## CONCEPT OF GRAM EQUIVALENTS

## Section - 3

In the method of gram-equivalents (milli-equivalents), there is no need to balance a redox reaction. For example, let us consider a *typical redox reaction* (consisting of two reactants and two main products):

$$\underset{\text{Reducing agent}}{X} + \underset{\text{Oxidising agent}}{Y} \longrightarrow A + B$$

According to concept of gram equivalent (or meq) ;

$$\mathbf{1\ gm.eq.\ of\ X \equiv 1\ gm.\ eq.\ of\ Y \equiv 1\ gm\ eq.\ of\ A \equiv 1\ gm\ eq.\ of\ B}$$

or meq. of an oxidising agent = meq. of a reducing agent

**Note :** gm.eq. $= \dfrac{g}{E}$ and meq $= \dfrac{g}{E} \times 1000$

Now to understand the method of gram-equivalents more clearly, let us understand the following illustrations very carefully.

### *Illustrating the concept :*

*What volume of a deci-normal solution of $K_2Cr_2O_7$ is required to oxidise 1.0 gm of Tin (Sn) in presence of HCl to its highest oxidation state ?*

The involved reaction is :

$$K_2Cr_2O_7 + Sn + H^+ \longrightarrow Sn^{4+} + Cr^{3+}$$ (The highest oxidation state of Sn is +4, so $SnCl_4$ will be formed here)

Here, Sn is a reducing agent (RA) and $K_2Cr_2O_7$ is an oxidant (OA).

So applying the concept of gram-equivalent, we have :

$$\text{gm.eq of OA } (K_2Cr_2O_7) = \text{ gm.eq of RA (Sn)}$$

Using the definition : $N\ V_L = \dfrac{g}{E}$, we get :

$$\frac{1}{10} \times V_L = \frac{1}{E_{Sn}} \quad ; \quad E_{Sn} = \frac{119}{x}$$ where $x$ = electron transfer / mol (*also known as valence or n-factor*)

$$Sn \longrightarrow Sn^{4+} + 4e^- \quad \Rightarrow \quad x = 4$$

$$\Rightarrow \quad \frac{V}{10} = \frac{1}{119/4} \quad \Rightarrow \quad V = \frac{40}{119} = 0.336\,L = 336\,mL$$

**Note :** From definition of gram-equivalents $\left(\text{gmeq.} = \dfrac{g}{E} = N\ V_L\right)$, it is very clear that, we must be very confident about the concept of equivalent weight (E), for an oxidising or reducing agent. So first let us learn intricacies of the concept of equivalent weight for an oxidant(s) and reductant(s) in detail.

## Equivalent Weight (E) of Oxidant and Reductant :

The equivalent weight of an oxidising or reducing agent is defined as :

Equivalent weight = $E = \dfrac{M_0}{x}$

where $M_0$ = molecular mass ; $x$ = number of electrons transferred (loss or gain) by one mole of oxidising or reducing agent as given by their balanced ionic half reaction.

### *Illustrating the concept :*

- **$FeSO_4$** (Ferrous Sulphate)

As reducing agent :

$Fe^{2+} \longrightarrow Fe^{3+} + 1e^-$

1 mole of e's are transferred / mole of $FeSO_4$, so $x = 1$

$\Rightarrow \quad E_{FeSO_4} = \dfrac{M_O}{x} = \dfrac{152}{1} = 152$

As an oxidising agent :

$Fe^{2+} + 2e^- \longrightarrow Fe$

$\Rightarrow \quad x = 2$

$\Rightarrow \quad E_{FeSO_4} = \dfrac{M_O}{x} = \dfrac{152}{2}$

***Note :*** $FeSO_4$ (ferrous sulphate) acts both as an oxidant and as a reductant

- **$H_2C_2O_4$** (Oxalic acid) or **$C_2O_4^{2-}$** (Oxalate ion)

As reducing agent

$C_2O_4^{2-} \longrightarrow 2\,CO_2 + 2e^-$ or $H_2C_2O_4 \longrightarrow 2\,CO_2 + 2\,H^+ + 2\,e^-$

