

- 1.(D)** $4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$
Atomic weight of Al = 27
Thus $4 \times 27\text{ g}$ of Al reacts with oxygen = $3 \times 32\text{ g}$
 $\therefore 27\text{ g}$ of Al reacts with oxygen = $\frac{3 \times 32}{4 \times 27} \times 27 = 24\text{ g}$
- 2.(A)** In both cases, the same volume of hydrogen is evolved for the same amount of zinc reacted.
 $\text{Zn} + \text{H}_2\text{SO}_4 \longrightarrow \text{ZnSO}_4 + \text{H}_2 \uparrow$
 $\text{Zn} + 2\text{NaOH} \longrightarrow \text{Na}_2\text{ZnO}_2 + \text{H}_2 \uparrow$
- 3.** Equivalent of KMnO_4 = Equivalents of $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$
 $5.4\text{ ml of } 0.1\text{ N } \text{KMnO}_4 = \frac{5.4 \times 0.1}{1000} = 5.4 \times 10^{-4}\text{ equivalents}$
Amount of $\text{FeSO}_4 = 5.4 \times 10^{-4} \times \text{Mol. wt. of } \text{FeSO}_4 \cdot 7\text{H}_2\text{O}$
 $= 5.4 \times 10^{-4} \times 278 = 0.150\text{ g}$
Total weight of mixture = 5.5 g
Amount of ferric sulphate = $5.5 - 0.150\text{ g} = 5.35\text{ g}$
Hence moles of ferric sulphate = $\frac{\text{Mass}}{\text{Mol. wt.}} = \frac{5.35}{562} = 9.5 \times 10^{-3}\text{ gram-mole.}$
- 4.(A)** The change involved is $\text{MnO}_4^- + \text{e}^- \longrightarrow \text{MnO}_4^{2-}$
i.e. it involves only one electron
Eq. wt. = $\frac{\text{Mol. wt.}}{\text{number of } \text{e}^- \text{ involved}} = \frac{M}{1} = M$ [$\because \text{Mol. wt.} = M$]
- 5.** The complete oxidation under acidic conditions can be represented as follows :
 $5\text{H}_2\text{O}_2 + 2\text{MnO}_4^- + 6\text{H}^+ \rightarrow 5\text{O}_2 + 2\text{Mn}^{2+} + 8\text{H}_2\text{O}$
Since $34\text{ g of } \text{H}_2\text{O}_2 = 2000\text{ ml of } 1\text{N } \text{H}_2\text{O}_2$ $\left\{ \because \text{Eq. wt. of } \text{H}_2\text{O}_2 = \frac{34}{2} \right\}$
 $\therefore 34\text{ g of } \text{H}_2\text{O}_2 = 2000\text{ ml of } 1\text{N } \text{KMnO}_4$ [$\because \text{N}_1\text{V}_1 = \text{N}_2\text{V}_2$]
or $\frac{X}{10}\text{ g of } \text{H}_2\text{O}_2 = \frac{2000 \times X}{100 \times 34}\text{ ml of } 1\text{N } \text{KMnO}_4$
Therefore the unknown normality = $\frac{2000 \times X}{34 \times 100 \times X} = \frac{10}{17}$ or 0.588 N
- 6.(C)** $\text{N}_2\text{H}_4 \rightarrow \text{Y} + 10\text{e}^-$, O.S. of N in N_2H_4 : $2x + 4 = 0 \Rightarrow x = -2$
The two nitrogen atoms will balance the charge of 10e^-
Hence oxidation state of N will increase by +5, i.e. from -2 to $+3$
- 7.** Given : $2\text{NH}_2\text{OH} + 4\text{Fe}^{3+} \rightarrow \text{N}_2\text{O} + \text{H}_2\text{O} + 4\text{Fe}^{2+} + 4\text{H}^+$ (i)
and $\text{MnO}_4^- + 5\text{Fe}^{2+} + 8\text{H}^+ \rightarrow \text{Mn}^{2+} + 5\text{Fe}^{3+} + 4\text{H}_2\text{O}$ (ii)
 $\therefore 10\text{NH}_2\text{OH} + 4\text{MnO}_4^- + 12\text{H}^+ \rightarrow 5\text{H}_2\text{O} + 21\text{H}_2\text{O} + 4\text{Mn}^{2+}$
[On multiplying (i) by 5 and (ii) by 4 and then adding the resulting equations]
Molecular weight of $\text{NH}_2\text{OH} = 33$

