

Name of the Teacher- Sutapa Chakrabarty

Subject: Chemistry

Class: Semester-2

Paper: C3T: Inorganic Chemistry

Topic: Redox and Precipitation Reaction

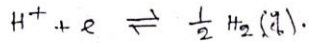
PART 4

Comments: Go through the whole lesson as much as possible for conception. But some topics such **formal potential, application of formal potential, Z-R solution, problems of Latimer diagram** are very important for preparation of examination. Also complete the assignment given below.

[N.B.- Acknowledgement of indebtedness to Mr. Sibshankar Das, my respected Teacher regarding collection of study materials in Inorganic Chemistry]

iv) Potentiometric determination of pH

The fundamental rxn for hydrogen electrode is



Its single electrode potential at 1 atm hydrogen pressure is,

$$E_H = E_H^0 + \frac{RT}{nF} \ln \frac{C_{H^+}}{p_{H_2}^{1/2}}$$

Where E_H^0 = standard potential of the hydrogen electrode = 0, At 1 atm hydrogen pressure $p_{H_2}^{1/2} = 1$.

For dilute solⁿ it can be written as

$$E_H = 2.303 \frac{RT}{F} \log C_{H^+}$$

$$= 0.059 \log C_{H^+} \text{ at } 25^\circ C.$$

Where,

R = molar gas constant

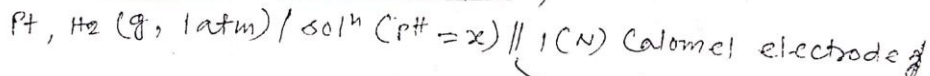
T = absolute temp^r

n = No. of electrons involves = 1

F = Faraday.

$$\therefore E_H = -0.059 \text{ pH} \quad (\because \text{pH} = -\log C_{H^+})$$

This eqⁿ is utilised in determining the pH of a solⁿ. For this, hydrogen electrode is dipped into the solⁿ and combine with a Calomel electrode through a salt bridge of 1(N) KCl solⁿ. The cell set up is,



→ salt bridge of 1(N) KCl solⁿ

The EMF of the cell is measured with a potentiometer. The potential of the 1(N) Calomel electrode = 0.280 volt

$$\text{Hence, } E_{\text{cell}} = E_{\text{cat}} - E_{\text{H}} = 0.280 - (-0.059\text{pH})$$

$$\therefore E_{\text{cell}} = 0.280 + 0.059\text{pH}$$

$$\therefore \text{pH} = \frac{E_{\text{cell}} - 0.280}{0.059}$$

From this equation the pH of the unknown solution can be calculate easily.

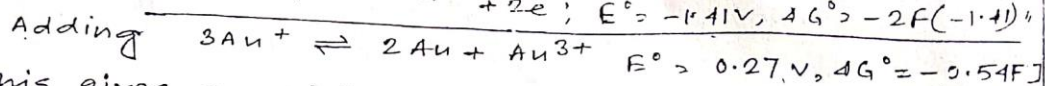
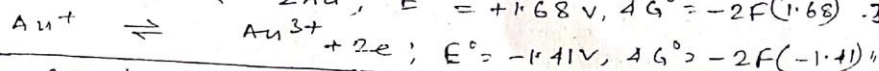
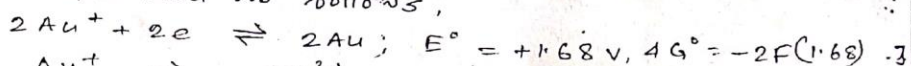
v) Relative Stability of Oxidation states and disproportionation.

The relative stability of different oxidation states of an element may be predicted from the knowledge of standard Potentials.

a) The disproportionation of Au^+ takes place according to the equation -



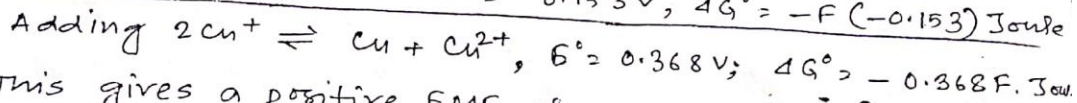
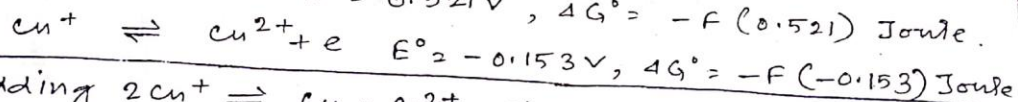
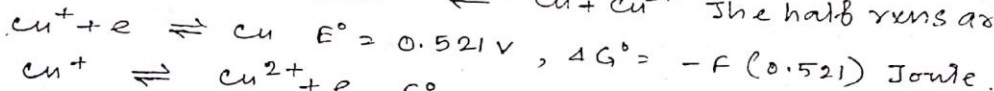
This is explained as follows,



This gives a positive of $E^\circ = 0.27\text{V}$ for the total rxn. and the free energy change of the process is negative ($-0.54F\text{J}$). Hence Au^+ disproportionates to Au^0 and Au^{3+} in solution.

