Chapter

Electrochemistry

Interconversion of electrical & chemical energy.

Terminologies:

⊛ electrolyte:

- strong
  - * all strong acids & basis
  - may be both ionic or covalent
- Weak
  - * all weak acids & bases
  - may be both ionic or covalent
- Non-electrolyte
  - Covalent compounds.
  - C_6H_6
  - $C_{12}H_{12}O_{11}$

⊛ Conductivity:

electrolytic Conductor	Metallic Conductor
* I due to ions	* Current due to free e^-
* Mass transfer occur	* Mass not transfer
* Energy transfer occur.	* Energy transfer occur.
* Physical + chemical change occur	* only physical change occur.
* $E \cdot C \propto T$	* No chemical change.
* Conduct current in molten or solution form	* $M \cdot C \propto \frac{1}{T}$
	* Conduct current in solid or molten

⊛ Factors affecting Ionic Conductivity

* Ion in Graseous state:

Conductivity ∝ charge

Cond ∝ 1/size

$Na^{+} < Mg^{+2} < Al^{+3}$ (Good Cond in G. state)

* in Solution state:

Same charge

$Li^{+1} < Na^{+1} < K^{+1} < Rb^{+1} < Cs^{+1}$

* Small size
* attract more H_2O molecules

hydration sphere

Cond ∝ 1/charge

Conduc ∝ size

$Na^{+}_{(aq)} > Mg^{+2}_{(aq)} > Al^{+3}_{(aq)}$

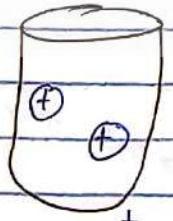
* Concentration:

Conduct ∝ 1/concent

Conduc ∝ dilut

5M NaCl < 1M NaCl

5 moles in $1dm^3$



dilute ⇓ ions separate ⇑ not combine

conce ⇓ ion sepon ⇓ → combine

ions separate from each other.

dielectric const ⇑

* Conduc ∝ T
* Condu ∝ polarity of solve
* Conduc ∝ amount of solvent

⊛ Transport Number:

condu ki fraction cation, anion

* Friction of Total current carried by cation / anion.

T.N of cation ← $t_+ = \frac{I_+}{I_+ + I_-}$ $t_- = \frac{I_-}{I_+ + I_-}$

* Transpor No# of Anion > cation (size ⇓ more molecules is attracted)

T.N ∝ dilution | T.N ∝ 1/conce

T.N ∝ amount of solvent | T.N → depend on Nature of ions

T.N ∝ 1/common ion

Q# Cl^- transport max current in:

(a) KCl (aqu) ⇓ common ion present, Backward chance ⇑

(b) HNO_3 (aqu)

Note

* Accurate T.N = 0.5
* T.N of cation < 0.5 → T ⇑ T.N ⇓
* T.N of anion > 0.5 → T ⇑ T.N ⇑

* Temperature $\propto$ t.N (cation)

* Temperature $\propto \frac{1}{t.N}$ (anion)

(#) New Lec

- **Oxidation State:**

* apparent or real charge in an atom or molecule.
* +ve, -ve, zero, fraction
* variable * depend on Compound.
* Max O.S = +9 (Irridium)
* Min O.S = -4 (Carbon)
* Show how much es is gain or loss by an atom.

Rules:

* element:

O.S = 0 $P_4 = 0$ $Na = 0$ $S_8 = 0$

* O.S b/w Same atom:

O.S = 0 $P_4 = 0$ $Cl_2 = 0$ $C_6H_{12}O_6$ ⇓ B/W Carbon = 0

* Atomic ions:

O.S = charge

Na^{+1} => O.S = +1

O^{-2} => O.S = -2

* Polyatomic Ion:

O.S = charge

* NH_4^{+1} O.S = +1
* CO_3^{-2} O.S = -2

* Neutral Molecule:

O.S = 0

H_2O O.S = 0
NH_3 O.S = 0

* Atom O.S in a Molecule:

(by algebric Sum)

HNO_3 => N = +5

* Maximum O.S:

Max O.S = valence es

. S => Max = +6 . N => Max = +5
. C => Max = +4 . Mn => Max = +7

Oxygen => Max O.S = +2 → exception

* Min O.S:

define for Non-metal

Min O.S = unpaired es → n = upaired es

Min = n - 8

O => Min = 6 - 8 = (-2)

Cl => Min = 7 - 8 = (-1)

C => Min = 4 - 8 = (-4)

N → Max = +5
N → Min = -3
difference b/w its => 8

* Max & Min O.S difference = 08

Cl → Max = +7
Cl → Min = -1
↓ not show -2, -3, -4 ...

