REDOX REACTION

Reduction	Oxidation
1) Removal of Oxygen	Removal of Hydrogen.
2) $2MgO \rightarrow Mg + O_2$	
2) Addition of Hydrogen	Addition of oxygen.
3) Gain of e^-	loss of e^-
$Mn^{+7} + 3e^- \rightarrow Mn^{+4}$	$Fe^{+2} \rightarrow Fe^{+3} + e^-$
4) Decrease in oxidation no.	Increase in oxidation no.

→ Na, K, Rb Alkali metal → +1

→ Mg, Be, B Alkali Earth metal → +2

Oxygen = −2 for

exception

1) $(O-O)$ (O_2^{-2}) peroxide = −1 $\qquad$ $Fe_2O = +2$

$O_2^- \rightarrow$ superoxide $\rightarrow -\frac{1}{2}$ $\qquad$ $Fe_2O_3 = +1$

$O_3^- \rightarrow -\frac{1}{3}$

2) Hydrogen → metal − −1 $\quad$ NaH

→ Nonmetal − +1 $\quad$ HCl

3) self-bonding

fix = 0 $\quad$ ⇒ O_2, H_2, N_2, I_2, S_8, O_3, P_4

Na, Cu, Mg, Fe = 0

* Compounds oxidation no.

$SO_4 \rightarrow -2$ $\qquad$ $NH_4 = +1$

$NO_3 \rightarrow -1$ $\qquad$ $SO_3 = -2$

$PO_4 \rightarrow -3$

To solve the linear equation :

$$x + 2 = -3$$
$$x = -3 - 2$$
$$x = -5$$

Que) 1) MgO
$+2 + x = 0$
$x = -2$

2) $KMnO_4$
$+1 + x - 8 = 0$
$x = +7$

3) $K_2Cr_2O_7$
$+1 \times 2 + 2x + (-2 \times 7) = 0$
$x = 6$

Redox Reaction :- A reaction in which oxidation & reduction takes place simultaneously.

$$\underset{+2}{CuSO_4} + \underset{+2}{Zn} \longrightarrow \underset{+3}{ZnSO_4} + \underset{+2}{Cu}$$

ox

reduction

if decreasing ~~charge~~ oxi. no → Red
if increasing ox no. → oxidation

red

$$\underset{0}{H_2} + \underset{0}{O_2} \longrightarrow \underset{+1\ \ -2}{H_2O}$$

ox

ox

$$\underset{0}{Mg} + \underset{0}{O_2} \longrightarrow \underset{+2\ \ -2}{MgO}$$

red

* AGENTS

1) oxidising agent / oxidents (IN) -
other oxidation self Reduction.

$$\underset{+7}{KMnO_4} + \underset{0}{Fe} \longrightarrow \underset{+4}{MnO_4} + \underset{+2}{FeO}$$

oxidising agent.

compounds which are present there in higher oxidising state are HNO_3, H_2SO_4, $KMnO_4$, $K_2Cr_2O_7$, Cr_2O_3.

ii) Reducing Agent - Reductance

→ self get oxidised & other Reduction.

eg) H_2S, KI etc.

1) $O_2 \longrightarrow O^{-2}$ → oxi. agent

$\underbrace{\quad\quad\quad\quad}_{\text{Red}}$

2) $F_2 \longrightarrow F^-$ → oxi. agent

$\underbrace{\quad\quad\quad\quad}_{\text{Reduction}}$

3) $Li^0 \longrightarrow Li^{0+}$ → redu. ag.

$\underbrace{\quad\quad\quad\quad}_{\text{ox}}$

4) $K \longrightarrow K^+$ → redu. agent.