The potentials in the series $\text{Au}^{3+} \xrightarrow{+1.41\text{V}} \text{Au}^+ \xrightarrow{+1.68\text{V}} \text{Au}^0$ increases from left to right, the species in the middle (Au^+) in such a case is unstable and disproportionates.

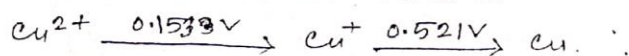
b) For the rxn $2\text{Cu}^+ \rightleftharpoons \text{Cu} + \text{Cu}^{2+}$ The half rxns are



This gives a positive EMF of 0.368V for the overall cell rxn and the free energy change of the process is negative ($-0.368F$).

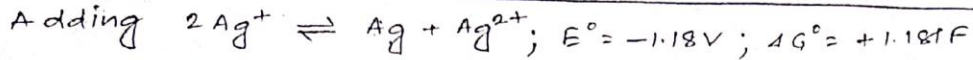
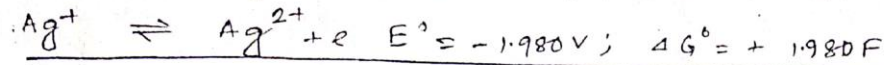
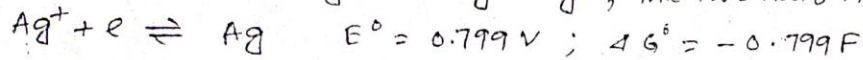
Hence the disproportionation of Cu^+ to Cu^0 and Cu^{2+} takes place spontaneously.

The potential in the series --



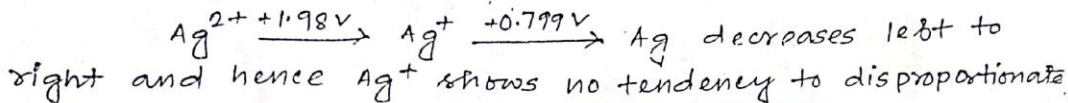
increases from left to right, this makes the species Cu^+ in the middle is unstable and hence it disproportionate.

1) For the rxn $2Ag^+ \rightleftharpoons Ag + Ag^{2+}$, The two half rxns are-



This gives an overall cell potential of $E^\circ = -1.181V$ and consequently Ag^+ will not disproportionate. The free energy change in the overall process is also positive and hence the rxn cannot take place spontaneously.

The potentials in the series -



• Formal Potential

In the process $Ox + ne \rightleftharpoons Red$, according to Nernst equation the potential is given by

$$E = E^\circ + \frac{0.059}{n} \log \frac{a_{Ox}}{a_{Red}} \text{ at } 25^\circ C$$

When $a_{Ox} = a_{Red}$, Then $E = E^\circ$

Here it is believed that all the involved species will remain in their simple forms and the activity equals to concentration. Such a situation may only exist in a very dilute condition. In practice, the solution is sufficiently concentrated to disallow the condition activity = concentration. Besides this, the ions may not exist in simple forms in the presence of ~~the~~ ^{other} electrolytes. Under these conditions, the activity and concentration widely differ. In such cases, the concept of standard potential is of no practical use and it requires another parameter called formal potential which takes care the actual situation existing in solution.

The formal potential of a system is the experimentally observed potential of the system in solution where the formal concentration (formality) of the oxidised and a reduced forms are the same. The formal potential covers the effects of the conc. of the species, pH of the solution, the possibility of complexation and other factors.

