

GASEOUS STATE

1. Introduction

A given substance may occur in solid, liquid or gaseous phase depending upon the relative value of two tendencies namely Mutual Attraction (MA) and Escaping Tendency (ET)

- (i) if MA is greater than ET then substance will occur in solid state
- (ii) if MA is slightly greater than ET then substance will occur in liquid state
- (iii) if MA is very much less than ET then substance will occur in gaseous state

Out of the three states of matter, the most simplest one is the gaseous state.

The state is characterized by sensitivity of volume change with change of pressure and temperature. It is due to large distance between molecules as compared to their own dimensions. There exist weak Vander Waal's forces, molecules move independent of each other with speed about 400 ms^{-1} .

Gases show maximum equality in their behavior irrespective of their nature.

2. Measurable properties of gases

1. Mass

Def. The gases do possess mass. The mass of gas is generally used in the form of number of moles which is related as

$$(i) \quad \text{no. of moles} = \frac{\text{wt. in gm}}{\text{molecular mass of gases}} \left(n = \frac{w}{M} \right)$$

Two other useful formulae to calculate number of moles of gas are –

$$(ii) \quad \text{number of moles} = \frac{\text{no. of molecules of given gas}}{\text{Avogadro's number of molecules}} \left(n = \frac{N}{N_A} \right)$$

$$(iii) \quad \text{no. of moles} = \frac{\text{volume of given gas in litres at STP}}{22.4 \text{ L}}$$

When container contains more than one gas then molecular mass of mixture is termed as effective molecular mass (EMM) which is intermediate between molecular masses of all the gases present in the container.

$$\text{Effective molecular mass} = \frac{\text{Total mass of all the gases present in the container}}{\text{Total no. of moles of all gases present in the container}}$$

2. Volume

Def. Volume of gas is nothing but volume of the container in which it is present.

Relation between different units of volume

$$1 \text{ m}^3 = 10^3 \text{ dm}^3 = 10^3 \text{ litre} = 10^6 \text{ cm}^3 = 10^6 \text{ ml} = 10^9 \text{ mm}^3.$$

3. Temperature

Def. Degree of hotness or coldness of a body is measured by temperature

$$\frac{C}{100} = \frac{K - 273}{100} = \frac{F - 32}{180}$$

C – Celcius scale, K – Kelvin scale, F – Fahrenheit scale

Note : In all the problems of gaseous state (i.e. in all gas law equations), temperature must be expressed in kelvin scale. i.e., $t^{\circ}\text{C} + 273 = \text{TK}$

4. Pressure

Def. Force acting per unit area

$$P = \frac{F}{A}$$

Units :

CGS : dyne/cm²
MKS : Newton/m² (1N/m² = 1Pa)
Relation : 1 N/m² = 10 dyne/cm²

Units of pressure :

1 atm = 76 cm of Hg
= 760 mm of Hg
= 760 torr
= 1.01325 × 10⁵ N/m²
= 101.325 kPa
= 1.01325 bar
= 14.7 lb/in² (Psi)
= 10.33 meters of H₂O

5. Density

Def. Mass per unit volume

$$d = \frac{m}{V}$$

Units :

CGS : g/cm³
MKS : kg/m³
Relation : 1 kg/m³ = 10⁻³ g/cm³

Density of gases

Absolute density

(mass per unit volume)

(i) $d = \frac{m}{V}$

(ii) unit : g/l

(iii) function of temp., pressure, no. of moles

Relative density

(Relative to hydrogen turned as vapour density)

(i) $VD = \frac{\text{molecular mass}}{2}$

(ii) No unit

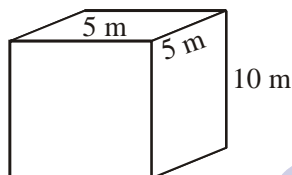
(iii) independent of Pressure, Temperature

Note : Mass, volume and no. of moles are extensive properties that depend on mass hence then all diverly additive in nature.

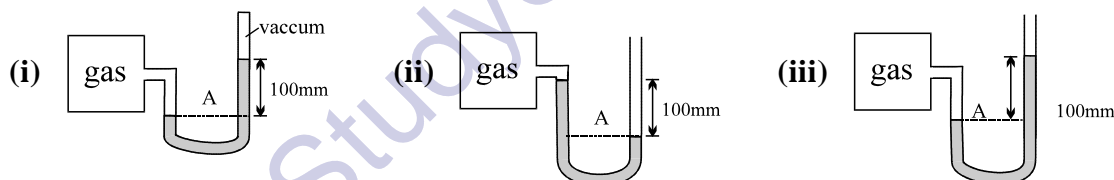
Note : Density, Pressure and Temperature are intensive properties they does not depend on mass hence they are non-additive in nature.

1. Daily Practice Problem Sheet

- Q.1** Normal body temperature (98°F) on Celsius scale will be equal to
(A) 188.8°C (B) 36.67°C (C) 54.5°C (D) 98°C
- Q.2** 413K temperature on Celsius scale Fahrenheit scale will manifest at, respectively
(A) 140°C , 252°F (B) 140°C , 284°F (C) 40°C , 381°F (D) 486°C , 140°F
- Q.3** 1 bar pressure in is equal to
(A) 0.987 Atm (B) 101.325 Atm (C) 101.325 mmHg (D) 76 mmHg
- Q.4** No. of mole of CH_4 gas present in 2.4 g of gas are:
(A) .15 (B) 1.5 (C) 15 (D) .015
- Q.5** How high a column of air would be necessary to cause the barometer to read 76 cm of Hg, if the atmosphere were of uniform density 1.2 kg/m^3 ? The density of Hg = $13.6 \times 10^3\text{ Kg/m}^3$.
(A) 8613 meters (B) 8613 centimeter (C) 8.613 meter (D) 8613 mm
- Q.6** How much water can be occupied in the following cube .

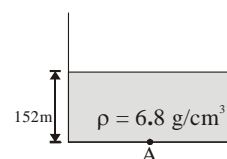


- (A) 250 L (B) $2.5 \times 10^5\text{ L}$ (C) $2.5 \times 10^5\text{ mL}$ (D) 2.5 L
- Q.7** How much pressure will be felt by a gas balloon which at 100 m depth in a sea from sea level.
(A) 100 Pa (B) 98 Pa (C) $9.8 \times 10^5\text{ Pa}$ (D) $1.08 \times 10^6\text{ Pa}$
- Q.8** Calculate the pressure of the gas in cm of Hg in each case (liquid used is Hg)

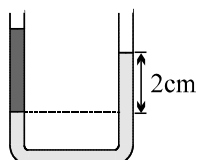


- Q.9** Chlorine has two isotopes Cl^{35} and Cl^{37} , where % abundance are 75 and 25. Calculate the average relative atomic mass of chlorine.

- Q.10** Calculate the pressure at the point A in :



- (A) N/m^2 (B) mm of Hg (C) $\text{m of H}_2\text{O}$ (D) torr (E) atm
- Q.11** A beaker of circular cross section radius 4 cm is filled with Hg upto height 10 cm . Find the force exerted by the mercury on the bottom of the beaker.
- Q.12** The liquids shown in the fig. in the two arms are mercury (specific gravity = 13.6) and water. If the difference of heights of the mercury columns in 2 cm . Find the height h of the water column.



3. The Gas Laws

(i) Boyle's Law :

It relates the volume and the pressure of a given mass of a gas at constant temperature.

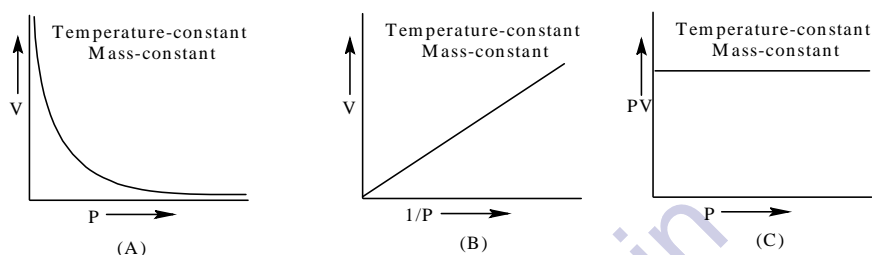
Boyle's law states that, "at constant temperature, the volume of a sample of a gas varies inversely with the pressure".

$$\therefore P \propto \frac{1}{V} \text{ (when temperature and number of moles are kept constant)}$$

The proportionality can be changed into an equality by introducing a constant k , i.e.,

$$P = \frac{k}{V} \text{ or } PV = k$$

Boyle's law can be verified by any one of the following three ways graphically.



Alternatively, Boyle's law can also be stated as follows :

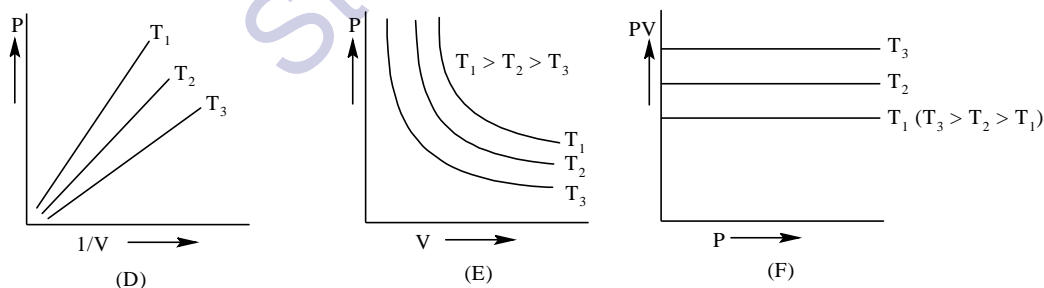
"Temperature remaining constant, the product of pressure and volume of a given mass of a gas is constant".

The value of the constant depends upon the amount of a gas and the temperature.

Mathematically, it can be written as,

$$P_1 V_1 = P_2 V_2 = P_3 V_3 = \dots\dots\dots$$

Location of straight line and curve changes with temperature in the **isotherm** shown in the following figure.



According to Boyle's law, $PV = \text{Constant}$ at constant temperature

$$\therefore \log P + \log V = \text{constant}$$

$$\log P = -\log V = \text{constant}$$

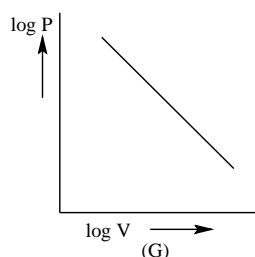


Illustration 1: A gas is present at a pressure of 2 atm. What should be the increase in pressure so that the volume of the gas can be decreased to $1/4^{\text{th}}$ of the initial value if the temperature is maintained constant.

