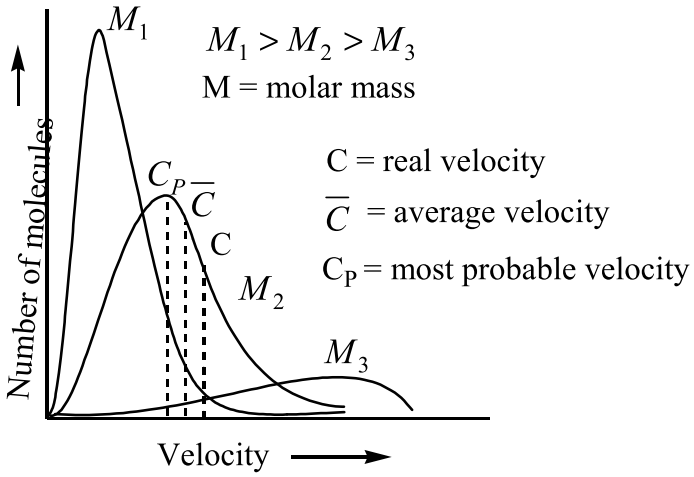
A. According to kinetic gas equation it was assumed that all the molecules in a gas have the same velocity. But it is not correct, when any two molecules collide exchange of energy takes place and hence their velocities keep on charging. At any instant few molecules may have zero velocity, a few molecules may be at high velocities and some may be with low velocities.

Maxwell worked out the distribution of speeds between different molecules by applying probability considerations.

If one plots a graph between fraction of molecules $\frac{\Delta N}{N}$ vs velocity one gets distribution curve of the type.



The Maxwell distribution of velocities and its variation with the temperature



The Maxwell distribution of velocities
and its variation with molar mass

These curves are shown at different temperatures $T_1, T_2 \left(T_1 < T_2\right)$

The graph reveals that

1) There are no molecules with zero velocity and only very few molecules possess the highest velocity
2) The velocities of most of molecules lie near a mean value
3) As the temperature of the gas is increased, the curve becomes more flattened and shifts towards higher velocity

It means that at higher temperature the number of molecules possessing higher velocities is more than at lower temperature.

The peak point corresponds to the most probable velocity. It is the velocity possessed by maximum number of molecules.

The average velocity of the molecules is slightly higher than the most probable velocity. The RMS velocity is slightly higher than the average velocity.

CHAPTER IV - STATES OF MATTER (GASES AND LIQUIDS)

UNIT: 3

1 Behaviour of real gases, derivation of from ideal behaviour, compressibility factor Vs Pressure – diagrams of real gases.

2 Conditions of liquefaction of gasses, critical temperature.

3 Liquid state – properties of liquids in terms of inter molecular attractions.

4 Vapour pressure, viscosity and surface tension

Very Short Answer Questions:

1. Define compression factor. What is its value for a perfect gas?

A. It is the ratio of the actual molar volume of a gas to the molar volume of a perfect gas under the same conditions.

$$Z = \frac{V_m}{\frac{RT}{P}} = \frac{PV_m}{RT}$$

Compression factor value is one for a perfect gas

2. Explain the different terms in Vander Waals equation.

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

P = pressure of the gas, V = volume of the gas, a = correction factor in pressure, b = correction factor in volume, n = number of moles, R = Universal gas constant. T = Absolute temperature

3. Define critical temperature. Why were hydrogen like gases considered as permanent gases in the beginning?

A. It is the highest temperature at which a gas can be liquefied by the application of external pressure. Above critical temperature a gas can not be liquefied even by applying high pressure.

Since the critical temperatures are much lower than the room temperature, Hydrogen like gases are called permanent gases.

4. What is Joule – Thomson effect? How does it help in the liquefaction of gases?

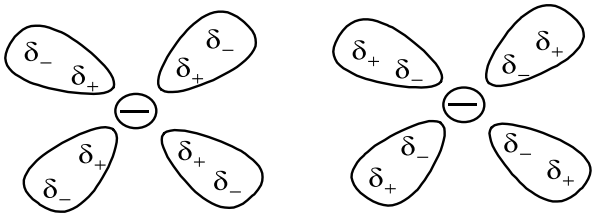
A. When a gas is allowed to expand from high pressure to low pressure, cooling takes place. This is known as Joule – Thomson effect.

When a gas is allowed to expand through a narrow opening called throttle without supplying any heat from outside. It has to do some work in over coming the attractions of their neighbours and in doing so, the molecules convert some of their kinetic energy into potential energy. Thus temperature decreases.