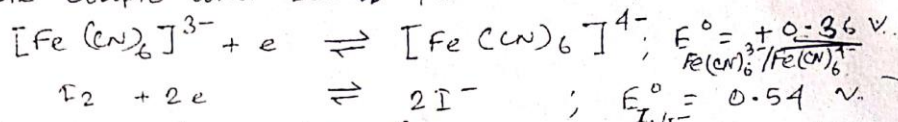
The modified form of the Nernst eqⁿ is given by

$$E = E^\circ + \frac{0.059}{n} \log \frac{[Ox]}{[Red]} \text{ at } 25^\circ C.$$

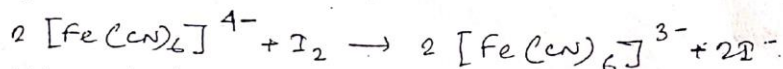
where E° is the formal potential.

• The Utility of Formal Potential.

The standard potentials of Ferricyanide / Ferrocyanide couple and Iodine / Iodide couple are;



Since the iodine / Iodide couple has a greater +ve potential than $[\text{Fe}(\text{CN})_6]^{3-} / [\text{Fe}(\text{CN})_6]^{4-}$ couple, the I_2 is a better oxidant than $[\text{Fe}(\text{CN})_6]^{3-}$ ion. Therefore, it would be expected that I_2 would oxidise $[\text{Fe}(\text{CN})_6]^{4-}$ to $[\text{Fe}(\text{CN})_6]^{3-}$ ion.



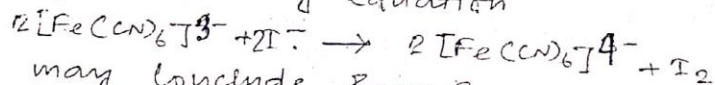
In reality, it is the $[\text{Fe}(\text{CN})_6]^{3-}$ ion that oxidises I^- to free I_2 in acid medium. It is explained in terms of formal potential.

The expression for Redox potential of $[\text{Fe}(\text{CN})_6]^{3-} / [\text{Fe}(\text{CN})_6]^{4-}$ couple can be written as according to Nernst equation as follows;

$$E = E^\circ + \frac{0.059}{1} \log \frac{[\text{Fe}(\text{CN})_6]^{3-}}{[\text{Fe}(\text{CN})_6]^{4-}} \text{ at } 25^\circ\text{C.}$$

$$= 0.36 + 0.059 \log \frac{[\text{Fe}(\text{CN})_6]^{3-}}{[\text{Fe}(\text{CN})_6]^{4-}}$$

In the solⁿ of low pH (1M H⁺) $[\text{Fe}(\text{CN})_6]^{4-}$ and $[\text{Fe}(\text{CN})_6]^{3-}$ ions combined with one or more H⁺ ions to form hydroferricyanic acid and hydro ferrocyanic acid. $[\text{Fe}(\text{CN})_6]^{4-}$ ion is a stronger base than $[\text{Fe}(\text{CN})_6]^{3-}$ ion and hence the reactivity of $[\text{Fe}(\text{CN})_6]^{4-}$ ions with H⁺ ions is more than the reactivity of $[\text{Fe}(\text{CN})_6]^{3-}$ ion with H⁺ ion. Thus the concentration of $[\text{Fe}(\text{CN})_6]^{4-}$ is appreciably diminished in acid medium. Hence the potential is raised from +0.36 V (in neutral solⁿ) to +0.77 volt in 1M HCl, H₂SO₄ or HClO₄. The formal potential is now above that of I_2 / I^- couple. Hence the oxidation of I^- to I_2 by $[\text{Fe}(\text{CN})_6]^{3-}$ ions take place by the following equation

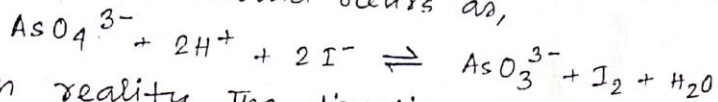


Hence we may conclude formal potential command over normal standard potential.

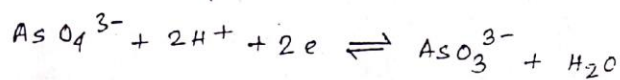
Influence of different factors on Redox potential.

i) Influence of pH of the medium.

a) The standard reduction potential of $\text{AsO}_4^{3-}/\text{AsO}_3^{3-}$ and I_2/I^- systems are 0.56V and 0.54V respectively. Hence the rxn should occur as,



But in reality the direction of the reaction depends on the pH of the solⁿ. The potential of the $\text{AsO}_4^{3-}/\text{AsO}_3^{3-}$ couple depends on pH.