Solution : $PV = \text{constant}$ for a given mass of gas at constant pressure
 $\Rightarrow P_1 V_1 = P_2 V_2$ $P_1 = 2 \text{ atm}$ $V_1 = V$ $V_2 = V/4$ $P_2 = ?$
 Now, $2 \times V = P_2 \times \frac{V}{4} \Rightarrow P_2 = 8 \text{ atm}$
 \therefore Pressure should be increased from 2 to 8 atm
 \therefore total increase = $8 - 2 = 6 \text{ atm}$.

Illustration 2: A sample of gas occupies 10 litre under a pressure of 1 atmosphere. What will be its volume if the pressure is increased to 2 atmospheres? Assume that the temperature of the gas sample does not change.

Solution: Mass and temperature is constant hence we can use Boyle's law
 $P_1 V_1 = P_2 V_2$
 Given $P_1 = 1 \text{ atm}$, $V_1 = 10 \text{ litre}$, $P_2 = 2 \text{ atm}$, $V_2 = ?$
 $1 \times 10 = 2 \times V_2 \Rightarrow V_2 = 5 \text{ litre}$

2. Daily Practice Problem Sheet

- Q.1** A vessel of 120 cm^3 contains a certain mass of a gas at 20°C and 750 torr pressure. The gas was transferred to a vessel of volume 180 cm^3 . Calculate the pressure of the gas at the same temperature.
- Q.2** At what pressure, the volume would be $1/3^{\text{rd}}$ of the volume occupied by a gas under 875 mm pressure of Hg at 15°C ? Assume T constant.
- Q.3** A sample of gas at room temperature is placed in an evacuated bulb of volume 0.51 dm^3 and is found to exert a pressure of 0.24 atm. The bulb is connected to another evacuated bulb whose volume is 0.63 dm^3 . What is the new pressure of the gas at room temperature ?
- Q.4** A gaseous system has a volume of 580 cm^3 at a certain pressure. If its pressure is increased by 0.96 atm, its volume becomes 100 cm^3 . Determine the pressure of the system.
- Q.5** Volume of a cylinder containing 10 marbles and gas is 1 litre at 2 atm pressure. Now pressure on the cylinder is increased to 4 atm, at which volume becomes 725 ml. Calculate volume of each marble.

(ii) Charles' Law :

It relates the volume and temperature of a given mass of a gas at constant pressure. For each degree change of temperature, the volume of a sample of a gas changes by the fraction $\frac{1}{273}$ of its volume at 0°C .

Let the volume of a given amount of gas be V_0 at 0°C . The temperature is increased by $t^\circ\text{C}$ and the new volume becomes V_t

$$\text{Thus, } V_t = V_0 + \frac{V_0}{273} \times t = V_0 \left(1 + \frac{t}{273} \right) \quad \text{or} \quad V_t = V_0 \left(\frac{273 + t}{273} \right)$$

A new temperature scale was introduced known as Kelvin scale or absolute scale (named after the British physicist and mathematician Lord Kelvin). The lower limit of the scale is called absolute zero which corresponds to -273°C .

At absolute zero or -273°C , all molecular motions would stop and the volume of the gas would become zero. The gas would become a liquid or solid. Thus, absolute zero is that temperature at which no substance exists in the gaseous state. The temperature in absolute scale is always obtained by adding 273 to the temperature expressed in $^{\circ}\text{C}$.

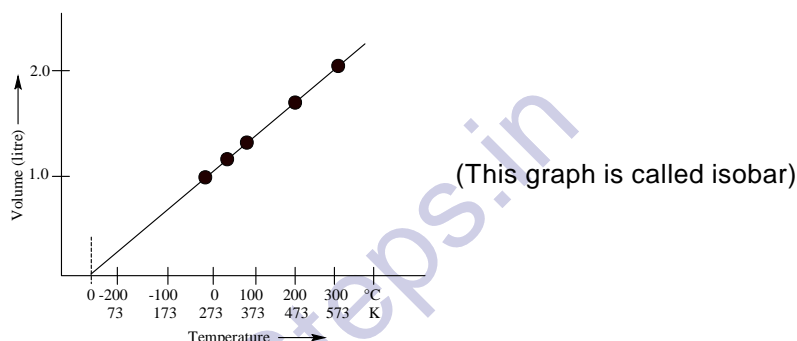
$$K = (t^{\circ}\text{C} + 273)$$

This new temperature scale may be used for deducing Charles' law.

By substituting T for $273 + t$ and T_0 for 273 in Eq. (i).

$$V_t = \frac{V_0 \times T}{T_0} \quad \text{or} \quad \frac{V_t}{T} = \frac{V_0}{T_0}$$

$$\text{or} \quad \frac{V}{T} = \text{constant, if pressure is kept constant}$$



Alternatively, Charles' law can be stated as follows :

“The volume of a given amount of a gas at constant pressure varies directly as its absolute temperature”.

$$V \propto T \quad (\text{if pressure is kept constant})$$

Illustration 3: 1 mole of an ideal gas at constant atmospheric pressure is heated to increase its volume by 50% of initial volume. The change in temperature made was 300 K to T K. Then calculate final temperature.

Solution : At constant pressure for a given amount of gas

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow \frac{V}{300} = \frac{1.5V}{T} \quad (V_1 = V, \quad T_1 = 300 \text{ K}, \quad V_2 = 1.5V)$$

$$T = 450 \text{ K}$$

(iii) Pressure-Temperature Law : (Gaylussac's Law)

It relates the pressure and absolute temperature of a given mass of a gas at constant volume. Volumes remaining constant, the pressure of given mass of a gas increases or decreases by $\frac{1}{273}$ of its pressure at 0°C per degree change of temperature.

$$P_t = P_0 + \frac{P_0 \times t}{273} \quad \text{or} \quad P_t = P_0 \left(1 + \frac{t}{273}\right) \quad \text{or} \quad P_t = P_0 \left(\frac{273 + t}{273}\right) = P_0 \frac{T}{T_0}$$

$$\text{or} \quad \frac{P_t}{T} = \frac{P_0}{T_0} \quad \text{or} \quad P \propto T \quad (\text{if volume and number of moles are kept constant})$$

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

Illustration 4: Pressure of gas contained in a closed vessel is increased by 0.4%, when heated by 1°C. Calculate its final temperature. Assume ideal nature.

Solution : Since no information is given about volume hence we can assume that volume is constant

$$\text{and use } \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

$$\text{given } P_1 = P, \quad T_1 = T, \quad P_2 = P + \frac{0.04P}{100} = 1.004P, \quad T_2 = T + 1$$

$$\therefore \frac{P}{T} = \frac{1.004P}{T+1} \Rightarrow T = 250 \text{ K}$$

3. Daily Practice Problem Sheet

- Q.1** A balloon blown up with 1 mole of gas has a volume of 480 mL at 5°C. the balloon is filled to (7/8)th of its maximum capacity suggest,
- (a) Will the balloon burst at 30°C ?
- (b) The minimum temperature at which it will burst.
- Q.2** A gaseous system occupies the volume 800 cm³ at 22°C at a certain pressure. To what temperature, the gas should be heated in order to increase its volume by 20% keeping the pressure constant ?
- Q.3** A gas cylinder containing cooking gas can withstand a pressure of 14.9 atmosphere. The pressure gauge of the cylinder indicates 12 atmosphere at 27°C. Due to a sudden fire in the building the temperature starts rising. At what temperature will the cylinder explode ?
- Q.4** A chamber of constant volume contains hydrogen gas. When the chamber is immersed in a bath of melting ice (0°C) the pressure of the gas is 800 torr. What pressure will be indicated when the chamber is brought to 100°C ?

(iv) Avogadro's Law

(i) **For Solid, liquid and gas**

1 mole of any substance contains Avogadro's number (N_A) of molecules/atoms/particles etc.

$$N_A = 6.023 \times 10^{23}$$

(ii) **For gases:**

In 1812, Amadeo Avogadro stated that samples of different gases which contain the same number of molecules (any complexity, size, shape) occupy the same volume at the same temperature and pressure. It follows from Avogadro's hypothesis that $V \propto n$ (T and P are constant).

$$V \propto n \quad (T, P \text{ constant}) \Rightarrow \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

STP: 273.15 K 1 atm

SATP: 298.15 K 1 bar

4. Daily Practice Problem Sheet

- Q.1** At similar conditions of temperature and pressure three gases He, N₂ and O₂ has same volume. Arrange all three gases in increasing order of No. of atoms present in the given volume.
(A) O₂ > N₂ > He (B) He > N₂ > O₂ (C) He = N₂ = O₂ (D) He < N₂ = O₂
- Q.2** Calculate weight in gram present in $1.12 \times 10^{-3} \text{ m}^3$ of CH₄ gas at STP.
(A) 0.8 g (B) 1.6 g (C) 1.12 g (D) 8 g
- Q.3** Calculate volume occupied by 8.8 g CO₂ gas, if 5.6 g CO gas occupies 6.24 L at same Temperature and pressure.
(A) 22.4 L (B) 1.12 L (C) 4.48 L (D) 6.24 L
- Q.4** A container has SO₂ gas at 2 atm pressure in a vessel of V L capacity, if no. of moles of SO₂ are doubled in the same container at same temperature and volume. Calculate new pressure in the container.
(A) 1 atm (B) 4 atm (C) 2 atm (D) 8 atm
- Q.5** Calculate No. of molecules of CH₄ and SO₂ gas present in 1.12 L and 2.24 L volume, respectively at STP.
- Q.6** From 5.61 L of SO₂ gas at STP, 6.02×10^{22} molecules are removed, than calculate mass of remaining gas .
- Q.7** (i) Calculate no. of molecule in 500 g CaCO₃ ($N_A = 6 \times 10^{23}$)
(ii) Calculate no. of Ca, C, O atom in it
(iii) If 1.5×10^{24} atoms, Ca atoms are removed then calculate no. of O and C atoms remained and weight of remaining CaCO₃.
- Q.8** What mass of NaCl would contain the same total number of ions as 245 g of MgCl₂?
- Q.9** Calculate the total number of atoms in 0.5 mole of K₂Cr₂O₇.
- Q.10** Arrange following in increasing order of mass.
(i) 1 g Ca (ii) 1 g atom Ag
(iii) 1 mole of Na (iv) 3×10^{24} atoms of O
- Q.11** Calculate the total number of electrons present in 18 mL of water.
- Q.12** Find the mass of the nitrogen contained in 1 kg of
(i) KNO₃, (ii) NH₄NO₃ and (iii) (NH₄)₂HPO₄.