Short Answer Questions:

1. Explain ion – dipole attractions with examples.

A. **ION DIPOLE FORCES:**



A. Ion- Dipole forces:

These forces are mainly present in aqueous solutions of ionic substances.

Ex: NaCl in water solution.

Water is a polar molecule and in it 'H' atoms possess partial +Ve charge and 'O' atoms

Possess partial –ve charge. When ionic compounds like NaCl dissolve in water, they dissociate into component ions like $Na^{+}\ and\ Cl^{-}$. Now, the water molecules orient in the presence of ions in such a way that the + ve end of the dipole is near an anion and the –ve end of the dipole is near a cation. The magnitude of interaction energy depends on the charge on the ion (Z), the strength of the dipole (μ) and on the inverse square of the distance (r) between the ion and the dipole.

It can be expressed mathematically as, $E = Z\mu / r^2$.

2. Explain induced dipole – dipole attraction.

A. To explain the intermolecular forces among individual atoms or non-polar molecules, London dispersion forces have been proposed. These forces result from the motion of electrons in atoms. At a given instant the electron distribution in an atom may be unsymmetrical giving the atom a short-lived dipole moment. This instantaneous dipole on one atom can affect the electron distribution in neighbouring atoms and induct temporary dipoles in its neighbours. As a result of which weak attractive forces develop among the molecules. Thy are known as London forces or dispersion forces. These forces are smaller and are in the range-1- 10 KJ/mole. The exact magnitude depends on a property known as polarisability. A smaller molecular or atom is less polarisable and has smaller dispersion forces. A larger molecule or heavier atom is more polarisable and has large dispersion forces.

These forces are always attractive and inversely proportional to the sixth power of distance between the two interacting particles $\left(\propto \frac{1}{r^6} \right)$

3. Why experimental isotherms of real gases resemble the perfect – gas isotherms at high temperature and low pressures?

A. A) **Isotherms of CO_2 and critical temperature:**

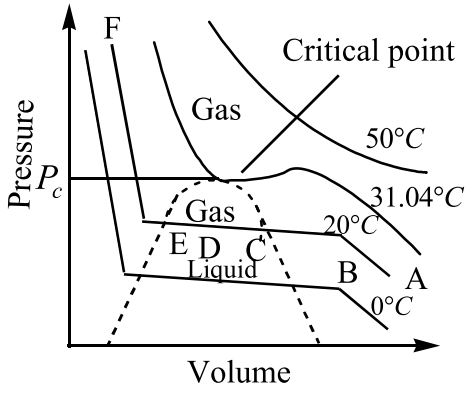
i) Isotherms of CO_2 have shapes different from those implied by Boyle's law, particularly at high pressures and low temperature.

ii) Isotherm of CO_2 at $50°C$:

It is a rectangular hyperbola and approximates to the isotherm of ideal gas

iii) Isotherm of CO_2 at $20°C$:

At point 'A' the sample of CO_2 is gas. As the sample is compressed to B by pressing in a piston, the pressure increases and it continues until the sample reaches C (according to Charles law)



After this, the piston can be pressed in without any further increase in pressure through D to E because the gas is completely condensed to compact liquid at E, where the piston rests the surface of liquid. From E to F the compression is very small through there is very high pressure increase.

C to E corresponds to the molecules being so close and they attract each other and cohere into a liquid.

E to F shows that as already molecules are in contact still they are pushed closer, repulsive forces operate and to overcome these forces very high pressure is required.

At D the container contains liquid and gas separated by a sharp surface

iv) Isotherm of CO_2 at $30°C$: At $30°C$ slightly higher temperature the liquid is formed at higher pressure. It is difficult to make out the surface because the remaining gas is such at a high pressure that its density is almost equal to that of the liquid.

v) At $31.04°C$ the gases in state transforms continuously into the condensed state and the surface between liquid and gas cannot be visible. This is considered as critical temperature.

vi) The gas can not be liquefied by the application of only pressure without cooling it below the critical temperature.

vii) Ex: Liquid N_2 can't liquefied by increase in pressure unless it is cooled below (126k (or) $-147°C$)

viii) The dense fluid obtained by compressing a gas above its critical temperature is not a true liquid because no surface that separates it from vapour phase. It is called super critical fluid. It can be used as solvent.

4. Explain the deviation of real gas behaviour from perfect gas using compressions factor of value.

A. The deviation of real gases from ideal behaviour can be studied in terms of quantity Z which is called compression factor.

