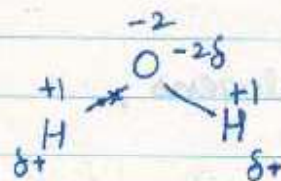
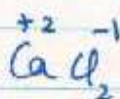
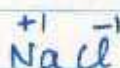
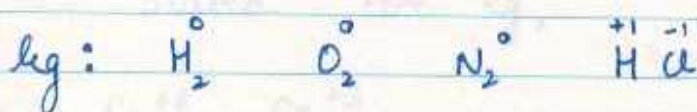


## Redox & Equivalent Concept

• Oxidation state : actual charge on an atom or ion if it exists in monoatomic state or it is hypothetical charge assigned to an atom if it is in combined state.

It arises due to EN difference.



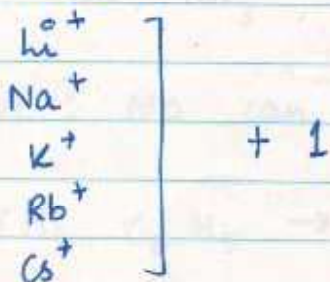
### \* Rules to assign O.S. :->

1. O.S. of an element in its free state is taken as zero.



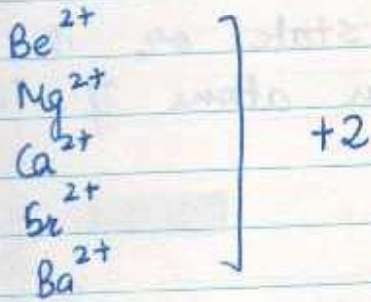
2. In combined state,

a) Alkali metals are given O.S. +1

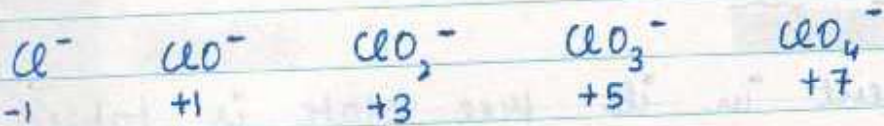
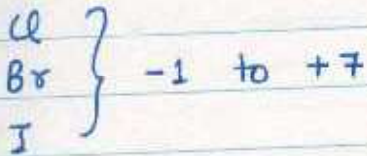
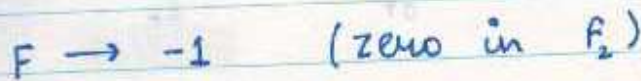




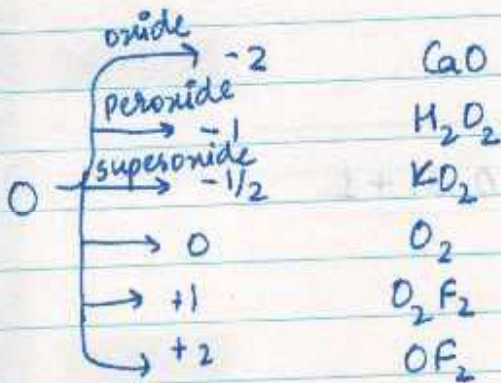
b) Alkaline earth metals are given O.S. of +2



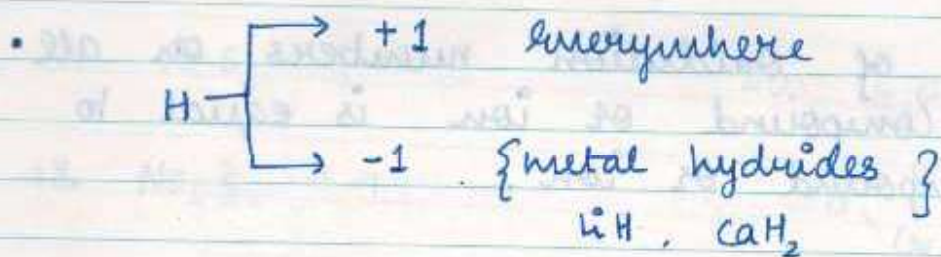
c) For halogens,



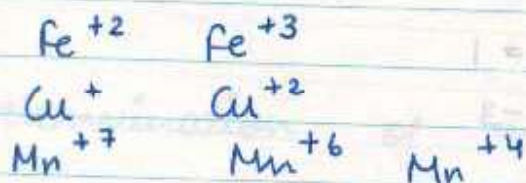
3. For some important elements







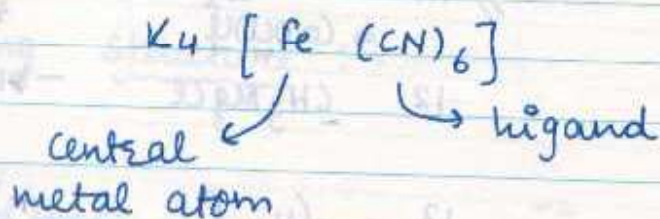
4. d-block elements can have variable oxidation states. For eg,



5. O.S. can be +ve, -ve, zero or fractional

6. Maximum possible O.S. is +8  $\rightarrow$  Os & Ru

7. In case of co-ordination compounds, neutral ligands are given O.S. of zero



CO, NO,  $\text{NH}_3$ ,  $\text{H}_2\text{O} \rightarrow$  neutral  $\Rightarrow$  OS = 0

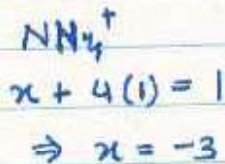
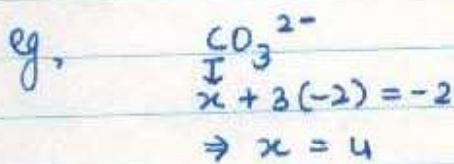
except : NO can be neutral or +1

$\text{CH}_3, \text{C}_6\text{H}_5 \rightarrow +1$

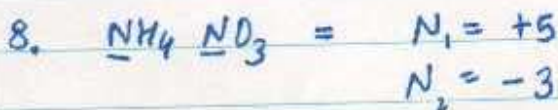
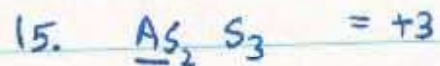
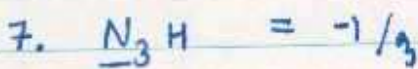
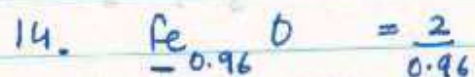
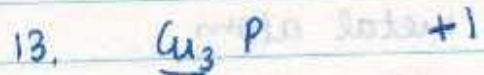
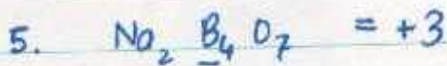
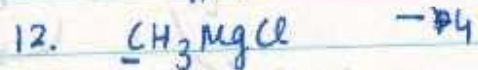
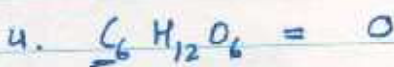
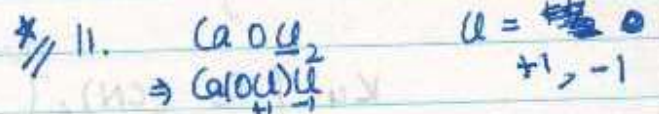
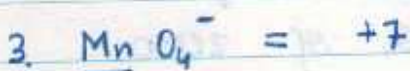
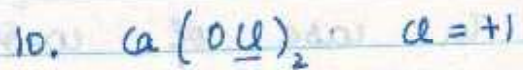
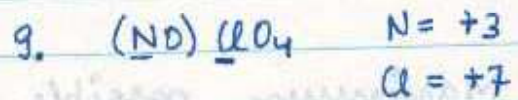
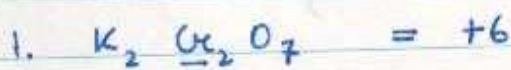


8. Algebraic sum of oxidation numbers on all atoms in a compound or ion is equal to charge on compound or ion.

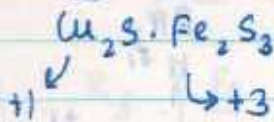
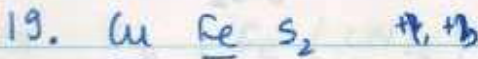
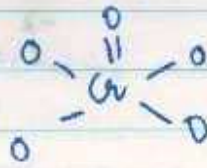
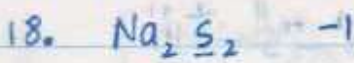
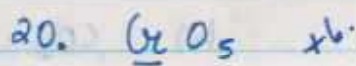
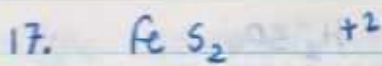
$\sum$  o.s. of each element in cpd / ion = charge on ion



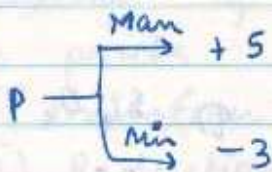
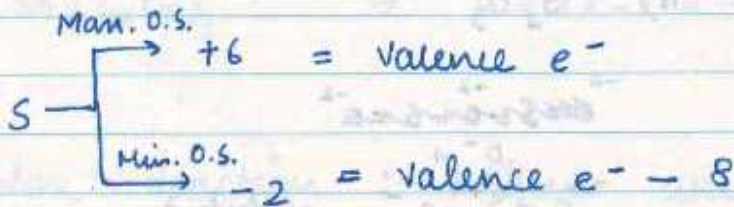
9. Find O.S. on underlined atoms



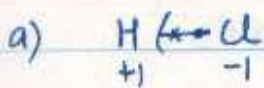




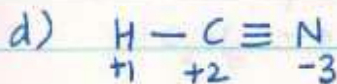
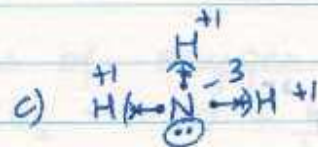
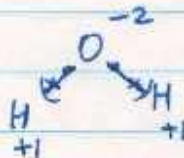
• Determination of Max. & Min. O.S.



