Fill Ups of States of Matter

Q.1. The total energy of one mole of an ideal monatomic gas at 27ºC is .... calories. (1984 - 1 Mark)

Ans. 900

Sol. 900; Energy of one mole of an ideal monoatomic gas **3/2** RT

R = 8.314 J K^{-1} $mole^{-1}$ = 1.99 cal K^{-1} $mole^{-1}$ T = 27°C = 273 + 27 = 300 K

⇒ E = **3/2** × 1.99 × 300 = 900 cal

Q.2. C_p – C_v for an ideal gas is .... (1984 - 1 Mark)

Ans. R

Sol. R. [∵ $C_p - C_v = R$]

Q.3. The rate of diffusion of gas is .... proportional to both .... and square root of molecular mass. (1986 - 1 Mark)

Ans. inversely, time

Sol. Inversely, time; $\left[\frac{r_1}{r_2} = \sqrt{\frac{d_2}{d_1}} = \frac{t_2}{t_1} = \sqrt{\frac{M_2}{M_1}}\right]$.

Q.4. The value of PV for 5.6 litres of an ideal gas is .... RT, at N.T.P. (1987 - 1 Mark)

Ans. 0.25

Sol.

0.25; $\left(\because 5.6l \equiv \frac{1}{4} \text{ mole of an ideal gas}\right)$

Using ideal gas equation

$$PV = nRT = \frac{1}{4}\text{RT} = 0.25\,\text{RT}$$

Q.5. Eight gram each of oxygen and hydrogen at 27ºC will have the total kinetic energy in the ratio of .... (1989 - 1 Mark)

Ans. 1 : 16

Sol. 1 : 16; $\frac{K.E._{O_2}}{K.E._{H_2}} = \frac{(3/2)n_{O_2}RT}{(3/2)n_{H_2}RT} = \frac{n_{O_2}}{n_{H_2}} = \frac{8/32}{8/2} = \frac{1}{16}$

True False of States of Matter

Q.1. Kinetic energy of a molecule is zero at 0°C. (1985 - ½ Mark)

Ans. F

Sol. False : K.E. = $\frac{3}{2}KT$, and cannot be zero at 0°C or 273 K.

Q.2. A gas in a closed container will exert much higher pressure due to gravity at the bottom than at the top. (1985 - ½ Mark)

Ans. F

Sol. False : The pressure exerted by the gas is the result of collisions of the molecules on the walls of the container.

Q.3. In the van der Waal's equation $(V - nb) = nRT$ the constant 'a' reflects the actual volume of the gas molecules. (1993 - 1 Mark)

Ans. F

Sol. False : The constant 'a' reflects the intermolecular attraction between gaseous molecules. The constant 'b' reflects the actual volume of one mole of gaseous molecules.

Q.4. A mixture of ideal gases is cooled upto liquid helium temperature (4.22 K) to form an ideal solution. (1996 - 1 Mark)

Ans. F

Sol. False : An ideal gas cannot be liquefied as there exists no intermolecular attraction between molecules.

Subjective questions of States of Matter

Q.1. Calculate density of NH_3 at 30°C and 5 atm pressure. (1978)

Ans. Sol. TIPS/Formulae :

$$PV=\frac{m}{M}RT \Rightarrow P=\frac{m}{V}\times\frac{RT}{M}=d\frac{RT}{M} \quad \therefore d=\frac{MP}{RT}$$

Substituting the value, we get

$$d=\frac{17\times5}{0.082\times303}=3.42\text{ g/litre}$$

Q.2. 3.7 g of a gas at 25°C occupied the same volume as 0.184g of hydrogen at 17°C and at the same pressure. What is the molecular weight of the gas? (1979)

Ans. Sol. Given, moles $=\frac{\text{mass}}{\text{mol.wt}}$

mass of gas = 3.7g,

mass of hydrogen = 0.184g $T_1 = 298K$, $T_2 = 17°C = 273 + 17 = 290K$

Moles of $H_2 = n_1 = \frac{\text{Mass}}{\text{M. wt.}}=\frac{0.184}{2}=0.092$

Moles of gas $= n_2 = \frac{\text{Mass}}{\text{M. wt.}}=\frac{3.7}{M}$

For hydrogen $P_1V_1 = n_1RT_1$ ....(i)

For gas $P_1V_1 = n_2RT_2$ ....(ii)

($\because$ Pressure and volume of gas are same)

$\therefore$ From equation (i) and equation (ii)

$$\frac{P_1V_1}{P_1V_1}=\frac{n_1RT_1}{n_2RT_2} \text{ or } 1 \ \frac{0.092\times 298}{n_2\times 290}$$

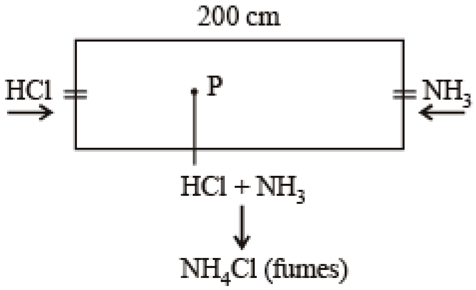
$$\text{or } n_2=\frac{0.092\times 298}{290} \quad \text{or } \frac{3.7}{\text{M}}=\frac{0.092\times 298}{290}$$

$$\text{or } \frac{3.7}{\text{M}}=0.0945 \ \therefore \text{M}=\frac{3.7}{0.0945}=\mathbf{39.15}$$

Q.3. A straight glass tube has two inlets X and Y at two ends. The length of tube is 200 cm. HCl gas through inlets X and NH_3 gas through inlet Y are allowed to enter the tube at the same time. What fumes appear at point P inside the tube. Find distance of P from X. (1980)

Ans. Sol. Let NH_3 diffuse through = x cm

HCl diffuses through = y cm



According to Graham's law of diffusion

$$\frac{x}{y}=\sqrt{\frac{\text{Mol. wt HCl}}{\text{Mol. wt of } NH_3}}=\sqrt{\frac{36.5}{17}}=\sqrt{2.14}=1.465$$

x = 1.465 y ... (1)

x + y = 200 cm ... (2)

From these equations; y = 85.2 cm

Distance between P and X = y = 85.2 cm.