It is the ratio of the actual molar volume of a gas to the molar volume of a perfect gas under the same conditions.

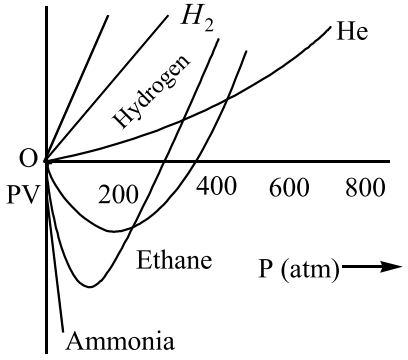
A convenient way of showing the deviation of real gases from ideal behavior is to plot a graph between 'Z' and 'P'

In the case of a perfect gas

i) Z = 1 at all temperatures and pressures because PV = nRT

ii) The graph of Z vs P will be a straight line parallel to pressure axis and attractive interactions are dominant.

In the case of real gases:



(a)

In the case of a perfect gas

i) Z = 1 at all temperature and pressures because PV = nRT

ii) The graph of Z vs P will be a straight line parallel to pressure axis

In the case of real gases:

i) H_2 always show positive deviation at all pressures i.e, Z > 1. It means it is less compressible than expected from ideal behaviour

All other gases show

i) Negative deviation (Z < 1) at low pressures

ii) Positive deviation (Z > 1) at high pressures

The negative deviations are a result of the attractive interactions between molecules and the positive deviations are as a result of the repulsive interactions.

In general,

i) At very low pressure all gases have almost ideal behaviour

ii) At high pressure, (Z > 1) it is difficult to compress

iii) At intermediate pressures most gases have Z < 1. They are more compressible than expected from ideal behaviour. Eg CH_4 and CO_2 (at low and moderate pressures)

iv) The extent of deviation from ideal behaviour becomes less and less with increase in temperature and decrease in pressure.

5. Explain the viscosity of liquids?

A. It is well known that all liquids do not flow with the same speed. Some liquids like water, alcohol, ether etc. flow very rapidly, while some one like glycerine, honey, castor oil etc. flow slowly. This indicates that every liquid has some internal resistance to flow. This internal resistance to flow possessed by a liquid is called its viscosity.

Liquids which flow rapidly have low internal resistance. So their viscosity is less. Liquids which flow slowly have high internal resistance. So their viscosity is high.

Coefficient of viscosity is the force per unit area required to maintain unit difference of velocity between two parallel layers in the liquid, one unit apart.

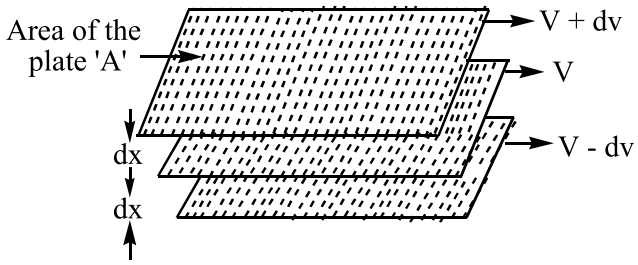
Units: Dyne cm^{-2} (C.G.S) or Poise

Millipoise = 10^{-3} Poise.

$Ns\,m^{-2}$ or Pa s (pascal second) is S.I unit.

Examples: 1) Glass is not a solid. It is a super –cooled liquid with a very high viscosity

2) H_2SO_4 is viscous, due to H – bonding.



Gradation of velocity in the laminar flow

Long Answer Questions:

1. Define Critical pressure and Critical volume?

A. **Critical pressure (P_c):**

The minimum pressure required to liquefy a gas at the critical temperature is known as critical pressure.

$$P_c = \frac{a}{27b^2}$$

Critical volume (V_c):

The volume occupied by 1 mole of a gas at critical temperature and critical pressure is known as critical volume V_c

$$V_c = 3b$$

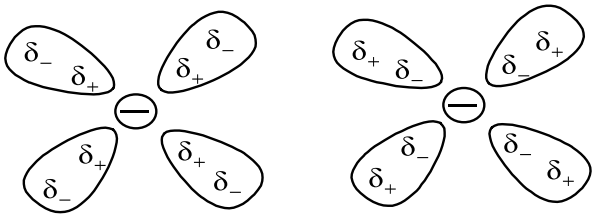
Note: T_c, P_c, V_c are collectively called critical constants

2. Write a note on intermolecular forces.

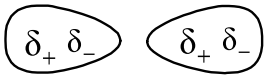
A. Inter molecular forces:

a) ION DIPOLE FORCES: Ion dipole forces are mainly important in aqueous solutions of ionic substances such as NaCl in which dioolar water molecules surround the ions.