(v) Ideal Gas Equation

Combining all these gas laws, a simple equation can be derived at, which relates P, V, n and T for a gas

$$PV = nRT \quad (\text{for } n \text{ moles of gas})$$

$$\Rightarrow \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad (\text{Combined gas law})$$

P is the pressure of the gas and can be expressed in atm or Pa. Correspondingly, the volume must be expressed in litres or m³ respectively. n is the number of moles and T is the temperature in Kelvin. R is called the universal gas constant.

Numerical Values of R

- (i) In litre atmosphere = $0.0821 \text{ litre atm deg}^{-1} \text{ mole}^{-1}$
- (ii) In ergs = $8.314 \times 10^7 \text{ erg deg}^{-1} \text{ mole}^{-1}$
- (iii) In joules = $8.314 \text{ joules deg}^{-1} \text{ mol}^{-1}$
- (iv) In calories = $1.937 \text{ cal deg}^{-1} \text{ mole}^{-1}$

Illustration 5 : An open vessel at 27°C is heated until $3/5^{\text{th}}$ of the air in it has been expelled. Assuming that the volume of the vessel remains constant find

(A) the air escaped out if vessel is heated to 900K .

(B) temperature at which half of the air escapes out.

Solution : One should clearly note the fact that on heating a gas in a vessel, there are the number of moles of gas which go out, the volume of vessel remains constant.

Let initial no. of moles of gas at 300 K be 'n'. On heating $3/5$ moles of air escape out at temperature T.

$$\therefore \text{Moles of air left at temperature } T = \left(n - \frac{3}{5}n \right) = \frac{2n}{5}$$

(A) On heating vessel to 900 K , let n_1 moles be left,

$$n_1 T_1 = n_2 T_2 \quad \Rightarrow \quad n_1 \times 900 = 300 \times n \quad \Rightarrow \quad n_1 = \frac{1}{3}n$$

$$\therefore \text{moles escaped out} = n - \frac{n}{3} = \frac{2}{3}n \text{ moles}$$

(B) Let $n/2$ moles escape out at temperature T, then

$$n_1 T_1 = n_2 T_2 \quad \Rightarrow \quad \frac{n}{2} \times T = n \times 300 \quad \Rightarrow \quad T = 600 \text{ K}$$

Illustration 6 : 5g of ethane is confined in a bulb of one litre capacity. The bulb is so weak that it will burst if the pressure exceeds 10 atm . At what temperature will the pressure of gas reach the bursting value?

Solution: $PV = nRT \Rightarrow 10 \times 1 = \frac{5}{30} \times 0.082 \times T$

$$T = \frac{60}{0.082} = 60 \times 12.18 = 730.81 \text{ K} = 457.81^\circ\text{C}$$

(vi) Relation between Molecular Mass and Gas Densities

(A) **Actual density :** For an ideal gas $PV = nRT$ or $PV = \frac{w}{M}RT$, where w = mass of the gas in gms and M = Molecular wt. in gms.

$$\therefore PM = \frac{w}{V}RT \quad \text{or} \quad PM = \rho RT, \text{ (where } \rho \text{ is the density of the gas} = \frac{w}{V} \text{)}$$

$$\therefore = \frac{PM}{RT}$$

$$(i) \quad \frac{d_1 T_1}{P_1} = \frac{d_2 T_2}{P_2} \quad \text{(for same gas at different temperature and pressure)}$$

$$(ii) \quad \frac{d_1}{d_2} = \frac{M_1}{M_2} \quad (\text{for different gases at same temperature \& pressure})$$

(Where d = density of gas)

Illustration 7: The density of an unknown gas at 98°C and 0.974 atm is $2.5 \times 10^{-3} \text{ g/ml}$. What is the mol wt. of gas?

Solution: Density = $2.5 \times 10^{-3} \text{ g/ml} = 2.5 \text{ g/L}$

$$\therefore PM = \rho RT$$

$$0.974 \times M = 2.5 \times 0.0821 \times 371 \Rightarrow M = 78.18.$$

(B) Vapour Density: For gases another term which is often used is vapour-density. Vapour density of a gas is defined as the ratio of the mass of the gas occupying a certain volume at a certain temperature and pressure to the mass of hydrogen occupying the same volume at the

same temperature and pressure i.e. $W(\text{gas}) = \frac{PVM}{RT}$.

$$\text{and } W_{\text{H}_2} = \frac{PV \times 2}{RT} \quad (\because \text{mol. wt. of hydrogen is } 2)$$

$$\therefore \frac{W_{\text{gas}}}{W_{\text{H}_2}} = \frac{M}{2} = (\text{Vapour density of gas})$$

Vapour density of a gas is same at any temperature, pressure and volume.

Illustration 8: When 3.2 g of sulphur is vapourised at 450°C and 723 mm pressure, the vapours occupy a volume of 780 ml . What is the molecular formula of sulphur vapours under these conditions? Calculate the vapour density also.

$$\text{Solution: } PV = nRT \Rightarrow \frac{723}{760} \times \frac{780}{760} = \frac{3.2}{M} \times 0.082 \times 723$$

$$M = 255.9$$

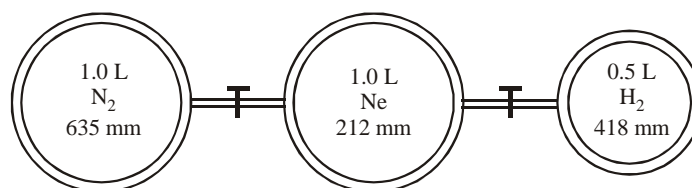
$$\text{no. of atoms of sulphur in one molecule} = \frac{M}{32} = 8$$

$$\therefore \text{Molecular formula of sulphur} = \text{S}_8$$

$$\text{V. D.} = \frac{M}{2} = \frac{255.9}{2} \approx 127.99$$

5. Daily Practice Problem Sheet

- Q.1** Calculate the weight of CH_4 in a 9 litre cylinder at 16 atm and 27°C . ($R = 0.08 \text{ L atm K}^{-1}$)
- Q.2** Calculate the moles of hydrogen present in a 500 cm^3 sample of hydrogen gas at a pressure of 760 mm of Hg and 27°C .
- Q.3** Calculate the volume occupied by 4.0245×10^{23} molecules of O_2 at 27°C having pressure of 700 torr.
- Q.4** A gas occupies 300 mL at 27°C and 730 mm pressure. What would be its volume at STP ?
- Q.5** A gas at 0°C and 1 atm pressure occupies 2.5 litre. What change in temperature would be necessary if the pressure is to be adjusted to 1.5 atm and the gas has been transferred to a 2.0 litre container ?
- Q.6** 3.7 gm of a gas at 25°C occupied the same volume as 0.184 gm of hydrogen at 17°C and at the same pressure. What is the molecular mass of the gas ?
- Q.7** An under water bubble with a radius of 0.5 cm at the bottom of tank, where the temperature is 5°C and pressure is 3 atm rises to the surface, where temperature is 25°C and pressure is 1 atm. What will be the radius of bubble when it reaches to surface ?
- Q.8** The density of a gas is 0.259 g/mL at 400 K and 190 torr. Find its molar mass.
- Q.9** Calculate the density of CO_2 at 100°C and 800 mm Hg pressure.
- Q.10** A gas occupies 0.418 litre at 740 mm of Hg and 27°C . Calculate :
 (a) its volume at STP
 (b) molecular weight if gas weighs 3.0 g
 (c) new pressure of gas if the weight of gas is increased to 7.5 g and temperature becomes 280 K
 (d) the volume of vessel at 300 K
- Q.11** A balloon filled with helium raises to a certain height at which it gets fully inflated to a volume of 1×10^5 litre. If at this altitude temperature and atmospheric pressure is 268 K and 2×10^{-3} atm respectively, what weight of helium is required to fully inflated the balloon ?
- Q.12** Consider the arrangement of bulbs shown below:



What is the pressure of the system when all the stopcocks are opened?

- Q.13** A glass bulb of volume 400 cm^3 is connected to another bulb of volume 200 cm^3 by means of a tube of negligible volume. The bulbs contain dry air and are both at a common temperature and pressure of 20°C and 1.00 atm. The larger bulb is immersed in steam at 100°C ; the smaller, in melting ice at 0°C . Find the final common pressure.

(vii) Dalton's Law of Partial Pressures:

The total pressure of a mixture of non-reacting gases is equal to the sum of their partial pressures.

$$\text{By Dalton's Law} \quad P_T = P_1 + P_2 + \dots$$

By the partial pressure of a gas in a mixture is meant, the pressure that the gas will exert if it occupies alone the total volume of the mixture at the same temperature.

Derivation: $n = n_1 + n_2 + \dots$

$$n = \frac{PV}{RT} \Rightarrow \frac{PV}{RT} = \frac{P_1V}{RT} + \frac{P_2V}{RT} + \dots \Rightarrow P = P_1 + P_2 + \dots$$

Assumption: Volume of all the gases is same as they are kept in same container.

Relationship between partial pressure and number of moles

Important formula

$$(i) \quad P_1 = \left(\frac{n_1}{n_1 + n_2} \right) P = x_1 P \quad \text{where } x_1 = \text{mole fraction of gas}$$

$$(ii) \quad \text{Partial pressure of a gas in the mixture} = \frac{\text{Volume of the gas}}{\text{Total Volume}} \times P$$

Partial pressure and aqueous tension: Dalton's law is used to calculate the pressure of a dry gas when it is collected over water at atmospheric pressure.

By Dalton's law,

$$\text{Pressure of dry gas} = \text{atmospheric pressure} - \text{aqueous tension}$$

Aqueous tension depends on temperature. It increases with temperature and becomes 760 mm at 100°C.

Illustration 9: A gaseous mixture contains 55% N_2 , 20% O_2 , and 25% CO_2 by mass at a total pressure of 760 mm. Calculate the partial pressure of each gas.