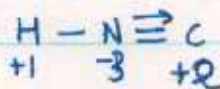
• Using structure  $\rightarrow$



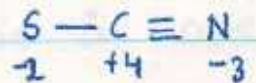
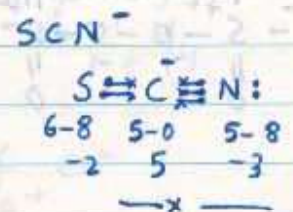
b)

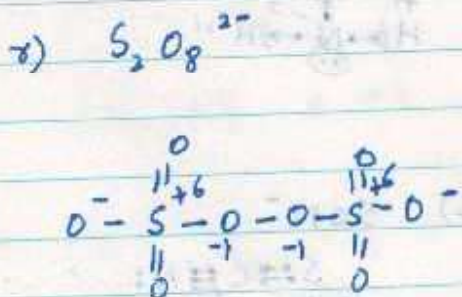
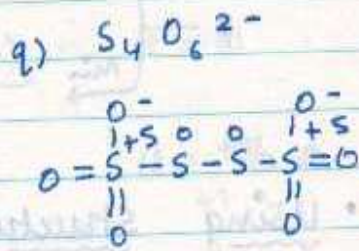
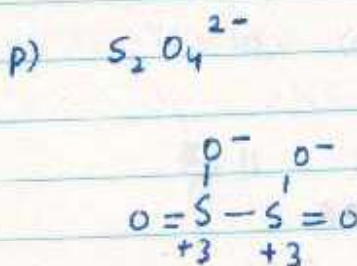
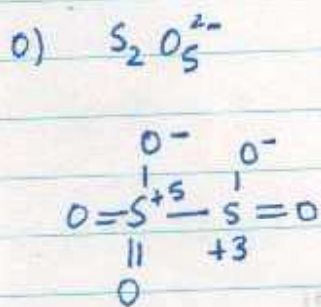
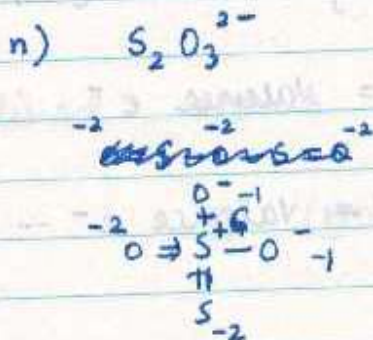
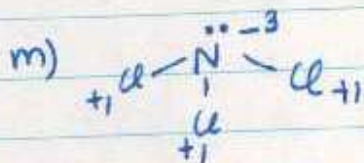
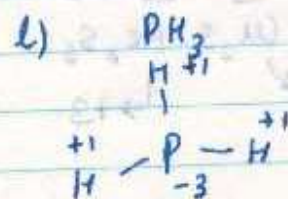
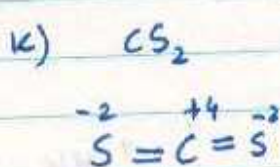
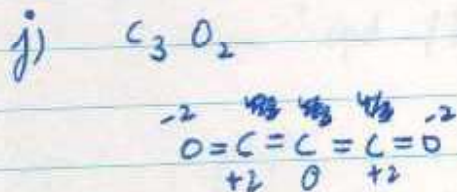
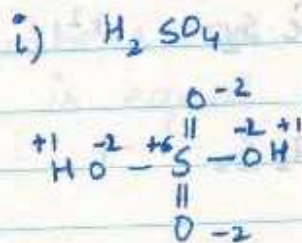
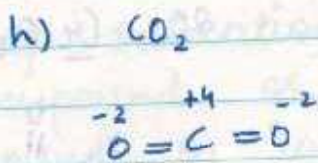
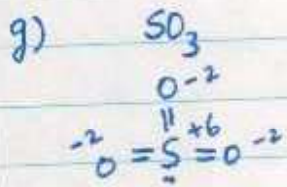


e)



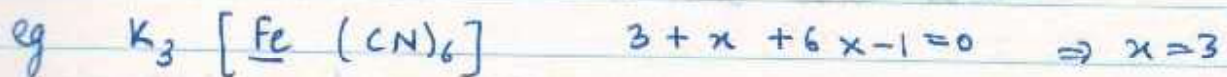
f)







• For co-ordination compounds,



\* Redox Reaction  $\Rightarrow$  These are the rxn in which exchange of  $e^-$  take place. It consists of two parts

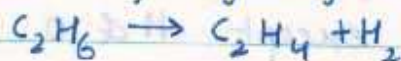
- i) Oxidation half
- ii) Reduction half

### Oxidation Half

• gain of oxygen



• loss of hydrogen



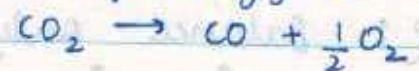
• loss of  $e^-$



- increase in O.S.

### Reduction Half

• loss of oxygen



• gain of hydrogen



• gain of  $e^-$



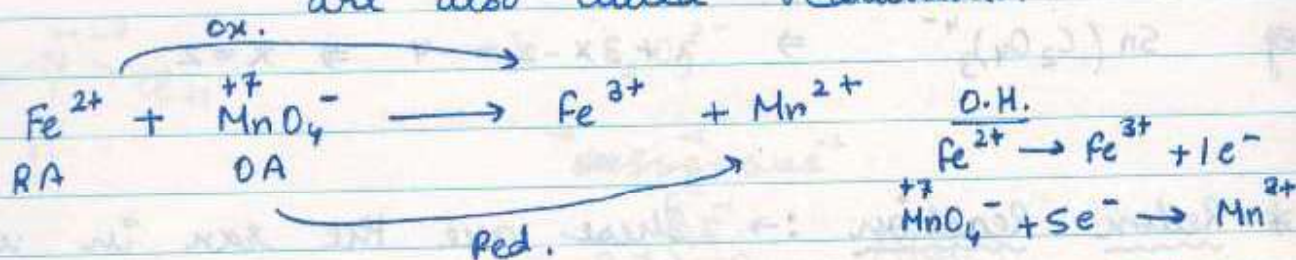
- decrease in O.S.



Note: Only oxidation half or only reduction half is not possible these will occur simultaneously.

Oxidising agent: Substances which oxidise others but themselves get reduced. They are also called oxidants.

Reducing agents: Substances which reduce others but themselves get oxidised. They are also called reductants.



### \* Balancing of Redox Reaction :->

Rule-1: Assign O.N. to the elements which undergo change in oxidation states

Rule-2: Balance only those atoms for which O.S. is changing in separate half rxns.

Rule-3: Balance all other atoms except H & O by inspection

Rule-4: Now balance H & O





i) In Acidic medium  $\rightarrow$

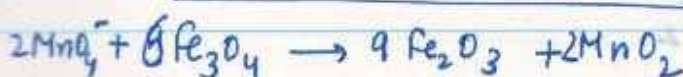
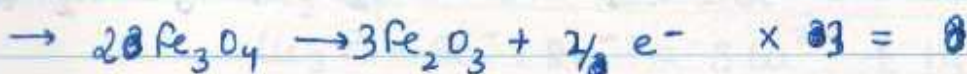
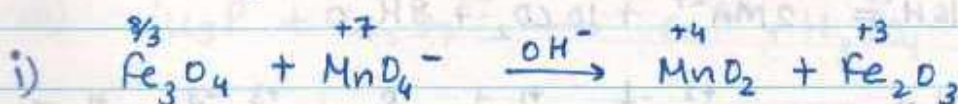
For one excess 'O' atom add one 'H<sub>2</sub>O' on the other side and 2 H<sup>+</sup> on same side.

ii) In Alkaline medium  $\rightarrow$

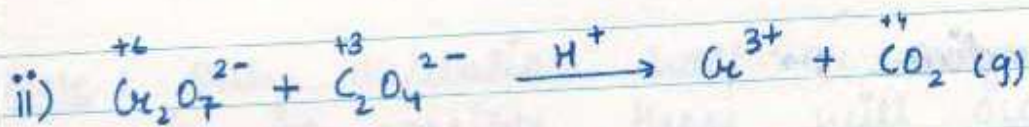
Add one H<sub>2</sub>O on the same side & 2OH<sup>-</sup> on other side (For one extra 'O' atom).

Note: Sometimes just by adding H<sub>2</sub>O, rxn gets balanced so there is no need to add H<sup>+</sup> ion or OH<sup>-</sup> ion.

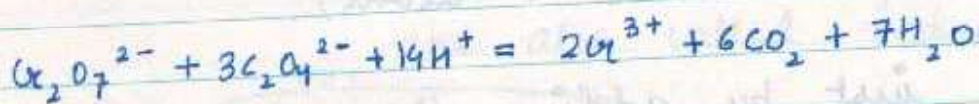
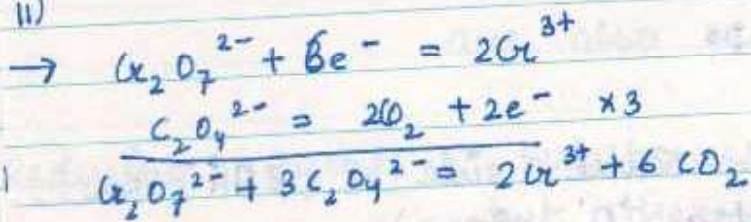
In a balanced rxn, total no. of atoms as well as charge remains balanced.



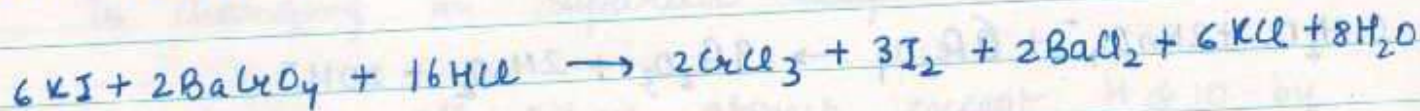
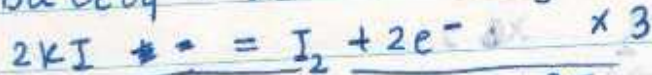
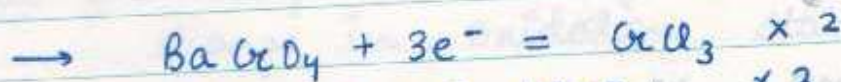
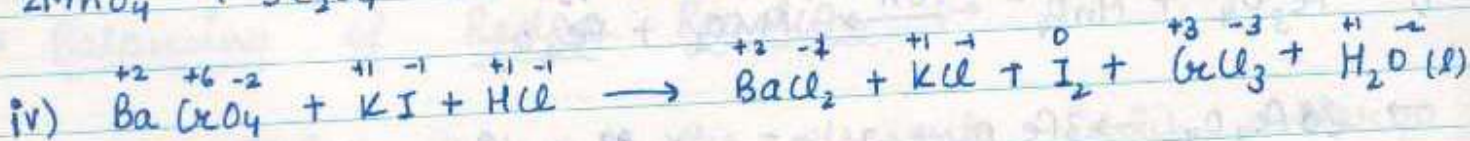
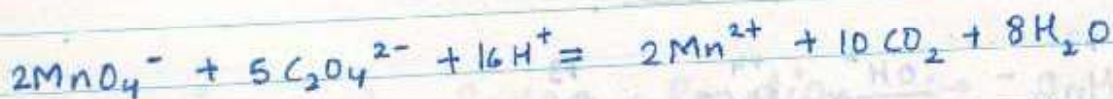
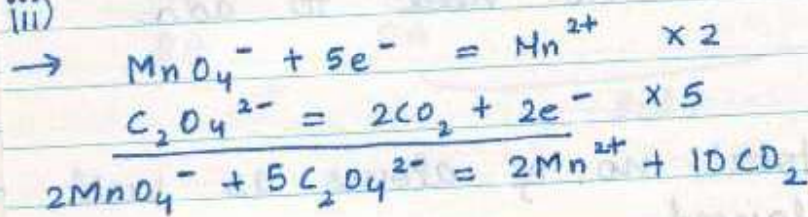