$x = 2$ or $E_{H_2C_2O_4} = \dfrac{90}{2}$

***Note :*** $E_{Na_2C_2O_4} = \dfrac{134}{2}$ ; $E_{CaC_2O_4} = \dfrac{128}{2}$ as n–factor for $C_2O_4^{2-}$ is always 2

Sodium oxalate Calcium oxalate

- **$SO_2$** (Sulphur Dioxide)

As reducing agent

$SO_2 + 2H_2O \longrightarrow SO_4^{2-} + 4H^+ + 2e^-$

$x = 2 \quad \Rightarrow \quad E_{SO_2} = \dfrac{64}{2}$

- **HI** (Hydrogen Iodide)

As reducing agent only

$2HI \longrightarrow I_2 + 2H^+ + 2e^- \quad \Rightarrow$ 2 mole of e's / 2 mole of HI

$x = \dfrac{2}{2} = 1$ (for 1 mole of HI) $\quad \Rightarrow \quad E_{HI} = \dfrac{128}{1}$

- **$Fe_2(SO_4)_3$** (Ferric sulphate)

  As oxidising agent only

  $Fe^{3+} + 1e^- \longrightarrow Fe^{2+}$

  Since 1 mole of $Fe_2(SO_4)_3 \equiv$ 2 moles of $Fe^{3+}$

  $\Rightarrow \quad x = 2$ for 1 mole of $Fe_2(SO_4)_3$

  $\Rightarrow \quad E_{Fe_2(SO_4)_3} = \frac{400}{2}$

- **$K_2Cr_2O_7$** (Potassium Dichromate)

  As oxidising agent only

  $Cr_2O_7^{2-} + 6e^- + 14\,H^+ \longrightarrow 2\,Cr^{3+} + 7H_2O$

  $x = 6 \quad \Rightarrow \quad E_{K_2Cr_2O_7} = \frac{294}{6}$

- **$KMnO_4$** (Potassium Permanganate)

  As oxidising agent in acidic medium

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

  $x = 5 \quad \Rightarrow \quad E_{KMnO_4/H^+} = \frac{158}{5}$

  As oxidising agent in strong alkaline medium

  $MnO_4^- + 1e^- \longrightarrow MnO_4^{2-}$

  $x = 1 \quad \Rightarrow \quad E_{KMnO_4} = \frac{158}{1}$

  As oxidising agent in dilute alkaline medium

  $MnO_4^- + 2H_2O + 3e^- \longrightarrow MnO_2 + 4\,OH^-$

  $x = 3 \quad \Rightarrow \quad E_{KMnO_4/OH^-} = \frac{158}{3}$

- **$MnO_2$** (Managanese dioxide)

  As oxidising agent in acidic medium

  $MnO_2 + 4H^+ + 2e^- \longrightarrow Mn^{2+} + 2\,H_2O$

  $x = 2 \quad \Rightarrow \quad E_{MnO_2} = \frac{87}{2}$

- **$Na_2S_2O_3$** (Sodium Thiosulphate)

  As reducing agent in acidic medium

  $2\,S_2O_3^{2-} \longrightarrow S_4O_6^{2-} + 2e^-$

  $x = 2/2 = 1$ for 1 mole of $Na_2S_2O_3$

  $\Rightarrow \quad E_{Na_2S_2O_3/H^+} = \frac{158}{1}$

  As reducing agent in alkaline medium

  $S_2O_3^{2-} + 10OH^- \longrightarrow 2SO_4^{2-} + 5H_2O + 8e^-$

  $x = 8$

  $\Rightarrow \quad E_{Na_2S_2O_3/OH^-} = \frac{158}{8}$

***Note :*** From the above case, you should observe carefully that the equivalent weight depends upon medium.

- **$HNO_3$** (Nitric Acid)

  As oxidising agent (conc. $HNO_3$)

  $NO_3^- + 2H^+ + 1e^- \longrightarrow NO_2 + H_2O$

  $x = 1$

  $\Rightarrow \quad E_{HNO_3\,(conc.)} = \frac{63}{1} = 63$

  As oxidising agent (dil. $HNO_3$)