Q.4. 1 litre of mixture of CO and CO_2 is taken. The mixture is passed through a tube containing red hot charcoal. The volume now becomes 1.6 litre. The volumes are measured under the same conditions. Find the composition of mixture by volume. (1980)

Ans. Sol. Following reaction takes places in tube $C + CO_2 \longrightarrow 2CO$

Volume of mixture of CO and CO_2 = 1L

Let volume of CO_2 in mixture = x

∴ Volume of CO in mixture = 2x

∴ Original volume of CO in mixture = 1 – x

Total volume of CO after reaction = (1 – x) + 2x = 1 + x 1 + x = 1.6

(∵ It is given total volume after reaction = 1.6L)

∴ x = 0.6 L

∴ Volume of CO_2 = 0.6 L Volume of CO = 0.4 L CO_2 : CO = 3 : 2

Q.5. At room temperature, ammonia gas at 1 atm pressure and hydrogen chloride gas at P atm pressure are allowed to effuse through identical pin holes from opposite ends of a glass tube of one metre length and of uniform cross-section. Ammonium chloride 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? (1982 - 4 Marks)

Ans. Sol. Since the pressures of gases are different, and the temperature is constant, the rate at which molecules of the two gases diffuse is directly proportional to the pressure. This rate of diffusion is also directly proportional to the distance travelled by the gas.

Hence r_1 (of HCl gas) at pressure P = $60 = \frac{kP}{\sqrt{36.5}}$ .....(i)

and r_2 (of NH_3) at 1 atm. pressure P = $40 = \frac{k \times 1}{\sqrt{17}}$ .....(ii)

From (i) and (ii)

$$\frac{r_1}{r_2} = \frac{60}{40} = \frac{kP}{\sqrt{36.5}} \times \frac{\sqrt{17}}{k \times 1}$$

$$P = \frac{60}{40} \times \frac{\sqrt{36.5}}{\sqrt{17}} = 2.197 \text{ atm}$$

Q.6. Calculate the average of kinetic energy, in Joules of the molecules in 8.0 g of methane at 27°C. (1982 - 2 Marks)
Ans. Sol. TIPS/Formulae : Total kinetic energy = n (3/2 RT)

where n = Number of moles of the gas

R = Gas constant

T = Absolute temperature Molecular weight of methane, $CH_4 = 12 + 4 \times 1 = 16$

$\therefore$ Number of moles of methane in 8.0 gm of methane

$$= \frac{8.0}{16.0} = 0.5$$

R = 8.314 joules/K/mole, T = 27 + 273 = 300 K

$\therefore$ Total kinetic energy of the molecules in 8.0 gm of methane at 27°C = $n \times 3/2\ RT = 0.5 \times 3/2 \times 8.314 \times 300 = 1870.65$ joules

$$\therefore \text{Average kinetic energy} = \frac{1870.65}{6.023 \times 10^{23} \times 0.5}$$

$$= 6.21 \times 10^{-21} \text{ joules/molecule}$$

Q.7. Oxygen is present in 1 litre flask at a pressure of 7.6×10^{-10} mm of Hg. Calculate the number of oxygen molecules in the flask at 0°C. (1983 - 2 Marks)

Ans. Sol. NOTE THIS STEP : First we should calculate the number of moles of the gas under the given conditions by the relation PV = nRT

Here $P = 7.6 \times 10^{-10}$ mm Hg

$$= \frac{7.6 \times 10^{-10}}{760} \text{atm.} = 1 \times 10^{-12} \text{ atm.}$$

V = 1 litre, T = 273 + 0 = 273K, R = 0.082 litre atm./K/mol Putting the values in equation

$$n = \frac{PV}{RT} = \frac{1\times 10^{-12}\times 1}{0.082\times 273}\text{moles}$$

Now since 1 mole = 6.023×10^{23} molecules

$$\frac{10^{-12}}{0.082\times 273}\text{moles} = \frac{6.023\times 10^{23}\times 10^{-12}}{0.082\times 273}\text{molecules}$$

$= 2.7 \times 10^{10}$ molecules

Q.8. When 2 gm of a gas A is introduced into an evaluated flask kept at 25°C, the pressure is found to be one atmosphere. If 3 gm of another gas B is then added to the same flask, the total pressure becomes 1.5 atm. Assuming ideal gas behaviour, calculate the ratio of the molecular weights MA : MB. (1983 - 2 Marks)

Ans. Sol. From ideal gas equation,

$$PV = nRT \Rightarrow PV = \left(\frac{m}{M}\right) RT \text{ or } M = m\frac{RT}{PV}$$

Let the molecular wt. of A and B be M_A and M_B respectively.