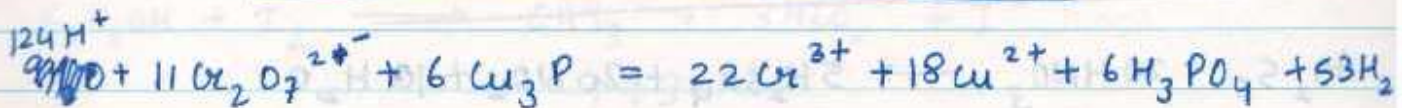
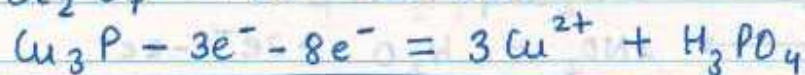
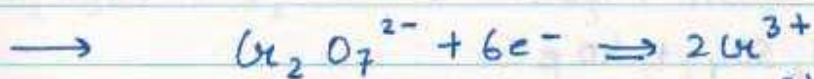
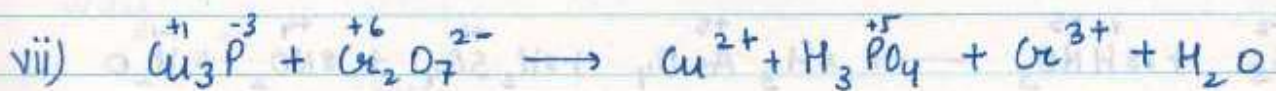
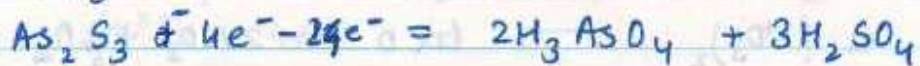
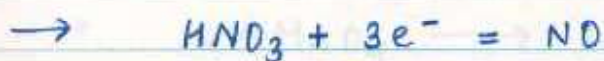
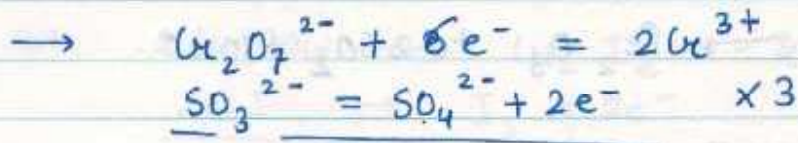
ii)



iii)

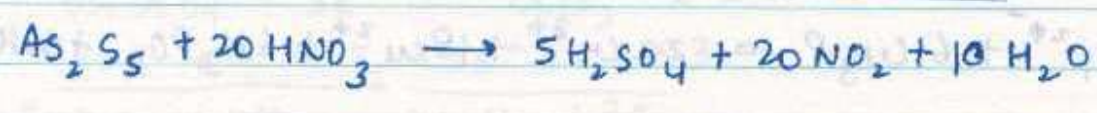
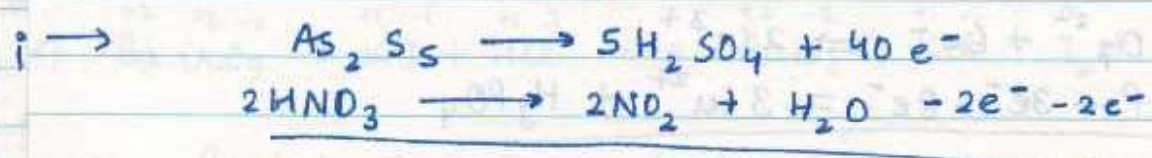
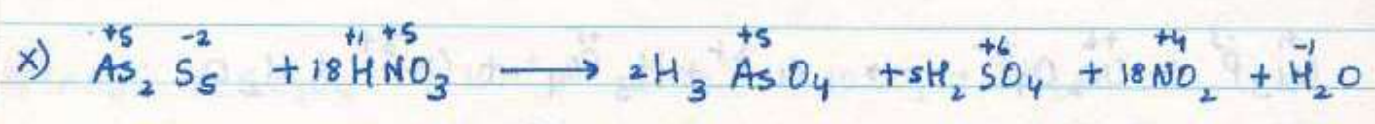
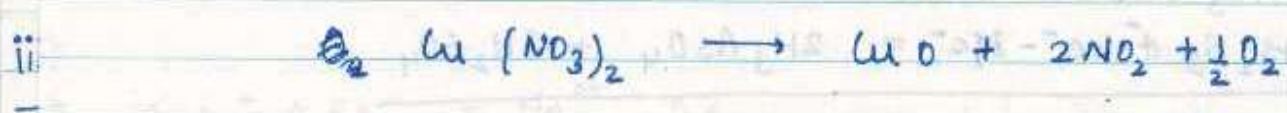
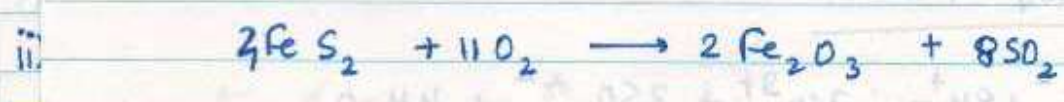
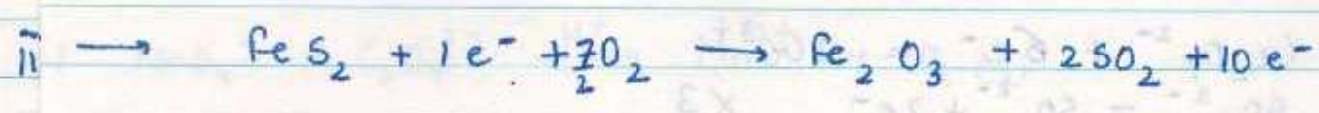
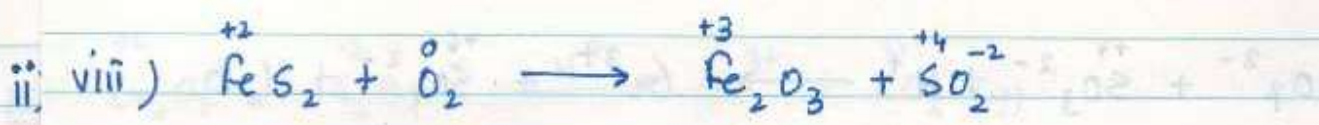




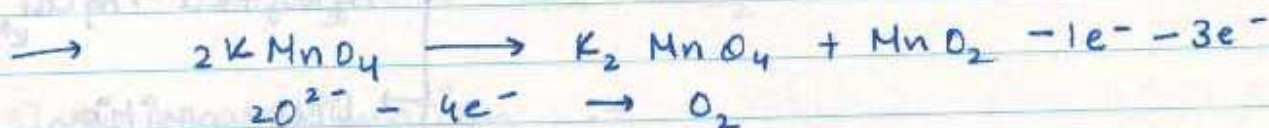
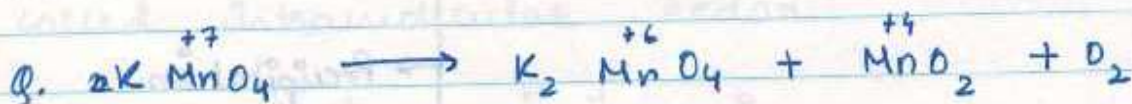
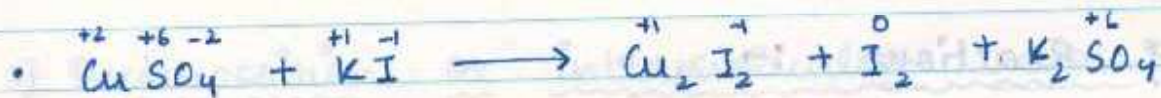




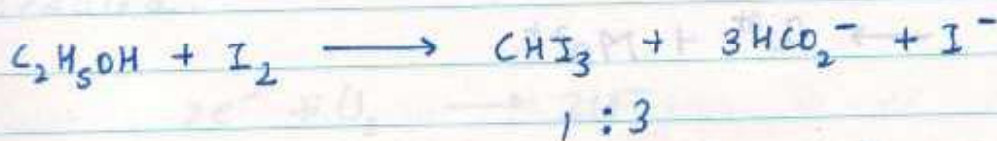
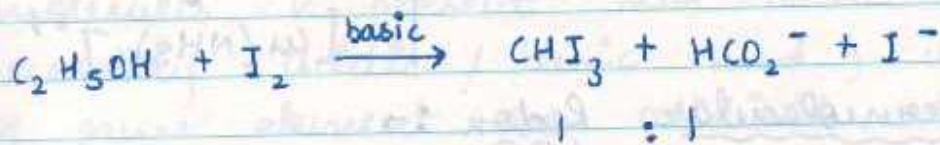
$-1 \rightarrow +4$   $+5 \rightarrow +6$





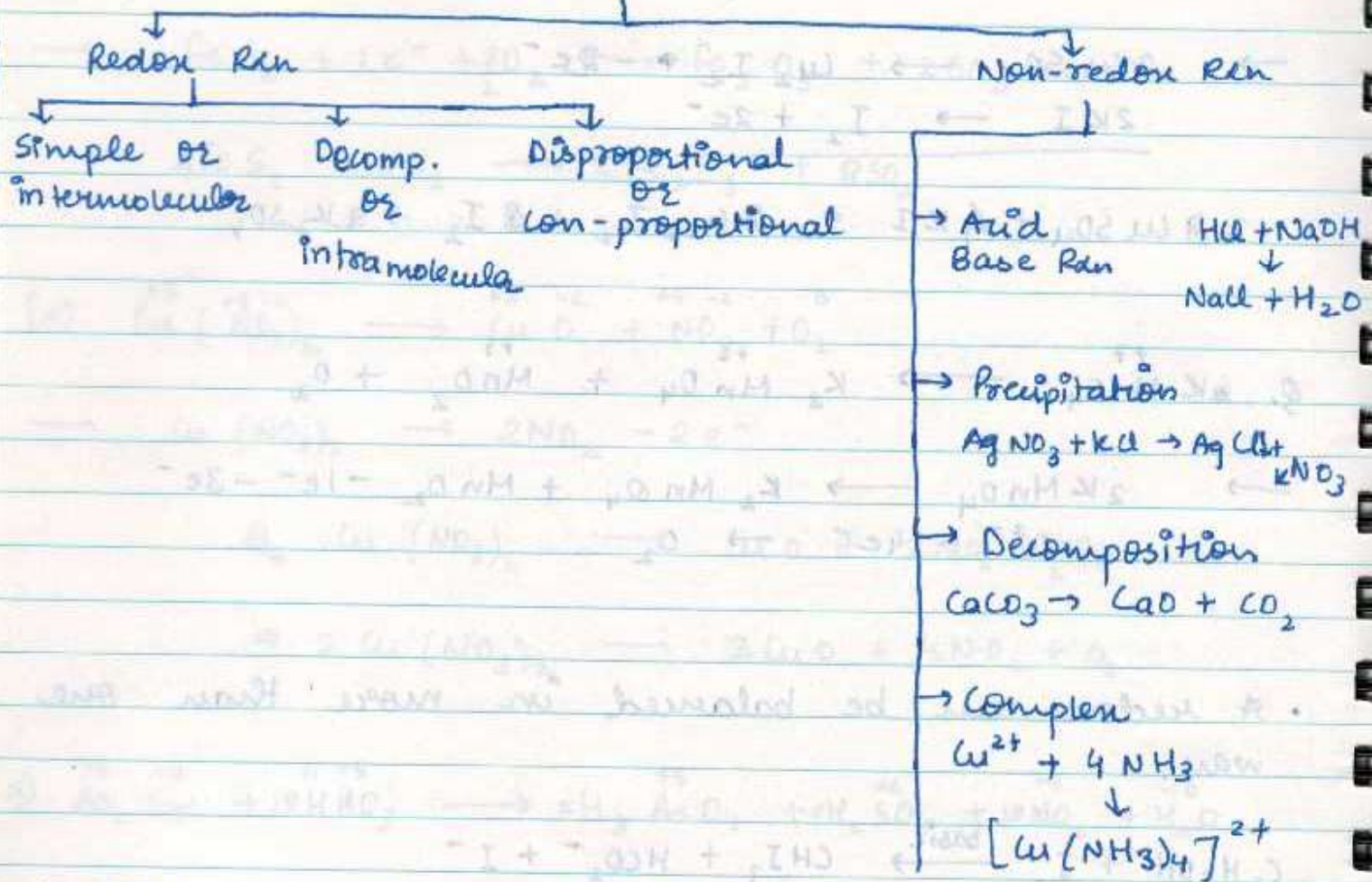


• A redox can be balanced in more than one way.