  $NO_3^- + 4\,H^+ + 3e^- \longrightarrow NO + 2\,H_2O$

  $x = 3$

  $\Rightarrow \quad E_{HNO_3\,(dil)} = \frac{63}{3} = 21$

***Note :*** Equivalent weight depends on dilution for $HNO_3$.

*Try to Calculate 'E'*

(a) $HNO_3$ (very dilute) $\longrightarrow NH_4^+$ (b) $HNO_3$ (cold dilute) $\longrightarrow NH_2OH$

**Note :** Some substances act both as OA (oxidants) and reductants (RA) in different reactions. Like $FeSO_4$, $H_2O_2$, $HNO_2$ etc

**$H_2O_2$** (Hydrogen Peroxide) *as OA* : $H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$

*as RA* : $H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-$

**$FeSO_4$** (Ferrous Sulphate) *as OA* : $Fe^{2+} + 2e^- \longrightarrow Fe$

*as RA* : $Fe^{2+} \longrightarrow Fe^{3+} + 1e^-$

**$HNO_2$** (Nitrous Acid) *as OA* : $HNO_2 + H^+ + 1e^- \longrightarrow NO + H_2O$

*as RA* : $HNO_2 + H_2O \longrightarrow NO_3^- + 3H^+ + 2e^-$

Some other type of oxidants and reductants are also given below.

- $(NH_4)_2Cr_2O_7 \longrightarrow N_2 + Cr_2O_3$

  (Ammonium Dichromate)

  *RA* : $NH_4^+ \longrightarrow N_2$ *OA* : $Cr_2O_7^{2-} \longrightarrow Cr_2O_3$

- $Fe\,C_2O_4 \rightarrow Fe^{3+} + CO_2$

  (Ferrous Oxalate)

  *RA* : $Fe^{2+} \longrightarrow Fe^{3+} + 1e^-$ ; $C_2O_4^{2-} \longrightarrow 2CO_2 + 2e^- \Rightarrow x = 3$

**Note :** In $FeC_2O_4$ both cations and anions are oxidised.

- $I_2 \longrightarrow I^- + IO_3^-$ (disproportionation reaction)

  *OA* : $I_2 \longrightarrow 2I^-$ *RA* : $I_2 \longrightarrow IO_3^-$

  (same compound as OA/RA)

**Note :** Some compounds act as Acids and RA or OA in separate experiments.

$H_2C_4O_4$ …

$$H_2C_2O_4 \begin{cases} \longrightarrow \text{as acid } (x=2) \\ \longrightarrow \text{as RA } (x=2) \end{cases} ; \quad HNO_3 \begin{cases} \longrightarrow \text{as acid } (x=1) \\ \longrightarrow \text{as OA } (x=1,3,6,8) \end{cases} ; \quad NaHC_2O_4 \begin{cases} \longrightarrow \text{as acid } (x=1) \\ \longrightarrow \text{as RA } (x=2) \end{cases}$$

**Illustration - 16** *Complete and balance the following.*

(a) $H_2O_2 + I_2 \longrightarrow I^- + ??$ (b) $H_2O_2 + Sn^{2+} \longrightarrow Sn^{4+} + ??$

(c) $MnO_4^{2-} + H^+ \longrightarrow Mn^{2+} + ??$ (d) $NO_2 + H_2O \longrightarrow NO + ??$

**SOLUTION :**

(a) $I_2$ is reduced to $I^-$, so $H_2O_2$ must be oxidised, i.e., it acts as a reducing agent. Hence final product will be $O_2$ gas.

The complete chemical equation is :

$$H_2O_2 + I_2 \longrightarrow I^- + O_2$$

The balanced equation will be :

$$H_2O_2 + I_2 \longrightarrow 2I^- + O_2 + 2H^+$$

**(b)** $Sn^{2+} \longrightarrow Sn^{4+}$, i.e., $Sn^{2+}$ is oxidised. So $H_2O_2$ acts as an oxidising agent, which means it must be reduced to $H_2O$.

$$H_2O_2 + Sn^{2+} \longrightarrow Sn^{4+} + H_2O$$

The balanced equation will be :

$$H_2O_2 + Sn^{2+} + 2H^+ \longrightarrow Sn^{4+} + 2H_2O$$

### Important Note :

$H_2O_2$ is chemical species acting both as a reducing agent and as an oxidising agent.