Water molecules are polar and in them hydrogen atoms possess partial positive charges and oxygen atoms possess partial negative charges due to electronegativity difference between hydrogen and oxygen atoms. When ionic compounds like NaCl dissolve in water, they dissociate into component ions like Na^+ and Cl^-. Now the water molecules orient in the presence of ions in such a way that the positive end of the dipole is near an anion and the negative end of the dipole is near a cation.



b) DIPOLE FORCES: Neutral but polar molecules experience dipole – dipole forces. These are due to electrical interactions among dipoles on neighbouring molecules. These forces are again attractive between unlike poles and repulsive between like poles and depend on the orientation of the molecules. The net force in a large collection of molecules results from many individual interactions of both types. The forces are generally weak and are significant only when the molecules are in close contact.



c) LONDON./DISPERSION FORCES: These forces result from the motion of electrons around atoms. Take, for example, atoms of helium. The electron distribution around a helium atom is for averaged over time spherically symmetrical. However, at a given instant the electron distribution in an atom may be unsymmetrical giving the atom a short – lived dipole moment. The instantaneous dipole on one atom can affect the electron distribution is neighbouring atoms and induce temporary dipoles in those neighbours. As a result, weak attractive forces develop known as London forces or dispersion forces. London forces are generally small. Their energies are in the range $1-10\,kj\,mol^{-1}$



d) DIPOLE – INDUCED DIPOLE FORCES: These forces are between polar molecules with permanent dipole moments and the molecules with no permanent dipole moment. Permanent dipole of the polar molecules induces dipole on the electrically neutral molecule by deforming into electronic cloud.

Magnitude of these forces depends on the magnitude of the dipole moment of permanent dipole and polarisatricity of neutral molecule. This interaction is proportional to $\left(\frac{1}{r^2}\right)$,

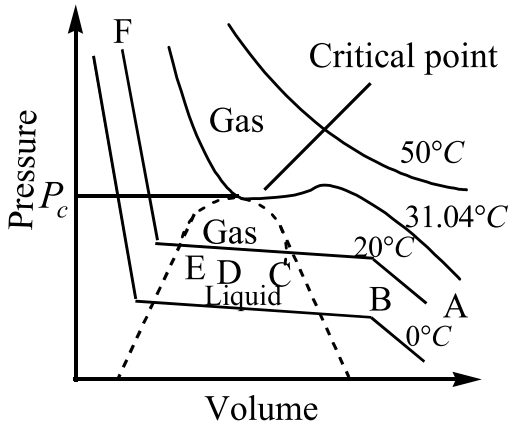
where r = distance between molecules.

3. Explain the isotherms of carbon dioxide? What is the critical temperature of a gas?

A. Consider the isotherm of $CO_2\,at\,20^oC$. At point A the sample is a gas. By pressing in a piston the sample is compressed to B and the pressing is continued till the sample reaches to

C. After this the sample is converted to D and further to E by pushing the piston in without increase in pressure. Now the sample E requires as per expectation high pressures, for conversion into E. Practically the sample starts condensing at C and the condensation gets completed when the piston is pushed into E.

At E, the piston rests on the surface of the liquid. From E to F the compression is very small though there is very high-pressure increase. The conversion from C to E corresponds to the molecules being so close that they attract each other and cohere into a liquid At D the container contains both the liquid and gas separated by a sharp surface. At 30^0C the liquid forms at a higher pressure. At 31.04^oC the gaseous state transforms continuously into the condensed state and at no stage we find visible surface between the liquid and the gas. This temperature is known as critical temperature (T_C) and above that temperature a single state of matter (gas) exists in the container. If may be concluded that a gas cannot be liquefied by the application of pressure alone, without cooling it below its critical temperature.