$\underbrace{\quad\quad\quad\quad}_{\text{ox}}$

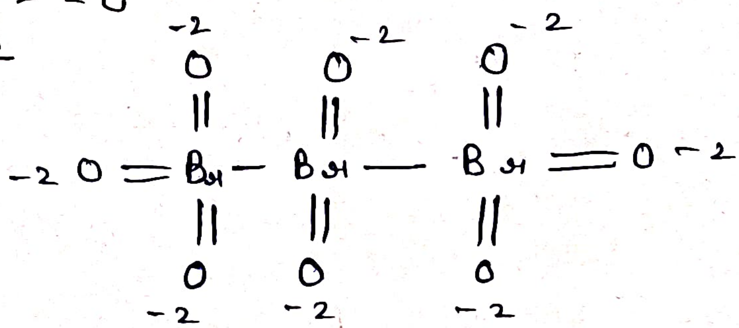
* **STOCK NOTATION** - compounds having Roman number in their formula helps in identifying the oxidising agents.

* **PARADOX OF OXI. NUMBER** -

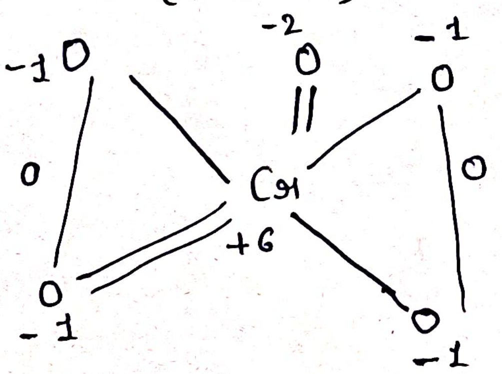
1) $Br_3O_8 \longrightarrow 3x + 8x - 2 = 0$

$x = \frac{16}{3}$

Tri bromo octaoxide.



2) CrO_5 $(x = +10)$

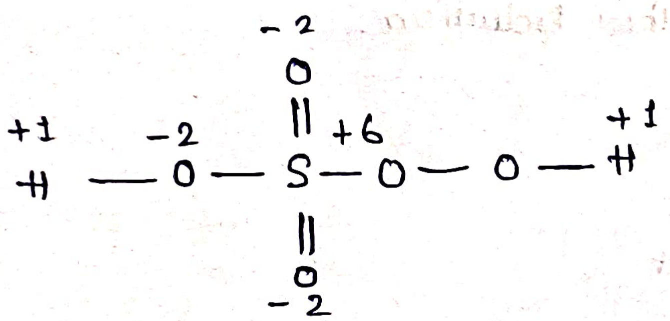


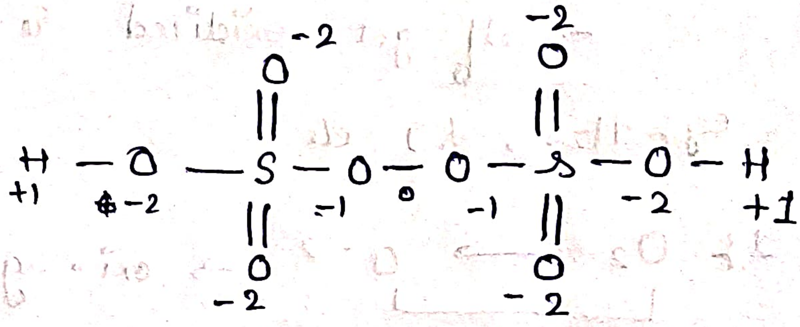
Carbon suboxide

C_3O_2 $\quad c = \frac{4}{3}$

$\overset{-2}{O} = \overset{+2}{C} = \underset{0}{C} = \underset{0}{C} = \overset{+2}{C} = \overset{-2}{O}$

3) H_2SO_5 – peroxo sulphuric acid. 4) $H_2S_2O_8$





* Types of Redox Reaction :-

1) Combination Redox Reaction.