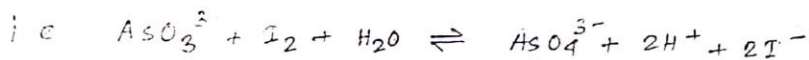


Therefore,

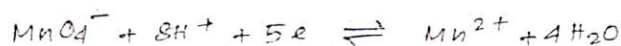
$$\begin{aligned} E &= E^\circ + \frac{0.059}{2} \log \frac{[\text{AsO}_4^{3-}][\text{H}^+]^2}{[\text{AsO}_3^{3-}]} \\ &= E^\circ + \frac{0.059}{2} \left[\log [\text{H}^+]^2 + \log \frac{[\text{AsO}_4^{3-}]}{[\text{AsO}_3^{3-}]} \right] \\ &= E^\circ - 0.059 \text{pH} + \frac{0.059}{2} \log \frac{[\text{AsO}_4^{3-}]}{[\text{AsO}_3^{3-}]} \\ &= E^{\circ'} + \frac{0.059}{2} \log \frac{[\text{AsO}_4^{3-}]}{[\text{AsO}_3^{3-}]} \end{aligned}$$

Here $E^{\circ'}$ = formal potential.

Thus the formal potential ($E^{\circ'}$) of the system decreases with increasing of pH. In presence of excess NaHCO_3 solⁿ (pH \approx 8) $E^{\circ'}$ becomes 0.088V. Which is less than 0.54V (standard potential of I_2/I^- system) under the comparable condition. Thus at pH \approx 8, AsO_4^{3-} can not oxidise I^- to iodine (I_2) but I_2 can oxidise AsO_3^{3-} to AsO_4^{3-} .



b) The redox potential of $\text{MnO}_4^-/\text{Mn}^{2+}$ couple largely depends on the pH of the solⁿ.



$$\begin{aligned} E &= E^\circ + \frac{0.059}{5} \log \frac{[\text{MnO}_4^-][\text{H}^+]^8}{[\text{Mn}^{2+}]} \\ &= E^\circ + \frac{0.059}{5} \log [\text{H}^+]^8 + \frac{0.059}{5} \log \frac{[\text{MnO}_4^-]}{[\text{Mn}^{2+}]} \end{aligned}$$

$$= 1.51 - 0.024\text{pH} + \frac{0.059}{5} \log \frac{[\text{MnO}_4^-]}{[\text{Mn}^{2+}]}$$

$$= E^\circ' + \frac{0.059}{5} \log \frac{[\text{MnO}_4^-]}{[\text{Mn}^{2+}]}$$

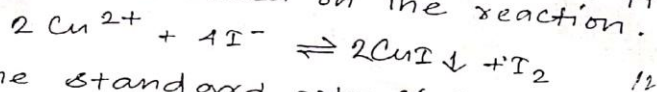
Here E° = standard potential

E°' = formal potential

Though the standard potential (E°) at $\text{pH} = 0$ is 1.51V, At $\text{pH} = 6$, $E^\circ' = 0.95\text{V}$, hence among the halides only I^- can be oxidised, $E^\circ_{\text{I}_2/\text{I}^-} = 0.54\text{V}$, $E^\circ_{\text{Br}_2/\text{Br}^-} = 1.07\text{V}$, $E^\circ_{\text{Cl}_2/\text{Cl}^-} = 1.36\text{V}$. At $\text{pH} = 3$, $E^\circ' = 1.23\text{V}$ and Br^- can be oxidised but Cl^- cannot be oxidised. Thus to oxidise Cl^- a much higher concentration of H^+ is required ($\text{pH} = 1.5$).

ii) Influence of precipitation on Redox potential.

a) The volumetric estimation of copper by an iodometric method is based on the reaction.



But the standard potential of the systems $\text{Cu}^{2+}/\text{Cu}^+$ ($E^\circ = 0.15\text{V}$) and I_2/I^- ($E^\circ = 0.54\text{V}$) suggest that reverse rxn should take place. Now let us consider the effect of precipitation of CuI on the $\text{Cu}^{2+}/\text{Cu}^+$ system, $\text{Cu}^{2+} + e \rightleftharpoons \text{Cu}^+$,

$$E = E^\circ + 0.059 \log \frac{[\text{Cu}^{2+}]}{[\text{Cu}^+]}$$

In presence of I^- , Cu^+ is precipitated as CuI , having the low solubility product 10^{-12} i.e. $\text{Cu}^+ + \text{I}^- \rightleftharpoons \text{CuI} \downarrow$.