Then $M_A = 2\frac{RT}{1\times V}$; $M_B = \frac{3\times RT}{0.5\times V}$

$$\therefore \frac{M_A}{M_B} = \frac{2RT}{V}\times\frac{0.5V}{3RT} = \frac{2\times 0.5}{3} = \frac{1}{3}$$

Therefore, the ratio $M_A : M_B = 1 : 3$

Q.9. Calculate the root mean square velocity of ozone kept in a closed vessel at 20°C and 82 cm mercury pressure. (1985 - 2 Marks)

Ans. Sol.

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

Given T = 20°C = 20 + 273 = 293K R = 8.314×10^7 erg per degree per mol M (of O_3) = 48

$$\therefore U_{rms} = \sqrt{\frac{3\times 8.314\times 10^7\times 293}{48}} = 3.9 \times 10^4 \text{ cm sec}^{-1}$$

Q.10. A spherical balloon of 21 cm diameter is to be filled up with hydrogen at N.T.P. from a cylinder containing the gas at 20 atmospheres at 27°C. If the cylinder can hold 2.82 litres of water, calculate the number of balloons that can be filled up. (1987 - 5 Marks)

Ans. Sol. Volume of ballon = 4.851 L (as calculated above)

Let no. of balloons to be filled = n

∴ Total volume occupied by n balloons = 4.851× n

Volume of H_2 present in cylinder = 2.82 L (given)

∴ Total volume of H_2 at NTP = (4.851n + 2.82)L

P_1 = 1 atm P_2 = 20 atm

V_1 = 4.85 × n + 2.82 L V_2 = 2.82 L

T_1 = 273 K T_2 = 300 K

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \text{ or } \frac{1 \times (4.85 \times n + 2.82)}{273} = \frac{20 \times 2.82}{300}$$

$$\therefore n = \frac{48.504}{4.851} \simeq 10$$

Q. 11. The average velocity at T_1K, and the most probable velocity at T_2K of CO_2 gas is 9.0 × 104 cm sec^{-1}. Calculate the value of T_1 and T_2. (1990 - 4 Marks)

Ans. Sol. TIPS/Formulae :

Average velocity = $\sqrt{\frac{8RT}{\pi M}}$

and Most probable velocity = $\sqrt{\frac{2RT}{M}}$

Given -For CO_2 Average velocity at

T_1 = Most probable velocity at T_2

$= 9 \times 10^4 \text{ cm/sec} = \frac{9 \times 10^4}{100} \text{ m/sec.}$

$= 9 \times 10^2$ m/sec.

$$\therefore 9 \times 10^2 = \sqrt{\frac{8 \times 8.314 \times T_1}{3.14 \times 44 \times 10^{-3}}} \ ...(A)$$

[Average velocity at T_1K]

and $9 \times 10^2 = \sqrt{\frac{2 \times 8.314 \times T_2}{44 \times 10^{-3}}}$ ...(B)

[Most probable velocity at T_2K] On solving, $T_1 = 1682.5$ K, $T_2 = 2143.4$ K

Q.12. Calculate the volume occupied by 5.0 g of acetylene gas at 50°C and 740 mm pressure. (1991 - 2 Marks)

Ans. Sol. Applying the general gas equation

$$PV = nRT = \frac{m}{M} RT$$

Here, Mol. wt. of acetylene i.e., C_2H_2 (M) = 26, P = $\frac{740}{760}$ atm,

T = 50°C = 50 + 273 = 323 K

$$\therefore \quad V = \frac{mRT}{MP} \text{ or } V = \frac{5 \times 0.082 \times 323 \times 760}{26 \times 740} = 5.23 \text{ L}$$

Q. 13. At 27°C, hydrogen is leaked through a tiny hole into a vessel for 20 minutes. Another unknown gas at the same temperature and pressure as that of H_2 is leaked through the same hole for 20 minutes. After the effusion of the gases the mixture exerts a pressure of 6 atmosphere. The hydrogen content of the mixture is 0.7 mole. If the volume of the container is 3 litres, what is the molecular weight of the unknown gas? (1992 - 3 Marks)

Ans. Sol. Using gas equation; PV = nRT

Total no. of moles of gases in the mixture (n)

$$= \frac{PV}{RT} = \frac{6 \times 3}{0.0821 \times 300} = 0.7308 \text{ mol.}$$

Thus no. of moles of unknown gas = 0.7308 – 0.7 = 0.0308 mol.

Now we know that

$$\frac{r_1}{r_2} = \frac{\text{moles of hydrogen gas}}{\text{moles of unknown gas}} = \frac{0.7}{0.0308}$$

Also we know that $\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$

$$\therefore \quad M_2 = \left(\frac{r_1}{r_2}\right)^2 M_1 \quad \text{or} \quad M_2 = \left(\frac{0.7}{0.0308}\right)^2 \times 2 \; = 1033$$

Q.14. At room temperature the following reactions proceed nearly to completion : (1992 - 4 Marks)

$2NO + O_2 \rightarrow 2NO_2 \rightarrow N_2O_4$

The dimer, N_2O_4, solidifies at 262 K. A 250 ml flask and a 100 ml. flask are separated by a stop-cock. At 300 K, the nitric oxide in the larger flask exerts a pressure of 1.053 atm. and the smaller one contains oxygen at 0.789 atm. The gases are mixed by opening the stopcock and after the end of the reaction the flasks are cooled at 220K. Neglecting the vapour pressure of the dimer, find out the pressure and composition of the gas remaining at 220 K. (Assume the gases to behave ideally).