1) red

$$\underset{0}{H_2} + \underset{0}{O_2} \longrightarrow \underset{+1}{H_2}\underset{-2}{O}$$

ox

2) red

$$\underset{0}{Ca} + \underset{0}{O_2} \longrightarrow \underset{+2}{Ca}\underset{-1}{O}$$

ox

3) red

$$\underset{0}{Mg} + \underset{0}{O_2} \longrightarrow \underset{+2}{Mg}\underset{-1}{O}$$

ox

4) red

$$\underset{-4}{CH_4} + \underset{0}{O_2} \longrightarrow \underset{+4}{CO_2} + 2H_2\underset{-2}{O}$$

ox

Redox

- Combination
- Decomposition
- Displacement
  - Metal Dis.
  - Non-metal Dis.
- Disproportionation

2) Decomposition Redox Reaction.

No Redox

$$\underset{+2}{Mg}\underset{+6}{S}\underset{-2}{O_4} \longrightarrow \underset{+2}{Mg}\underset{-2}{O} + \underset{+6}{SO_3}$$

No Redox

$$\underset{+2}{Ca}\underset{+4}{C}\underset{-2}{O_3} \xrightarrow{\Delta} \underset{+2}{Ca}\underset{-2}{O} + \underset{+4}{C}\underset{-2}{O_2}$$

Red

$$\underset{+1}{K}\underset{+5}{Cl}\underset{-2}{O_3} \longrightarrow \underset{+1}{K}\underset{-1}{Cl} + \underset{0}{O_2}$$

ox

red

$$\underset{+2}{Mg}\underset{-2}{O} \xrightarrow{\text{electrolysis}} \underset{0}{Mg} + \underset{0}{O_2}$$

oxidation.

3) Displacement Reaction.

Metal displacement.

ox

$$\underset{+2}{CuSO_4} + \underset{0}{Zn} \longrightarrow \underset{+2}{ZnSO_4} + \underset{0}{Cu}$$

red.

Non-metal disp.

$$\underset{+1}{H_2SO_4} + \underset{0}{Zn} \longrightarrow \underset{+2}{ZnSO_4} + \underset{0}{H_2(g)}$$

mainly Hydrogen gas evolved in Non-metal displacement redox.

4) Disproportionation Redox Rxn

same compound → oxidation & Reduction both.

→ min → 3 oxidation number

→ not in max / min ox. no. state should be Intermediate

} CONDITIONS

eg) ✗ $F \begin{cases} \rightarrow -1 \\ \rightarrow 0 \end{cases}$ ✓ $O \begin{cases} \rightarrow -2 \\ \rightarrow -1 \\ \rightarrow 0 \end{cases}$ ✓ $N \begin{cases} \rightarrow -3 \\ \rightarrow +3 \\ \rightarrow 0 \end{cases}$

eg) $KMnO_4 \longrightarrow$ max $= +7$ can't show.

$H_2S \longrightarrow -2$ ✗

$K_2MnO_4 \longrightarrow +6$ ✓

+7

$$\underset{+6}{MnO_4^{-2}} \longrightarrow \underset{+7}{MnO_4^-} + \underset{+4}{MnO_2}$$

red. ($+6 \to +4$)

ox ($+6 \to +7$)

* BALANCING :-

Oxidation Number Bal. Method

Ion-electron balancing method or Half Reaction method.

Que) $\underset{+6\ \ -2}{Cr_2O_7^{-2}} + \underset{+4\ \ -2}{SO_3^{-2}} \longrightarrow \underset{+3}{Cr^{+3}} + \underset{+6}{SO_4^{-2}}$

ox ($+4 \to +6$)

red ($+6 \to +3$)

Step - I) Write the skeletal Reaction & oxidation no. of all entities.

Step - II) Write half Reaction of Both.

1) oxidation half Reaction.

$$3\left(\underset{+4}{SO_3^{-2}} \rightarrow \underset{+6}{SO_4^{-2}}\right)$$ increased by +2

2) Reduction half Reaction.

$$\underset{+6}{Cr_2O_7^{-2}} \longrightarrow \underset{+3}{2Cr^{+3}}$$ decreased by 6

Step - III → Balance entities which is in the half Rxn and also inc / dec there oxidation no.

