Daily Tutorial Sheet-1	JEE Advanced (Archive)

1.(D) $4\,Al + 3\,O_2 \rightarrow 2\,Al_2O_3$

Atomic weight of Al = 27

Thus $4 \times 27\,g$ of Al reacts with oxygen $= 3 \times 32\,g$

$\therefore$ $27\,g$ of Al reacts with oxygen $= \dfrac{3 \times 32}{4 \times 27} \times 27 = 24\,g$

2.(A) In both cases, the same volume of hydrogen is evolved for the same amount of zinc reacted.

$$Zn + H_2SO_4 \longrightarrow ZnSO_4 + H_2\uparrow$$

$$Zn + 2NaOH \longrightarrow Na_2ZnO_2 + H_2\uparrow$$

3. Equivalent of $KMnO_4$ = Equivalents of $FeSO_4 \cdot 7H_2O$

5.4 ml of 0.1 N $KMnO_4 = \dfrac{5.4 \times 0.1}{1000} = 5.4 \times 10^{-4}$ equivalents

Amount of $FeSO_4 = 5.4 \times 10^{-4} \times$ Mol. wt. of $FeSO_4 \cdot 7H_2O$

$= 5.4 \times 10^{-4} \times 278 = 0.150\,g$

Total weight of mixture = 5.5 g

Amount of ferric sulphate = 5.5 – 0.150 g = 5.35 g

Hence moles of ferric sulphate $= \dfrac{\text{Mass}}{\text{Mol. wt.}} = \dfrac{5.35}{562} = 9.5 \times 10^{-3}$ gram-mole.

4.(A) The change involved is $MnO_4^- + e^- \longrightarrow MnO_4^{2-}$

i.e. it involves only one electron

Eq. wt. $= \dfrac{\text{Mol. wt.}}{\text{number of } e^- \text{ involved}} = \dfrac{M}{1} = M$ $\quad [\because \text{Mol. wt.} = M]$

5. The complete oxidation under acidic conditions can be represented as follows :

$$5H_2O_2 + 2MnO_4^- + 6H^+ \rightarrow 5O_2 + 2Mn^{2+} + 8H_2O$$

Since 34 g of $H_2O_2 = 2000\,\text{ml of } 1N\,H_2O_2$ $\quad \left\{\because \text{Eq. wt. or } H_2O_2 = \dfrac{34}{2}\right\}$

$\therefore$ 34 g of $H_2O_2 = 2000\,\text{ml of } 1\,N\,KMnO_4$ $\quad [\because N_1V_1 = N_2V_2]$

or $\dfrac{X}{10}$ g of $H_2O_2 = \dfrac{2000 \times X}{100 \times 34}$ ml of 1 N $KMnO_4$

Therefore the unknown normality $= \dfrac{2000 \times X}{34 \times 100 \times X} = \dfrac{10}{17}$ or $0.588\,N$

6.(C) $N_2H_4 \rightarrow Y + 10\,e^-$, O.S. of N in N_2H_4: $2x + 4 = 0 \quad \Rightarrow \quad x = -2$

The two nitrogen atoms will balance the charge of $10\,e$

Hence oxidation state of N will increase by $+5$, i.e. from -2 to $+3$

7. Given : $2\,NH_2OH + 4\,Fe^{3+} \rightarrow N_2O + H_2O + 4Fe^{2+} + 4\,H^+$ .... **(i)**

and $MnO_4^- + 5Fe^{2+} + 8\,H^+ \rightarrow Mn^{2+} + 5\,Fe^{3+} + 4H_2O$ .... **(ii)**

$\therefore$ $10\,NH_2OH + 4\,MnO_4^- + 12\,H^+ \rightarrow 5\,H_2O + 21\,H_2O + 4\,Mn^{2+}$

[On multiplying (i) by 5 and (ii) by 4 and then adding the resulting equations]

Molecular weight of $NH_2OH = 33$

Thus 4000 ml of 1 M MnO_4^- would react with $NH_2OH = 330\,g$

$\therefore$ 12 ml of 0.02 M MnO_4 would react with

$$NH_2OH = \frac{330 \times 12 \times 0.02}{4000}\,g$$

$\therefore$ Amount of NH_2OH present in 1000 ml of diluted solution $= \dfrac{330 \times 12 \times 0.02 \times 1000}{4000 \times 50}\,g$

Since 10 ml of sample of hydroxylamine is diluted to one litre

$\therefore$ Amount of hydroxyl amine in one litre of original solution

$$= \frac{330 \times 0.02 \times 12 \times 1000}{4000 \times 50} \times \frac{1000}{10}\,g = 39.6\,g$$

8.(C) The sum of oxidation states of all atoms in compound is zero

O.S. of C in CH_2O

$x + 2 + (-2) = 0 \quad \Rightarrow \quad x = 0$

9.(B) Equivalent weight $= \dfrac{\text{molecular weight}}{\text{n - factor}}$

If n factor is 1, then equivalent weight will be equal to its molecular weight

In $MnSO_4$, the oxidation state of Mn is $+II$, In Mn_2O_3, the oxidation state of Mn is $+III$

In MnO_2, the oxidation state of Mn is $+IV$, In MnO_4^-, the oxidation state of Mn is $+VII$

In MnO_4^{2-}, the oxidation state of Mn is $+VI$

Thus, when $MnSO_4$ is converted into MnO_2, then the n factor is 2, and the equivalent weight of $MnSO_4$ will be half of its molecular weight