Ans. Sol. $2NO + O_2 \rightarrow 2NO_2 \rightarrow N_2O_4$

Calculating the number of moles of NO and O_2 by applying the formula,

$n = \frac{PV}{RT}$

Moles of NO in the larger flask = $\frac{1.053 \times 0.250}{0.082 \times 300} = 0.0107$

[250 mL = 0.250 L]

Moles of O_2 in the smaller flask = $\frac{0.789 \times 0.100}{0.082 \times 300} = 0.0032$

[100 mL = 0.100 L]

The reaction takes place as follows.

$$\text{Mols before} \quad \begin{array}{ccccc} 2NO & + & O_2 & \longrightarrow & N_2O_4 \\ 0.0107 & & 0.0032 & & 0 \end{array}$$

reaction Mole after (0.0107 – 0 0.0032 reaction 2 × .0032)

Hence moles of NO reacting completely with 0.0032 moles of O_2 = 2 × 0.0032 = 0.0064 Moles of NO left = 0.0107 – 0.0064 = 0.0043

NOTE : Oxygen will be completely changed into NO_2 which in turn is completely converted into N_2O_4which solidifies at 262 K. Hence at 220 K, the dimer is in the solid state and only NO present in excess will remain in the gaseous state occupying volume equal to 350 ml.

Hence pressure (P) of NO gas left

$$= \frac{nRT}{V} = \frac{0.0043 \times 0.082 \times 220}{0.350} = 0.221 \text{ atm}$$

[Total volume = 0.250 + 0.100 = 0.350 L]

Q.15. A gas bulb of 1 litre capacity contains 2.0 × 10^{21} molecules of nitrogen exerting a pressure of 7.57 × 10^3 Nm^{-2}. Calculate the root mean square (r.m.s) speed and the temperature of the gas molecules. If the ratio of the most probable speed to the root mean square speed is 0.82, calculate the most probable speed for these molecules at this temperature. (1993 - 4 Marks)

Ans. Sol. Given V= 1L = 10^{-3} m^3, P = 7.57 × 10^{-3} Nm^{-2}, R = 8.314J, n = 2 × $10^{21}/6.023 \times 10^{23}$ moles

$$PV = nRT \text{ or } T = \frac{PV}{nR}$$

$$= \frac{7.57 \times 10^{-3} \times 10^{-3} \times 6.023 \times 10^{23}}{2 \times 10^{21} \times 8.31} = 274.13 \text{ K}$$

$$U_{rms} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3 \times 8.314 \times 274.13}{28 \times 10^{-3}}} \text{ m/s} = 494.15 \text{ m/s}$$

(Given U)

$$\frac{U_{mp}}{U_{rms}} = \text{(given)}$$

$$\therefore U_{mp} = 0.82 \times U_{rms} = 0.82 \times 494.15 = 405.2 \text{ m/sec}$$

Q.16. A 4 : 1 molar mixture of He and CH_4 is contained in a vessel at 20 bar pressure . Due to a hole in the vessel, the gas mixture leaks out. What is the composition of the mixture effusing out initially? (1994 - 2 Marks)

Ans. Sol. TIPS/Formulae : Partial pressure = Mole fraction × Total pressure

$$\therefore p_{He} = x_{He} \times P = \frac{4}{5} \times 20 = 16 \text{ bar}$$

$$\left[\text{mole fraction of He} = \frac{4}{5}\right]$$

$$\therefore pCH_4 = 20 - 16 = 4 \text{ bar}$$

Now applying the formula

$$\frac{r_{He}}{r_{CH_4}} = \frac{P^0_{He}}{P^0_{CH_4}}\sqrt{\frac{M_{CH_4}}{M_{He}}} = \frac{16}{4}\sqrt{\frac{16}{4}} = \frac{16}{4}\sqrt{4}$$

$\therefore$ rHe :rCH_3 = 8 :1

$\therefore$ Composition of the mixture (He : CH_4) effusing out = 8 :1

Subjective questions of States of Matter (Part -2)

Q.17. An LPG (liquefied petroleum gas) cylinder weighs 14.8 kg when empty. When full, it weighs 29.0 kg and shows a pressure of 2.5 atm. In the course of use at 27º C, the weight of the full cylinder reduces to 23.2 kg. Find out the volume of the gas in cubic meters used up at the normal usage conditions, find the final pressure inside the cylinder.

Assume LPG to be n-butane with normal boiling point of 0ºC. (1994 - 3 Marks)

Ans. Sol. Calculation of volume of gas : Weight of cylinder with gas = 29.0 kg

Weight of empty cylinder = 14.8 kg

∴ Weight of gas in the cylinder = 14.2 kg Pressure in cylinder = 2.5 atm

∴ No. of moles (n) in 14.2 kg (14.2×10^3g) of butane

$$n = \frac{\text{Wt. of butane}}{\text{Mol. wt. of butane}} = \frac{14.2 \times 10^3}{58} = 244.83 \text{ mol}$$

Applying gas equation,

$$V = \frac{nRT}{P} = \frac{244.83 \times 0.0821 \times 300}{2.5} = 2412 \text{ litres}$$

[27°C = 273 + 27 = 300]

Calculation of pressure in cylinder after use.

