MLL CHAPTER 1 NUMERICALS

Q1Calculate the mass percent of different elements present in sodium sulphate (Na2SO4).

A1

The molecular formula of sodium sulphate is Na_2SO_4 .

Molar mass of Na_2SO_4 = [(2 × 23.0) + (32.066) + 4 (16.00)]
= 142.066 g

Mass percent of an element $= \dfrac{\text{Mass of that element in the compound}}{\text{Molar mass of the compound}} \times 100$

∴ Mass percent of sodium:

$= \dfrac{46.0\,g}{142.066\,g} \times 100$

$= 32.379$

$= 32.4\%$

Mass percent of sulphur:

$= \dfrac{32.066g}{142.066g} \times 100$

$= 22.57$

$= 22.6\%$

Mass percent of oxygen:

$= \dfrac{64.0g}{142.066g} \times 100$

$= 45.049$

$= 45.05\%$

Q2Calculate the amount of carbon dioxide that could be produced when (i) 1 mole of carbon is burnt in air. (ii) 1 mole of carbon is burnt in 16 g of dioxygen

A2 The balanced reaction of combustion of carbon can be written as: (i) As per the balanced equation, 1 mole of carbon burns in1 mole of dioxygen (air) to produce1 mole of carbon dioxide. (ii) According to the question, only 16 g of dioxygen is available. Hence, it will react with 0.5 mole of carbon to give 22 g of carbon dioxide. Hence, it is a limiting reactant.

Q3Determine the molecular formula of an oxide of iron in which the mass per cent of iron and oxygen are 69.9 and 30.1 respectively. Given that the molar mass of the oxide is 159.69 g mol–1.

A3

Mass percent of iron (Fe) = 69.9% (Given)

Mass percent of oxygen (O) = 30.1% (Given)

Number of moles of iron present in the oxide $= \frac{69.90}{55.85}$

= 1.25

Number of moles of oxygen present in the oxide $= \frac{30.1}{16.0}$

= 1.88

Ratio of iron to oxygen in the oxide,

$$= 1.25 : 1.88$$

$$= \frac{1.25}{1.25} : \frac{1.88}{1.25}$$

= 1 : 1.5

= 2 : 3

∴ The empirical formula of the oxide is Fe_2O_3.

Empirical formula mass of Fe_2O_3 = [2(55.85) + 3(16.00)] g Molar mass of Fe_2O_3 = 159.69 g

$$\therefore n = \frac{\text{Molar mass}}{\text{Emprical formula mass}} = \frac{159.69 \text{ g}}{159.7 \text{ g}}$$

$$= 0.999$$

$$= 1 (\text{approx})$$

Molecular formula of a compound is obtained by multiplying the empirical formula with n.

Thus, the empirical formula of the given oxide is Fe_2O_3 and n is 1.

Hence, the molecular formula of the oxide is Fe_2O_3.

Q4In three moles of ethane (C2H6), calculate the following: (i) Number of moles of carbon atoms. (ii) Number of moles of hydrogen atoms. (iii) Number of molecules of ethane

A4 (i) 1 mole of C2H6 contains 2 moles of carbon atoms. Number of moles of carbon atoms in 3 moles of C2H6 = 2 × 3 = 6

(ii) 1 mole of C2H6 contains 6 moles of hydrogen atoms. Number of moles of carbon atoms in 3 moles of C2H6 = 3 × 6 = 18

(iii) 1 mole of C2H6 contains 6.023 × 1023 molecules of ethane. Number of molecules in 3 moles of C2H6 = 3 × 6.023 × 1023 = 18.069 × 1023

Q5 What is the concentration of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2 L?