Step - IV → Equal the difference in oxi. no.

$$3SO_3^{-2} \longrightarrow 3SO_4^{-2}$$

$$Cr_2O_7^{-2} \longrightarrow 2Cr^{+3}$$

$$3SO_3^{-2} + Cr_2O_7^{-2} \rightarrow 3SO_4^{-2} + 2Cr^{+3}$$

Step - V → write the balanced eq^n.

Step - VI → Balance no. of oxy by adding H_2O in oppo. side of oxygen and then balance Hydrogen by adding H^+. (acidic medium)

$$3SO_3^{-2} + CrO_7^{-2} + 8H^+ \rightarrow 3SO_4^{-2} + 2Cr^{+3} + 4H_2O$$

2) $\underset{+7\ -2}{MnO_4^-} + \underset{-1}{Br^-} \longrightarrow \underset{+4}{MnO_2} + \underset{+5}{BrO_3^-}$

Step I. ox. half Rxn

$$\underset{-1}{Br^-} \longrightarrow \underset{+5}{BrO_3^-}$$ ↑ by 6

Red half Rxn

$$2\left(\underset{+7}{MnO_4^-} \rightarrow \underset{+4}{MnO_2}\right)$$ dec by 3

$Br^- \longrightarrow BrO_3^-$

$2MnO_4^- \longrightarrow 2MnO_2$

$Br^- + 2MnO_4^- \longrightarrow BrO_3^- + 2MnO_2$

$1H_2O \qquad\qquad + 2OH^-$

eg 7.10) $\quad MnO_4 + I^- \longrightarrow I_2 + MnO_2$

$\quad +7 \qquad -1 \qquad\quad 0 \qquad +4$

ox Rxn

$(2I^- \longrightarrow I_2 + 2e^-) \times 3$

$\quad -1 \qquad\quad 0$

Red Rxn

$(MnO_4^- + 3e^- \longrightarrow MnO_2) \times 2$

$\quad +7 \qquad\qquad\quad +4$

$6I^- \longrightarrow 3I_2 + \cancel{6e^-}$

$2MnO_4^- + \cancel{6e^-} \longrightarrow 2MnO_2$

$6I^- + 2MnO_4 + 4H_2O \longrightarrow$

$3I_2 + 2MnO_2 + 8OH^-$

3) $Cl_2 \longrightarrow Cl^- + ClO_3^-$

$\quad 0 \qquad -1 \qquad +5$

ox. Rxn

$Cl_2 \longrightarrow 2ClO_3^- + 10e$

Red Rxn

$5 \times (Cl_2 + 2e^- \longrightarrow 2Cl^-)$

$\qquad 0 \qquad\qquad -1$

$Cl_2 \longrightarrow 2ClO_3^- + \cancel{10e^-}$

$5Cl_2 + \cancel{10e^-} \longrightarrow 10Cl^-$

$6Cl_2 \longrightarrow 2ClO_3^- + 10Cl^-$

$3Cl_2 \longrightarrow ClO_3^- + 5Cl + 3H_2O$

$+ 6OH^-$

4) $NO_2^- + Al \longrightarrow NH_3 + Al^{+3}$

$\quad +3 \qquad 0 \qquad -3 \qquad +3$

$(Al \longrightarrow Al^{+3} + 3e^-) \times 2$

$NO_2^- + 6e^- \longrightarrow NH_3$

$2Al + NO_2^- + 7H^+ \longrightarrow 2Al^{+3} +$

$NH_3 + 3H_2O$

6) $MnO_4^- + S^{-2} \longrightarrow MnO_2 + S$

$\quad +7 \qquad -2 \qquad +4 \quad 0$

$(S^{-2} \longrightarrow S + 2e^-) \times 3$

$(MnO_4^- + 3e^- \longrightarrow MnO_2) \times 4$

$\quad +7 \qquad\qquad +4$