$$H_2O_2 \begin{array}{l} \xrightarrow{\text{as oxidant}} H_2O \\ \xrightarrow{\text{as reductant}} O_2 \end{array}$$

**(c)** $MnO_4^{2-}$ is reduced to $Mn^{2+}$, i.e it is acting as an oxidising agent. Now $H^+$ is already in the maximum oxidation state, so $MnO_4^{2-}$ must also be reduced to $Mn^{7+}$ state i.e. $MnO_4^-$ will be formed. Hence it is a case of disproportionation reaction.

$$MnO_4^{2-} + H^+ \longrightarrow Mn^{2+} + MnO_4^-$$

The balanced equation will be :

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

**(d)** $NO_2$ (+4 state) disproportionate (like in last case) to NO (+2 state) and $NO_3^-$ (+5 state).

$$NO_2 + H_2O \rightarrow NO + NO_3^-$$

The balanced equation will be :

$$3\,NO_2 + H_2O \longrightarrow NO + 2NO_3^- + 2\,H^+$$

### Equivalent weight of a compound undergoing disproportionation :

Let a compound '*A*' contains an ion '*X*' undergoing disproportionation.

Then, $$E_A = \text{Equivalent weight of A} = \frac{\text{Effective molecular weight of A}}{\text{Number of } e^- \text{ transfer}}$$

where $$\text{Effective molecular weight of A} = \begin{pmatrix}\text{Total weight of A in} \\ \text{Oxidation half reaction}\end{pmatrix} + \begin{pmatrix}\text{Total weight of A in} \\ \text{Reduction half reaction}\end{pmatrix}$$

## 1. When number of e⁻ transfer are same in oxidation and reduction half reactions

### *Illustrating the concept :*

*$H_2O_2$ undergoes disproportionation in acidic medium given by the following reaction :*

$$H_2O_2 \longrightarrow H_2O + O_2$$

**Oxidation half reaction :** $H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-$ $[2e^-$ transfer$]$ . . . (i)

**Reduction half reaction :** $H_2O_2 + 2H^+ + 2e^- \longrightarrow 2H_2O$ $[2e^-$ transfer$]$ . . . (ii)

Thus, Effective molecular mass of $H_2O_2 = \underset{\text{from (i)}}{M_{H_2O_2}} + \underset{\text{from (ii)}}{M_{H_2O_2}} = 2M_{H_2O_2}$

$$\Rightarrow \quad E_{H_2O_2} = \frac{2M_{H_2O_2}}{2} = 34$$

## 2. When number of $e^-$ transfer are different in oxidation and reduction half reactions

***Illustrating the concept :***

*$Br_2$ undergoes disproportionation in basic medium given by the following reaction:*

$$Br_2 + OH^- \longrightarrow Br^- + BrO_3^- + H_2O$$

**Oxidation half reaction :** $Br_2 + 12OH^- \longrightarrow 2BrO_3^- + 6H_2O + 10e^-$ $\left[10e^- \text{ transfer}\right]$ ... (i)

**Reduction half reaction :** $Br_2 + 2e^- \longrightarrow 2Br^-$ $[2e^- \text{ transfer}]$ ... (ii)

Now, Effective molecular mass of $Br_2 = \underset{\text{from (i)}}{M_{Br_2}} + \underset{\text{from (ii)}}{5M_{Br_2}} = 6M_{Br_2}$

***Note :*** $5M_{Br_2}$ is taken because we have to equalize the number of electrons in both oxidation and reduction.

Thus, $E_{Br_2} = \dfrac{6M_{Br_2}}{10} = 0.6\ M_{Br_2} = 96$

This means n-factor of $Br_2$ is $\dfrac{10}{6}$

***Note :*** However, we should use "Mole concept" while solving problems based on such reactions.

**Illustration - 17** *1.0 gram of a sample of $H_2O_2$ solution containing y% $H_2O_2$ by weight requires y mL of $KMnO_4$ solution for complete titration under acidic condition. Find the molarity of $KMnO_4$ solution.*

**SOLUTION :**

This question involves a redox reaction between $H_2O_2$ (RA) and $KMnO_4$ (OA). Let us apply the concept of gram (milli) equivalents to solve this problem.

i.e., meq of $H_2O_2$ = meq of $KMnO_4$

Let us find n-factor (valence factor) of $H_2O_2$ and $KMnO_4$ as follows.