HNO_4 → O.S of N = +7
Max for N = +5

Peroxide linkage

— O — O —

$\Rightarrow O_2^{-2}$

when

O.S > valence e⁻

(peroxide linkage present)

HNO_4 ⟶ N = +7

HNO_2 ⟶ O_2^{-2} ⟶ Remove

$x - 4 - 2 + 1 = 0$

$x = +5$

H_2SO_5 S = +8

H_2SO_3 ⟶ O_2 (-2)

This charge will be considered

$+2 + x - 6 - 2$

$x = +6$

6 v.e⁻ ← Cr_2O_6 ⟶ Cr = ?

Cr_2O_8 ⟶ O_2^{-2}

Cr = 10

Cr = 9

Cr_2O_6 ⟶ O_2^{-2}, O_2^{-2}

Cr = +6

Cr_2O_4 ⟶ O_2^{-2}, O_2^{-2}, O_2^{-2}

$-8 - 6$

$x = +7$

Cr_2O_7 ⟶ $O_2^{-2}, O_2^{-2}, O_2^{-2}, O_2^{-2}$

$2x - 4 - 8 = 0$

$x = \frac{12}{2}$ $\boxed{x = +6}$

Note:

$CaOCl_2$ ⟶ $\boxed{Cl = +1}$ $\boxed{Cl = -1}$

$Ca(OCl)Cl$

$OCl^- =$

$x - 2 = -1$

$\boxed{x = +1}$

Fe_3O_4 ⟶ $\boxed{Fe = +2, +3}$

$^{+2}_{FeO}$ $^{+6\ -6}_{Fe_2O_3}$

give ionic compound clue

NH_4NO_3 ⟶ $\boxed{N = -3, +5}$

NH_4^{+1} NO_3^{-1}

$\boxed{x = -3}$ $\boxed{x = +5}$

Oxidation	Reduction
* Loss of es	* Gain of es
* Loss of hydrogen	* Gain of hydrogen
* Gain of Oxygen	* Loss of Oxygen.

$\overset{0}{Na} + \overset{0}{Cl_2} \longrightarrow \overset{+1\ -1}{NaCl}$

* Na = 0 → +1 (oxidized)
* Cl = 0 → -1 (Reduced)

⊛ Redox Rxn

* Both oxidation & Reduction occur.

⊛ Non-Redox Rxn:

* Acid & Base Rxn
* Neutralization
* double displacement rxn
* Precipitation rxn.

$\overset{+1\ -1}{AgNO_3} + \overset{+\ -1}{NaCl} \longrightarrow \downarrow \overset{+1\ -1}{AgCl} + \overset{+\ -1}{NaNO_3}$ (PPt rxn) or (displa rxn)

⊛ Disproportionation Rxn:

* Same element → oxidized
* Same element → Reduced

e.g. Canizaro's rxn.

⊛ Transfer of e⁻s

$\overset{+7\ -1}{MnO_4^-} \longrightarrow \overset{+4\ -4}{MnO_2}$

Rxn occurs by how many e⁻s transfer?

$\Delta O.S = (+7) - (+4) = 3$

Change in O.S

No # es⁻ transfer = initial O.S − Final O.S

Q#

$Cr_2O_3 \rightleftharpoons Cr_2O_2$

+2 (Large) → +4 (small)

$\Delta O.S = +4 - (+2)$

$\Delta O.S = 2\ es^-$

Reduction potential ↓

ability of gain of es⁻.

(+) Oxidizing agent

* es⁻ acceptor
* also called reductant

=> Oxidizing strength ∝ E.N

=> O.S ∝ 1/Size

=> O.S ∝ +ve charge

=> O.S → variable

=> O.S ∝ 1/(-ve charge)

=> O.S ∝ Reduction potential

Reducing agent

* es⁻ donor.
* also called oxidant

=> Reducing stre ∝ size

=> R.S ∝ -ve

=> R.S ∝ 1/(+ve)

=> R.S ∝ O.Potential

=> $Na^+ > Mg^{+2} > Al^{+3}$

=> Li < Na < K < Rb < Cs

Q#

$3H_2 + N_2 \longrightarrow 2NH_3$

0 0 → -3 → +1

* N = O. agent
* H = R. agent

Imp MCQs point

how we will find that given Compound is oxidizing or Reducing

① oxidation = valence es

then Compound only be Oxidizing agent

e.g

* S = +6 only O.A

* $KMnO_4$ → Mn = +7 (O.A)