A5

Molarity (M) of a solution is given by,

$$= \frac{\text{Number of moles of solute}}{\text{Volume of solution in Litres}}$$

$$= \frac{\text{Mass of sugar/molar mass of sugar}}{2\ \text{L}}$$

$$= \frac{20\text{g} / \left[(12\times12)+(1\times22)+(11\times16)\right]\text{g}}{2\ \text{L}}$$

$$= \frac{20\text{g} / 342\ \text{g}}{2\ \text{L}}$$

$$= \frac{0.0585\ \text{mol}}{2\ \text{L}}$$

$= 0.02925$ mol L^{-1}

$\therefore$ Molar concentration of sugar $= 0.02925$ mol L^{-1}

Q6 Express the following in the scientific notation: (i) 0.0048 (ii) 234,000 (iii) 8008 (iv) 500.0 (v) 6.0012

A6 (i) 0.0048 = 4.8×10^{-3} (ii) 234, 000 = 2.34×10^{5} (iii) 8008 = 8.008×10^{3} (iv) 500.0 = 5.000×10^{2} (v) 6.0012 = 6.0012

Q7 How many significant figures are present in the following? (i) 0.0025 (ii) 208 (iii) 5005 (iv) 126,000 (v) 500.0 (vi) 2.0034

A 7(i) 0.0025 There are 2 significant figures.

(ii) 208 There are 3 significant figures.

(iii) 5005 There are 4 significant figures.

(iv) 126,000 There are 3 significant figures.

(v) 500.0 There are 4 significant figures.

(vi) 2.0034 There are 5 significant figures

Q8 Round up the following upto three significant figures: (i) 34.216 (ii) 10.4107 (iii) 0.04597 (iv) 2808

A8 (i) 34.2 (ii) 10.4 (iii) 0.0460 (iv) 2810

Q9What will be the mass of one ^{12}C atom in g?

A9 1 mole of carbon atoms = 6.023 × 1023 atoms of carbon = 12 g of carbon Mass of one 12C atom

= 1.993 × 10–23 g

Q10 Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation:

$N_2(g) + H_2(g) \rightarrow 2NH_3(g)$ (i) Calculate the mass of ammonia produced if 2.00 × 103 g dinitrogen reacts with 1.00 × 103 g of dihydrogen. (ii) Will any of the two reactants remain unreacted? (iii) If yes, which one and what would be its mass?

A 10 (i) Balancing the given chemical equation, $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

From the equation, 1 mole (28 g) of dinitrogen reacts with 3 mole (6 g) of dihydrogen to give 2 mole (34 g) of ammonia.

⇒ 2.00×10^3 g of dinitrogen will react with $\frac{6\text{ g}}{28\text{ g}} \times 2.00 \times 10^3$ g dihydrogen

i.e., 2.00×10^3 g of dinitrogen will react with 428.6 g of dihydrogen.

Given,

Amount of dihydrogen = 1.00×10^3 g Hence,

N_2 is the limiting reagent.

∴ 28 g of N_2 produces 34 g of NH_3.

Hence, mass of ammonia produced by 2000 g of N_2 $= \frac{34\text{ g}}{28\text{ g}} \times 2000$ g

= 2428.57 g

(ii) N_2 is the limiting reagent and H_2 is the excess reagent. Hence, H_2 will remain unreacted.

iii) Mass of dihydrogen left unreacted = 1.00 × 103 g – 428.6 g = 571.4 g

MLL FOR CHAPTER 7 EQUILIBRIUM (NUMERICALS)

CLASS XI

1. The following concentrations were obtained for the formation of NH3 from N2 and H2 at equilibrium at 500K. [N2] = 1.5 × 10–2M. [H2] = 3.0 ×10–2 M and [NH3] = 1.2 ×10–2M. Calculate equilibrium constanT.

Sol. The equilibrium constant for the reaction

N2(g) + 3H2(g) 2NH3(g) can be written as

$Kc = [NH_3]^2/[N_2][H_2]^3$

$= (1.2x10^{-2})^2/(1.5x10^{-2})(3.0x10^{-2})^3$

$= 0.106x10^4 = 1.06x10^3$

2. For the equilibrium, 2NOCl(g) 2NO(g) + Cl2(g) the value of the equilibrium constant, Kc is 3.75 × 10–6 at 1069 K. Calculate the Kp for the reaction at this temperature?