\* Types of Reactions :-



1] Simple or intermolecular Redox :-



It involves more than one reactant molecules in which atoms of one element are oxidised while atoms present in another molecule gets reduced.

Becoz exchange of e<sup>-</sup> occur b/w two diff molecules that's why it is called intermolecular Redox.

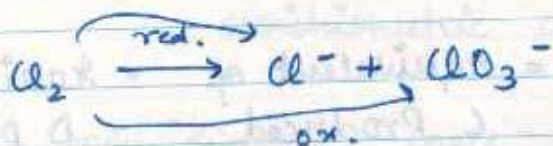


## 2] Decomposition or intramolecular redox :-

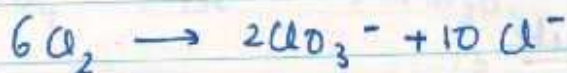
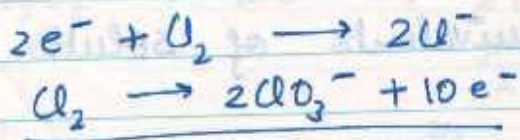
It involves only one reactant molecule in which atoms of one element get oxidised while atoms of other element present in same molecule gets reduced becoz exchange of  $e^-$  occur within same molecule, it is also called intramolecular redox.



## 3] Disproportionation :-



It involves only one molecule (if other molecules  $\times r$  present then those will provide necessary acidic / basic med.) in which atoms of same element are oxidised as well as reduced.



H.W.

DPP # 1, 2

Ans  $\rightarrow$  1





H.W.

D-block elements :  $\text{KMnO}_4$ ,  $\text{K}_2\text{Cr}_2\text{O}_7$  preparation & properties

\* Law of Chemical Equivalence  $\Rightarrow$

Acc<sup>n</sup> to it, during a chemical reaction no. of equivalents of each reactant reacted will be equal to no. of equivalents of each product produced irrespective of stoichiometric coefficients of reactants & products.



Equivalent of A reacted = Equivalent of B reacted = Equivalent of C Produced = Equivalent of D produced

$$\text{No. of equivalents} = \frac{W}{E} = \frac{W}{M} \times n_{\text{factor}} = \text{mole} \times n_{\text{factor}}$$

$$\text{Equivalent wt (E)} = \frac{M}{n_{\text{factor}}}$$

• Normality (N)  $\Rightarrow$  No. of equivalents of solute present per litre of solution.

$$N = \frac{\text{no. of eq. of solute}}{V_s(L)}$$

$$N = \frac{n_B}{V_s(L)} \times n_{\text{factor}} = M \times n_f$$





No. of equivalents =  $NV$  (L)  
 No. of milliequivalents =  $NV$  (ml)

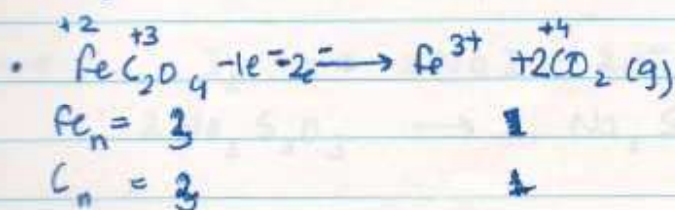
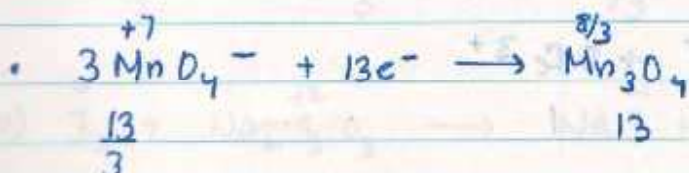
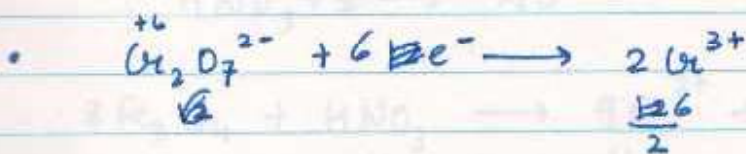
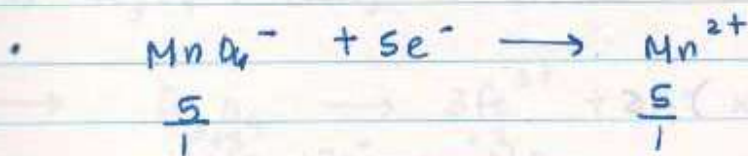
\* calculation of n factor  $\Rightarrow$

A] For Redox Reaction  $\Rightarrow$

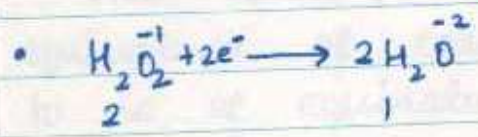
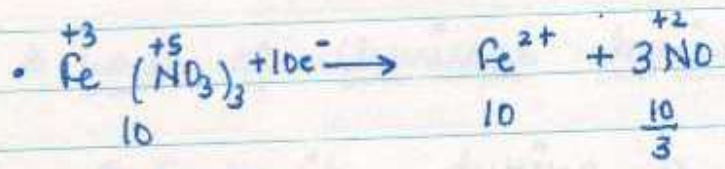
$n_f$ : No. of  $e^-$  lost or gained per molecule or ion of reactants or products.

$$n \text{ factor} = \frac{\text{no. of } e^- \text{ lost / gained}}{\text{stoichiometric coeff. of reactant / products}}$$

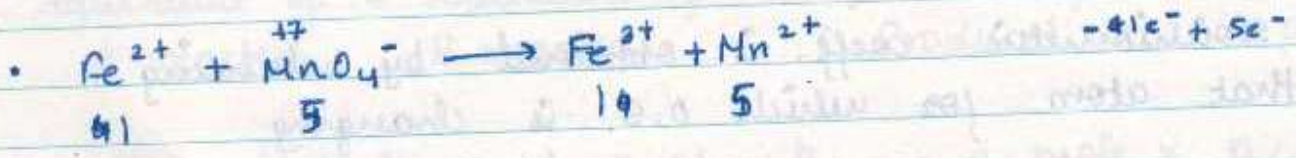
Note: Stoichiometric coeff. is obtained by balancing that atom for which O.S. is changing



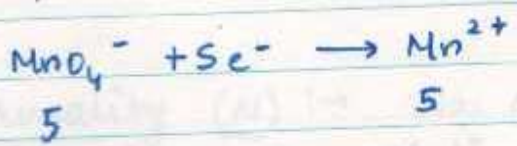




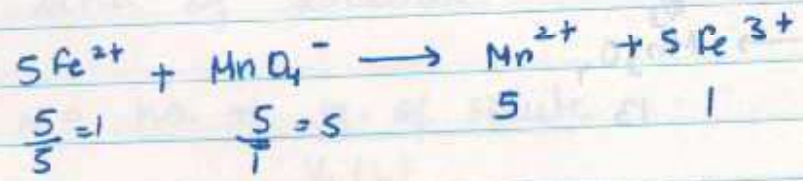
i) For simple redox rxn



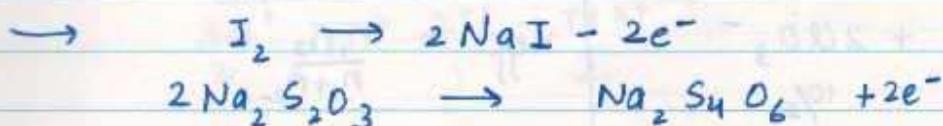
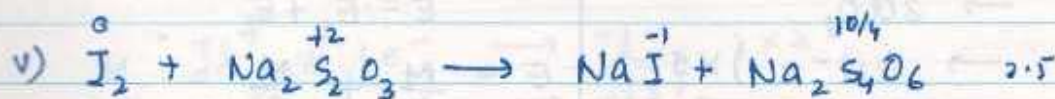
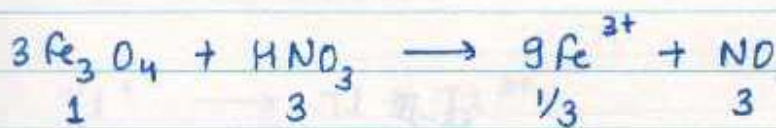
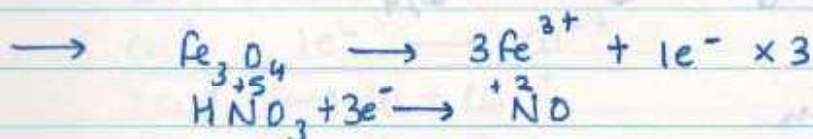
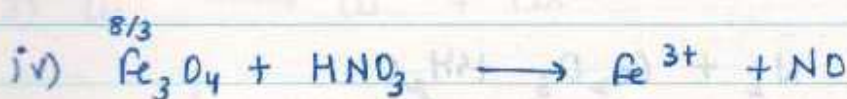
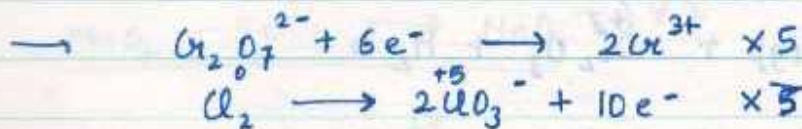
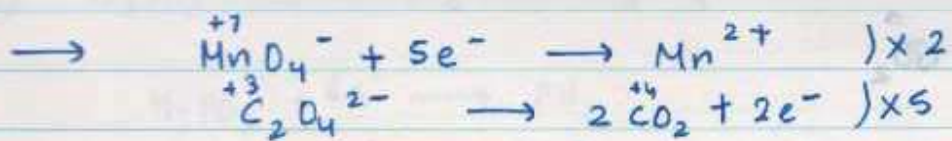
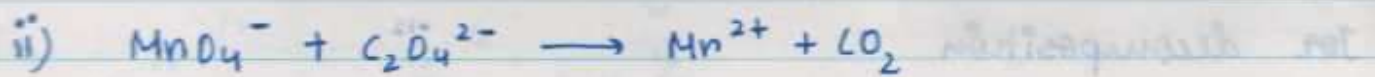
M-I