$3S^{-2} + 2MnO_4^- \longrightarrow 3S + 2MnO_2$

$+ 4H_2O \qquad\qquad + 8OH^-$

5) $K_2Cr_2O_7 + HCl \longrightarrow KCl + CrCl_3 + Cl_2 + H_2O$

$\quad (2HCl \longrightarrow Cl_2 + 2e^-) \times 3$

$(K_2Cr_2O_7 + 6e^- \longrightarrow 2CrCl_3)$

$14HCl + K_2Cr_2O_7 \longrightarrow 3Cl_2 + 2CrCl_3 + 2KCl + 7H_2O$

7) $Zn + NO_3^- + H^+ \longrightarrow Zn^{+2} + NH_4^+ + H_2O$

$\quad 0 \quad\; +5 \qquad\qquad\qquad +2 \qquad\quad -3$

$(Zn \rightarrow Zn^{+2} + 2e^-) \times 4$

$NO_3^- + 8e^- \longrightarrow NH_4^+$

$4Zn + NO_3^- + 10H^+ \longrightarrow 4Zn^{+2} + NH_4^+ + 3H_2O$ //

8) $I^- + ClO^- \longrightarrow I_2 + Cl^-$

$2I^- \rightarrow I_2 + 2e^-$

$ClO^- + 2e^- \longrightarrow Cl^-$

$2I^- + ClO^- + H_2O \longrightarrow I_2 + Cl^- + 2OH^-$

9) $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$

$\quad 0 \qquad +5 \qquad\quad +2 \qquad\quad +2$

$(Cu \rightarrow Cu(NO_3)_2 + 2e^-)^{\times 3}$

$(HNO_3 + 3e^- \longrightarrow NO) \times 2$

$3Cu + 8HNO_3 \longrightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

10) $C_2O_4^{-2} + MnO_4^- \longrightarrow CO_2 + Mn^{+2}$

$\quad +3 \qquad\quad +7 \qquad\qquad +4 \qquad +2$

$(C_2O_4^{-2} \longrightarrow CO_2 + 2e^-) \times 5$

$\quad +3 \qquad\qquad\quad +4$

$(MnO_4^- + 5e^- \longrightarrow Mn^{+2}) \times 2$

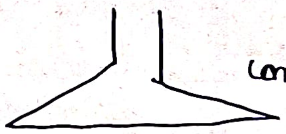
$\quad +7 \qquad\qquad\qquad +2$

$5C_2O_4^{-2} + 2MnO_4^- + 16H^+ \longrightarrow 10CO_2 + 2Mn^{+2} + 8H_2O$

* TITRATION → Redox Rxn.

Indicators	PH		Acidic colour	Basic colour
1) Self Indicator		→	Blue	Colourless
~~Eg~~ $K_2Cr_2O_7$ / $KMnO_4$		→	Pink Colour	
2) Congo Red	6.8 – 8.0	→	Yellow	Red
3) Methyl Blue	3.2 – 4.4	→	Yellow	Blue
4) Phenolphthalein	8.5 – 10	→	Colourless	Pink.

Titration



conical flask

Acidic titration
Basic Titration $\rightarrow$ $KMnO_4$ + oxalic acid $(COOH)_2$

$$\begin{matrix} COOH \\ | \\ COOH \end{matrix}$$

$\rightarrow$ Ferrous Ammonium sulphate
$FeSO_4(NH_4)_2SO_4$

$\rightarrow$ from 12th, $KMnO_4$ / $K_2Cr_2O_7$

~~K_2O~~ $K_2Cr_2O_7 + H^+ \rightarrow Cr^{+3}$

Equivalent weight $\rightarrow \dfrac{Mw}{\text{no. of } e^- \text{ transf}} = \dfrac{W_{K_2Cr_2O_7}}{3} = \dfrac{294}{3} = 98g/mol$