* CO_2 → C = +4 (only O.A)

* H_2SO_4 → S = +6 (only O.A)

* $K_2Cr_2O_7$ → Cr = +6 (only O.A)

② oxidation = unpaired es

only (Reducing agent) → compound be in min O.S

* H_2O → O = 2 = 2 → 2 unpaired es → only R.A

* H_2S → 2 = 2 → R.A

* HCl → 1 = 1 → R.A

③ Not min nor Max

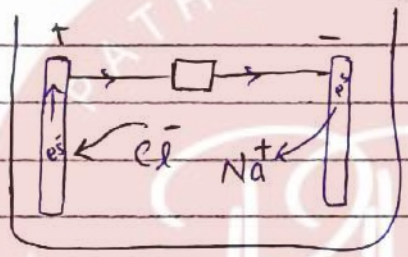
both O.A & R.A

* SO_3 → O.A

* SO_2 → both O.A & R.A.

Electrodes

* These are metallic Rods which Conduct Current.
* Exist in Solid, liquid or Gas
* if Liquid / Gas or = → deposit on solid.



Anode:

* oxidation Occur
* may be +ve (electrolytic cell)
  -ve (Galvanic cell)
* Current (es) enter solution.

⇓

only electrolytic Case

Cathode:

* Reduction Occur
* may be +ve (electrolytic cell)
  -ve (Galvanic cell)
* Current leave solution

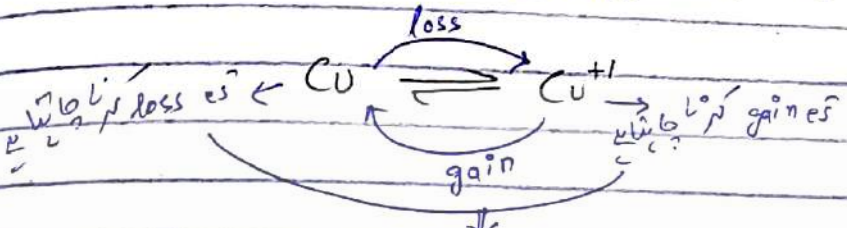
⇓

only electrolytic Case

when element is deposit in its ion

⊛ Electrode Potential

The P.D b/w element & its ions.



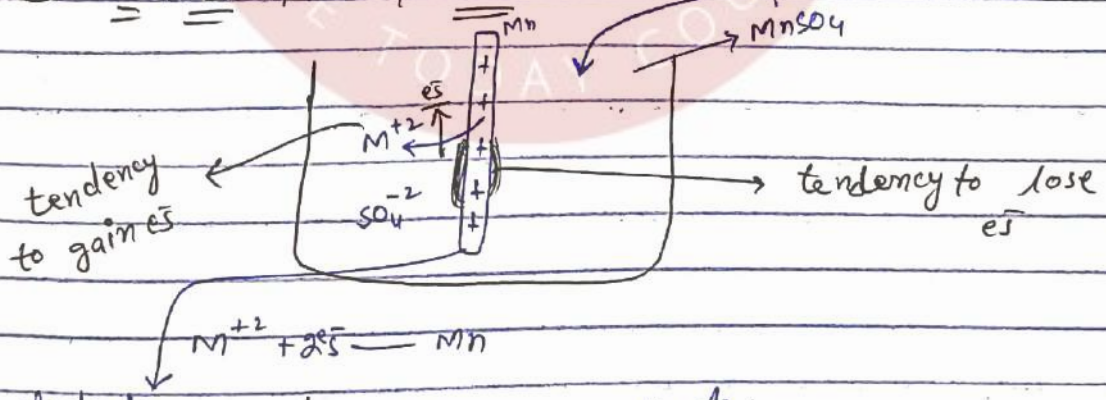
ان دونوں کے درمیان produce ہوتا ہے P.D

⊛ The P.D create b/w electrode & its own ions.

Possibility

* Gain es̄
* Loss es̄
* Not gain or loss es̄ → Inert electrode.