$$K_{sp} = [\text{Cu}^+][\text{I}^-]$$

$$= 10^{-12}$$

$$\text{Hence } [\text{Cu}^+] = \frac{10^{-12}}{[\text{I}^-]}$$

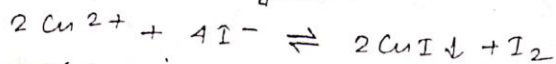
$$\text{Hence } E = 0.15 + 0.059 \log \frac{[\text{Cu}^{2+}][\text{I}^-]}{10^{-12}}$$

$$= 0.15 + 12 \times 0.059 + 0.059 \log [\text{Cu}^{2+}][\text{I}^-]$$

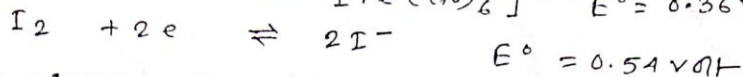
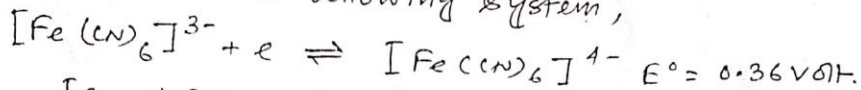
$$= 0.86 + 0.059 \log [\text{Cu}^{2+}][\text{I}^-]$$

Thus for the system $\text{Cu}^{2+} + \text{I}^- + e \rightleftharpoons \text{CuI} \downarrow$; we get the standard potential 0.86V. Now we see, though the $\text{Cu}^{2+}/\text{Cu}^+$ couple cannot oxidise I^- to I_2 ($E^\circ_{\text{I}_2/\text{I}^-} =$

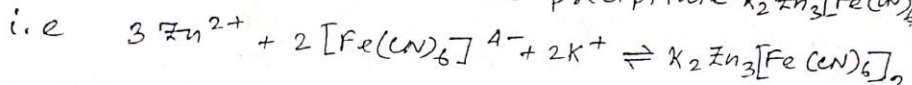
0.54V). But in the presence of I^- , the oxidation becomes possible and in iodometric estimation of copper, the following rxn is utilised.



Let us consider the following system,



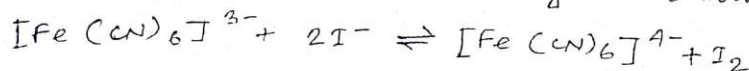
From the above values it is expected that $[Fe(CN)_6]^{3-}$ should not be able to oxidise I^- to I_2 . But in presence of Zn^{2+} ion $[Fe(CN)_6]^{4-}$ gets removed due to the formation of an insoluble precipitate $K_2Zn_3[Fe(CN)_6]_2$



Reduction potential of $[Fe(CN)_6]^{3-}/[Fe(CN)_6]^{4-}$ expressed as

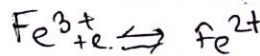
$$E = 0.36 + 0.059 \log \frac{[Fe(CN)_6]^{3-}}{[Fe(CN)_6]^{4-}}$$

Due to the removal of $[Fe(CN)_6]^{4-}$ through this precipitation the formal potential of the system $[Fe(CN)_6]^{3-}/[Fe(CN)_6]^{4-}$ increases and it exceeds that of I_2/I^- systems and it can liberate iodine by the following rxn.



iii) Influence of Complex formation.

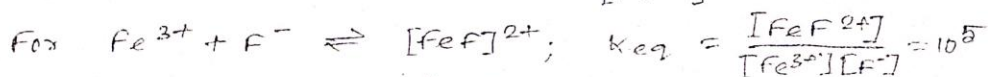
Fe^{3+} ion normally oxidises I^- ion to I_2 in aqueous solution, since the standard potential values of the two couples are $E^\circ_{Fe^{3+}/Fe^{2+}} = 0.77 \text{ volt}$ and $E^\circ_{I_2/I^-} = 0.54 \text{ V}$. But in presence of excess F^- ion, Fe^{3+} is removed as the complex ion $[FeF]^{2+}$ and the ratio $\frac{[Fe^{3+}]}{[Fe^{2+}]}$ increases.