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

and $H_2O_2 \longrightarrow O_2 + 2H^+ + 2e^-$

$x = 5$ for $KMnO_4$ and $x = 2$ for $H_2O_2$

$$\text{meq of } H_2O_2 = \frac{1 \times \left(\frac{y}{100}\right)}{E_{H_2O_2}} \times 1000 \quad \left[E_{H_2O_2} = \frac{34}{2}\right]$$

meq of $KMnO_4 = N(y)$

$$\Rightarrow \frac{1 \times \left(\frac{y}{100}\right)}{17} \times 1000 = N(y) \quad \Rightarrow \quad N = \frac{10}{17}$$

And Molarity $= \dfrac{N}{x}$; $x = 5$ for $KMnO_4/H^+$

$$\Rightarrow \quad M = \frac{10/17}{5} = \frac{2}{17} = 0.117$$

**Illustration - 18** *A solution of $KMnO_4/H^+$ containing 3 gm/L is used to titrate $H_2O_2$ solution containing 2 gm/L. What volume of $KMnO_4$ will be required to react with 20 mL of $H_2O_2$ solution ? Also find the strength of $H_2O_2$ in terms of available oxygen.*

**SOLUTION :**

This redox reaction is a typical one.

$$N_{KMnO_4} = \frac{\text{strength}}{E} = \frac{3}{158/5} = \frac{15}{158}$$

$$N_{H_2O_2} = \frac{\text{strength}}{E} = \frac{2}{17}$$

Now using equation of titration

$$\text{meq. of OA} = \text{meq of RA}$$

$$\Rightarrow \quad N_{OA}V_{OA} = N_{RA}V_{RA} \quad (\text{meq} = NV_{cc})$$

$$\frac{15}{158} \times V = \frac{2}{17} \times 20$$

$$\Rightarrow \quad V = 24.8\text{ mL}$$

**What is strength in terms of available oxygen ?**

It is also called as volume strength of $H_2O_2$.

Volume strength is the volume of $O_2$ given out at S.T.P. by 1 volume solution of $H_2O_2$.

$$2H_2O_2 \longrightarrow 2H_2O + O_2$$

$$2\text{ mol } H_2O_2 \equiv 1\text{ mol of } O_2 \equiv 22.4\text{ L of } O_2$$

$$1\text{ mol } H_2O_2 \equiv 1/2\text{ mol of } O_2 \equiv 11.2\text{ L of } O_2 \text{ per } 1.0\text{ L}$$

$$\Rightarrow \quad \mathbf{1\text{ M } H_2O_2 \equiv 11.2\text{ vol of } O_2 \text{ at STP}}$$

$$\text{or} \quad \mathbf{1N\ H_2O_2 \equiv 5.6\text{ vol of } O_2}$$

In the present question :

$$\frac{2}{17}\text{N } H_2O_2 \equiv \frac{2}{17} \times 5.6 = 0.66\text{ L of } O_2$$

It is written as : 0.66 vol $H_2O_2$

**Try yourself :** Express the volume strength of $H_2O_2$ sample if 100 cc of $H_2O_2$ solution were diluted with 100 cc of water. 10 cc of this diluted solution required 20 cc of M/50 $KMnO_4$ in presence of $H_2SO_4$ for complete titration. [Ans : 2.24 vol $H_2O_2$]

**Illustration - 19** *1.25 gm of a sample of bleaching powder ($CaOCl_2$) is dissolved in 100 mL of water. 25 mL of it are treated with KI solution. The iodine so liberated required 12.5 mL of M/25 hypo solution ($Na_2S_2O_3$ /$H^+$) for complete titration. Find the percentage of available chlorine from the sample of bleaching powder.*