$Fe \longrightarrow Fe^{+3}$

eq. wt. $= \dfrac{56}{3}$

$2I^- \longrightarrow I_2$

eq. wt. $= \dfrac{127}{1}$

$\underset{-1}{I^-} \longrightarrow \underset{+5}{IO_3^-}$

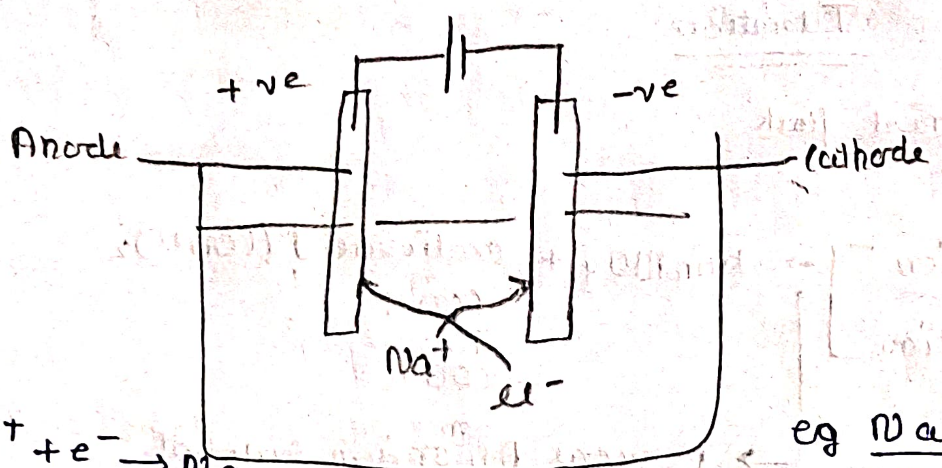
eq. wt of $I = \dfrac{127}{6}$

$\underset{+7}{KMnO_4}$ $\xrightarrow[\text{medium}]{\text{Acidic } H^+}$ $\underset{+2}{Mn^{+2}}$ $\rightarrow$ $\dfrac{158}{5}$

Basic $\rightarrow$ $\underset{+6}{MnO_4^{-2}}$ $\quad \dfrac{158}{1}$

Neutral $\rightarrow$ $\underset{+4}{MnO_2}$ $\quad \dfrac{158}{3}$

* Electrolysis :- P.T.O



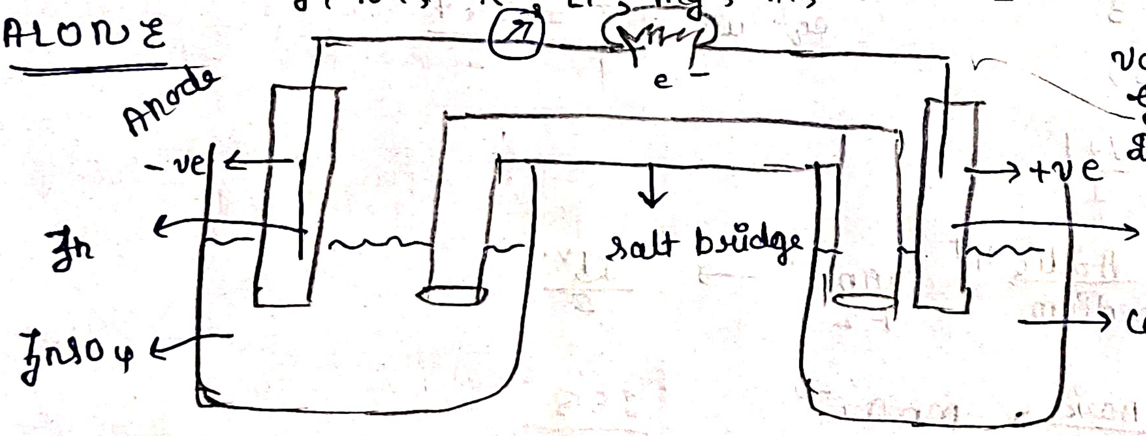
Red $Na^+ + e^- \rightarrow Na$
or
Ox $Cl^- \rightarrow Cl_2 + 2e^-$

eg NaCl electrolysis

* Uses of Electrolysis

1) Electrolysis
2) Electroplating. → inferior metal → superior metal
3) Purification of metal
4) Highly Reactive Elements extracting.
eg) Na, K, Li, Mg, Al, F_2, Cl_2