M-II

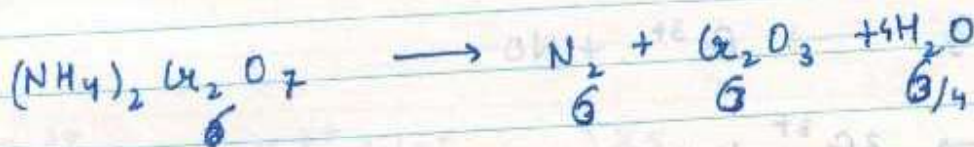
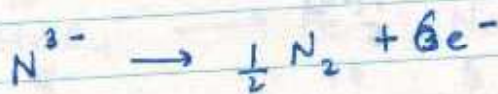
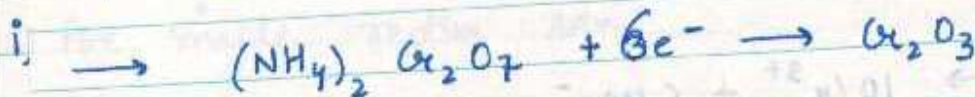
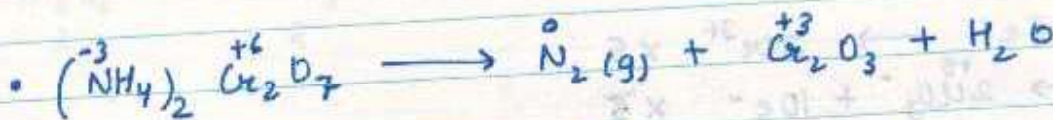
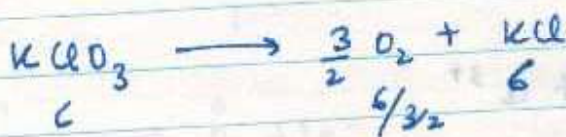
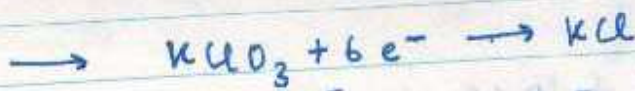
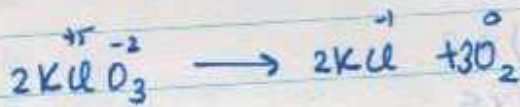




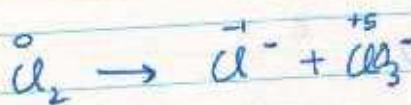




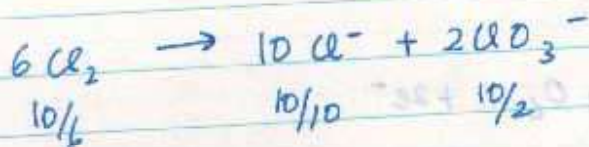
ii) For decomposition



iii) For disproportionation



M-I



M-II

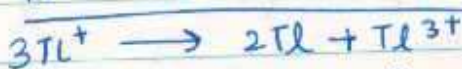
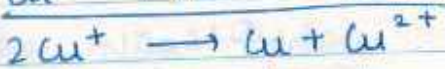
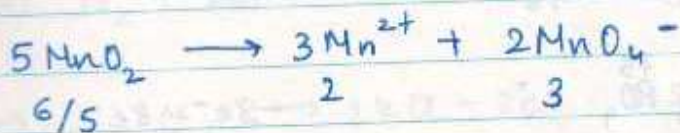
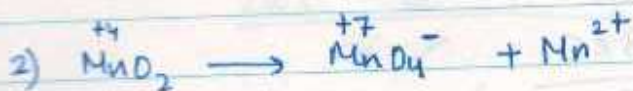
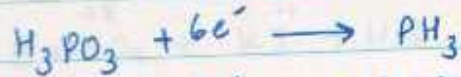
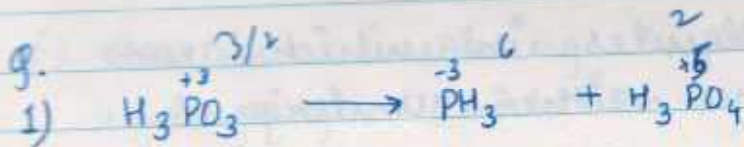
$$E = E_1 + E_2$$

$$E = \frac{M}{n_1} + \frac{M}{n_2} = \frac{M}{n_{\text{eff}}}$$

$$n_{\text{eff}} = \frac{n_1 n_2}{n_1 + n_2}$$

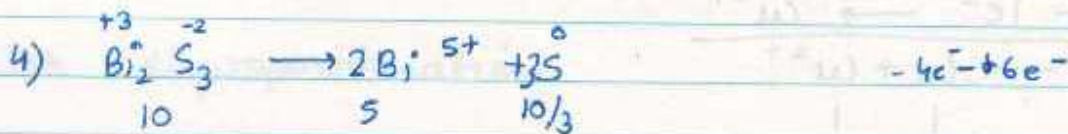
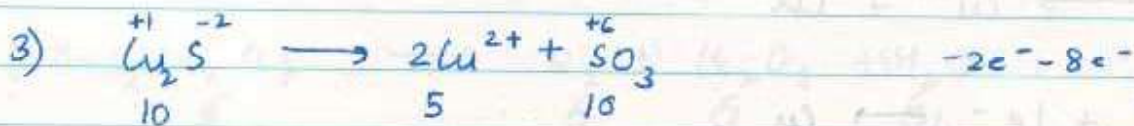
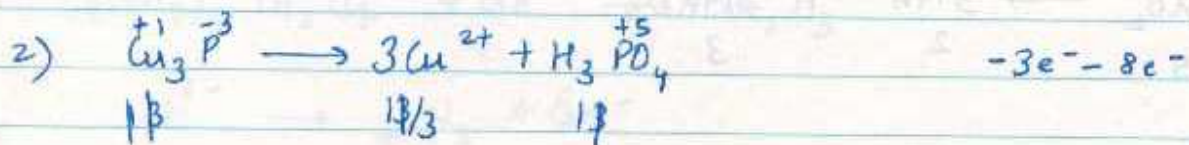
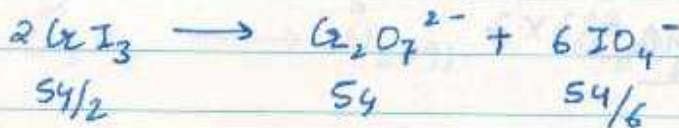
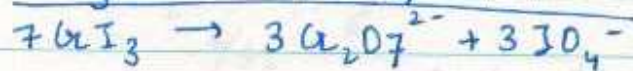
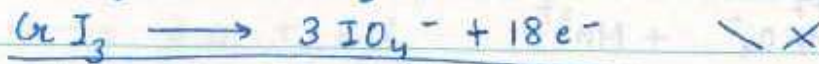
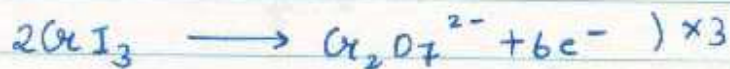
$$= \frac{2 \times 10}{12} = \frac{5}{3}$$







iv] Reactions involving molecules in which more than one element atoms either undergo oxidation / Reduction.



v] Reactions in which an element undergoes either oxidation or reduction in more than one oxidation states.





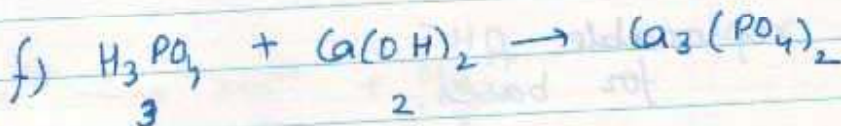
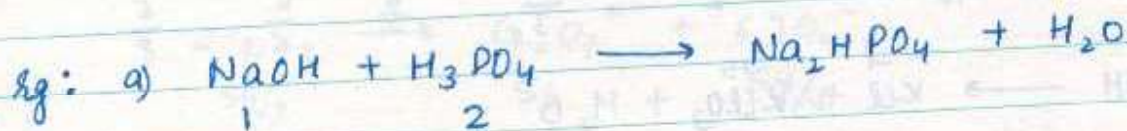




- $HCl \rightarrow 1$
- $H_2CO_3 \rightarrow 2$
- $H_3PO_4 \rightarrow 3$
- $H_3PO_3 \rightarrow 2$
- $H_3PO_2 \rightarrow 1$
- $Ca(OH)_2 \rightarrow 2$
- $NaOH \rightarrow 1$
- $Al(OH)_3 \rightarrow 3$

Case II:  $n_{\text{factor}}$  for acids / bases when rxn is given

$$n_f = \text{no. of } H^+ / OH^- \text{ replaced} \times | \text{charge on } H^+ / OH^- |$$





Case III :  $n_f$  of ions

$$n_f = |\text{magnitude of charge on ion}|$$

- $\text{Cl}^- = 1$
- $\text{SO}_4^{2-} = 2$
- $\text{PO}_4^{3-} = 3$
- $\text{NH}_4^+ = 1$
- $\text{CO}_3^{2-} = 2$

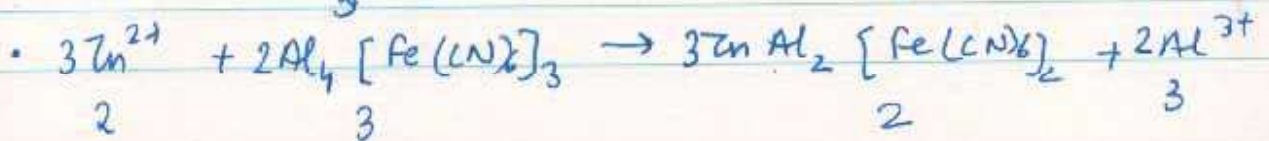
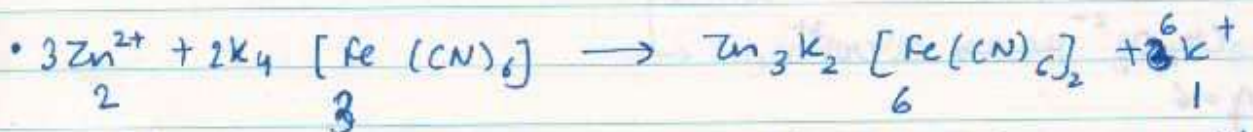
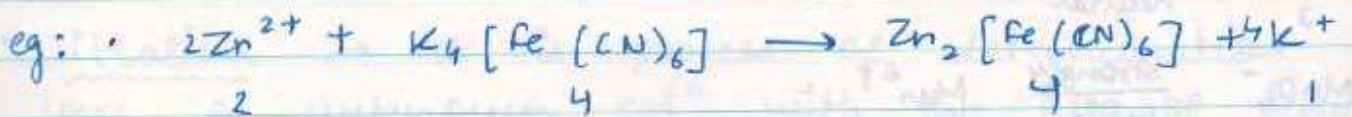
Case IV :  $n_{\text{factor}}$  for salts

$$n_f = \text{total +ve or total -ve charge}$$

- $\text{NaCl} = 1$
- $\text{CaCl}_2 = 2$
- $\text{Na}_3\text{PO}_4 = 3$
- $\text{Ca}_3(\text{PO}_4)_2 = 6$
- $\text{Al}_2(\text{SO}_4)_3 = 6$