Weight of cylinder after use = 23.2 kg

Weight of empty cylinder = 14.8 kg

∴ Wt. of unused gas = 8.4 kg $= \frac{8.4 \times 10^3}{58}$ moles of butane

Thus P = $\frac{nRT}{V} = \frac{8.4 \times 10^3 \times 0.0821 \times 300}{58 \times 2412}$ = 1.478 atm [V = 2412 L]

Calculation of volume of used gas at 2.5 atm and 27ºC.

Weight of used gas = 14.2 – 8.4 = 5.8 kg

Pressure under normal usage conditions = 1 atm

$$V=\frac{nRT}{P}=\frac{5.8\times10^{3}}{58}\times\frac{0.0821\times300}{1}\quad\left[\because n=\frac{5.8}{58}\right]$$

$= 2463$ litres $= 2.463\ m^3$

Q.18. A mixture of ethane (C_2H_6) and ethene (C_2H_4) occupies 40 litres at 1.00 atm and at 400 K. The mixture reacts completely with 130 g of O_2 to produce CO_2 and H_2O. Assuming ideal gas behaviour, calculate the mole fractions of C_2H_4 and C_2H_6 in the mixture. (1995 - 4 Marks)

Ans. Sol. Let the volume of ethane in mixture = x litre

∴ Volume of ethene = (40 – x) litre

Combustion reactions of ethane and ethene are :

(i) $C_2H_6(g)+3\frac{1}{2}O_2(g)\longrightarrow 2CO_2(g)+3H_2O(l)$

or $2C_2H_6(g) + 7O_2(g)$ ¾¾→ $4CO_2(g) + 6H_2O(l\ )$

(ii) $C_2H_4(g) + 3O_2(g)$ ¾¾→ $2CO_2(g) + 2H_2O(l\ )$

Volume of O_2 required for complete combustion of ethane

$=\frac{7x}{2}$ [For x litres]

Volume of O_2 required for complete combustion of ethene = (40–x) × 3 [For (40 – x) L]

∴ Total volume of O_2 required

$=\frac{7x}{2}+(40-x)3\ l$

Calculation of number of moles (n)

P = 1 atm, $V=\frac{7x}{2}+(40-x)3\ l$; R = 0.082 l atm K^{-1} mol^{-1};

T = 400 K

Since $n = \frac{PV}{RT} = \frac{1\times\left[\frac{7x}{2}+(40-x)3\right]}{0.082\times 400} = \frac{7x+(40-x)6}{2\times 0.082\times 400}$

Mass of n moles of $O_2 = \left[\frac{7x+(40-x)6}{2\times 0.082\times 400}\right]\times 32 = 130$

or $130 = \left[\frac{7x+240-6x}{65.6}\right]\times 32$

$\Rightarrow 8528 = 32x + 240 \times 32 \Rightarrow 32x = 848 \Rightarrow$ or $x = \frac{848}{32} = 26.5$

Hence mole fraction (%) of ethane $= \frac{26.5}{40}\times 100 = 66.25\%$

Mole fraction (%) of ethene = 33.75%

Q.19. The composition of the equilibrium mixture (Cl2 2Cl), which is attained at 1200°C, is determined by measuring the rate of effusion through a pin–hole. It is observed that at 1.80 mmHg pressure, the mixture effuses 1.16 times as fast as krypton effuses under the same conditions. Calculate the fraction of the chlorine molecules dissociated into atoms. (Relative atomic mass of Kr = 84.) (1995 - 4 Marks)

Ans. Sol.

Mixture	Kr ypton
$r_{mix} = 1.16$	$r_{Kr} = 1$
$M_{mix} = ?$	$M_{Kr} = 84$

We know that

$$\frac{r_{mix}}{r_{Kr}} = \sqrt{\frac{M_{Kr}}{M_{min}}} \text{ or } \frac{1.16}{1} = \sqrt{\frac{84}{M_{mix}}}$$

$$\text{or } (1.16)^2 = \frac{84}{M_{mix}} \Rightarrow M_{mix} = \frac{84}{(1.16)^2} = 62.426$$

Determination of the composition of the equilibrium mixture

Let the fraction of Cl_2 molecules dissociated at equlibrium $= x$

	$Cl_2 \rightleftharpoons$	$2Cl$	Total
Initially	1	0	1
At equilibrium	$1 - x$	$2x$	$1 - x + 2x = 1 + x$

$\therefore$ Total moles at equilibrium $= 1 - x + 2x = 1 + x$

$$\because \frac{\text{Normal molecular mass}}{\text{Experimental molecular mass}} = 1 + x$$

$$\therefore \frac{71}{64.426} = 1 + \alpha$$

$\therefore a = 0.137 = 13.7\%$.

Q. 20. A 20.0 cm^3 mixture of CO, CH_4 and He gases is exploded by an electric discharge at room temperature with excess of oxygen. The volume contraction is found to be 13.0 cm^3. A further contraction of 14.0 cm^3 occurs when the residual gas is treated with KOH solution. Find out the composition of the gaseous mixture in terms of volume percentage. (1995 - 4 Marks)

Ans. Sol. TIPS/Formulae :

(i) He does not react with oxygen.

(ii) KOH absorbs only CO_2.