4. Using the compression factor explain the behaviour of real gases.

A. Compression factor (Z) is very important to discuss the properties of real gases. It is the ratio of the actual molar volume of a gas to the molar volume of a perfect gas under the same conditions

Compression factor (Z) $\dfrac{Molar\ volume\ of\ the\ gas\ (V_m)}{Molar\ \ volume\ of\ \ perfect\ gas\ (V_m perfect)}$

Molar volume of a perfect gas $(V_m\ perfect) = \dfrac{RT}{P}$

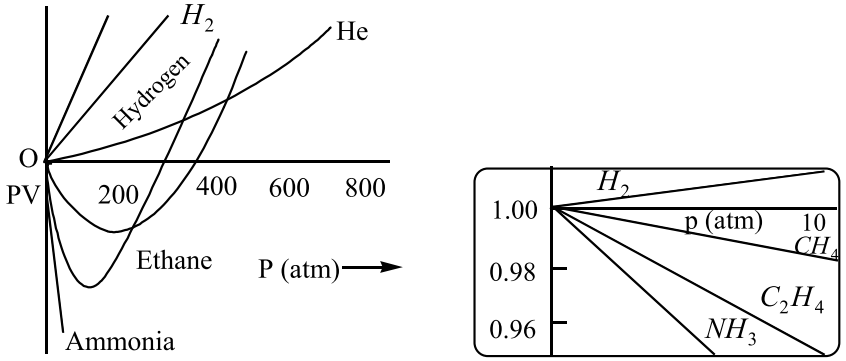
$$\therefore Z = \frac{\frac{V_m}{RT}}{P} = \frac{PV_m}{RT}$$

For a perfect gas, Z = 1 . So for other values of Z (other than 1) real gases differ from ideal behaviour.

For real gases, z varies with pressure. At low pressures some gases have Z<1. From this we can infer that their molar volumes are smaller than that of a perfect gas and the molecules cluster together slightly and attractive interactions are dominant. At high pressures Z>1 for almost all gases. Z>1 means that the molar volume of a gas is greater than that expected for a perfect gas. This is due to dominant repulsive forces which try to drive the molecules apart.

In general, at very low pressures all gases have almost ideal behaviour. At high pressures (Z=1). It is difficult to compress. At intermediate pressures most gases have Z < 1. Thus, gases show ideal behaviour when the volume occupied by them is large. Upto what pressure gases follow ideal behaviour, depends on the nature of the gas and its temperature.

The temperature at which a real gas obeys ideal gas laws over a wide range of pressure is called Boyle temperature or Boyle point and it depends on the nature of the gas. Above Boyle temperature real gases show +ve deviations from ideality and their Z values are greater than 1. The forces of attraction between the gas molecules are feeble. Below the Boyle temperature, all real gases first show Z < 1 with increase of pressure and reaches a minimum. On further increase of pressure Z continuously increases. Hence, It may be concluded that at low pressure and high temperature gases show ideal behaviour.



(a) (b)

5. Derive and explain Vander Waals equation of state.

A. **Vander Waals Equation:**

(Equation of state of real gases):

According to Vander Waal, the deviations of a gas from perfect gas equation are due to intermolecular interactions.

Vander Waals modified the perfect gas equation by introducing two correction terms, one for volume and the other for pressure to make the equation applicable to real gases as well.

(a) Volume correction:

The repulsive interactions between two molecules cannot allow them to get closer than a certain distance. Hence, for the gas molecules the available volume for the free movement is less than the volume of the container 'V'.

$\therefore$ In perfect gas equation, a volume correction is made by changing V to (V – nb)

'b' is the correction factor in volume. It is also known as excluded volume or covolume or Vander Waal constant. Excluded volume is four times to the actual volume of the molecules and it depends on the nature of the gas and 'n' is number of moles of the gas present in the vessel of volume 'V'

Therefore ideal volume Vi = (V – nb)

Note: 'b' is a measure of effective size of gas molecules

Pressure correction:

The effect of attractive interactions between molecules is to reduce the pressure that the gas exerts. As the attractions slow down the molecules, they strike the walls less frequently and the impact is weak. Thus there is reduction in pressure.

Reduction in pressure is directly proportional to square of the molar concentration (n/v)

$\therefore$ Reduction in pressure $\propto \left(\frac{n}{v}\right)^2$

or reduction in pressure $P = \frac{an^2}{v^2}$

where 'a' is a constant whose value depends upon the nature of the gas. It measures the force of attraction between the molecules of a gas.

Greater the value of 'a' greater is the strength of Vander Waal's interactions.

The real gas equation or Vander Waal's equation after making corrections in volume and pressure becomes.