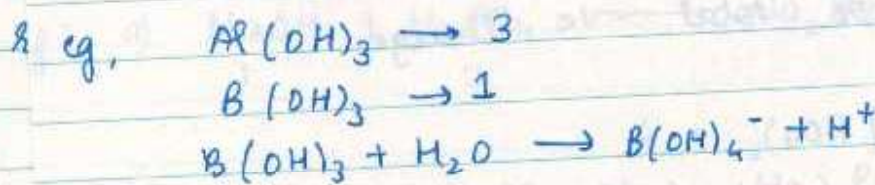
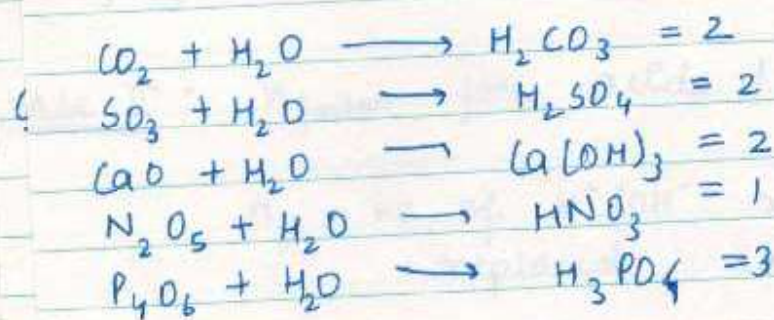
Case V :  $n_{\text{factor}}$  for Rxn involving complex compounds

$$n_f = \frac{\text{no. of ions replaced in complex}}{\text{ion}} \times |\text{charge on ion}| = \text{charge replaced per mol}$$



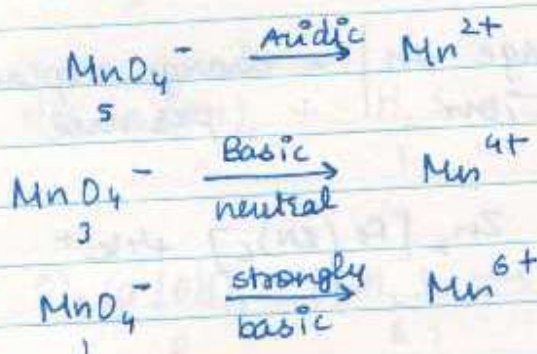


Case VI :  $n_f$  of acidic / basic oxides is equal to  $n_f$  of acid or base obtained by dissolving that oxide in water.



Some important oxidising agents  $\rightarrow$

1] KMnO<sub>4</sub>



2] K<sub>2</sub>Cr<sub>2</sub>O<sub>7</sub>





3]  $\underline{\underline{H_2O_2}}$

(RA)

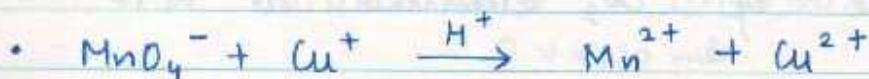
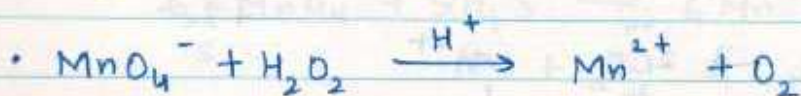


$n_f = 2$

(OA)



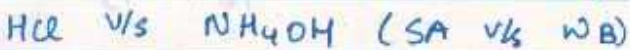
$n_f = 2$



\*  $HNO_3$  &  $H_2SO_4$  have 'N' & 'S' in their max. O.S.  
So they can also act as O.A.

• Titration  $\Rightarrow$  It is process of determination of conc. of unknown sol<sup>n</sup> with the help of known solution. We will mainly deal with the following types of titration:

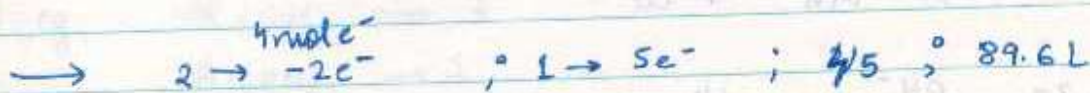
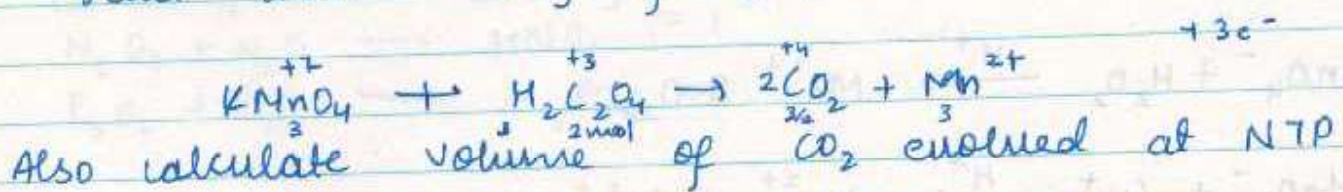
a) Acid - Base Titration



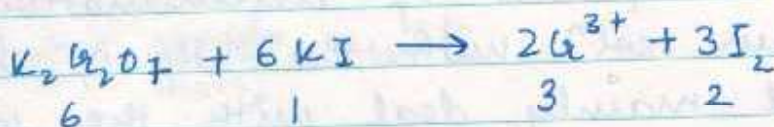
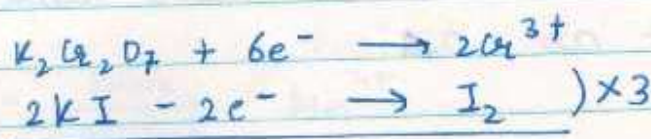


- b) Redox Titration
- Titration with  $\text{KMnO}_4$
  - " "  $\text{K}_2\text{Cr}_2\text{O}_7$
  - Iodometric titration
  - Iodimetric titration

g. Calculate no. of moles of  $\text{KMnO}_4$  which will react with 180 g of  $\text{H}_2\text{C}_2\text{O}_4$  as



g. Calculate moles of  $\text{K}_2\text{Cr}_2\text{O}_7$  required to produce 254 g of  $\text{I}_2$  from  $\text{KI}$  sol<sup>n</sup>.

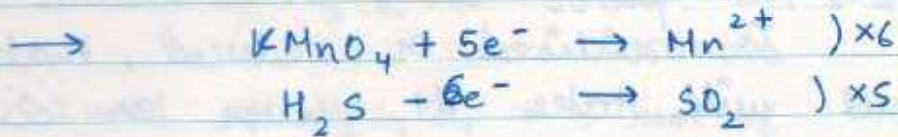


$$6 \times 2 = 3 \times 4 = 3 \times x = 2 \times 2$$

$$\therefore z = \frac{4}{6} = \frac{2}{3}$$



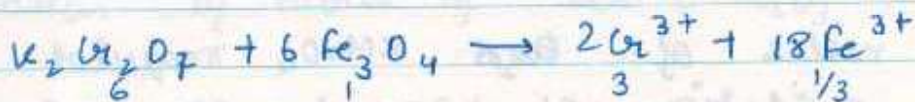
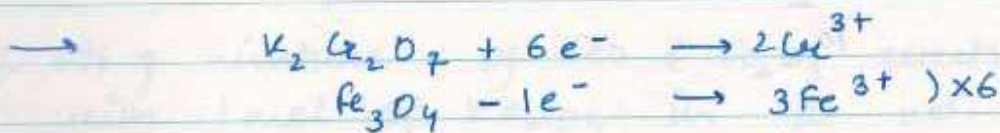
8. Calculate volume of 0.05 M  $\text{KMnO}_4$  which will react with 50 ml of 0.1 M  $\text{H}_2\text{S}$  in  $\text{H}^+$  med.  
( $\text{H}_2\text{S} \rightarrow \text{SO}_2$ )



$$\Rightarrow 0.05 \times V \times 5 = 0.1 \times 50 \times 6$$

$$\Rightarrow V = 120 \text{ ml}$$

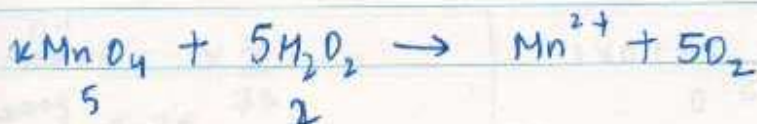
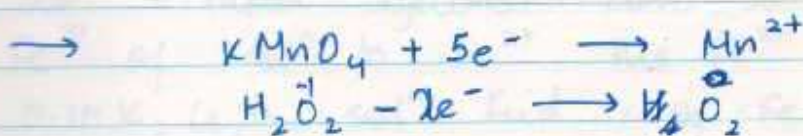
9. Calculate mm of  $\text{Fe}_3\text{O}_4$  that reacts completely with 25 ml of 0.3 M  $\text{K}_2\text{Cr}_2\text{O}_7$



$$\Rightarrow 25 \times 0.3 \times 6 = x$$

$$\therefore x = 45 \text{ mm}$$

10. Calculate conc. of  $\text{H}_2\text{O}_2$  sol<sup>n</sup> if 20 ml of  $\text{H}_2\text{O}_2$  sol<sup>n</sup> reacts completely with 10 ml 2M  $\text{KMnO}_4$

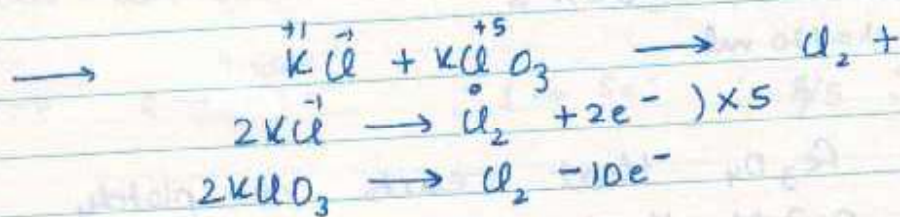


$$\left. \begin{array}{l} 5 \times 10 \times 2 = x \times 20 \times 2 \\ \therefore x = \frac{5 \text{ M}}{2} \end{array} \right\}$$

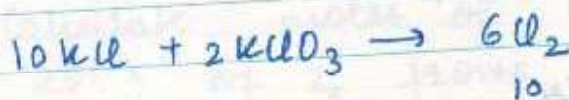


Note : w.r.t. decomposition or disproportionation,  $N$  is not defined coz to find normality we need  $n$  factor which can be obtained titration. eg, for  $H_2O_2$ ,  $n$  factor is 2 whether it is oxidised or reduced, that's why we will take its  $n$  factor as 2.