ALONE



Volt-C-cell / Galvanic Daniell Cell

$Zn \rightarrow Zn^{+2} + 2e^-$ $\qquad$ $Cu^{+2} + 2e^- \rightarrow Cu$

Electro Chemical cell

Salt Bridge → circuit Comple → KCl, KNO_3, NH_4Cl added
sometimes agar agar gel.

Case-I → Discharging → bulb → glow

Volt → 1.1 volt

e^- flow → Zn → Cu ~~HH~~

Cu → Zn

$Zn \rightarrow ZnSO_4$

$CuSO_4 \rightarrow Cu$

Electrochemical cell.

Case-II → when we Remove salt bridge

Ecell = 0 Circuit Break.

Case- III → After some time No $CuSO_4$ Rest in vessels.

$E_{cell} = 0$ $CuSO_4 \rightarrow Cu$

→ no ions are present for Reduction.

Case- IV → external voltage greater than 1.1 volt.

Electrolytic cell } e- flow → $Cu \rightarrow Zn$ $Cu \rightarrow CuSO_4$

I flow → $Zn \rightarrow Cu$ $ZnSO_4 \rightarrow Zn$.

Case- V Ext volt = 1.1 volt

No charging
No discharge } cell Not work.

Other All Rxn of Daniell cell

R.H.R $CuSO_4 \rightarrow Cu$

L.H.R $Zn \rightarrow ZnSO_4$

$CuSO_4 + Zn \rightarrow Cu + ZnSO_4$.

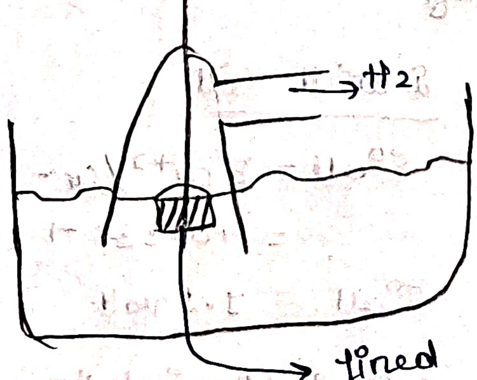
* CELL REPRESENTATION :-

$$\underbrace{Zn \mid ZnSO_4\,(1M)}_{ox} \parallel \underbrace{CuSO_4\,(0.5M) \mid Cu}_{red}$$

* SHE → Stand Hydr. Electric potential

$2H^+ + 2e^- \rightarrow 2H_2$

Calomel cell
Hg



1bar pressure

298K temp.

fined divided particle of Pt.

* Standard Reduction Potential.

1) Top of the series are good Red. Agent & they are high electropositive.

2) Down of the series are good oxi. Agent & they are high electronegative.

3) ऊपर वाले नीचे वाले को Replace करते है

Feasible → possible to Do
संभव

$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

$Cu + AgCl \rightarrow CuCl_2 + Ag$

→ Best Reducing Anglent → value more −ve Right

$Mg^{+2}/Mg \rightarrow -2.74$

$Cu^{+2}/Cu \rightarrow +0.34$

→ Bes oxidising agent → value more +ve left

$Mg^{+2}/Mg \rightarrow -2.74$

$Cu^{+2}/Cu \rightarrow +0.34$

* $E^{\circ}_{cell} = E_{cat} - E_{an.}$

$= E_{red} - E_{ox}$

$= E_{right} - E_{left}$

Ox || Red

$Zn \rightarrow Zn^{+2}$ $Cu^{+2} \rightarrow Cu$

Daniell Cell

$E^{\circ}_{cell} = E^{\circ}_{Cu^{+2}/Cu} - E^{\circ}_{Zn^{+2}/Zn}$

$= +0.337 + 0.763$

$E^{\circ}_{cell} = 1.1$ volt

$E^{\circ}_{cell} = ? = +ve \rightarrow$ cell → Chemical Energy → Electrical Energy

Ag | Cu → cell = ?