⊛ Reduction Potential:



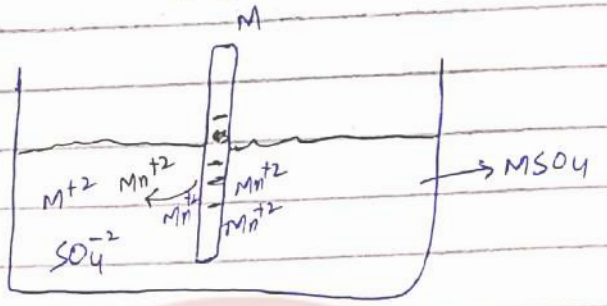
electrode

* acts as a cathode.
* Size of electrode ⇑
* double electric layer.
* in solution => -ve charge ⇑
* so Reducing strength of solut ⇑

New Lec

Oxidation Potential

$M_{(s)} \longrightarrow M^{+2}_{(aq)} + 2e^-$



* acts as anode (-ve)
* -ve charge on anode
* +ve ⇑ in solution
* solution 0 strength ⇑
* double electric layer
* -ve charge in center
* size or weight of electrode ⇓

Factors :

Electrode potential depend upon

* Temperature
* pressure
* Concentration
* Nature of electrode

ability of e^- gaining ✓

* cathode = Reduction potential ⇑ value acts as O.A
* Anode = oxidation potential ⇑ value

Q#

A $E_{red} = 0.8$ (anode)

B $E_{red} = 0.5$ (Cathode)

Q#

A $E_{red} = -0.5$ (cathode)

B $E_{red} = -0.8$ (Anode)

Q#

A $E_{oxid} = 0.5$ - cathode

B $E_{oxid} = 1.1$ Anode

Q#

A $E_{oxid} = -1.1$ cathode

B $E_{oxid} = -0.5$ Anode

Press → 1 atm
Tem → 25°C
conce → 1M

Standard Electrode Potential

The P.D at Nat * Temperature = 25°C * P = 1 atm * Conc = 1M

(#) Standard Hydrogen electrode

* Gas ion electrode } S. name
* Reference electrode

* Stand oxid pote = 0 volt
* Stand oxid pote = 0 volt

Actual potential ≠ 0

* PH of SHE = Zero

Construction:

T = 25°C , P = 1 atm (H_2 gas) , Conc = 1M HCl

$\log [H^+]$ $1M\ HCl \longrightarrow H^+ + Cl^-$

$1M \quad 1M$

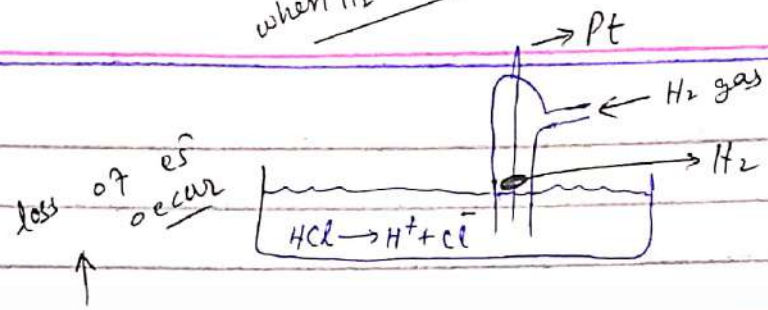
* $[H^+] = 1M$

gas ↑ H_2 deposit over it.

Large S.A for rxn. ← * Pt or Pd is used as inert electrode.

Catalyst.

H_2 gas produced when H_2 act as Cathode



* Anode half equation :

H_2 (at electrode present) $\longrightarrow 2H^+ + 2e^-$ (comes to soluti)

* Cathodic half rxn:

$2H^+ + 2e^- \longrightarrow H_2$ (toward electrode)
(Solut)
bubbles produced.

(*) Measurement :

Reduction potential:

S.H.E = 0 (Anode)

other electrode = +ve (cathode)
(Red)

bubbles produced ← S.H.E = 0 (cathode)

other = -ve (Anode
(Red)

Oxidation potential:

bubbles ← S.H.E = 0 (cathode)

other = +ve (Anode)

S.H.E = 0 (Anode)

other = -ve (cathode)

(*) metal & ion → electrode potential.