In the system $Fe^{3+} + e \rightleftharpoons Fe^{2+}$

$$E = E^\circ + \frac{0.059}{1} \log \frac{[Fe^{3+}]}{[Fe^{2+}]}$$

$$= 0.77 + 0.059 \log \frac{[Fe^{3+}]}{[Fe^{2+}]} \quad \text{--- (1)}$$



$$\therefore [Fe^{3+}] = \frac{[FeF^{2+}]}{10^5 \cdot [F^-]}$$

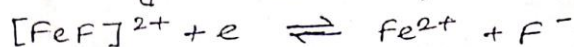
Substituting the value of $[Fe^{3+}]$ in equation (1) we get,

$$E = 0.77 + 0.059 \log \frac{[FeF^{2+}]}{10^5 [Fe^{2+}] [F^-]}$$

$$= 0.77 + 0.059 \log 10^{-5} + 0.059 \log \frac{[FeF^{2+}]}{[Fe^{2+}] [F^-]}$$

$$= 0.475 + 0.059 \log \frac{[FeF^{2+}]}{[Fe^{2+}] [F^-]}$$

Therefore the system becomes,



As the reduction potential decreased, the oxidising power of Fe^{3+} decreases and it can no more oxidise iodide ion. On the other hand $Cu(II)$ does not form any such stable complex with F^- ion. Thus Cu may be estimated selectively in a iron-copper mixture by titrating the iodine liberated from KI in presence of F^- ion.

• The E° value for Co^{3+}/Co^{2+} couple is 1.82 volt at 298K. Calculate the E° value for $[Co(CN)_6]^{3-}/[Co(CN)_6]^{4-}$ couple. Given stability constant for the species $[Co(CN)_6]^{3-}$ and $[Co(CN)_6]^{4-}$ are 10^{64} and 10^{20} respectively.

In the system $Co^{3+} + e \rightleftharpoons Co^{2+}$; $E^\circ = 1.82$ volt
According to Nernst equation we can write

$$E = 1.82 + 0.059 \log \frac{[Co^{3+}]}{[Co^{2+}]} \quad \text{--- (1)}$$

On addition of CN^- ion both Co^{3+} and Co^{2+} form complexes.

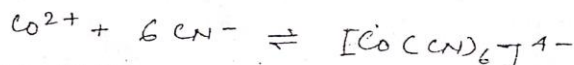


Given the stability constant of $[Co(CN)_6]^{3-}$ ion (K_1)

$$\text{Hence } K_1 = \frac{[Co(CN)_6]^{3-}}{[Co^{3+}] [CN^-]^6} = 10^{64}$$

$$\text{or, } [Co^{3+}] = \frac{[Co(CN)_6]^{3-}}{10^{64} \cdot [CN^-]^6}$$

Similarly



Given the stability constant of $[Co(CN)_6]^{4-}$ ion (K_2) = 10^{20} .

Hence

$$K_2 = \frac{[\text{Co}(\text{CN})_6^{4-}]}{[\text{Co}^{2+}][\text{CN}^-]^6} = 10^{20}$$

$$\text{or, } [\text{Co}^{2+}] = \frac{[\text{Co}(\text{CN})_6^{4-}]}{10^{20} [\text{CN}^-]^6}$$

Putting the values of $[\text{Co}^{3+}]$ and $[\text{Co}^{2+}]$ in equation-
①, we get,

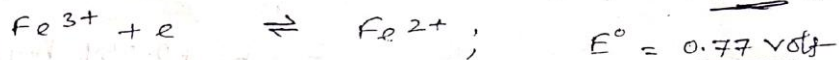
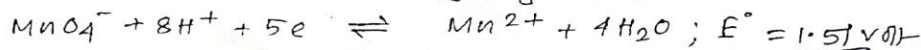
$$\begin{aligned} E &= 1.82 + 0.059 \log \frac{[\text{Co}(\text{CN})_6^{3-}] \times 10^{20} \times [\text{CN}^-]^6}{[\text{Co}(\text{CN})_6^{4-}] \times 10^{64} \times [\text{CN}^-]^6} \\ &= -0.776 + 0.059 \log \frac{[\text{Co}(\text{CN})_6^{3-}]}{[\text{Co}(\text{CN})_6^{4-}]} \end{aligned}$$

Therefore the E° value for $\text{Co}(\text{CN})_6^{3-}/\text{Co}(\text{CN})_6^{4-}$ couple is -0.776 volt.

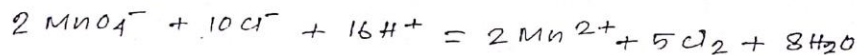
• Zimmermann - Reinhardt (Z-R) solution.

KMnO_4 is a stronger oxidant and is used in the estimation of Fe^{2+} ion.