NOTE : When the mixture of CO, CH_4 and He gases (20 ml) are exploded by an electric discharge with excess of O_2, He gas remains as such and the other reactions involved are :

$$CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g) \qquad ...(i)$$

$CH_4 (g) + 2O_2 (g) \rightarrow CO_2 (g) + 2H_2O(l)$ ...(ii)

Let the volumes of CO and CH_4 to be ‘a’ ml and ‘b’ ml in the mixture then

Volume of He gas = [20 – (a + b)] ml

For the initial contraction of 13 ml,

Volume of left hand side in the above reactions – 13 = Volume of right hand side

$$\therefore \quad [20-(a+b)]+\left(a+\frac{1}{2}a\right)+(b+2b)-13$$

= [20 – (a + 2b)] + a + b [neglect the volume of H_2O (l)](Since for gases, volume a no. of moles)

∴ 1/2 a + 2 b = 13 or a + 4 b = 26 ...(iv)

NOTE THIS STEP : The CO_2 produced above in reactions (ii) & (iii), (a + b) ml, reacts with KOH sol for a further contraction of 14 ml.

$CO_2(g) + 2KOH(l) \rightarrow K_2CO_3(l) + H_2O(l)$ (a + b) ml

∴ a + b = 14 ...(v)

Solving (iv) & (v) we get, a = 10 ml & b = 4 ml

$$\therefore CH_4=\frac{4}{20}\times 100=\mathbf{20\%}, \quad CO=\frac{10}{20}\times 100=\mathbf{50\%}$$

& He = 100 – (20 + 50) = 30%

Q. 21. An evacuated glass vessel weighs 50.0 g when empty, 148.0 g when filled with a liquid of density 0.98 g mL^{-1} and 50.5 g when filled with an ideal gas at 760 mmHg at 300K. Determine the molar mass of the gas. (1998 - 3 Marks)

Ans. Sol. Weight of liquid = 148 – 50 = 98 g

Volume of liquid = $\frac{98}{0.98}$ = 100 ml = volume of vessel It means, vessel of 100 ml contains ideal gas at 760 mm Hg at 300 K

Weight of gas = 50.5 – 50 = 0.5g

using, $PV=nRT=\frac{w}{m}RT$

$$\frac{760}{760}\times\frac{100}{1000}=\frac{0.5}{m}\times 0.082\times 300 \quad \left[n=\frac{0.5}{m}\right]$$

∴ Molecular weight of gas (m) = 123

Q.22. The degree of dissociation is 0.4 at 400 K and 1.0 atm for the gaseous reaction $PCl_5PCl_3 + Cl_2$. Assuming ideal behaviour of all gases, calculate the density of equilibrium mixture at 400 K and 1.0 atmosphere. (Relative atomic mass of P = 31.0 and Cl = 35.5) (1998 - 3 Marks)

Ans. Sol.

	PCl_5	$\rightleftharpoons$	PCl_3	$+$	Cl_2
Initial moles	1		0		0
Moles at eq.	1–0.4		0.4		0.4

$\therefore$ Total moles at equilibrium = 1 – 0.4 + 0.4 + 0.4 = 1.4

$$\text{Also } \frac{\text{Normal mol. wt. of } PCl_5}{\text{Exp. mol. wt. of } PCl_5} = 1 + a = 1.4$$

$$\text{or } \frac{208.5}{\text{Exp. mol. wt. of } PCl_5} = 1.4$$

$$\therefore \text{ Exp. mol. wt. of } PCl_5 \text{ or m. wt. of mixture} = \frac{208.5}{1.4}$$

Now using, $PV = \frac{w}{m}RT$ for mixture

$$d = \frac{w}{V} = \frac{Pm}{RT} = \frac{1 \times 208.5}{1.4 \times 0.082 \times 400} = 4.53 \text{ g/litre}$$

Q.23. Using van der waal's equation, calculate the constant, 'a' when two moles of a gas confined in a four litre flask exerts a pressure of 11.0 atmospheres at a temperature of 300 K.

The value of 'b' is 0.05 L mol^{-1}. (1998 - 4 Marks)

Ans. Sol. van der Waals equation for n moles of gas is

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

Given V = 4 litre; P = 11.0 atm, T = 300 K; b = 0.05 litre mole^{-1}, n =2

$$\text{Thus, } \left[11 + \frac{2^2 a}{4^2}\right][4 - 2 \times 0.05] = 2 \text{ x } 0.082 \text{ x } 300$$

$\therefore$ a = 6.46 atm litre^2 mol^{-2}

Q.24. For the reaction, $N_2O_5(g) \rightarrow 2NO_2(g) + 0.5\ O_2(g)$, calculate the mole fraction of $N_2O_5(g)$ decomposed at a constant volume and temperature, if the initial pressure is 600 mm Hg and the pressure at any time is 960 mm Hg. Assume ideal gas behaviour. (1998 - 3 Marks)

Ans. Sol. $N_2O_5(g) \rightleftharpoons 2NO_2(g) + \frac{1}{2}O_2(g)$

Initial pressure	600	0	0
Final pressure	600-P	2P	P/2

P ∝ moles when V and T are constant (where moles equivalent to pressure P are decomposed)

Total pressure = 600 – P + 2 P + P/2 = 960 mm of Hg

∴ P = 240 mm Hg

Thus moles of N_2O_5 decomposed = $\frac{240}{600} = 0.4$

Q.25. One mole of nitrogen gas at 0.8 atm takes 38 s to diffuse through a pinhole, whereas one mole of an unknown compound of xenon with flourine at 1.6 atm takes 57 s to diffuse through the same hole. Calculate the molecular formula of the compound. (1999 - 5 Marks)