Q. Calculate moles of  $KCl$  required to produce 10 mol  $Cl_2$  by rxn with  $KClO_3$ .

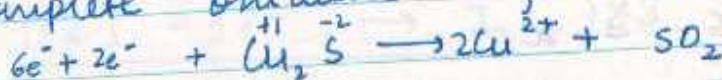


$$6 \times \frac{10}{6} ; 10 \times \frac{10}{6}$$



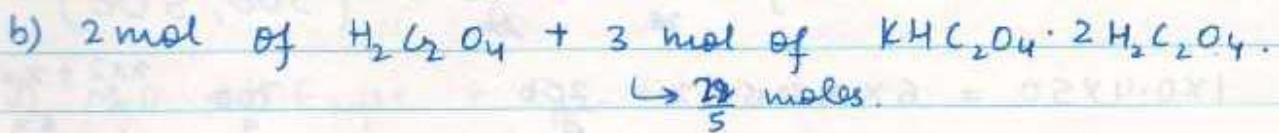
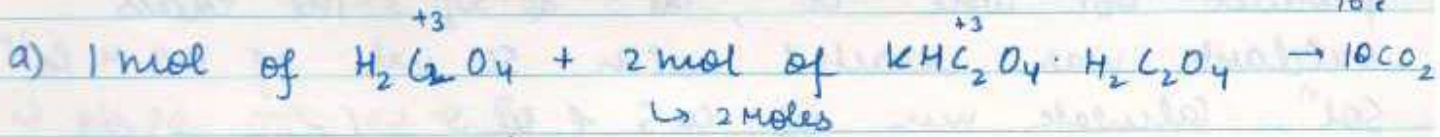
$100/6$  moles.

Q. Calculate moles of  ~~$KMnO_4$~~   $KMnO_4$  required for complete oxidation of 1.25 mol of  $Cu_2S$





Q. Calculate moles of  $\text{KMnO}_4$  required to react with a mixture of



Q. NaOH

a)  $\rightarrow$  ~~8~~

b)  $\rightarrow$  ~~19~~

H.W.

DPP #3

EX 1: 6, 7, 8, 9, 10, 11, 12, 13, 14, 15, 16, 17, - 23.

Q. 696 g mixture of  $\text{FeO}$  &  $\text{Fe}_2\text{O}_3$  reacts completely with 1 mol of  $\text{KMnO}_4$  in ac. med. Calculate composition by moles of  $\text{FeO}$  &  $\text{Fe}_2\text{O}_3$ .

$$\rightarrow 1 \times 5 = \frac{x}{72} \times 1 \Rightarrow x = 360 \text{ g} = 5 \text{ moles}$$

$$\frac{336}{160} = 2.1$$

Q. A min. of  $\text{FeO}$  &  $\text{Fe}_2\text{O}_3$  is reacted with acidified  $\text{KMnO}_4$  sol<sup>n</sup> having a concentration of 0.2 M, 100 ml of which was used. The sol<sup>n</sup> was then titrated against zinc dust which converted  $\text{Fe}^{3+}$  of sol<sup>n</sup> to  $\text{Fe}^{2+}$ . The  $\text{Fe}^{2+}$  requires 1000 ml of 0.1 M  $\text{K}_2\text{Cr}_2\text{O}_7$  sol<sup>n</sup>. Find % of  $\text{FeO}$  &  $\text{Fe}_2\text{O}_3$ .

$$\rightarrow 0.2 \times 100 \times 5 = 1 \times \frac{x}{72}$$

~~2000g.~~ 7.2g

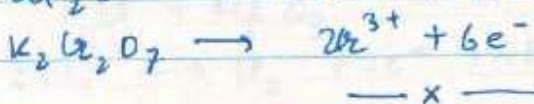
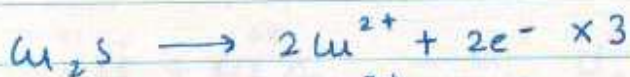
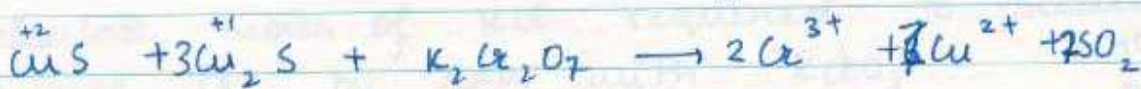
$$1 \times 0.1 \times 6 = 1 \times \frac{x}{0.6}$$

$$0.6 \times 160 = 96 \text{ g.}$$



Q. A mixture cont. equal moles of  $\overset{+2}{\text{CuS}}$  &  $\overset{+1}{\text{Cu}_2\text{S}}$  was treated with 100 ml of 0.15 M  $\text{K}_2\text{Cr}_2\text{O}_7$ . The product obt. were  $\text{Cr}^{3+}$ ,  $\text{Cu}^{2+}$  &  $\text{SO}_2$ . The excess oxidant was reacted with 50 ml of 0.4 M  $\text{Fe}^{2+}$  sol<sup>n</sup>. Calculate mm of  $\overset{+2}{\text{CuS}}$  &  $\overset{+1}{\text{Cu}_2\text{S}}$ . (500, 500)

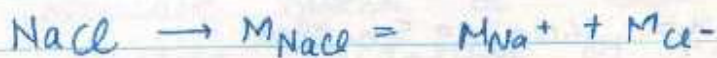
$$\rightarrow 1 \times 0.4 \times 50 = 6 \times x \times 0.15 \Rightarrow x = \frac{200}{9} \quad \frac{700}{9} \quad \begin{matrix} x \times 2 + x \times 2 \times 2 \\ 6x \end{matrix}$$



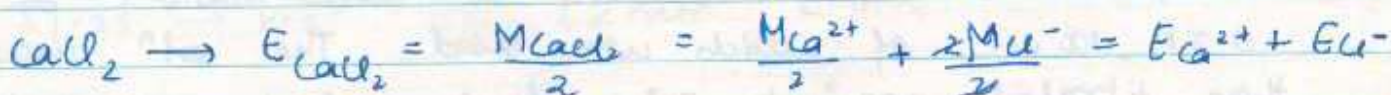
$$\frac{700}{9} \times 0.15 \times 6 = 2 \times x + 2 \times 2 \times 2 \times x$$



\*\*\*



$$\hookrightarrow E_{\text{NaCl}} = \frac{M_{\text{NaCl}}}{n_f} = \frac{M_{\text{Na}^+}}{1} + \frac{M_{\text{Cl}^-}}{1} = E_{\text{Na}^+} + E_{\text{Cl}^-}$$



$$\begin{aligned} \text{Q. } \text{H}_2\text{SO}_4 \rightarrow E &= E_{\text{H}^+} + E_{\text{SO}_4^{2-}} \\ &= 1 + 48 \\ &= \underline{49} \end{aligned}$$



$$\begin{aligned} \text{b) } \text{Na}_2\text{SO}_4 &\rightarrow E_{\text{Na}^+} + E_{\text{SO}_4^{2-}} \\ &= 23 + 48 = 71 \end{aligned}$$

$$\text{c) } \text{Na}_3\text{PO}_4 \rightarrow 23 + \frac{95}{3}$$

$$\begin{aligned} \text{d) } E_{\text{MgO}} &= E_{\text{Mg}^{2+}} + E_{\text{O}^{2-}} \\ &= \frac{24}{2} + \frac{16}{2} = 20 \end{aligned}$$

$$\begin{aligned} 1 \text{ mol MgO} &= 2 \text{ eq. of MgO} \\ \frac{1}{2} \text{ mol MgO} &= 1 \text{ eq. of MgO} \end{aligned}$$

$$\text{e) } E_{\text{Na}_2\text{O}_2} = \begin{matrix} E_{\text{Na}^+} & E_{\text{O}^{2-}} \\ 23 & 16 \end{matrix}$$

Q. A metal oxide contains 20% metal. Calculate eq. wt. of metal. (by mass)



$$\text{eq. of metal} = \text{eq. of oxide}$$

$$\frac{w_{\text{metal}}}{E_{\text{metal}}} = \frac{w_{\text{oxide}}}{E_{\text{oxide}}} \Rightarrow \frac{20}{x} = \frac{80}{8}$$

$$\therefore E_{\text{metal}} = 2$$

Q. A super oxide contains 84% by mass of metal. Calculate  $E_{\text{metal}}$ .

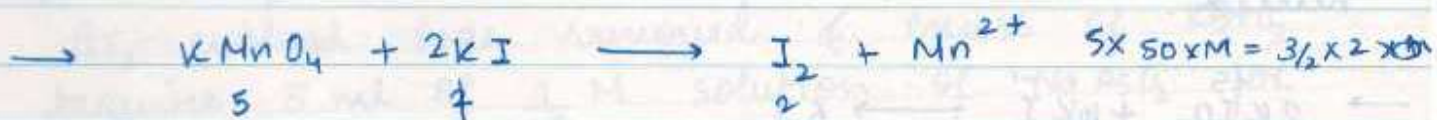
$$\rightarrow \frac{84}{E} = \frac{16}{32} \Rightarrow E = 168$$







Q. 50 ml of  $\text{KMnO}_4$  is mixed with excess of  $\text{KI}$ . The  $\text{I}_2$  liberated required 30 ml of 0.1 M  $\text{Na}_2\text{S}_2\text{O}_3$  sol<sup>n</sup>. Calculate molarity of  $\text{KMnO}_4$ .



$$\Rightarrow 30 \times 0.1 \times 1 = 2 \times x \quad \Rightarrow x = \frac{3}{2}$$

$$\Rightarrow M = \frac{3}{50}$$

- In this case 'n' factor of Iodine when it is formed is same as its 'n' factor when it again reacts. So no. of eq. of  $\text{I}_2$  will be same in both rxn. We can say,  
eq. of  $\text{KMnO}_4 = \text{eq. of hypo.}$

- If 'n' factor is diff. when  $\text{I}_2$  is formed & when it again reacts then milli eq. of  $\text{I}_2$  will be diff. So we will calculate mm in one rxn & put them in other.

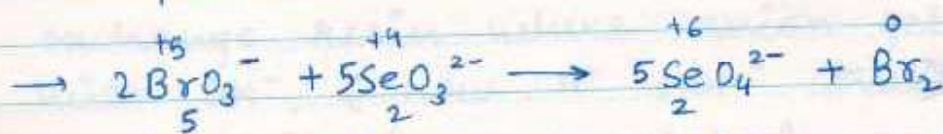






Q. Calculate mm of  $\text{SeO}_3^{2-}$  in solution on the basis of following data.

20 ml of  $\frac{1}{60}$  M  $\text{KBrO}_3$  was added to definite volume of  $\frac{60}{60}$  selenide ion ( $\text{SeO}_3^{2-}$ ) solution. The  $\text{Br}_2$  evolved was removed & excess of  $\text{KBrO}_3$  require 5 ml of  $\frac{1}{25}$  M solution of  $\text{NaAsO}_2$  for complete destn.

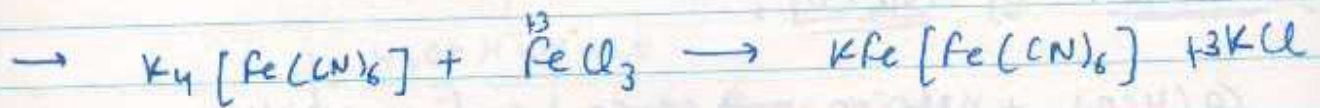


2/3  
Ans.



$$5 \times \frac{1}{25} \times 2 = \frac{3}{2} \times MV = \frac{1}{15}$$

Q. 20 ml of  $\text{K}_4[\text{Fe}(\text{CN})_6]$  sol<sup>n</sup> is divided into two equal parts. One part is completely reacted with 10 ml of  $\text{FeCl}_3$  to form  $\text{KFe}[\text{Fe}(\text{CN})_6]$  &  $\text{KCl}$ . Another part of  $\text{K}_4[\text{Fe}(\text{CN})_6]$  sol<sup>n</sup> is reacted with 20 ml of 0.5 M  $\text{FeCl}_2$  sol<sup>n</sup> to produce  $\text{K}_2\text{Fe}[\text{Fe}(\text{CN})_6]$  &  $\text{KCl}$ . Molarity of  $\text{FeCl}_3$  is 1 M.