Let us consider a following systems;



If the acidity of the solⁿ is maintained by the use of H_2SO_4 , no complicating side rxns occur. On the contrary if HCl is used complecations arise due to oxidation of Cl^- ion to Cl_2 .



Since for the system $\text{Cl}_2 + 2e \rightleftharpoons 2\text{Cl}^-$; $E^\circ = 1.36$ volt
A part of the MnO_4^- solution is used for the formation of Cl_2 from Cl^- ion. In order to cancel this side rxn in HCl medium a solution is used, called Z-R solution. It is added to the Fe^{2+} solution before titration. It contains MnSO_4 , H_2SO_4 and H_3PO_4 .

* Preparation of Z-R solution

Z-R solⁿ is prepared by dissolving 50 gm of pure $\text{MnSO}_4 \cdot 4\text{H}_2\text{O}$ in 250 ml distilled water, adding it to a cold mixture of 300 ml of water and 100 ml of con. H_2SO_4 followed by addition of 100 ml of syrupy H_3PO_4 .

• Function of Z-R solution.

* Role of $MnSO_4$

The redox potential of MnO_4^- / Mn^{2+} couple is given by

$$E = E^\circ + \frac{0.059}{5} \log \frac{[MnO_4^-][H^+]^8}{[Mn^{2+}]}$$

$$= 1.51 + \frac{0.059}{5} \log [H^+]^8 + \frac{0.059}{5} \log \frac{[MnO_4^-]}{[Mn^{2+}]}$$

Under standard conditions, the concentrations of both MnO_4^- and Mn^{2+} are 1M, their ratio being 1. On the addition of Z-R solution, there is also much Mn^{2+} present in the Fe^{2+} solution. So that the ratio $[MnO_4^-] / [Mn^{2+}]$ is considerably lower than one. The redox potential will be lower than the E° value of 1.51 volt and the formal potential becomes less than 1.36 volt. Thus $KMnO_4$ acts as a weaker oxidant and thereby tendency of the side rxn i.e. oxidation of Cl^- to Cl_2 is prevented.

* Functions of H_2SO_4

H_2SO_4 is used mainly to control the H^+ ion concentration i.e. pH of the solution. (about 2N at which the titration is carried out).

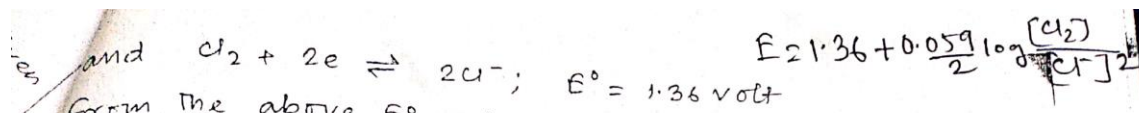
* Functions of H_3PO_4

$$E = E^\circ + 0.059 \log \frac{[Fe^{3+}]}{[Fe^{2+}]}$$

In presence of H_3PO_4 , Fe^{3+} complexes as colourless $[Fe(HPO_4)]^+$ ion, effectively lowering the concentration of free Fe^{3+} in solution. Hence redox potential of Fe^{3+} / Fe^{2+} couple is also reduced so that Fe^{2+} turns out to be a stronger reducing agent than it appears from its E° value i.e. oxidation of Fe^{2+} becomes thermodynamically facilitated. Due to formation of colourless soluble complex, sharp detection of the end point becomes possible.

• From the standard E° values of the following couples it is found that MnO_2 should not oxidise HCl to Cl_2 . But in fact the rxn is brought about in laboratory. Given $E^\circ_{MnO_2/Mn^{2+}} = 1.23V$, $E^\circ_{Cl_2/Cl^-} = 1.36V$.

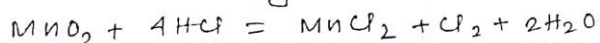
For the system $MnO_2 + 4H^+ + 2e \rightleftharpoons Mn^{2+} + 2H_2O$; $E^\circ = 1.23V$



From the above E° values of the couples it is clear that Cl_2 should oxidise Mn^{2+} to MnO_2 . But in fact reverse is true i.e. MnO_2 oxidises Cl^- ion to Cl_2 . This is explained as follows.