Ans. Sol. We know that

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}} \times \frac{P_1}{P_2} \text{ or } \frac{n_1}{t_1} \times \frac{t_2}{n_2} = \sqrt{\frac{M_2}{M_1}} \times \frac{P_1}{P_2}$$

$$\text{or } \frac{1}{38} \times \frac{57}{1} = \sqrt{\frac{M}{28}} \times \frac{0.8}{1.6} \ \therefore M = 252$$

$$\left[\begin{array}{l} Xe + (F)_x = 252 \\ 131 + 19x = 252 \ ; \quad \therefore x = 6 \end{array}\right.$$

Thus compound of xenon with fluorine is XeF_6

Q.26. The pressure exerted by 12 g of an ideal gas at temperature t°C in a vessel of volume V litre is one atm. When the temperature is increased by 10 degrees at the same volume, the pressure increases by 10%. Calculate the temperature t and volume V. (Molecular weight of the gas = 120.) (1999 - 5 Marks)

Ans. Sol. (I) Given P = 1 atm, w = 12 g; T = (t + 273)K; V = V litre

(II) If T = t + 10 + 273 = t + 283 K ; V = V litre,

$$P = 1 + \frac{10}{100} = 1.1 \text{ atm}$$

Using gas equation, $PV = \frac{w}{m}RT$

Case I. $1 \times V = \frac{12}{m}R(t+273)$ ... (1)

Case I. $1.1 \times V = \frac{12}{m}R(t+283)$ ...(2)

From (1) and (2), t = −173°C or t = 100 K

Also from (1), on substituting t and m (120), V = 0.82 litre

Q.27. Calculate the pressure exerted by one mole of CO_2 gas at 273 K if the van der Waal's constant a = 3.592 dm^6 atm mol^{-2}. Assume that the volume occupied by CO_2 molecules is negligible. (2000 - 2 Marks)

Ans. Sol.

vander Waals' equation for one mole of a gas is

$$\left[P + \frac{a}{V^2}\right](V - b) = RT \qquad ...(1)$$

Given that volume occupied by CO_2 molecules, 'b' = 0

Hence, (1) becomes $\left[P + \frac{a}{V^2}\right]V = RT$ or $P = \frac{RT}{V} - \frac{a}{V^2}$

Using R = 0.082, T = 273K, V = 22.4 l for 1 mole of an ideal gas at 1 atm pressure.

$$\therefore P = \frac{0.082 \times 273}{22.4} - \frac{3.592}{(22.4)^2} = 0.9922 \text{ atm.}$$

Q. 28. The compression factor (compressibility factor) for one mole of a van der Waals gas at 0°C and 100 atmospheric pressure is found to be 0.5. Assuming that the volume of a gas molecule is negligible, calculate the van der Waals constant a. (2001 - 5 Marks)

Ans. Sol. We know that, Compressibility factor , $Z = \frac{PV}{RT}$

$0.5 = \frac{100 \times V}{0.082 \times 273}$ $\therefore$ V = 0.1119L

NOTE : Further when volume of a gas molecule is negligible, van der Waal's equation becomes

$$\left(P + \frac{a}{V^2}\right)(V - 0) = RT$$

or PV = RT - $\frac{a}{V}$ or a = RTV- PV^2

Substituting the values a = (0.082 × 0.1119 × 273) – (100 × 0.1119 × 0.1119)

= 1.253 atm L^2 mol^{-2}

Q.29. The density of the vapour of a substance at 1 atm pressure and 500 K is 0.36 kgm^{-3}. The vapour effuses through a small hole at a rate of 1.33 times faster than oxygen under the same condition. (2002 - 5 Marks)

(a) Determine (i) molecular weight, (ii) molar volume, (iii) compression factor (Z) of the vapour and (iv) which forces among the gas molecules are dominating, the attractive or the repulsive?

(b) If the vapour behaves ideally at 1000 K, determine the average translational kinetic energy of a molecule.

Ans. Sol. (a) d = 0.36 kg m^{-3} = 0.36 g/L (i)

From Graham's Law of diffusion

$$\frac{r_v}{r_{O_2}} = \sqrt{\frac{M_{O_2}}{M_v}}; \; 1.33 = \sqrt{\frac{32}{M_v}}$$

$$\therefore M_v = \frac{32}{(1.33)^2} = 18.09;$$

where M_V = MW of the vapour

(ii) Thus, 0.36g = $\frac{0.36}{18.09}$mol

$\frac{0.36}{18.09}$mol occupies 1 L volume, so 1 mol occupies

$\frac{18.09}{0.36}L = 50.25L$

Thus, molar volume of vapour = 50.25 L

Assuming ideal behaviour the volume of the vapour can be calculated by

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \Rightarrow V_2 = 22.4 \times \frac{500}{273} = 41.025L$$

(iii) Compressibility factor (Z)

$$= \frac{(PV)_{obs}}{(PV)_{ideal}} = \frac{1 \times 50.25}{1 \times 41.025} = 1.224$$

(iv) Z is greater than unit y, hence it is the shor t range repulsive force that would dominate. (∵ actual density is less than given density)

(b) $E = \frac{3}{2}KT = \frac{3}{2} \times \frac{8.31}{6.02 \times 10^{23}} \times 100$

$= 2.07 \times 10^{-20}$ J per molecule

(∵ K, Boltzmann constant = R/N)

Q.30. The average velocity of gas molecules is 400 m/sec. Calculate its rms velocity at the same temperature. (2003 - 2 Marks)