Sol- We know that, $Kp = Kc(RT^{)\Delta n}$ For the above reaction,

Δn = (2+1) – 2 = 1

$Kp = 3.75 \times 10^{-6}$ (0.0831X1069)

kP = 0.033

3. At equilibrium, the concentrations of N2=3.0 × 10–3M, O2 = 4.2 × 10–3M and NO= 2.8 × 10–3M in a sealed vessel at 800K.What will be Kc for the reaction
N2(g) + O2(g) ↔ 2NO(g)
Sol.
$Kc=[NO]^2/[N2]\ [O2]$

=(2.8x10-3)/(3x10-3) (4.2x10-3)

=0.622

4. Hydrolysis of sucrose gives
Sucrose $+H_2O$↔Glucose+Fructose
Equilibrium constant Kc for the reaction is $2x10^{13}$ at 300K.Calculate ΔG^0 at 300K.
Sol. $\Delta G^0 = -2.303RTlogKc$
$= -2.303x8.314x300xlog2x10^{13}$
$= -7.64x10^4$J/mol
5. The concentration of hydrogen ion in a sample of soft drink is 3.8 × 10–3M. what is its pH ?
Sol. is its pH ?
Solution pH = – log[3.8 × 10–3]
= – {log[3.8] + log[10–3]}
= – {(0.58) + (– 3.0)} = – { – 2.42} = 2.42 Therefore, the pH of the soft drink is 2.42 and it can be inferred that it is acidic.

6. The pKa of acetic acid and pKb of ammonium hydroxide are 4.76 and 4.75 respectively. Calculate the pH of ammonium acetate solution.
 Solution pH = 7 + ½ [pKa – pKb]
 = 7 + ½ [4.76 – 4.75]
 = 7 + ½ [0.01] = 7 + 0.005
 = 7.005
7. The pKa of acetic acid and pKb of ammonium hydroxide are 4.76 and 4.75 respectively. Calculate the pH of ammonium acetate solution.
 Solution
 pH = 7 + ½ [pKa – pKb]
 = 7 + ½ [4.76 – 4.75]
 = 7 + ½ [0.01]

 = 7 + 0.005 = 7.005

8. Calculate the solubility of A_2X_3 in pure water, assuming that neither kind of ion reacts with water. The solubility product of A_2X_3, Ksp = 1.1 × 10–23.
 Solution $A_2X_3 \rightarrow 2A^{3+} + 3X^2 = {}^-1.1 \times 10–23$
 S = solubility of A2X3, then [A3+] = 2S; $[X^{2–]} = 3S$
 Ksp = $(2S)^2(3S)^3 = 108S^5 = 1.1 \times 10–23$ thus,
 $S^5 = 1 \times 10–^{25}$ $S = 1.0 \times 10^{–5}$ mol/L.

9. For the following equilibrium,kc=$6.3x10^{14}$ at 1000K
 $NO(g) + O_3(g) \leftrightarrow NO_2(g) + O_2(g)$
 Both the forward and reverse reaction in the equilibrium are elementary bimolecular reactions.What is Kc for the reverse reaction?
 Sol K'c =1/Kc
 $=1/6.3x10^{14}=1.59x10^{-15}$

10 . Assuming complete dissociation calculate the pH of 0.003MHCl
 Sol. $HCl \rightarrow H^+ + Cl^-$

 $[H^+]=[HCl]=3x10^{-3}M$

 Ph=-log ($3X10^{-3}$) = 2.52

THERMODYNAMICS:

EXPECTED QUESTIONS WITH ANSWERS(NUMERICALS)

Q1. Define molar heat capacity. Calculate number of KJ required to raise the temperature of 60g Al from 35º C to 55º C

Ans- quantity of heat required to raise the temperature of 1 mol of substance by 1ºC or 1 K (C = 24 J/mol K) n= 60/ 27 = 2.22 mol
q= CnΔT
q= 24 x 2.22 x 20 = 1065.6 J
= 1.0656 KJ

Q2. Enthalpy of combustion of carbon is -393.5 KJ/mol. Find the heat released in the formation of 35.2g of CO2

Ans- C + O2 → CO2 (44gm)
ΔH = -393.5 KJ/mol
heat released in the formation of 44g CO2 = -393.5
∴ heat released in the formation of 35.2 g CO2 = 393.5 x 35.2/ 44
= 314.8 KJ

Q3. Find the enthalpy of formation of NH3 (g)
N2 + 3H2 →2NH3 ΔrHº = -92.4 KJ/mol

Ans- ΔfHº of NH3 = ΔrHº/2 = -92.4/2 = -46.2 KJ/mol

Q4. The equilibrium constant for the reaction is 10. Calculate the value of ΔGº at 300 K.