10. Since in acidic solution

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4\,H_2O$$

MnO_4^- gains $5\,e^-$

So, its equivalent wt $= \dfrac{\text{molecular weight}}{5}$

It means 1.61×10^{-3} moles of $MnO_4^- = 5 \times 1.61 \times 10^{-3}$ equivalents

$A^{n+} \rightarrow AO^{3-}$ (oxidation number of A in $AO_3^- = 5$)

Loss of electron $= (5 - n)$

$\therefore$ 2.68×10^{-3} moles of the solution containing A^+ ions $= (5 - n) \times 2.68 \times 10^{-3}$ equivalents

Equivalents of oxidised 'A' = Equivalents of reduced A

$$(1.61 \times 10^{-3}) \times 5 = (5 - n) \times 2.68 \times 10^{-3}$$

$$(5 - n) = \frac{1.61 \times 10^{-3} \times 5}{2.68 \times 10^{-3}} = 3 \qquad (5 - n) = 3$$

$\therefore$ $n = 5 - 3 = 2$

11.(A) Sum of oxidation of all atoms in neutral compound is zero. Let the oxidation state of iron in the complex ion

$[Fe(H_2O)_5(NO)]^{2+}SO_4^{2-}$ be x; then $x + 5 \times 0 + 1 = +2$

$\Rightarrow \quad x = +1$

12. Meq. of sodium bromate $= 85.5 \times 0.672 = 57.456$

(i) $\because$ Meq. of $NaBrO_3 = 57.456$

$\therefore \quad \frac{w}{E} \times 1000 = 57.456$

$\therefore \quad \frac{w}{151/6} \times 1000 = 57.456 \qquad \left(\because E_{NaBrO_3} = \frac{M}{6} = \frac{151}{6}\right)$

$\therefore \quad w = \frac{57.456 \times 151}{6 \times 1000} = 1.446\,g$

Also, molarity $= \frac{\text{normality}}{\text{n factor}} = \frac{0.672}{6} = 0.112\,M$

(ii) Similarly use n factor 5 in place of 6 in this problem,

Hence $w = \frac{57.456 \times 151}{5 \times 1000} = 1.735\,g$ and molarity $= \frac{0.672}{5} = 0.1344\,M$

13. The redox changes are

For $FeCl_3$: $Fe^{3+} + e^- \rightarrow Fe^{2+}$

For $N_2H_6SO_4$: $N_2^{4-} \rightarrow N_2 + 4e^-$

For $KMnO_4$: $Mn^{7+} + 5e^- \rightarrow Mn^{2+}$

Meq. of $N_2H_6SO_4$ in 10 mL solution

= Meq. of $FeCl_3$ reacting with $N_2H_6SO_4$ = Meq. of $KMnO_4$

$\therefore$ Meq. of $N_2H_6SO_4$ in 10 mL solution $= 20 \times \frac{1}{50} \times 5 = 2$

$\therefore \quad \frac{w}{130/4} \times 1000 = 2 \qquad (\therefore$ Equivalent of $N_2H_6SO_4 = \frac{130}{4})$

$\therefore \quad w = \frac{2 \times 130}{4 \times 1000} = 0.065\,g$

$\therefore$ Weight of $N_2H_6SO_4$ in 10 mL $= 0.065\,g$

$\therefore$ Thus wt. of $N_2H_6SO_4$ in $1000\,mL = 6.5\,g/L$

14.(B) $BaO_2 + H_2SO_4 \rightarrow BaSO_4 + H_2O_2$

Oxygen is the most electronegative element in the reaction and has the oxidation state of -1 (in H_2O_2) and -2 (in $BaSO_4$).

15. Let V mL of reducing agent be used for $KMnO_4$ in different medium which act as oxidant

Acid medium, $Mn^{7+} + n_1e^- \rightarrow Mn^{a+}$

$\therefore \quad n_1 = 7 - a$

Neutral medium, $Mn^{7+} + n_2e^- \rightarrow Mn^{b+}$

$\therefore \quad n_2 = 7 - b$

Alkaline medium, $Mn^{7+} + n_3e^- \rightarrow Mn^{c+}$

$\therefore \quad n_3 = 7 - c$

$\therefore$ Meq. of reducing agent

= Meq. of $KMnO_4$ in acid medium = Meq. of $KMnO_4$ in neutral

= Meq. of $KMnO_4$ in alkaline $= 1 \times n_1 \times 20 = 1 \times n_2 \times 33.3 = 1 \times n_3 \times 100$

Since n_1, n_2, n_3 are integers and $n_1 < 7$,

$\therefore \quad n_1 = 5, n_2 = 3$ and $n_3 = 1$

Therefore, different oxidation states of Mn are :

Acid media $Mn^{7+} + 5e^- \rightarrow Mn^{a+} \quad \therefore \quad a = +2$

Neutral media $Mn^{7+} + 3e^- \rightarrow Mn^{b+} \quad \therefore \quad b = +4$

Alkaline media $Mn^{7+} + 1e^- \rightarrow Mn^{c+} \quad \therefore \quad c = +6$

Now same volume of reducing agent is treated with $K_2Cr_2O_7$ and therefore,

Meq. of reducing agent = Meq. of $K_2Cr_2O_7$

$$20 \times 5 = 1 \times 6 \times V$$

$$V = \frac{100}{6} = 16.66\,\text{mL}$$

$\therefore \quad 6e^- + Cr_2^{12+} \rightarrow 2Cr^{3+} \quad \therefore \quad 1M = 6 \times 1N$

It is important to note that the conditions are valid only when Mn in each medium exist as monoatomic, i.e., not as Mn_2.