Solution:

Total mass of the gases = 100 g

$$W_{N_2} = 55\text{g}, \quad W_{O_2} = 20\text{g}, \quad W_{CO_2} = 25\text{g}$$

$$n_{N_2} = 55/28 = 1.964, \quad n_{O_2} = 20/32 = 0.625, \quad n_{CO_2} = 25/44 = 0.568$$

Total moles = 3.157

$$P_{N_2} = \frac{1.964 \times 760}{3.157} = 472.8 \text{ mm}$$

$$P_{O_2} = \frac{0.625 \times 760}{3.157} = 150.46 \text{ mm}$$

$$P_{CO_2} = \frac{0.568 \times 760}{3.157} = 136.74 \text{ mm}$$

Illustration 10: A mixture containing 1.6 g of O_2 , 1.4g of N_2 and 0.4 g of He occupies a volume of 10 litre at 27°C. Calculate the total pressure of the mixture and partial pressure of each compound.

Solution:

$$PV = nRT \quad V = 10 \text{ litre} \quad T = 27^\circ\text{C} = 300\text{K}$$

$$n_{He} = \frac{0.4}{4} = 0.1, \quad n_{N_2} = \frac{1.4}{28} = 0.05, \quad n_{O_2} = \frac{1.6}{32} = 0.05$$

$$\text{Total no. of moles} = 0.1 + 0.05 + 0.05 \Rightarrow 0.2$$

$$PV = nRT \Rightarrow P \times 10 = 0.2 \times 0.082 \times 300 = 0.04926 \Rightarrow P = 0.492 \text{ atm}$$

$$\text{Partial pressure} = \text{Total pressure} \times \text{molefraction}$$

$$P_{\text{He}} = 0.492 \times \frac{0.1}{0.2} = 0.246 \text{ atm}$$

$$P_{\text{N}_2} = 0.492 \times \frac{0.05}{0.2} = 0.123 \text{ atm}$$

$$P_{\text{O}_2} = 0.492 \times \frac{0.05}{0.2} = 0.123 \text{ atm}$$

(viii) Graham's Law of Diffusion

Diffusion is the tendency of any substance to spread throughout the space available to it. Diffusion will take place in all direction and even against gravity.

The streaming of gas molecules through a small hole is called effusion.

According to Graham, the rate of diffusion (or effusion) of a gas at constant pressure and temperature is inversely proportional to the square root of its molecular mass.

$$r \propto \sqrt{\frac{1}{M}}, \text{ at constant P and T}$$

$$\therefore \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}, \text{ at constant P and T}$$

Since molecular mass of gas = $2 \times$ vapour density, $\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}}, \text{ at constant P and T}$

The rate of diffusion (or effusion) r of two gases under different pressure can be given by

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \times \frac{P_1}{P_2} \text{ at constant T only.}$$

$$r = \frac{\text{Volume diffused (V)}}{\text{time taken}} = \frac{\text{moles diffused (n)}}{\text{time taken}}$$

$$r = \frac{\text{distance travelled in a narrow tube of uniform cross-sectional area (x)}}{\text{time taken}}$$

Therefore, according to Graham's law of diffusion (effusion) at constant P and T.

$$\frac{V_1}{t_1} \times \frac{t_2}{V_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$$

d_1 and d_2 are the respective densities and V_1 and V_2 are volumes diffused (effused) in time t_1 and t_2 .

$$\frac{n_1}{t_1} \times \frac{t_2}{n_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$$

where n_1, n_2 are moles diffused (effused) in time t_1 and t_2 .

$$\frac{x_1}{t_1} \times \frac{t_2}{x_2} = \sqrt{\frac{d_2}{d_1}} = \sqrt{\frac{M_2}{M_1}}$$

where x_1 and x_2 are distances travelled by molecules in narrow tube in time t_1 and t_2 .

$$\begin{aligned}
 r &= \frac{\text{moles diffused}}{\text{time taken}} \\
 &= \frac{\text{distance travelled in a narrow tube}}{\text{time taken}} \\
 &= \frac{\text{pressure drop I}}{\text{pressure drop II}}
 \end{aligned}$$

Note: It should be noted that the rate of diffusion or effusion actually depends on pressure difference of the gas and not simply on its pressure. Moreover the pressure difference is to be measured for this gas only i.e. if a container holds [He] at a pressure of 0.1 atm and if a small pin-hole is made in the container and if the container is placed in a room, then the rate of effusion of He gas from the container to outside depends only on its pressure difference, which is 0.1-0 (as there is no He in the atmosphere). This implies that the diffusion of a gas is not dependent on the diffusion of any other gas.

Whenever we consider the diffusion of gas under experimental conditions, we always assume that the gas diffuses in vacuum and during the time period for which the diffusion is studied the rate of diffusion (or the composition of diffusing or effusing mixture of gases) remains constant.

Illustration 11: Pure O_2 diffuses through an aperture in 224 seconds, whereas mixture of O_2 and another gas containing 80% O_2 diffuses from the same in 234 sec under similar condition of pressure and temperature. What is molecular wt. of gas?

Solution: The gaseous mixture contains 80% O_2 and 20% gas.

$$\therefore \text{Average molecular weight of mixture } (M_{\text{mix}}) = \frac{32 \times 80 + 20 \times m_{\text{mix}}}{100} \quad \dots (i)$$

Now for diffusion of gaseous mixture and pure O_2

$$\frac{r_{O_2}}{r_m} = \sqrt{\frac{M_{\text{mix}}}{M_{O_2}}} \quad \text{or} \quad \frac{V_{O_2}}{V_{\text{mix}}} \times \frac{t_{\text{mix}}}{t_{O_2}} = \sqrt{\frac{M_{\text{mix}}}{32}}$$

$$\text{or} \quad \frac{1}{224} \times \frac{234}{1} = \sqrt{\frac{M_{\text{mix}}}{32}} \quad \dots (ii)$$

$$\therefore M_{\text{mix}} = 34.92$$

By (i) and (ii) mol weight of gas (m) = 46.6.

Illustration 12: Calculate the relative rates of diffusion of $^{235}\text{UF}_6$ and $^{238}\text{UF}_6$ in the gaseous state (Atomic mass of F = 19).

Solution $^{235}\text{UF}_6$ and $^{238}\text{UF}_6$ (at F = 19)

$$\begin{aligned}
 \frac{r_{\text{UF}_6^{235}}}{r_{\text{UF}_6^{238}}} &= \sqrt{\frac{M_{\text{UF}_6^{238}}}{M_{\text{UF}_6^{235}}}} \\
 &= \sqrt{\frac{352}{349}} = 1.0042
 \end{aligned}$$

6. Daily Practice Problem Sheet

- Q.1** Equal weights of CH_4 and O_2 are mixed in an empty container of one litre at 27°C . Calculate
 (a) the fraction of total pressure exerted by O_2
 (b) the total pressure if the weights of gases are 32 g each
- Q.2** Two gases A and B having molecular weights 60 and 45 respectively are enclosed in a vessel. The wt. of A is 0.50 g and that of B is 0.2 g. The total pressure of the mixture is 750 mm. Calculate partial pressure of the two gases.
- Q.3** 5 litre of N_2 under a pressure of 2 atm, 2 litre of O_2 at 5.5 atm and 3 litre of CO_2 at 5 atm are mixed. The resultant volume of the mixture is 15 litre. Calculate the total pressure of the mixture and partial pressure of each constituent.
- Q.4** A 10 litre flask at 298 K contains a gaseous mixture of CO and CO_2 at a total pressure of 2 atm. If 0.20 mole of CO is present, find its partial pressure and also that of CO_2 .
- Q.5** 1500 mL flask contains 400 mg O_2 and 60 mg H_2 at 100°C .
 (a) What is the total pressure in the flask?
 (b) If the mixture is permitted to react to form water vapour at 100°C , what will be their partial pressures?
- Q.6** 32 cc of hydrogen diffuses through a fine hole in 1 minute. What volume of CO_2 will diffuse in 1 minute under the same conditions?
- Q.7** 180 mL of a hydrocarbon diffuses through a porous membrane in 15 minutes while 120 mL of SO_2 under identical conditions diffuses in 20 minutes. What is the molecular mass of the hydrocarbon?
- Q.8** At room temperature, ammonia gas at one atmospheric pressure and hydrogen chloride at P atmospheric pressure are allowed to effuse through identical pinholes from opposite ends of a glass tube of one metre length and of uniform cross reaction. NH_4Cl is first formed at a distance of 60 cm from the end through which HCl gas is sent in. What is the value of P?
- Q.9** The rate of effusion of an unknown gas (X) through a pin hole is found to be 0.279 times the rate of effusion of hydrogen gas through the same pinhole, if both are at STP. What is the molecular mass of the unknown gas?
- Q.10** If a 2 m long narrow tube, HCl is allowed to diffuse in the tube from one end and NH_3 from the other end. If diffusion is started at the same time, predict at what point the white fumes of NH_4Cl will form?

4. The Kinetic Theory of Gases

In order to derive the theoretical aspect of the various gas laws based on simple experiment facts, Maxwell proposed the following postulates under the heading of kinetic theory of gases: The postulates of kinetic theory of gas are

- Each gas is made up of a large number of small (tiny) particles known as molecules.
- The volume of a molecule is so small that it may be neglected in comparison to total volume of gas.
- The molecules are never in stationary state but they are believed to be in chaotic (random) motion. They travel in straight line in all possible directions with altogether different but constant velocities. The direction of motion is changed by the collision with container or with the other molecules.

- (iv) The collision between molecules is perfectly elastic i.e., there is no change in the energies of the molecules after collision.
- (v) The effect of gravity on molecular motion is negligible.
- (vi) The kinetic energy of the gases depends on the temperature.
- (vii) The pressure of the gas arises due to collision of molecules with the walls of the container.