Ans- ΔGº = -2.303 RT log K

= -2.303 x 8.31 x 300 x log10

= - 55.27 J

= - 5.527 KJ/mol

Q5. At what temperature the reaction will be spontaneous.

2A + B → C

ΔH = 400 KJ/mol

ΔS = 2.0 KJ/mol K

Ans- ΔG= ΔH-TΔS, ΔG= 0 at equilibrium

T= ΔH/ΔS= 400/200 = 200K

HOTS QUESTIONS

Q1. Specific heat of gas is 0.313 J at constant volume, if molar mass is 40g/molfind atomicity.

Q2. A+B →C+D, ΔH = -10,000 J/mol, ΔS= -33.3 J/mol K

(i) At what temperature reaction is spontaneous?
(ii) At what temperature reaction will reverse?

Q3. In what way internal energy is different from enthalpy.

Q4. $C_2H_4 + C_{l2} \rightarrow C_2H_4C_{l2}$

ΔH = -270.6 KJ/mol, ΔS = -139 J/K/mol is the reaction favored by entropy, enthalpy, both or none.

Find ΔG*,* if T = 300K

Q5. For an isolated system $\Delta U = 0$ what will be ΔS?

SOLUTION TO HOTS QUESTIONS

Answer 1. Molar heat capacity (Cp) = 0.313 x 40 = 12.52 J/mol

Cp = R + Cv = 8.31 + 12.52 = 20.83 J/mol

Cp/Cv = 20.83/12.52 = 1.66

Hence gas is monoatomic.

Answer 2. $\Delta G = \Delta H - T\Delta S$, $\Delta G = 0$, at equilibrium

T = $\Delta H/\Delta S$ = *-1000/-33.3 = 300.3 K*

*Answer 3.I*nternal energy is sum of all forms of energy stored in a substance. The ΔU refers to heat change in a process if it does not change Δn, ΔT, ΔV. enthalpy is sum of internal energy and PV energy refers to heat change in a process if it is carried out at constant temperature and pressure.

*Answer 4.*as ΔH = -ve reaction is favored by enthalpy, as ΔS is –ve it is not favored by entropy.

$\Delta H = \Delta H - T\Delta S$

= -270.6 – 300 x 139 x 10-3

= -228.9 KJ/mol

*Answer 5.*for an isolated system with ΔU= 0, the spontaneous change will occur if $\Delta S>0$.

For eg, in isolated system involving intermixing of gases $\Delta U = 0$ but $\Delta S>0$ because of increase of disorder.

STATES OF MATTER(NUMERICALS)

1. A balloon is filled with hydrogen at room temperature. It will burst if pressure exceeds 0.2 bar. If at 1 bar pressure the gas occupies 2.27 L volume, up to what volume can the balloon be expanded?

Ans: $P_1V_1=P_2V_2$

$1 \times 2.27=0.2 \times V_2$

$V_2=2.27/0.2$

$=11.35L$

2.What will be the minimum pressure required to compress 500 dm3 of air at 1 bar to 200 dm3 at 30°C?

Ans: $P_1V_1=P_2V_2$

$1 \times 500=P_2 \times 200$

$P_2=500/200$

$= 2.5bar$

3.A vessel of 120 mL capacity contains a certain amount of gas at 35 °C and 1.2 bar pressure. The gas is transferred to another vessel of volume 180 mL at 35 °C. What would be its pressure?

Ans: $P_1V_1=P_2V_2$

$1.2 \times 120 =P_2 \times 180$

$P_2=0.8bar$

4.On a ship sailing in pacific ocean where temperature is 23.4 °C , a balloon is filled with 2 L air. What will be the volume of the balloon when the ship reaches Indian ocean, where temperature is 26.1°C ?

Ans: $V_1 /T_1=V_2 /T_2$

$2/296.4=V_2 /299.1$

$V_2=2.018l$

5.A student forgot to add the reaction mixture to the round bottomed flask at 27 °C but instead he/she placed the flask on the flame. After a lapse of time he realized his mistake, and using a pyrometer he found the temperature of the flask was 477 °C. What fraction of air would have been expelled out?