(#) Electrochemical Series

Arrangement of elements on the basis of standard Reduction potential

Application:

→ cell potential → P.D b/w cathode & Anode

① E_{cell}:

$$E_{cell} = E_{cath} - E_{anode}$$

② Feasibility of Rxn:

* E_{cell} = +ve (Rxn is feasible)

* E_{cell} = -ve (Rxn is not feasible)

Q#

A → Ered = 0.2 (anode)

B → Ered = 0.5 (cathode)

$E_{cell} = 0.5 - 0.2$

$E_{cell} = 0.3$ → Rxn is feasible

Q#

A → Ered = -0.5 (Anode)

B → Ered = -0.2 (cathode)

$E_{cell} = (-0.2) - (0.5)$

$E_{cell} = +0.3$ → Rxn feasible

Q# A $E_{oxid} = 2.1 \longrightarrow$ Cathode

B $E_{oxid} = 3.1 \longrightarrow$ Anode

Rxn not feasible

$E_{cell} = (2.1) - (3.1) = \boxed{-1.1}$

Q# $E_{oxid} = -2.1 \longrightarrow$ Anode

$E_{oxid} = -3.1 \longrightarrow$ Cathode

$E_{cell} = -3.1 + 2.1$

$\boxed{E_{cell} = -1.1} \longrightarrow$ Rxn not feasible

③ Displacement Rxn:

$\boxed{\text{displace} \propto \text{o.potential}}$

$\boxed{\text{displace} \propto \frac{1}{\text{R.potential}}}$

Q# $Cu + ZnSO_4 \longrightarrow$ No Rxn

oxid of Zn > Cu

Q# $Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$ (Rxn feasible)

Q# $A + CB \longrightarrow \boxed{AC} + B$

A $E_{red} = 2.1$ B $E_{red} = 3.1$

can displace A / B

Q# $A + XY \longrightarrow AY + X$

o.p ↑↑ than B ← A $E_{oxid} = -2.1$

B $E_{oxid} = -3.1$

④ Reactivity:

Reactivity of metal	$\propto$	$\frac{O.P}{R.P}$

Reactivity of Non-metal	$\propto$	$\frac{R.P}{O.P}$

Balancing Redox Rxn

(imp)

① Half Rxn Method: or ion electron Method

* For redox Rxn.
* in Solution form (Rxn)

Steps:

- write Redox Rxn.
- Split the equation → Anodic, cathodic
- Balance atoms
  - ① other atoms
  - ② Balance oxygen.
  - ③ Balance Hydrogen.
- Medium
  - ↳ Acidic

acidic ல் இது பயன்படும்

⇓

oxygen is balance by adding H_2O

Hydrogen → → → → H^+

Charge → → → → e^-

Basic Medium

* Same method as acidic
* add OH^- on each side
* No # OH^- added = No # H^+

Neutral Medium:

* Add H^+ & H_2O on either side.

* Add both equation

Q#:

$$\overset{+6}{Cr_2O_7^{-2}} + \overset{0}{I_2} \longrightarrow \overset{+3}{Cr^{+3}} + \overset{+5\ -1}{IO_3}$$

How ~~many~~ much H_2O should be added to balance the reduction Rxn.

(Reduction Rxn)

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

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

$$Cr_2O_7^{-2} \longrightarrow 2Cr^{+3} + 7H_2O$$

How much H^+ is needed to balance above rxn?

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

Q# Consider a rxn

$$\overset{+3}{Cr(OH)_3} + H_2O_2 \xrightarrow{NaOH} \overset{+6}{CrO_4^{-3}} + H_2O$$

in the half oxidation Rxn.

* How much H_2O is added (1 H_2O add)
* How much H^+ is needed (5 H^+ add)

3 e⁻

* ~~How many~~ e⁻ is added to balance Rxn.
* How many H^+ is neutralized in the Rxn. → $5H^+$

H ⟶ balance by H^+

* How many moles of H_2O is formed → (4)

$Cr(OH)_3 \xrightarrow{NaOH} CrO_4^{-2}$

$\Rightarrow Cr(OH)_3 + H_2O \longrightarrow CrO_4^{-2}$

$\Rightarrow Cr(OH)_3 + H_2O \longrightarrow CrO_4^{-2} + 5H^+$

$+5-2 = +3$

$\Rightarrow Cr(OH)_3 + H_2O \longrightarrow CrO_4^{-2} + 5H^+ + 3e^-$

$\Rightarrow 5OH^- + Cr(OH)_3 + H_2O \rightarrow CrO_4^{-2} + 5H^+ + 3e^- + 5OH^-$

$\Rightarrow 5OH^- + Cr(OH)_3 + H_2O \longrightarrow CrO_4^{-2} + 5H_2O$

$4H_2O$

(#) Oxidation Number Method

* write equation
* Identify O.Agent
* mention O.state
* change in O.S
* equalize e^- gain or lost.
* By multiplying with suitable number.
* Balance atom
* other atom → Groups → NO_3^{-1} by HNO_3 → element already present in rxn.
  → SO_4^{-2} by H_2SO_4
* Combine other atoms.