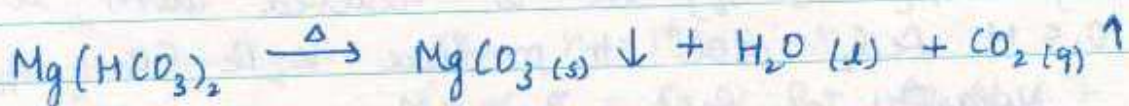
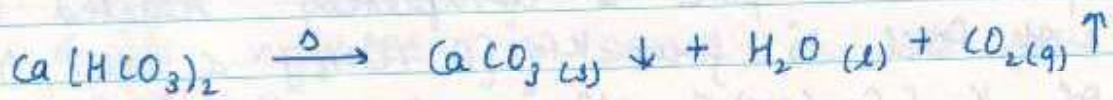
\* Hard Water : Presence of  $\text{Ca}^{2+}$  &  $\text{Mg}^{2+}$  salts in form of  $\text{HCO}_3^-$ ,  $\text{Cl}^-$  &  $\text{SO}_4^{2-}$  makes water hard.  
- Hard water does not produce lather with soap.

\* Soft Water : Water free from soluble salts of  $\text{Ca}^{2+}$  &  $\text{Mg}^{2+}$ .  
- It produces lather with soap solution.

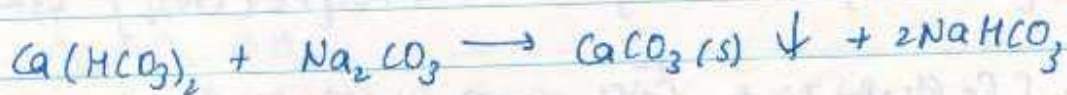
Types of Hardness :->

1] Temporary hardness : It is due to presence of  $\text{MgHCO}_3$  &  $\text{Ca(HCO}_3)_2$  and it can be removed by following ways,

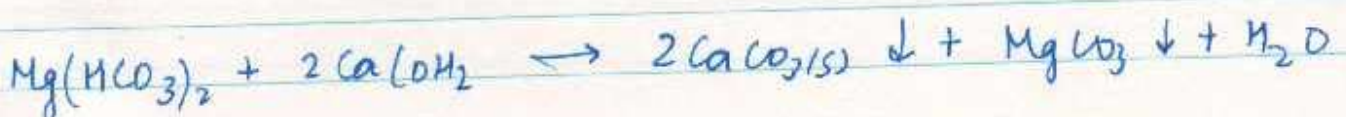
A] Boiling, on boiling soluble hardness causing salts are converted into insoluble precipitates.



B] Addition of  $\text{Na}_2\text{CO}_3$ ,



C] Clark's Method : In it calculated amt of lime is added which ppt hardness causing ions.

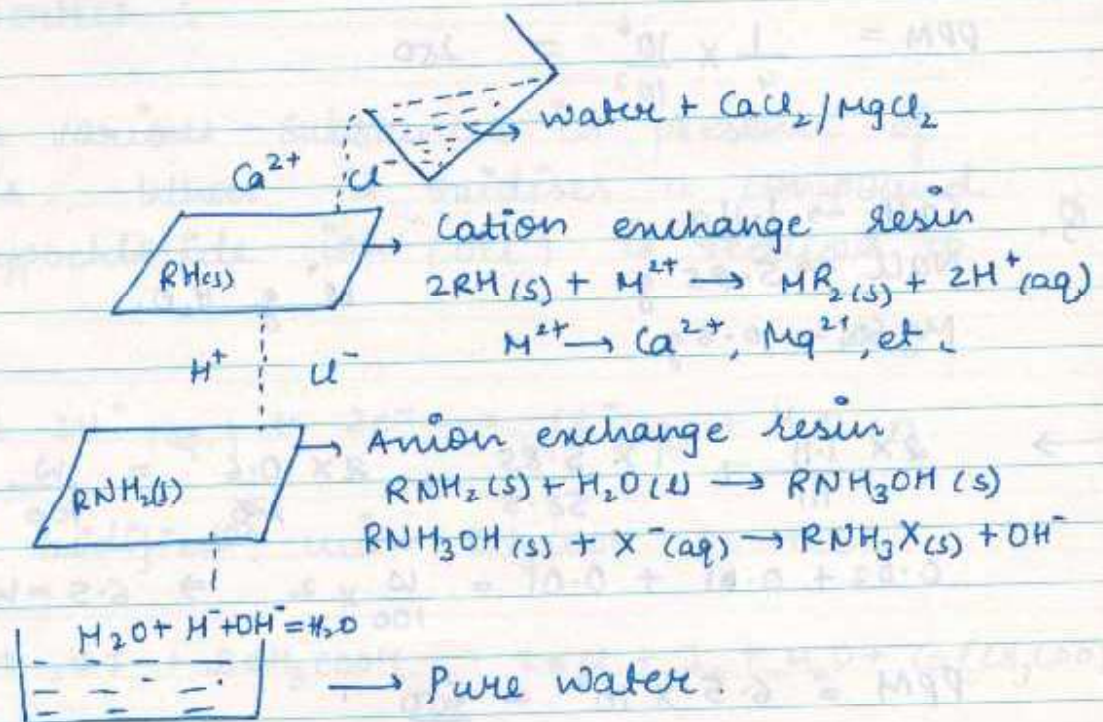




2] Permanent Hardness : due to presence of soluble salts of  $\text{Ca}^{2+}$  &  $\text{Mg}^{2+}$  in form of  $\text{Cl}^-$  &  $\text{SO}_4^{2-}$ . It cannot be simply removed by boiling.

→ To remove it, water is successively passed through cation exchange resin where cation gets exchanged with  $\text{H}^+$  & sol<sup>n</sup> becomes acidic.

→ Now this acidic water is passed through anion exchange resin where anion gets exchanged with  $\text{OH}^-$ , & now  $\text{H}^+$  &  $\text{OH}^-$  combine to form water & it becomes neutral.



• Measurement of Hardness of Water :->

Hardness of water is measured in terms of equivalent amt. of  $\text{CaCO}_3$  in PPM.

$$\text{PPM} = \frac{W_{\text{CaCO}_3}}{W_{\text{solution}}} \times 10^6$$



8. Calculate hardness of water containing 0.3 g of  $MgSO_4$  in 1000 g of water.

$$\rightarrow PPM = \frac{0.3 \times 2}{1000} \times 10^6$$

$$\text{Eq of } MgSO_4 = \text{Eq of } CaCO_3$$

$$2 \times \frac{0.3}{120} = \frac{W_{CaCO_3}}{100} \times 2$$

$$\Rightarrow W_{CaCO_3} = \frac{1}{4}$$

$$PPM = \frac{1}{4} \times \frac{10^6}{10^3} = 250$$

9.  $CaCl_2 \rightarrow 1.11g$   
 $NaCl \rightarrow 5.85g$   
 $MgSO_4 \rightarrow 0.6g$   
 $10^4 g H_2O$

$$\rightarrow 2 \times \frac{1.11}{111} + 1 \times \frac{5.85}{58.5} + 2 \times \frac{0.6}{120} = \frac{W}{100} \times 2$$

$$0.02 + 0.01 + 0.01 = \frac{W}{100} \times 2 \Rightarrow 6.5 = W$$

$$PPM = 6.5 \times 10^2 = \underline{650}$$

10.  $CaCl_2 \rightarrow 1mm$   
 $NaCl \rightarrow 2mm$   
 $MgSO_4 \rightarrow 2mm$

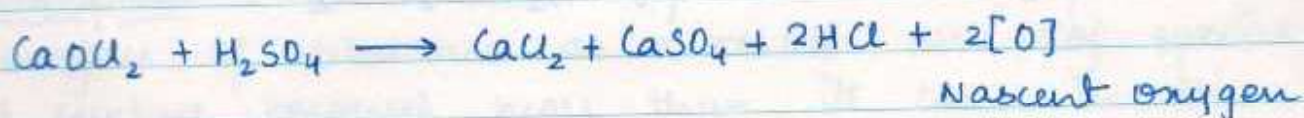
$$\rightarrow \frac{4 \times 10^{-3} \times 100}{10^4} \times 10^6 = \underline{40}$$





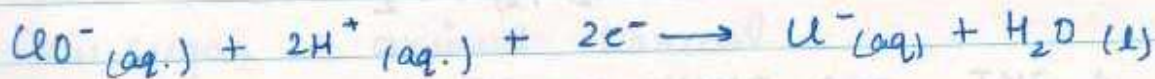
## \* Bleaching Powder ( $\text{CaOCl}_2$ ) :

In presence of slight amt. of dilute acids,  $\text{CaOCl}_2$  loses nascent oxygen & due to this [O] it shows oxidising as well as bleaching properties.

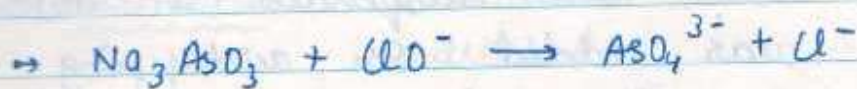
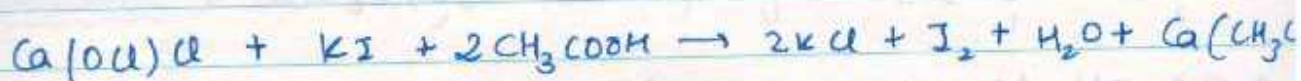


### Oxidising Properties :

→ It oxidises various substances in presence of dilute acids. When it oxidises a compound then its hypochlorite ion ( $\text{OCl}^-$ ) is reduced to  $\text{Cl}^-$  ion.

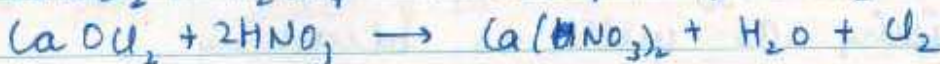


→ KI solution acidified with  $\text{CH}_3\text{COOH}$  or  $\text{HCl}$





Q. → Action of excess of dilute acids :-



Amnt. of  $\text{Cl}_2$  liberated by action of excess of dilute acids on sample of bleaching powder is called % available chlorine of that sample.

$$\% \text{ available chlorine} = \frac{W_{\text{Cl}_2}}{W_{\text{sample}}} \times 100$$



Q. 0.5 g of bleaching powder was suspended in water and excess of KI was added. On acidifying with dil.  $\text{H}_2\text{SO}_4$ ,  $\text{I}_2$  was liberated which reqd 50 ml of  $\frac{1}{10}$  N  $\text{Na}_2\text{S}_2\text{O}_3$ . Calculate % of chlorine

$$\rightarrow 50 \times \frac{1}{10} = 2 \times M \times V \Rightarrow \text{moles} = 2.5 \text{ mm}$$

$$\% \text{ Cl} = \frac{71 \times \frac{5}{2.5} \times 10^{-3}}{0.5} \times 100 = 35.5\%$$