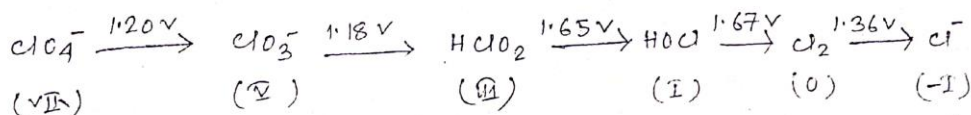
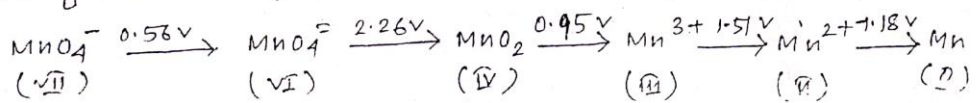
The used HCl is heavily concentrated and its molarity is beyond 1M . So the standard potential of the couple $\text{MnO}_2/\text{Mn}^{2+}$ will tend to increase beyond 1.23 volt and exceeds 1.36 volt which is the standard potential for Cl_2/Cl^- couple. If we reduce the Cl_2 gas pressure below 1 atm , by keeping the rxn mixture boiling, the E° value of the Cl_2/Cl^- couple will fall below 1.36 volt .

Thus MnO_2 can successfully oxidise HCl to Cl_2 according to the following rxn.



* Latimer Diagram.

Latimer diagram is a compact form of the different species of an element in varying oxidation states. The different species have arrows in between them and the standard potential involving two species are written over the arrows. The oxidation states of the element concerned is specified under each species. The species with highest oxidation state is placed at the extreme left and the lowest is placed at the extreme right. Two examples are given below;



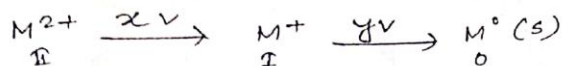
* Applications

The application of Latimer Diagram are given below;

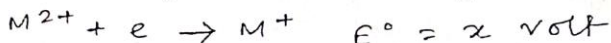
a) It is possible to predict whether a species will disproportionate from an inspection of the Latimer

diagram covers the redox species. If the potential to the right of the species is higher than the potential on the left, then the species will undergo disproportionation.

Let us consider the following case



The disproportionation rxn will be as,



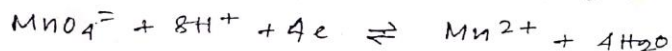
If we subtract the left hand side potential from the right hand side potential of the Latimer diagram we get the disproportionation rxn, whose standard potential is given by -

$$E^\circ = (y - x)V$$

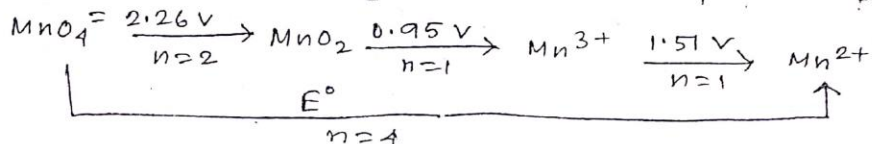
when $(y - x)$ is positive, E° is positive and ΔG° is -ve and so the rxn is spontaneous i.e. the species will undergo disproportionation.

b) By using the Latimer diagram, the E° value of any step can be calculated. The basic principle is that ΔG° is a thermodynamic state function. Hence the total change in the Gibbs free energy does not depend on the path through which the process is carried out and it depends only on the final and initial states.

Let us calculate the potential for MnO_4^- / Mn^{2+} couple



Thus the process may occur in two possible pathways



As ΔG° for both the pathways are same hence we can write

$$-4FE^\circ = -F(2 \times 2.26 + 0.95 + 1.51)$$

$$\Rightarrow E^\circ = \frac{(2 \times 2.26 + 1 \times 0.95 + 1 \times 1.51)}{2+1+1}$$

$$= 1.745 \text{ volt}$$

Solve the problems:

1. What is Z-R solution? What is the function of Z-R solution?

2. Find out the standard electrode potential for $\text{MnO}_4^- / \text{MnO}_2$ redox couple in acid medium. Given $E^\circ_{\text{MnO}_4^- / \text{Mn}^{2+}} = 1.51 \text{ V}$

$$E^\circ_{\text{MnO}_2 / \text{Mn}^{2+}} = 1.23 \text{ V}$$

at 298K

3. What is Latimer Diagram?

4. Why Au^+ disproportionates to Au^0 and Au^{3+} in solution?

$$E^\circ_{\text{Au}^{3+} / \text{Au}^+} = 2.141 \text{ V}$$

$$E^\circ_{\text{Au}^+ / \text{Au}^0} = 1.68 \text{ V}$$