Thus 4000 ml of 1 M MnO_4^- would react with $\text{NH}_2\text{OH} = 330 \text{ g}$

\therefore 12 ml of 0.02 M MnO_4^- would react with

$$\text{NH}_2\text{OH} = \frac{330 \times 12 \times 0.02}{4000} \text{ g}$$

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

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

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

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

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 = $\frac{\text{molecular weight}}{n\text{-factor}}$

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

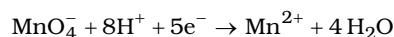
In MnSO_4 , the oxidation state of Mn is +II, In Mn_2O_3 , the oxidation state of Mn is +III

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

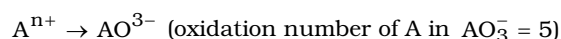
10. Since in acidic solution



MnO_4^- gains 5e^-

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

It means 1.61×10^{-3} moles of $\text{MnO}_4^- = 5 \times 1.61 \times 10^{-3}$ equivalents



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

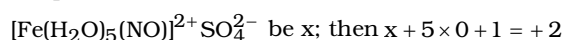
Equivalents of oxidised 'A' = Equivalents of reduced A

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

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

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

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



$$\Rightarrow x = +1$$

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

(i) \therefore Meq. of $\text{NaBrO}_3 = 57.456$

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

$$\therefore \frac{w}{151/6} \times 1000 = 57.456 \quad \left(\because E_{\text{NaBrO}_3} = \frac{M}{6} = \frac{151}{6} \right)$$

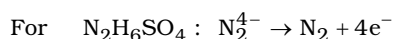
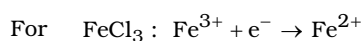
$$\therefore w = \frac{57.456 \times 151}{6 \times 1000} = 1.446 \text{ g}$$

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

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

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

13. The redox changes are



Meq. of $\text{N}_2\text{H}_6\text{SO}_4$ in 10 mL solution

$$= \text{Meq. of } \text{FeCl}_3 \text{ reacting with } \text{N}_2\text{H}_6\text{SO}_4 = \text{Meq. of } \text{KMnO}_4$$

$$\therefore \text{Meq. of } \text{N}_2\text{H}_6\text{SO}_4 \text{ in 10 mL solution} = 20 \times \frac{1}{50} \times 5 = 2$$

$$\therefore \frac{w}{130/4} \times 1000 = 2 \quad \left(\because \text{Equivalent of } \text{N}_2\text{H}_6\text{SO}_4 = \frac{130}{4} \right)$$

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

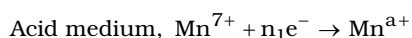
$$\therefore \text{Weight of } \text{N}_2\text{H}_6\text{SO}_4 \text{ in 10 mL} = 0.065 \text{ g}$$

$$\therefore \text{Thus wt. of } \text{N}_2\text{H}_6\text{SO}_4 \text{ in 1000 mL} = 6.5 \text{ g/L}$$

14.(B) $\text{BaO}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{BaSO}_4 + \text{H}_2\text{O}_2$

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



$$\therefore n_1 = 7 - a$$



$$\therefore n_2 = 7 - b$$



$$\therefore n_3 = 7 - c$$

\therefore Meq. of reducing agent

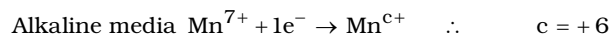
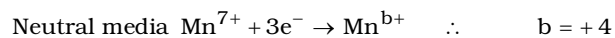
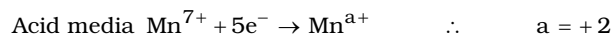
$$= \text{Meq. of } \text{KMnO}_4 \text{ in acid medium} = \text{Meq. of } \text{KMnO}_4 \text{ in neutral}$$

$$= \text{Meq. of } \text{KMnO}_4 \text{ in alkaline} = 1 \times n_1 \times 20 = 1 \times n_2 \times 33.3 = 1 \times n_3 \times 100$$

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

$$\therefore n_1 = 5, n_2 = 3 \text{ and } n_3 = 1$$

Therefore, different oxidation states of Mn are :



Now same volume of reducing agent is treated with $\text{K}_2\text{Cr}_2\text{O}_7$ and therefore,

Meq. of reducing agent = Meq. of $\text{K}_2\text{Cr}_2\text{O}_7$

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

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



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 .