$Cu | Cu^{+2} || Ag^{+} / Ag$

$E^\circ cell = E^\circ_{Ag^+/Ag} - E^\circ_{Cu^{+2}/Cu}$

$= 0.80 - 0.34$

$= +0.46\ V$ possible //

Li° Zn — not feasible

$Zn/Zn^{+2} \parallel Li^+/Li^\circ$ → we can change our electrode

$E^\circ cell = E^\circ_{Li^+/Li^\circ} - E^\circ_{Zn^{+2}/Zn}$

$= -3.0 + 0.34$

$= -2.66$ Volt

$E^\circ cell = E^\circ_{Zn^{+2}/Zn} - E^\circ_{Li^+/Li^\circ}$

$= +0.34 + 3.0$

$E^\circ cell = +3.34$

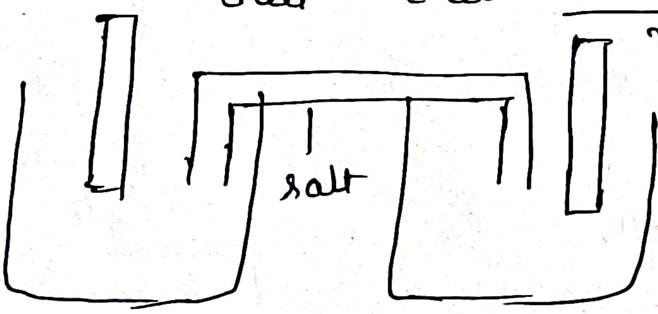
* NERNST Eq^n

Emf of cell → depends on con^n.

$$E_{cell} = E^\circ_{cell} - \frac{RT}{nF} \ln \frac{P}{R}$$

$$E_{cell} = E^\circ_{cell} - \frac{2.303\,T}{nF} \log_{10} \frac{P}{R}$$

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{n} \log_{10} \frac{P}{R}$$



$Zn \rightarrow Zn^{+2}$ Cu^{+2}/Cu

$Zn \rightarrow Zn^{+2} + 2e^-$

$Cu^{+2} + 2e^- \rightarrow Cu$

$Zn + Cu^{+2} \rightarrow Zn^{+2} + Cu$

$$E_{cell} = E^\circ_{cell} - \frac{0.0591}{n} \log_{10} \frac{P}{R}$$

$$E_{cell} = 1.1\,volt - \frac{0.0591}{2} \log_{10} \frac{[Zn^{+2}]}{[Cu^{+2}]}$$

$$\frac{[Zn^{+2}] \downarrow \text{low}}{[Cu^{+2}] \uparrow}$$

$[Cu^{+2}] \rightarrow$ high

other Rxn to Nernst eq^n.

$$\Delta G^\circ = -nF\,E^\circ_{cell}$$

$$\Delta G = -nF\,E_{cell}$$

at equilibrium

$$\Delta G = 0 = -nF\left(E^\circ_{cell} - \frac{RT}{nF}\ln K_c\right)$$

$$= -nF\,E^\circ_{cell} - \frac{RT \times \cancel{nF}}{\cancel{nF}}\ln K_c$$

~~at~~

$$-nF\,E^\circ_{cell} = \Delta G^\circ$$

$$= -nF\,E^\circ_{cell} + RT\ln K_c$$

$$nF\,E^\circ_{cell} = RT\ln K_c$$

$$E^\circ_{cell} = \frac{2.303\,RT}{nF}\log_{10} K_c$$

$$E^\circ_{cell} = \frac{0.0591}{n}\log_{10} K_c$$

$$\Delta G^\circ = 2.303\,RT\log_{10} K_c.$$