**SOLUTION :**

$$CaOCl_2 \longrightarrow Cl_2$$

$$Cl_2 + 2\,KI \longrightarrow I_2 + 2\,KCl$$

$$I_2 + 2\,Na_2S_2O_3 \longrightarrow 2\,NaI + Na_2S_4O_6$$

$$(2\,S_2O_3^{2-} \longrightarrow S_4O_6^{2-} + 2e^- ;\ x = 1)$$

In 25 ml sample :

meq of $Cl_2$ in bleaching powder

= meq of $I_2$ liberated = meq of hypo solution

$$\text{meq of Hypo} = NV_{cc} = \left(1 \times \frac{1}{25}\right) \times 12.5 = 0.5$$

$$\Rightarrow \quad \text{meq of } Cl_2/\,25\text{ml.} = 0.5$$

$$\Rightarrow \quad \text{meq of } Cl_2 \text{ in } 100\text{ mL} = (0.5 \times 4)$$

$$\Rightarrow \quad \frac{g}{E} \times 1000 = 2$$

$$(Cl_2 + 2e^- \longrightarrow 2Cl^- ;\ E = 71/2)$$

$$g = \frac{2}{1000} \times \frac{71}{2} = 0.071\text{ gms}$$

% available chlorine

$$= \frac{g_{Cl_2}}{g_{CaOCl_2}} \times 100 = \frac{0.071}{1.25} \times 100 = 5.68\,\%$$

**Try yourself :** 5.0 gm of $CaOCl_2$ is dissolved in water to make 500 mL solution. 20 mL of it is acidified with acetic acid and treated with KI solution. The iodine liberated required 20 mL of M/10 $Na_2S_2O_3$ solution. Find the percentage of available chlorine. [Ans: 35.5%]

**Illustration - 20** *50 ml of an aqueous solution of $H_2O_2$ was titrated with an excess of KI solution in dilute $H_2SO_4$, the liberated iodine required 20 ml of 0.1 N $Na_2S_2O_3$ solution for complete reaction. Calculate volume strength of $H_2O_2$ solution.*

**SOLUTION :**

Both the chemical equations involved in the problems are typical redox reactions.

$$H_2O_2 + KI + H_2SO_4 \longrightarrow I_2 + H_2O + K_2SO_4 \quad \ldots\text{(i)}$$

and $$I_2 + 2\,Na_2S_2O_3 \longrightarrow 2\,NaI + Na_2S_4O_6 \quad \ldots\text{(ii)}$$

Now : meq. of $I_2$ = meq. of $Na_2S_2O_3$ (in eq (ii))

$\Rightarrow$ meq. of $I_2 = 20 \times 0.1 = 2$

and meq. of $H_2O_2$ = meq. of $I_2$ (in eq (i))

$$N_{H_2O_2} \times V_{H_2O_2} = 2$$

$$\Rightarrow \quad N_{H_2O_2} = \frac{2}{50} = \frac{1}{25}$$

Now, 1 N $H_2O_2$ solution = 5.6 vol. of $O_2$

$$\Rightarrow \quad \frac{1}{25} N\ H_2O_2 \text{ solution} = 5.6 \times \frac{1}{25}$$

$$= 0.224 \text{ vol of } O_2$$

$\Rightarrow$ Volume Strength of $H_2O_2$ solution = 0.224 vol.

***Note :*** In above two problems, $Na_2S_2O_3$ is used as reducing agent (very common and important reductant). You must remember its action in neutral and basic both, though neutral medium (hypo solution) is more important.

$$Na_2S_2O_3 \xrightarrow{\text{neutral}} S_4O_6^{2-} \quad ; \quad Na_2S_2O_3 \xrightarrow{OH^-} SO_4^{2-}$$

## IN-CHAPTER EXERCISE - C

**1.** *An excess of acidified solution of potassium iodide was added to 20 millilitres of hydrogen peroxide solution and the iodine so liberated required 15 ml of 0.4 N sodium thiosulphate solution for iodometric titration. Calculate the volume strength of $H_2O_2$ solution.*

**2.** *0.4 g sample of a copper ore was converted into copper sulphate solution. The resulting solution was acidified with dilute acetic acid and excess KI added. The liberated iodine required 0.248 g $Na_2S_2O_3 . 5H_2O$ for complete reaction. What is the percentage of copper in the ore ?*

**3.** *0.367 g of a mixture containing sodium oxalate and potassium oxalate requires 125 ml of 0.04 N $KMnO_4$ for titration at the end -point. What is the percentage of sodium oxalate in the mixture ?*