Ans: $V1 / T1 = V2 / T2$

$V1 / 300 = V1 + x / 750$

$750/300 = (V1 + x)/x$

$2.5 = (V1 / x) + 1$

$1.5 = V1 / x$

$x / V1 = 0.67$

$x = 0.67V1$

The amount of air expelled out is 0.67 times the volume of the flask 0.67/1.67= 0.4011fraction of air is expelled.

6. At 25°C and 760 mm of Hg pressure a gas occupies 600 mL volume. What will be its pressure at a height where temperature is 10°C and volume of the gas is 640m L.

Ans $P1V1 / T1 = P2V2 / T2$

$760X600/298 = P2X640/283$

$P2 = 760X600X283/298X640$

= 676.6 mm of Hg

7. At 0°C, the density of a certain oxide of a gas at 2 bar is same as that of dinitrogen at 5 bar. What is the molecular mass of the oxide?

Ans: $PV = nRT$

$PV = (m/M) RT$

$d = M1P1 / RT = M2P2 / RT$

$M1P1 = M2P2$

$M1 = 28X5/2$

$= 70$

8. What will be the pressure exerted by a mixture of 3.2 g of methane and 4.4 g of carbon dioxide contained in a 9 dm3 flask at 27 °C ?

Ans: $PV = nRT$

$n = (3.2/16 + 4.4/44) = 0.3$

$P = 0.3moleX0.083bardm3 /K/moleX300K/9dm3$

$P = 0.83bar.$

9.What will be the pressure of the gaseous mixture when 0.5 L of H_2 at 0.8 bar and 2.0 L of dioxygen at 0.7 bar are introduced in a 1L vessel at 27°C?

Ans: PV=nRT

$P_1V_1=P_xV_2 \Rightarrow 0.5X0.8= P_x .1 \Rightarrow$ pressure of H_2 in the flask=P_x=0.40bar

$P_1V_1=P_xV_2$◊$0.7X2=P_xx1 \Rightarrow$ pressure of oxygen in the flask= 1.4bar

Pressure of the gaseous mixture=1.4 + 0.4 = 1.8bar.

10.Density of a gas is found to be 5.46 g/dm3 at 27 °C at 2 bar pressure. What will be its density at STP?

Ans: d= PM/RT $\Rightarrow$

M = dRT/P

= 5.46X 0.083X300/2

= 70

d= (1.013X70)/(0.083X273)

= 3.13 g/dm3

11. Calculate the temperature of 4.0 mol of a gas occupying 5 dm3 at 3.32 bar. (R = 0.083 bar dm3 K –1 mol–1).

Ans: PV=nRT

T=PV/nR

=(3.32X5)/(4X0.083)

= 50K

12. Calculate the total pressure in a mixture of 8 g of dioxygen and 4 g of dihydrogen confined in a vessel of 1 dm3 at 27°C. R = 0.083 bar dm3 K –1 mol–1 .

Ans: $n= nO_2+nH_2$

= 8/32 + 4/2

=0.25 +2

= 2.25moles

PV=nRT;

P=nRT/V

= 2.25X0.083X300/1

= 56.025 bar

13. Calculate the volume occupied by 8.8 g of CO_2 at 31.1°C and 1 bar pressure. R = 0.083 bar L K^{-1} mol^{-1} .

Ans: $PV=nRT$;

$n=8.8/44= 0.2$

$V= nRT/P$

$= 0.2X0.083X304.1/1$

$= 5.05L$

14. 2.9 g of a gas at 95 °C occupied the same volume as 0.184 g of dihydrogen at 17 °C, at the same pressure. What is the molar mass of the gas?

Ans: $PV= nRT$

$V=nRT/P= (0.184/2)XRX290/P=(2.9/M)XRX368/P$

$0.092X290=368X2.9/M$

$M= 368X2.9/0.092X290$

$= 40$

15. A mixture of dihydrogen and dioxygen at one bar pressure contains 20% by weight of dihydrogen. Calculate the partial pressure of dihydrogen.

Ans: let 20g H_2 + 80g O_2

$n= 20/2 + 80/32= 10 +2.5= 12.5$

Mole fraction of H_2= 10/12.5= 0.8

$pH_2= 0.8XP$