Vander Waal's equation for n moles gas is $\left(P + \frac{an^2}{v^2}\right)(v - nb) = nRT$

Units of Vander Waal's equation is $a = \frac{pV^2}{n^2} = atm\, litre^2\, mole^{-2}$

$\therefore\ b = \frac{V}{n} = litre\, mole^{-1}$

Vander Waal's equation for one mole of gas $\left(P + \frac{a}{V^2}\right)(v - b) = RT$

6. Explain the liquefaction of gases.

A. **Liquefaction of gases:**

When temperature of the gas is lowered, volume of the gas and also the kinetic energy of the molecules decreases. The molecules become closer and molecular motion becomes slow. As the process of decrease of temperature continues, the gas molecules go on moving closer and closer. Ultimately, at sufficiently low temperature, the gas changes into liquid state. This is also achieved by increasing the pressure. Thus, liquefaction of gas can be achieved by decrease of temperature and by increase of pressure.

7. What are bulk properties? Explain them with examples.

A. The bulk properties of matter are the observable properties of matter. Some of them are behaviour of gases, characteristic properties of solids and liquids and change of state.

Explanation of bulk properties with examples:

I) **The four observable properties of gases are** Pressure, volume, temperature and mass. These four properties are inter-related. These inter- relationships are understood through gas laws.

a) **Boyle's law** states that at a given temperature the volume of a given mass of gas is inversely proportional to its pressure.

$V\alpha\frac{1}{P}$; (n and T are constants)

b) **Gay- Lussac- Charles law** states that at constant pressure the volume of a given mass of a gas is directly proportional to its temperature on Kelvin scale.

$V\alpha T$; (n and P are constants)

c) **Avogardro's law** states that under the same conditions of temperature and pressure equal volumes of different gases contain equal number of molecules or moles.

$V\alpha n$; (P and T are constants)

2) Change of state: Liquefaction of gases.

A gas can be liquefied by cooling it below its critical temperature and then apply high pressures. For example, CO_2 gas can be liquesfied at $31.04^oC\left(T_C\right)$ by applying high pressure. Similarly N_2 gas is liquefied by cooling it to below - 147^oC and then apply pressure.

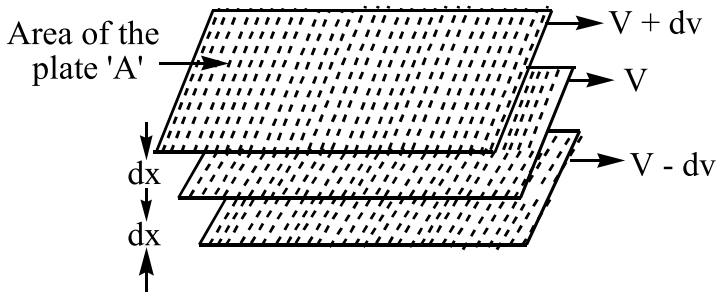
3) Characteristic properties of liquids:

a) **Surface tension (υ) :** It is the force acting along the surface of a liquid at right angles to any line of 1 unit length.

Because of surface tension, the water drops assume spherical shape and the water meniscus in a tube are curved.

b) **Viscosity (η) :** In a liquid (or a fluid) that flows, there will be a friction between the layers of the liquid. This friction resists the smooth flow of the liquid. Viscosity is a measure of this resistance to the flow of a liquid.

The more the viscosity, the more slowly the liquid moves. H – bonding and Van der Waal forces cause high viscosity. Ex: High viscosity of H_2SO_4 is due to H-bonding in the liquid



Gradation of velocity in the laminar flow

8. Explain the surface tension of liquids with suitable examples?

A. **Surface tension:** (γ)

It is defined as the force acting along the surface of a liquid at right angles to any time of 1 unit length. Surface tension is numerically and dimensionally equal to surface energy. It has dimensions kgs^{-2} and in SI unit Nm^{-1}

If we consider a molecule in the bulk of the liquid it experiences equal intermolecular forces in all directions. Hence, there is no net force acting on it. But a molecule at the surface has intermolecular forces from inside only. Therefore, there is a net attractive force on it towards the interior of the liquid. Due to this the surface area of the liquid tends to minimize. The liquid thus behave as if it was under a strain or tension.

Surface tension decreases with increase of temperature because of increase in kinetic energy of molecules and decrease in intermolecular forces.

Because of surface tension

1) The surface of the liquid in a tube is always curved. (Meniscus)

2) Liquids wet the things by spreading across their surface an thin fibres.

3) The rise or fall of liquids in capillary tube is due to surface tension.

4) The liquid has lowest energy state when the surface are is minimum. Spherical shape satisfies this condition. That’s why mercury drops are spherical.