The Kinetic Equation : Maxwell also derived an equation on the basis of above assumptions as

$$PV = \frac{1}{3} mnu^2 \quad \text{where}$$

P = Pressure of gas

V = Volume of gas

m = mass of one molecule of gas

n = no. of molecules of gas

u = root mean square velocity of molecules

For 1 mole $n = N$ (Avogadro number)

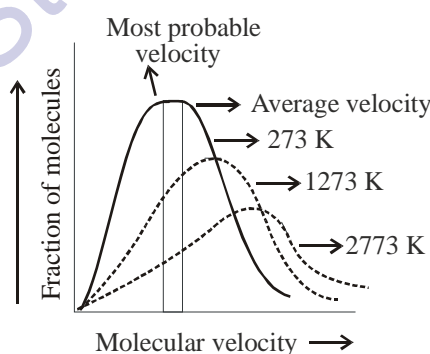
$m \times N = \text{Molecular mass } M$.

$$\therefore PV = \frac{1}{3} Mu^2 \quad \text{or} \quad u^2 = \frac{3PV}{M} = \frac{3RT}{M} \quad \text{or} \quad u = \sqrt{\frac{3PV}{M}} = \sqrt{\frac{3RT}{M}}$$

5. Distribution of Molecular Velocities

Maxwell and Boltzmann proposed that gas molecules are always in rapid random motion colliding with each other and with the walls of container. Due to such collisions, their velocities always changes. A fraction of molecules have a particular molecular velocity at a time. James Clark Maxwell calculated the distribution of velocity among fraction of total number of molecules, on the basis of probability.

The distribution of velocities of different gas molecules may be shown by the following curve.



From the curve it may be concluded that

- (i) Only a small fraction of molecules have either very low or very high velocity.
- (ii) Curve becomes flat when temperature is raised i.e. distribution around average velocity becomes wider. Average molecular velocity increases with rise in temperature.
- (iii) Most of the molecules have velocity close to most probable velocity represented by the top of curve.
- (iv) At higher temperature greater number of molecules have high velocity, while few molecules have lower velocity.

Average Velocity : As per kinetic theory of gases, each molecule is moving with altogether different velocity. Let 'n' molecules be present in a given mass of gas, each one moving with velocity $u_1, u_2, u_3, \dots, u_n$. The average velocity or U_{av} = average of all such velocity terms.

$$\text{Average velocity} = \frac{u_1 + u_2 + u_3 + \dots + u_n}{n}$$

$$U_{av} = \sqrt{\frac{8RT}{M}}$$

Root Mean Square Velocity : Maxwell proposed the term U_{rms} as the square root of means of square of all such velocities.

$$U_{rms}^2 = \frac{u_1^2 + u_2^2 + u_3^2 + \dots}{n}$$

Also $U_{rms} = \sqrt{\frac{3RT}{M}}$

Most probable velocity: It is the velocity possessed by maximum no. of molecules.

$$U_{mpv} = \sqrt{\frac{2RT}{M}}$$

$$\begin{aligned} \text{Furthermore } U_{mpv} : U_{av} : U_{rms} &:: \sqrt{\frac{2RT}{M}} : \sqrt{\frac{8RT}{\pi M}} : \sqrt{\frac{3RT}{M}} \\ &= \sqrt{2} : \sqrt{\frac{8}{\pi}} : \sqrt{3} = 1 : 1.128 : 1.224 \end{aligned}$$

Also $U_{av} = U_{rms} \times 0.9213$

Kinetic Energy of Gas: As per kinetic equation $PV = \frac{1}{3} mn u_{rms}^2$.

For 1 mole $m \times n = \text{Molecular Mass (M)}$

$$\therefore PV = \frac{1}{3} M u_{rms}^2 = \frac{2}{3} \times \frac{1}{2} M u^2 = \frac{2}{3} \times \text{K.E. / mole} \quad \text{or} \quad \frac{\text{K.E.}}{\text{mole}} = \frac{3}{2} RT$$

$$\text{Also KE per molecule} = \frac{3}{2} \frac{RT}{n} = \frac{3}{2} kT. \text{ Where } k \text{ is the Boltzmann constant } \left(k = \frac{R}{n} \right)$$

Kinetic Energy of gas sample :

(i) Average kinetic energy of a single molecule = $\frac{3}{2} \cdot \frac{R}{N} \cdot T = \frac{3}{2} kT$

$k = \text{boltzman constant} = 1.38 \times 10^{-23} \text{ J/deg}$

(ii) Total Kinetic Energy for one mole of gas = $\frac{3}{2} RT$

(iii) kinetic Energy for n mol of gas = $n \times \frac{3}{2} RT$

Illustration 13: Calculate rms speed of O_2 at 273 K and 1×10^5 Pa pressure. The density of O_2 under these conditions is 1.42 kg m^{-3} .

Solution: Data are given in SI units

$$C = \sqrt{\frac{3P}{d}} = \sqrt{\frac{3 \times 10^5}{1.42}} = 459.63 \text{ m sec}^{-1}$$

Illustration 14: At what temperature will the r.m.s. velocity of oxygen be one and half times of its value at N.T.P.?

Solution: $\frac{1}{2} mc^2 = \frac{3}{2} kT$

Suppose the temperature required is T' then the velocity will be $\frac{3}{2}C$

$$\therefore \frac{1}{2} mc'^2 = \frac{3}{2} kT'$$

$$\therefore \frac{3/2C}{C} = \sqrt{\frac{T'}{T}} \text{ or, } T' = \frac{9}{4} \times 273 = 614.25^\circ \text{ K}$$

Illustration 15: Calculate the average and total kinetic energy of 0.5 mole of an ideal gas at 0°C .

Solution: Average kinetic energy = $\frac{3}{2} KT = \frac{3}{2} \times 1.38 \times 10^{-23} \times 273 = 5.65 \times 10^{-21} \text{ J}$

$$\begin{aligned} \text{Total kinetic energy of } n \text{ mole of gas} &= n \times \frac{3}{2} RT = 0.5 \times \frac{3}{2} \times 8.314 \times 273 \\ &= 1.702 \text{ kJ} \end{aligned}$$

7. Daily Practice Problem Sheet

- Q.1** Calculate the temperature at which the root mean square speed of SO_2 molecules is the same as that of oxygen at 27°C .
- Q.2** Calculate the rms speed of ozone kept in a closed vessel at 20°C and 82 cm Hg pressure.
- Q.3** A gas has a density of $1.2504 \text{ g/27}^\circ\text{C}$ at 0°C and a pressure of 1 atm. Calculate the rms, average and the most probable speeds of its molecules at 0°C .
- Q.4** The density of hydrogen at 0°C and 760 mm Hg pressure is 0.00009 g/cc. What is the rms speed of hydrogen molecules ?
- Q.5** The kinetic molecular theory attributes an average kinetic energy of $3/2 KT$ to each particle. What rms speed would a mist particle of mass 10^{-12} gm have at room temperature (27°C) according to the kinetic molecular theory.
- Q.6** Compute rms speed of (i) O_2 at 15°C and 77 cm pressure, (ii) NH_3 molecules at NTP and (iii) average speed of CH_4 at 500°C .
- Q.7** At what temperature would the most probable speed of CO molecules be twice that at 0°C ?

- Q.8** Under 3 atm, 12.5 litre of a certain gas weigh 15 g, calculate the average speed of gaseous molecules.
- Q.9** Calculate the pressure exerted by 10^{23} gas molecules, each of mass 10^{-22} g in a container of volume one litre. The rms speed is 10^5 cm sec⁻¹.
- Q.10** Calculate the room mean square speed, total and average translational kinetic energy in joule of the molecules in 8 g methane at 27 °C.
- Q.11** Two bulbs A and B equal capacity are filled with He and SO₂ respectively, at the same temperature.
- If the pressure in two bulbs is same, calculate ratio of u_{rms} for them.
 - At what temperature, velocity of SO₂ becomes half of the speed of He molecules at 27 °C.
 - How will the speeds be effected if volume of B becomes four times.
 - How will the speeds be effected if half of the molecules of SO₂ are removed from B.
- Q.12** Two flasks A and B have equal volumes. Flask A contains H₂ at 27°C while B contains equal mass of C₂H₆ at 627°C. In which flask and by how many times are molecules moving faster ? Assume ideal gas nature for both.
- Q.13** Compare the rms speed of O₂ at 27°C and 54°C. Also find the temperature at which rms speed of O₂ is doubled to that of 27°C.
- Q.14** Calculate the root mean square speed of oxygen molecules having a kinetic energy of 2kcal mol⁻¹. At what temperature would the molecules have this value of u_{rms} ?
- Q.15** Calculate the temperature at which the molecules of the first two members of the homologous series C_nH_{2n+2} will have the same rms speed as CO₂ gas has 770 K.

6. Deviations from Ideal Behaviour

An ideal gas is one which obeys the gas laws of the gas equation $PV = RT$ at all pressure and temperatures. However no gas in nature is ideal. Almost all gases show significant deviations from the ideal behaviour. Thus the gases H₂, N₂ and CO₂ which fail to obey the ideal-gas equation are termed as non-ideal or real gases.

Compressibility Factor : The extent to which a real gas departs from the ideal behaviour may be depicted in terms of a new function called the compressibility factor, denoted by Z. It is defined as

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

The deviations from ideality may be shown by a plot of the compressibility factor Z, against P. For an ideal gas, $Z = 1$ and it is independent of temperature and pressure.

The deviations from ideal behaviour of a real gas will be determined by the value of Z being greater or less than 1.

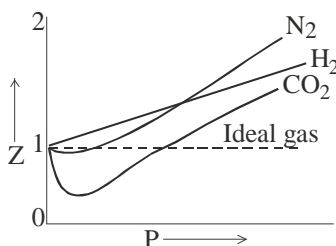
The difference between unity and the value of the compressibility factor of a gas is a measure of the degree of non ideality of the gas.

For a real gas, the deviations from ideal behaviour depends on :

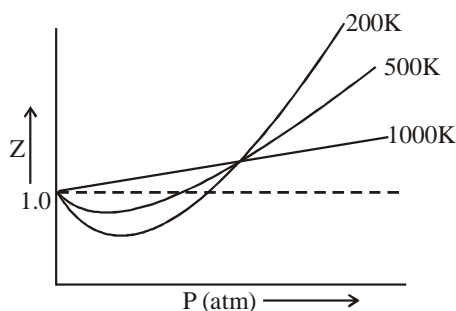
- pressure ; and
- temperature.

This will be illustrated by examining the compressibility curves of some gases discussed below with the variation of pressure and temperature.

Effect of Pressure Variation on Deviations:



Effect of Temperature on Deviations:



From the above curves we can conclude that:

1. At low pressure and fairly high temperatures, real gases show nearly ideal behaviour and the ideal-gas equation is obeyed.
2. At low temperatures and sufficiently high pressures, a real gas deviates significantly from ideality and the ideal-gas equation is no longer valid.
3. The closer the gas is to the liquefaction point, the larger will be the deviation from the ideal behaviour.