Ans. Sol. TIPS/Formulae :

$$C_{rms} = \sqrt{\frac{3RT}{M}}, C_{av} = \sqrt{\frac{8RT}{\pi M}}$$

$$\frac{C_{rms}}{C_{av}} = \sqrt{\frac{3RT}{M}} \times \sqrt{\frac{\pi M}{8RT}} = \sqrt{\frac{3\pi}{8}} = 1.085$$

$$C\ rms = 1.085 \times C_{av} = 1.085 \times 400 = 434\ ms^{-1}$$

Q. 31. A graph is plotted between PVm along Y-axis and P along X-axis, where Vm is the molar volume of a real gas. Find the intercept along Y-axis. (2004 - 2 Marks)

Ans. Sol. The van der Waal equation (for one mole) of a real gas is

$$\left(P+\frac{a}{V_m^2}\right)(V_m-b)=RT\ ;\ PV_m-Pb+\frac{a}{V_m}-\frac{ab}{V_m^2}=RT$$

$$PV_m=RT+Pb-\frac{a}{V_m}+\frac{ab}{V_m^2}\ \ldots(\text{i})$$

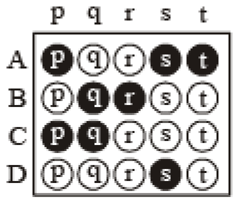
NOTE THIS STEP : To calculate the intercept P → 0, hence $V_m \rightarrow \infty$ due to which the last two terms on the right side of the equation (i) can be neglected.

$\therefore PV_m = RT + Pb$

When P = 0, intercept = RT

Match the following of States of Matter

Each question contains statements given in two columns, which have to be matched. The statements in Column-I are labelled A, B, C and D, while the statements in Column-II are labelled p, q, r, s and t. Any given statement in Column-I can have correct matching with ONE OR MORE statement(s) in Column-II. The appropriate bubbles corresponding to the answers to these questions have to be darkened as illustrated in the following example :



	p	q	r	s	t
A	●	○	○	●	●
B	○	●	●	○	○
C	●	●	○	○	○
D	○	○	○	●	○

If the correct matches are A-p, s and t; B-q and r; C-p and q; and D-s then the correct darkening of bubbles will look like the given.

Q.1. Match gases under specified conditions listed in Column I with their properties/laws in Column II. Indicate your answer by darkening the appropriate bubbles of the 4 × 4 matrix given in the ORS. (2007)

Column I	**Column II**
(A) Hydrogen gas (P = 200 atm, T = 273 K)	**(p) Compressibility factor ≠ 1**
(B) Hydrogen gas (P ~ 0, T = 273 K)	**(q) Attractive forces are dominant**
(C) CO_2 (P = 1 atm, T = 273 K)	**(r) PV = nRT**
(D) Real gas with very large molar volume	**(s) P(V – nb) = nRT**

Ans. (A) - (p, s); (B) - (r); (C) - (p, q); (D) - (r);

Sol. (A) : (p) and (s) Because 200 atm pressure is very large.

For H_2 gas, at very high pressure Z > 1.

(B) : (r) Since P ~ 0, it means very low presure, so ideal behaviour is observed.

(C) : (p) and (q) Since P is 1 atm, Z for CO_2 would be less than 1.

(D) : (r) In real gas with very high molar volume, molecules will be very far apart from each other due to which van der Waal's forces as well as actual volume occupied by molecules will be negligible.

Integer Type ques of States of Matter

Q.1. At 400 K, the root mean square (rms) speed of a gas X (molecular weight = 40) is equal to the most probable speed of gas Y at 60 K. The molecular weight of the gas Y is (2009)

Ans. 4

Sol.

$$v_{rms} \text{ of } X = \sqrt{\frac{3RT_x}{M_x}}\ ; \quad v_{mp} \text{ of } Y = \sqrt{\frac{2RT_y}{M_y}}$$

$$\text{Given } vr_{ms} = v_{mp} \Rightarrow \sqrt{\frac{3RT_x}{M_x}} = \sqrt{\frac{2RT_y}{M_y}}$$

$$\Rightarrow M_y = \frac{2RT_yM_x}{3RT_x} = \frac{2\times 60\times 40}{3\times 400} = 4$$

Q.2. To an evacuated vessel with movable piston under external pressure of 1 atm, 0.1 mol of He and 1.0 mol of an unknown compound (vapour pressure 0.68 atm. at 0°C) are introduced.

Considering the ideal gas behaviour, the total volume (in litre) of the gases at 0°C is close to (2011)

Ans. 7

Sol. $P_{He} = 1 - 0.68 = 0.32$ atm, $n = 0.1$

$$V = \frac{nRT}{P} = \frac{0.1\times 0.0821\times 273}{0.32} = 7$$

Q. 3. The diffusion coefficient of an ideal gas is proportional to its mean free path and mean speed. The absolute temperature of an ideal gas is increased 4 times and its pressure is increased 2 times. As a result, the diffusion coefficient of this gas increases x times. The value of x is (JEE Adv. 2016)

Ans. 4

Sol. Diffusion coefficient $\propto \lambda\ \mu$

Since $\lambda \propto \frac{T}{P}$ and $\mu \propto \sqrt{T}$

$\therefore$ Diffusion coefficient $\propto \frac{T\sqrt{T}}{P}$

Thus $\frac{D_i}{D_f} = \frac{\frac{T\sqrt{T}}{P}}{\frac{4T\sqrt{4T}}{2P}} : = \frac{1}{(4\times 2)/2} = \frac{1}{4}$ or $\frac{D_f}{D_i} = \frac{4}{1}$