**4.** *0.1 M $KMnO_4$ is used for the titrating hypo solution. What volume of the solution in ml will be required to react with 0.158 g of $Na_2S_2O_3$ ?*

**5.** *5.5 g of a mixture of $FeSO_4.7H_2O$ and $Fe_2(SO_4)_3 . 9H_2O$ requires 5.4 ml of 0.1 N $KMnO_4$ solution for complete oxidation. Calculate the gram mole of hydrated ferric sulphate in the mixture.*

**6.** *Find equivalent weight of $Fe(HC_2O_4)_2$ as : (a) an Acid (b) a reducing agent.*

**7.** *3.55 gm of bleaching powder are ground to a fine paste with water and made up to 100 mL. 25 mL of this after addition of KI and dilute acetic acid required 20 mL of 0.125 N thiosulpahte solution. Calculate the percentage of available chlorine in the sample of bleaching powder.*

**8.** *Calculate equivalent weight of underlined species in the following unbalanced reactions:*

**(i)** $\underline{MnO_4^{2-}} \longrightarrow MnO_2 + MnO_4^-$ (*acidic medium*)

**(ii)** $\underline{Cl_2} + OH^- \longrightarrow Cl^- + ClO_2^-$ (*basic medium*)

**(iii)** $\underline{HCuCl_2} \longrightarrow Cu + Cu^{2+} + 4Cl^- + 2H^+$ (*acidic medium*)

**9.** ***Choose the correct option for each of the following questions. Only one option is correct.***

**(i)** 10 *ml of a solution of* $H_2O_2$ *labelled* '10 *volume*' *just decolorizes* 100 *ml of* $KMnO_4$ *solution acidified with dilute* $H_2SO_4$. *Calculate the amount of* $KMnO_4$ *in the given solution.*

**(A)** 0.1563 *gm* **(B)** 0.563 *gm* **(C)** 5.63 *gm* **(D)** 0.256 *gm*

**(ii)** *The equivalent mass of* $Na_2S_2O_3$ *in the reaction with* $I_2$ *is equal to* :

**(A)** $\frac{molar\ mass}{2}$ **(B)** $\frac{molar\ mass}{3}$ **(C)** $\frac{molar\ mass}{4}$ **(D)** *molar mass*

**(iii)** In basic medium $CrO_4^{2-}$ reacts with $S_2O_3^{2-}$ resulting in formation of $Cr(OH)_4^-$ and $SO_4^{2-}$. *How many ml of* 0.1 *M* $Na_2CrO_4$ *are required to react with* 40 *ml of* 0.25 *M* $Na_2S_2O_3$?

**(A)** 240.2 *mL* **(B)** 24.02 *mL* **(C)** 266.67 *mL* **(D)** 26.67 *mL*

**(iv)** *KI reacts with* $H_2SO_4$ *producing* $I_2$ *and* $H_2S$. *The volume of 0.2 M* $H_2SO_4$ *required to produce* 3.4 *gm of* $H_2S$ *is* :

**(A)** 4*L* **(B)** 2.5*L* **(C)** 3.8*L* **(D)** 5*L*

**(v)** *A* 0.46 *gm sample of* $As_2O_3$ *required* 25.0 *ml of* $KMnO_4$ *solution for its titration. The molarity of* $KMnO_4$ *solution is* :

**(A)** 0.016 **(B)** 0.074 **(C)** 0.032 **(D)** 0.128

**(vi)** *Number of moles of* $K_2Cr_2O_7$ *reduced by* 1 *mole of* $Sn^{+2}$ *is :*

**(A)** 1/3 **(B)** 1/6 **(C)** 2/3 **(D)** 1

**(vii)** 10 *ml of* $H_2O_2$ *solution (volume strength = x) required* 10 *ml of N/*0.56 $MnO_4^-$ *solution in acidic medium. Hence x is* :

**(A)** 0.56 **(B)** 5.6 **(C)** 0.1 **(D)** 10

**(viii)** *Equivalent weight of* $H_3PO_2$ *(mol. wt. = M) when it disproportionates into* $PH_3$ *and* $H_3PO_3$ *is* :

**(A)** *M* **(B)** *M*/2 **(C)** *M*/4 **(D)** 3*M*/4