Greater is the departure of Z from unity, more is the deviation from ideal behaviour.

- (i) When $Z < 1$, this implies that gas is more compressible.
- (ii) When $Z > 1$, this means that gas is less compressible.
- (iii) When $Z = 1$, the gas is ideal.

Vander Waals Equation of State for a Real Gas: The equation of state generated by Vander Waals in 1873 reproduces the observed behaviour with moderate accuracy. For n moles of gas, the Vander Waals equation is

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

where **a** and **b** are constants characteristic of a gas. This equation can be derived by considering a real gas and converting it to an ideal gas.

Volume Correction : We know that for an ideal gas $P \times V = nRT$. Now in a real gas the molecular volume cannot be ignored and therefore let us assume that '**b**' is the volume excluded (out of the volume of container) for the moving gas molecules per mole of a gas. Therefore due to n moles of a gas the volume excluded would be **nb**.

\therefore a real gas in a container of volume V has only available volume of $(V - nb)$ and this can be thought of, as an ideal gas in a container of volume $(V - nb)$

Pressure Correction : Let us assume that the real gas exerts a pressure P . The molecules that exert the force on the container will get attracted by molecules of the immediate layer which are not assumed to be exerting pressure.

It can be seen that pressure the real gas exerts would be less than the pressure an ideal gas would have exerted. Therefore if a real gas exerts a pressure P , then an ideal gas would exert a pressure equal to $P + p$ (p is the pressure lost by the gas molecules due to attractions). This small pressure p would be directly proportional to the extent of attraction between the molecules which are hitting the container wall and the molecules which are attracting these.

Therefore $p \propto \frac{n}{V}$ (concentration of molecules which are hitting the container's wall)

$p \propto \frac{n}{V}$ (concentration of molecules which are attracting these molecules) $\Rightarrow p \propto \frac{n^2}{V^2}$

$$\therefore P = \frac{an^2}{V^2}$$

where a is the constant of proportionality which depends on the nature of gas. Higher value of 'a' reflects the increased attraction between gas molecules.

The Vander Waals constant b (the excluded volume) is actually 4 times the volume of a single molecule. i.e. $b = 4 N_A V$ where $N_A \rightarrow$ Avogadro number.

$$\therefore b = 4 \times 6.023 \times 10^{23} \frac{4}{3} \pi r^3, \text{ where } r \text{ is the radius of a molecule.}$$

The constant a and b : Vander Waals constant for attraction (A) and volume (B) are characteristic for a given gas. Some salient features of 'a' and 'b' are:

- (i) For a given gas Vander Waal's constant of attraction 'a' is always greater than Vander Waals constant of volume (B).
- (ii) The gas having higher value of 'a' can be liquified easily and therefore H_2 and He are not liquified easily.
- (iii) The units of $a = \text{litre}^2 \text{ atm mole}^{-2}$ and that of $b = \text{litre mole}^{-1}$
- (iv) The numerical values of a and b are in the order of 10^{-1} to 10^{-2} to 10^{-4} respectively.
- (v) Higher is the value of 'a' for a given gas, easier is the liquification.

Explanation of deviation by Van der Waals equation

- (i) **At lower pressure :** 'V' is large and 'b' is negligible in comparison with V. Then Vander Waals equation reduces to :

$$\left(P + \frac{a}{V^2}\right)V = RT \quad \Rightarrow \quad PV + \frac{a}{V} = RT$$

$$PV = RT - \frac{a}{V} \quad \Rightarrow \quad Z = \frac{PV}{RT} = 1 - \frac{a}{VRT}$$

or **$PV < RT$** at low pressure (below Boyle temperature)
this accounts for the dip in PV vs P isotherm at low pressure.

- (ii) *At fairly high pressures* : $\frac{a}{V^2}$ may be neglected in comparison with P.

The Vander Waals equation becomes

$$P(V-b) = RT \Rightarrow PV - Pb = RT$$

$$PV = RT + Pb \Rightarrow Z = \frac{PV}{RT} = 1 + \frac{Pb}{RT}$$

or **PV > RT** at higher pressure (above Boyle temperature)

This accounts for the rising parts of the PV vs P isotherm at high pressures.

- (iii) *At very low pressure* : V becomes so large that both b and $\frac{a}{V^2}$ become negligible and the Vander Waals equation reduces to **PV = RT**

$$Z = \frac{PV}{RT} = 1 ; Pb = \frac{a}{V} \text{ At extremely low pressure (at Boyle temperature)}$$

This shows why gases approach ideal behaviour at very low pressures.

- (iv) *Hydrogen and Helium* : These are two lightest gases known. Their molecules have very small masses. The attractive forces between such molecules will be extensively small. So $\frac{a}{V^2}$ is negligible even at ordinary temperatures. Thus **PV > RT**.

Dieterici Equation :

$$P(V - nb) = nRT e^{a/VRT} \text{ (for 'n' mole of a gas)}$$

Berthelot Equation :

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

Virial Equation Of State For 1 Mole Of Gas :

$$Z = \frac{PV}{RT} = 1 + B \frac{1}{V} + C \frac{1}{V^2} + D \frac{1}{V^3} + \dots$$

B = second virial co-efficient , temperature dependent = $b - \frac{a}{RT}$

C = third virial co – efficient, temperature dependent = b^2

Illustration 16 : The compressibility factor for 1 mole of a van der Waals gas at 0°C and 100 atm pressure is found to be 0.5. Assuming that the volume of a gas molecule is negligible, calculate the van der Waals constant, a .

Solution: For 1 mole of the gas,

$$Z = \frac{pV}{RT} \Rightarrow 0.5 = \frac{100 \times V}{.0821 \times 273}; V = 0.112 \text{ L}$$

Neglecting b , van der Waals equation reduces to $\left(P + \frac{a}{V^2}\right) V = RT$

$$\text{or } pV + \frac{a}{V} = RT$$

$$\text{or } 100 \times 0.112 + \frac{a}{0.112} = .0821 \times 273$$

$$a = 1.25 \text{ L}^2\text{atm mol}^{-2}$$

7 Critical phenomenon & Liquification of gases

The phenomena of converting a gas into liquid is known as liquifaction. The liquifaction of gas is achieved by controlling P and T as follows:

1. **Increasing pressure :** An increase in pressure results in an increase in attraction among molecules.
2. **Decreasing temperature :** A decrease in temperature results in decrease in kinetic energy of molecules.

Critical temperature (T_c) : It is defined as the characteristic temperature for a given gas below which a continuous increase in pressure will bring liquification of gas and above which no liquifaction is noticed although pressure may be increased e.g. T_c for CO_2 is 31.2°C .

$$T_c = \frac{8a}{27Rb}$$

Critical pressure (P_c) : It is defined as the minimum pressure applied on 1 mole of gas placed at critical temperature, to just liquefy the gas

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

Critical Volume (V_c) : The volume occupied by 1 mole of gas placed at critical conditions.

$$V_c = 3b \text{ (i.e. } P = P_c \text{ and } T = T_c)$$

8. Collision parameters

– **Mean free path**

$$\lambda = \frac{d_1 + d_2 + \dots + d_n}{n} = \eta \sqrt{\frac{3}{Pd}} \quad (\eta = \text{coefficient of viscosity})$$

$$\lambda = \frac{\text{Average velocity / RMS velocity}}{\text{collision number or frequency}} = \frac{kT}{\sqrt{2} \pi \sigma^2 P}$$

k = Boltzman constant ; σ = collision diameter.

- **Collision frequency (z)** : no. of collision taking place per second per unit volume.
- **Collision diameter** : Closest distance between the centre of two molecule which are participating in collision.
- **Relative Humidity (RH)** :
At a given temperature

$$RH = \frac{\text{Partial pressure of water vapour in air at } T}{\text{vapour pressure of water at } T}$$
- **Loschmidt Number**
The number of molecules present in 1 cc of gas (or) vapour at STP.
value : 2.617×10^{19} 1 cc

8. Daily Practice Problem Sheet

- I. Instructions :** From Question (1 to 4) a single statement is made. Write T if the statement is true & F if the statement is false.
1. A gas with $a = 0$ can be very easily liquefied.
 2. Mean free path increases with the increase in temperature at constant pressure.
 3. The behaviour of a real gas approach ideal gas at high temperature & low pressure.
 4. Real gases deviate from ideal behaviour because their molecules cause attractive interaction on one another.
- II. Instructions :** Questions (5 to 11) consist of a problem followed by several alternative answers, only one of which is correct. Mark the letter corresponding to the correct answer.
5. A real gas obeying vander waal equation will resemble ideal gas if the
 (A) constants a and b are negligibly small (B) a is large and b is small
 (C) a is small and b is large (D) constant a and b are large
 6. The compressibility factor of real gas is usually greater than one ($Z > 1$) at high temperature and high pressure. This is because
 (A) the constant a is negligible while b is not (B) the constant b is negligible while a is not
 (C) both a and b are negligible (D) both a and b are no negligible
 7. For three different gases values of vander Waal's constant 'a' and 'b' are given. What is the correct order of liquefaction of gases :

Gases	a	b
X_2	1.3	0.090
Y_2	4.1	0.023
Z_2	2.2	0.075

 (A) $X_2 > Y_2 > Z_2$ (B) $Y_2 > Z_2 > X_2$ (C) $Z_2 > Y_2 > X_2$ (D) $X_2 > Z_2 > Y_2$
 8. The value of vander Waal's constant 'a' for the gases O_2 , N_2 , NH_3 and CH_4 are 1.36, 1.39, 4.17 and $2.253 \text{ l}^2 \text{ atm, m}^{-2}$, respectively. The gas which can most easily be liquefied is
 (A) O_2 (B) N_2 (C) NH_3 (D) CH_4

9. What is not true about the vander waal constant 'b' among the statements given below ?
- I. It is called excluded volume II. It accounts for the interparticle forces
 III. Its units are mol dm^{-3} IV. Its value depends on molecular size
 (A) I, II (B) II, IV (C) II, III (D) I, IV
10. Among the four statements given below for real gases, which are incorrect :
- (A) The molecules causes attractive interactions.
 (B) They show deviation from ideal gas laws.
 (C) The molecules are almost volume less points.
 (D) The molecules have negligible mass.
11. Vander Waal's constant for three different gases are given
- | Gas | a | b |
|-----|------|-------|
| X | 3.0 | 0.025 |
| Y | 10.0 | 0.030 |
| Z | 6.0 | 0.035 |
- Which is correct ?
- (A) Maximum critical temperature - Y (B) Most ideal behaviour - X
 (C) Maximum molecular volume - Z (D) All are correct

SOLVED PROBLEMS

SUBJECTIVE

Problem 1 :

Calculate the total pressure in a 10 L cylinder which contains 0.4 g of helium, 1.6g of oxygen and 1.4 g of nitrogen at 27°C. Also calculate the partial pressure of helium gas in the cylinder. Assume ideal behaviour for gases.

Solution :

$$n_{\text{He}} = 0.4/4 = 0.10, \quad n_{\text{O}_2} = 1.6/32 = 0.05, \quad n_{\text{N}_2} = 1.4/28 = 0.05$$

$$\text{Total moles} = 0.2$$

$$PV = nRT \quad \Rightarrow \quad P \times 10 = 0.2 \times 0.082 \times 300 \quad \Rightarrow \quad P = 0.492 \text{ atm}$$

Problem 2 :

One litre flask contains O_2 at 0°C and 7.6×10^{-10} mm Hg. Calculate the number of molecules of O_2 present.

Solution :

$$PV = nRT$$

$$P = \frac{7.6 \times 10^{-10}}{760} \text{ atmosphere} = 10^{-12} \text{ atmosphere}, \quad V_1 = 1 \text{ litre} \quad n = ?$$

$$R = 0.082 \text{ litre L atm K}^{-1} \text{mol}^{-1}, \quad T = 273 \text{ K}$$

$$\therefore 10^{-12} \times 1 = n \times 0.082 \times 273$$

$$\therefore n = 4.46 \times 10^{-12} \text{ moles}$$

$$\text{Molecules} = 4.46 \times 10^{-12} \times 6.02 \times 10^{23} = 2.68 \times 10^{10}$$

Problem 3 :

Assuming that N_2 molecules is spherical and its radius is 2×10^{-10} meter, calculate the empty space in one mole of N_2 gas at NTP.

Solution :

$$\text{Volume of one molecule of } N_2 = \frac{4}{3} \pi r^3$$

$$= \frac{4 \times 3.14 (2 \times 10^{-10})^3}{3} = 3.35 \times 10^{-29} \text{ m}^3 = 3.35 \times 10^{-23} \text{ cm}^3$$

$$\text{Volume of one mole } N_2 \text{ molecules} = 6.02 \times 10^{23} \times 3.35 \times 10^{-23} \text{ cm}^3 = 20.167 \text{ cm}^3$$

$$\text{Volume of one mole } N_2 \text{ at NTP} = 22400 \text{ cc.}$$

$$\text{Hence empty space} = 22400 - 22379.833 \text{ cm}^3$$

$$\% \text{ empty space} = \frac{\text{Empty volume} \times 100}{\text{Total Volume}} = \frac{22379.833 \times 100}{22400} = 99.9\%$$

Problem 5 :

Using Vander Waals equation calculate the pressure exerted by one mole of CO_2 . Its volume at 373 K is 0.05 dm^3 . Given $a = 3.592 \text{ atm L}^2 \text{ mol}^{-2}$ and $b = 0.0426 \text{ L mol}^{-1}$.

Solution :

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

$$P = ?, n = 1, V = 0.05 \text{ litre,}$$

$$a = 3.592,$$

$$b = 0.04267,$$

$$T = 373 \text{ K,}$$

$$R = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}$$

$$\left(P + \frac{3.592}{(0.05)^2} \right) (0.05 - 0.04267) = 1 \times 0.0821 \times 373 \Rightarrow P = 1620.95 \text{ atm}$$

Problem 6 :

Density of a mixture of CO and CO_2 at 303 K and 73 cm of Hg is 1.5 gram/litre. What is the mole percent of two gases in the mixture.

Solution :

$$PV = \frac{wRT}{M} \quad \text{or} \quad P = \frac{wRT}{M \times V} \quad \text{or} \quad P = \frac{dRT}{M}$$

$$P = \frac{73}{76} \text{ atm, } d = 1.5 \text{ g/litre, } T = 303 \text{ K, } R = 0.0821 \text{ L atm K}^{-1} \text{ mol}^{-1}, \quad M = ?$$

$$\therefore \frac{73}{76} = \frac{1.5 \times 0.0821 \times 303}{M} \Rightarrow M = 38.85 \text{ (molecular mass of the mixture)}$$

$$\text{Suppose, Total moles} = 100, \text{ Moles of CO} = a, \quad \text{Moles of } CO_2 = 100 - a$$

$$\text{Mass of } CO_2 = (100 - a)44$$

$$\text{Total mass} = 28a + (100 - a)44$$

$$\text{Molecular mass of mixture} = \frac{28a + (100 - a)44}{100}$$

$$\text{But molecular mass of mixture is } 38.85$$

$$38.85 = \frac{28a + (100 - a)44}{100} \Rightarrow a = 32.19$$

Mole percent of CO = 31.19, Mole percent of CO₂ = 67.81

Problem 7 :

Two grams of gas A are introduced in a evacuated flask at 25°C. The pressure of the gas is 1 atm. Now 3g of another gas B is introduced in the same flask, the total pressure becomes 1.5 atm. Calculate (A) the ratio of molecular mass M_A and M_B (B) volume of the vessel, if A is O₂.

Solution :

For gas A : Its mass = 2g ; P_A = 1 atm , T = 298 K

For gas B : Its mass = 3g , P_B = 0.5 atm. T = 298 K

According to Dalton's law of partial pressure

$$P = p_A + p_B \quad 1.5 = 1 + p_B \quad \therefore p_B = 0.5 \text{ atm}$$

$$\text{Now, } PV = \frac{wRT}{m}$$

For gas A : p_A = 1 atm, m = M_A, w = 2g

$$1 \times V = \frac{2 \times RT}{M_A} \quad \dots (1)$$

for gas B : p_B = 0.5 atm, m = M_B, w = 3g

$$0.5 \times V = \frac{3 \times RT}{M_B} \quad \dots (2)$$

$$\text{Dividing the equation (1) and (2)} \quad 0.5 = \frac{3 \times M_A}{2 M_B} \quad \text{or} \quad \frac{M_A}{M_B} = \frac{1}{3}$$

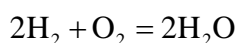
(B) If A is O₂ then M_A = 32, T = 298 K, V = ?

$$\text{Put these values in equation (1)} \quad V = \frac{2 \times 0.0821 \times 298}{32} = 1.53 \text{ litres}$$

Problem 8 :

One litre O₂ and one litre H₂ are taken in a vessel of 2 litre at STP. These gases are made to react to form water. Calculate (A) moles and weight of water formed. (B) amount of gas left in the vessel. (C) Total pressure of the gas at 100°C.

Solution :



Volume before reaction (L) 1 1 0

Volume after reaction (L) 0 0.5 1

(A) For H₂O : PV = nRT, At STP P = 1 atm, V = 1 litre,

$$\therefore \text{Moles of H}_2\text{O}(n) = 4.46 \times 10^{-2}$$

$$\text{Weight of H}_2\text{O} = 4.46 \times 10^{-2} \times 18 = 0.803 \text{ gram}$$

(B) Gas left in the vessel = 0.5 litre at STP

∴ 22.4 litre O_2 at STP weigh 32 g

∴ 0.5 litre O_2 at STP weigh $32g = \frac{32 \times 0.5}{22.4} = 0.714$ gram

(C) Moles of H_2O formed = 4.46×10^{-2}

O_2 left = 0.5 litre at STP = 2.23×10^{-2} moles

Total moles in the vessel = $4.46 \times 10^{-2} + 2.23 \times 10^{-2} = 6.69 \times 10^{-2}$

$PV = nRT$, $P = ?$ $V = 2$ litre, $n = 6.69 \times 10^{-2}$, $T = 373$ K

$P \times 2 = 6.69 \times 10^{-2} \times 0.0821 \times 373 \Rightarrow \therefore P = 1.02$ atm

Problem 9 :

At $27^\circ C$ vapour density of the mixture of NO_2 and N_2O_4 is 38.3. Calculate the moles of NO_2 in 100 g mixture.

Solution :

V.D. of mixture = 38.3, M.Wt. of mixture = 2×38.3 , Wt. of mixture = 100 g

Number of moles present = $\frac{100}{2 \times 38.3}$... (1)

Suppose mass of NO_2 = a gram Mass of N_2O_4 = (100 - a) gram

Moles of NO_2 = $w / M = a / 46$

Moles of N_2O_4 = $w / M = (100 - a) / 92$

Total moles $\frac{a}{46} + \frac{100 - a}{92}$... (2)

Equating equation (1) and (2)

$$\frac{100}{2 \times 38.3} = \frac{a}{46} + \frac{100 - a}{92}$$

∴ Moles of NO_2 = $w / M = 20.1 / 46 = 0.43$

OBJECTIVE

Problem 1 :

The ratio of the root mean square velocity of H_2 at 50 K and that of O_2 at 800 K is
 (A) 4 (B) 2 (C) 1 (D) 1/4

Solution : (C)

$$U_{\text{rms}} = \sqrt{\frac{3RT}{M}}; U_{\text{rms}} \propto \sqrt{\frac{T}{M}} \Rightarrow \frac{U_{\text{rms}}(H_2 \text{ at } 50 \text{ K})}{U_{\text{rms}}(O_2 \text{ at } 800 \text{ K})} = \sqrt{\frac{50}{2} \times \frac{32}{800}} = 1$$

Problem 2 :

If for two gases of molecular weights M_A and M_B at temperature T_A and T_B , $T_A M_B = T_B M_A$, then which property has the same magnitude for both the gases.
 (A) density (B) pressure (C) KE per mole (D) U_{rms}

Solution : (D)

(i) Density of a gas $(\rho) = \frac{PM}{RT}$. Since $\frac{M_B}{T_B} = \frac{M_A}{T_A}$.

\therefore at the same pressure $\rho_A = \rho_B$. But if pressure is different then $\rho_A \neq \rho_B$.

(ii) Pressure of the gases would be equal if their densities are equal other wise not.

(iii) KE per mole $= \frac{3}{2}RT$

\therefore it will be different for the two gases

(iv) $U_{\text{rms}} = \sqrt{\frac{3RT}{M}}$, Since $\frac{T_A}{M_A} = \frac{T_B}{M_B}$; U_{rms} of A = U_{rms} of B

Problem 3 :

What percent of a sample of nitrogen must be allowed to escape if its temperature, pressure and volume are to be changed from 220°C , 3 atm and 1.65 L to 100°C , 0.7 atm and 1 L respectively?

(A) 41.4 % (B) 8.18 % (C) 4.14 % (D) 81.8 %.

Solution : (D)

$$n_1 = \frac{3 \times 1.65}{R \times 493}; n_2 = \frac{0.7 \times 1}{R \times 383} \Rightarrow \text{Fraction remaining} = \frac{n_2}{n_1} = \frac{0.7 \times 1}{R \times 383} \times \frac{R \times 493}{3 \times 1.65} = 0.182$$

$$\text{Fraction escaped} = 1 - 0.182 = 0.818 \Rightarrow \text{Percentage escaped} = 81.8 \%$$

Problem 4 :

Assuming that air is 79 % by mole of N_2 , 20 % O_2 and 1 % Ar, the density of air at 25°C and 1 atm is :

(A) 1.18 g/lit (B) 1.08 g/lit (C) 1.28 g/lit (D) 1.0 g/lit.

Solution : (A)

$$PV = nRT \Rightarrow P = \frac{w}{M} \frac{RT}{V} = \frac{\rho RT}{M} \Rightarrow \rho = \frac{PM_{\text{eff}}}{RT}$$

$$M_{\text{eff}} = (0.79 \times 28) + (0.20 \times 32) + (0.01 \times 39.9) = 28.9 \Rightarrow \therefore \rho = \frac{28.9 \times 1}{0.0821 \times 298} = 1.18 \text{ g/lit}$$

Problem 5 :

0.2 g of a gas X occupies a volume of 0.44 L at same pressure and temperature. Under identical conditions of P and T, 0.1 g of CO₂ gas occupies 0.32 L. Gas X can be

- (A) O₂ (B) SO₂ (C) NO (D) C₄H₁₀

Solution : (B)

$$0.1 \text{ g CO}_2 = \frac{0.1}{44} \text{ moles of CO}_2 \text{ occupies} = 0.32 \text{ L}$$

$$\frac{0.2}{M} \text{ moles of X occupies} = 0.44 \text{ L}$$

$$PV = nRT \Rightarrow V \propto n \Rightarrow \frac{V(\text{CO}_2)}{V(\text{X})} = \frac{n(\text{CO}_2)}{n(\text{X})} \Rightarrow \frac{0.32}{0.44} = \frac{0.1 \times M}{44 \times 0.2} \Rightarrow M = 64$$

Problem 6 :

A flask containing 12 g of a gas of relative molecular mass 120 at a pressure of 100 atm was evacuated by means of a pump until the pressure was 0.01 atm. Which of the following is the best estimate of the number of molecules left in the flask

$$(N_0 = 6 \times 10^{23} \text{ mol}^{-1}).$$

- (A) 6×10^{19} (B) 6×10^{18} (C) 6×10^7 (D) 6×10^{13}

Solution : (B)

$$P_1 = \frac{n_1}{V} RT \quad \text{and} \quad P_2 = \frac{n_2}{V} RT, \quad \text{Since } V \text{ and } T \text{ remain constant hence } \frac{n_1}{n_2} = \frac{P_1}{P_2}$$

$$n_2 = n_1 \times \frac{P_2}{P_1} = \frac{12}{120} \times \frac{0.01}{100} = 10^{-5} \Rightarrow \text{Number of molecules left in the flask} = n_2 \times N_0 = 6 \times 10^{18}$$

Problem 7 :

The behaviour of a real gas is usually depicted by plotting compressibility factor Z versus P at a constant temperature. At high temperature and high pressure, Z is usually more than one. This fact can be explained by van der Waals equation when

- (A) the constant 'a' is negligible and not 'b'
 (B) the constant 'b' is negligible and not 'a'
 (C) both constant 'a' and 'b' are negligible
 (D) both the constant 'a' and 'b' are not negligible.

Solution : (A)

$$(P + n^2 a / V^2)(V - nb) = nRT$$

At low pressures, 'b' can be ignored as the volume of the gas is very high. At high temperature 'a' can be ignored as the pressure of the gas is high.

$$\therefore P(V - b) = RT \Rightarrow PV - Pb = RT \Rightarrow PV = RT + Pb \Rightarrow \frac{PV}{RT} = Z = 1 + \frac{Pb}{RT}$$

Problem 8 :

Under identical conditions of temperature, the density of a gas X is three times that of gas Y while molecular mass of gas Y is twice that of X. The ratio of pressures of X and Y will be

- (A) 6 (B) 1/6 (C) 2/3 (D) 3/2.

Solution : (A)

$$P\rho = \frac{PM}{RT} \Rightarrow P = \frac{\rho RT}{M} \Rightarrow P_x = \frac{\rho_x RT}{M_x}, \quad P_y = \frac{\rho_y RT}{2M_x}, \quad 6P_y = \frac{\rho_x RT}{M_x} \Rightarrow \frac{P_x}{P_y} = 6$$

Problem 9 :

X ml of H_2 gas effuses through a hole in a container in 5 seconds. The time taken for the effusion of the same volume of the gas specified below under identical conditions is :

- (A) 10 seconds : He (B) 20 seconds : O_2
 (C) 25 seconds : CO (D) 35 seconds : CO_2

Solution : (B)

$$\frac{r_{H_2}}{r_{He}} = \sqrt{\frac{4}{2}} = \sqrt{2} \Rightarrow \therefore (A) \text{ is incorrect} \quad \frac{r_{H_2}}{r_{O_2}} = \sqrt{\frac{32}{2}} = 4$$

$$(B) \text{ is correct} \quad \frac{r_{H_2}}{r_{CO}} = \sqrt{\frac{28}{2}} = \sqrt{14} \quad (C) \text{ is incorrect} \quad \frac{r_{H_2}}{r_{CO_2}} = \sqrt{\frac{2}{44}} = \sqrt{\frac{1}{22}} \quad (D) \text{ is incorrect}$$

Problem 10 :

One mole of $N_2O_4(g)$ at 300 K is kept in a closed container under one atmosphere pressure. It is heated to 600 K when 20% by mass of $N_2O_4(g)$ decomposes to $NO_2(g)$. The resultant pressure is

- (A) 1.2 atm (B) 2.4 atm (C) 2 atm (D) 1 atm.

Solution : (B)

$$\text{Initial pressure of } N_2O_4 \text{ at 600 K, } \Rightarrow P_{600K} = \frac{1}{300} \times 600 = 2 \text{ atm}$$



$$\text{At equilibrium} \quad 1 - \alpha \quad 2\alpha$$

$$\therefore \text{Moles at equilibrium} = 1 - \alpha + 2\alpha = 1 + \alpha$$

$$\therefore \text{The resultant pressure} = 2(1 + \alpha) = 2(1 + 0.2) = 2.4 \text{ atm}$$

ANSWERS

1. Daily Practice Problem Sheet

1. B 2. B 3. A 4. A 5. A 6. B 7. D
8. (i) 10 cm of Hg (ii) 66 cm of Hg (iii) 86 cm of Hg 9. 35.5
10. (A) 10129.28, (B) 76000, (C) 1033.6, (D) 76000, (E) 100 11. 67 N
12. 27.2 cm

2. Daily Practice Problem Sheet

1. 500 torr 2. 2625 mm 3. 0.1073 atm 4. 0.2 atm 5. 45 ml

3. Daily Practice Problem Sheet

1. (A) it will not burst (B) 44.67 °C 2. 81 °C 3. 99.5 °C 4. 1093 torr

4. Daily Practice Problem Sheet

1. D 2. A 3. D 4. B 5. 3.01×10^{22} , 6.02×10^{22}
6. 9.02×10^{22}
7. (i) 3.0155×10^{24} (ii) 3.0155×10^{24} , 3.0155×10^{24} , 9.0465×10^{24}
(iii) 4.5345×10^{24} , 1.5115×10^{24} , 250 g
8. 225 g 9. 3.31×10^{24} 10. (ii) > (iv) > (iii) > (i)
11. $10 \times 6.022 \times 10^{23}$ 12. (i) 138.5 g (ii) 350 g (iii) 212 g

5. Daily Practice Problem Sheet

1. 96 g 2. 2.03×10^{-2} mole 3. 17.97 dm³ 4. 262.2 mL
5. 327.6 K 6. 41.33 7. 0.74 cm 8. 34.02
9. 1.5124 g litre⁻¹ 10. (A) 0.37 litre (B) 181.55 (C) 2.27 atm (D) 0.418 litre
11. 36.36 g 12. 0.56 atm 13. 1.13 atm

6. Daily Practice Problem Sheet

1. (A) 1/3 (B) 73.89 atm 2. 489.23 mm (B) 260.77 mm
3. $P_M = 2.4$ atm, $P_{N_2} = 0.667$ atm, $P_{O_2} = 0.733$ atm, $P_{CO_2} = 1$ atm
4. PP of CO = 0.49 atm, PP of CO₂ = 1.51 atm 5. (A) 0.867 atm (B) 0.102 atm, 0.51 atm
6. 6.82 cc 7. M = 16 8. 2.198 atm 9. 25.69
10. NH₄Cl will form at 1.186 m from the NH₃ end of the tube 11. 79.3 g mol⁻¹

7. Daily Practice Problem Sheet

1. 600 K 2. 3.9×10^4 cm/sec
3. 4.93×10^4 cm/sec, 4.59×10^4 cm/sec, 4.03×10^4 cm/sec 4. 183.8×10^3 cm/sec

GASEOUS STATE

33

5. 0.35 cm/sec 6. 47.3×10^3 cm/sec, 6.8×10^4 cm/sec, 10.9 cm/sec 7. 819°C
8. 8.028×10^4 cm/sec 9. 3.3×10^7 dyne/cm²
10. 6.84×10^4 cm/sec, 1871.42 J, 62.14×10^{-22} J
11. (i) $\frac{u_1}{u_2} = 4$ (ii) T = 1200 K (iii) no effect (iv) no effect 12. 2.237 : 1
13. 0.958 : 1, 927°C 14. $u_{\text{rms}} = 7.20 \times 10^4$ cm sec⁻¹, T = 393.66°C 15. 280 K, 525 K

8. Daily Practice Problem Sheet

1. False 2. True 3. True 4. True
5. A 6. A 7. B 8. C 9. C 10. CD 11. D

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