Q#

$$\overset{0}{Zn} + \overset{+5}{HNO_3} \longrightarrow \overset{+2}{Zn(NO_3)_2} + \overset{+2}{NO} + \overset{x}{H_2O}$$

a) How many es⁻ is lost in the Rxn.

b) How many HNO_3 is added in the Rxn.

$Zn \longrightarrow Zn(NO_3)_2 + 2es^-$] x3 =>

$HNO_3 + 3es^- \longrightarrow NO$] x2 =>

No# es⁻ gain & lost = 06

$3Zn \longrightarrow 3Zn(NO_3)_2 + 6es^-$

$2HNO_3 + 6es^- \longrightarrow 2NO$

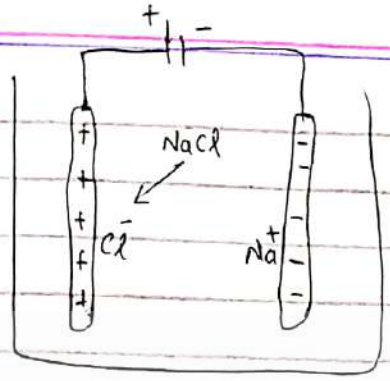
$3Zn + 2HNO_3 + 6es^- \longrightarrow 3Zn(NO_3)_2 + 2NO + 6es^-$

$6NO_3$ ← 06 HNO_3

So 06 HNO_3 will be added

New Lec

Electrolysis



The decomposition of a compound with the help of electricity.

* occurs in electrolytic cell.
* Endothermic
* Non- spontaneous.
* Discharging of ions occur (ions ↑→ stable) convert
* ΔG_f (Gibbs free energy) = +ve

imp xxxxx

Product of Electrolysis

* Nature of electrolyte
  - NaCl in H_2O
  - $MgCl_2$ in H_2O
* Preferssional discharge.

Cl^-
OH^-

→ That ion will be discharge which require less energy for discharge.

Cation Order

$Ag^{+1} > Cu^{+2} > H^+ > Pb^{+2} > Sn^{+2} > Fe^{+2} > Zn^{+2} > Al^{+3} > Ca^{+2} > Na^{+1} > K^{+1}$

Anionic Order:

$$OH^- > I^- > Br^- > Cl^- > NO_3^{-1} > SO_4^{-2} > F^{-1}$$

(3) State:

Molten NaCl (only NaCl)

$$NaCl \longrightarrow Na^+ + Cl$$

Loss of es ← Anode

$$2Cl^- \longrightarrow Cl_2 + 2es^-$$

Cathode

$$2Na^+ + 2es^- \longrightarrow 2Na$$

Aqueous NaCl

$$NaCl \longrightarrow Na^+ + Cl^-$$
$$H_2O \longrightarrow H^+ + OH^-$$

Anode: ✓ ✓

OH^- & Cl^-

product

$$4OH^- \longrightarrow O_2 + H_2O$$

Cathode: (Na^+ & H^+)

$$2H^+ + 2es^- \longrightarrow H_2$$

(4) Amount of solvent:

Conce	diluted
electrolyte ⇈	electrolyte ⇊
solvent ⇊	Solvent ⇈

when Anion are present in large access

Conce = dil

Cation:

when cation in excess.

H^+ Zn^{+2} → large

* Follow prefressional order.

Anion: when anion in excess.

* Not follow prefressional order.

Q#

(*) Concentrated $NaCl_{(aq)}$

⇓

electrolyte

$NaCl$ = More

water = less

Anode: $(Cl^- + OH^-)$ ↗ Max

$2Cl^- \longrightarrow Cl_2 + 2e^-$

$NaCl \longrightarrow Na^+ + Cl^-$

$H_2O \longrightarrow H^+ + OH^-$

cathode (H_2) Anode

Cathode

$H^+ + 2e^- \longrightarrow H_2$ ✓

(*) Dilute aqe NaCl

$NaCl$ = less

H_2O = More

Anode: $(Cl^- \& OH^-)$

$OH^- \longrightarrow O_2$

Cathode: $H^+ \& Na^+$

↓

prefessional order $H^+ \longrightarrow H_2$

⑤ Nature of electrode:

- inert electrode → Pt, Pd, graphite → Not take part in the rxn.
- active electrode → Cu, Zn → Take part in the rxn.

Laws of Electrolysis

① Faraday's 1st Law: → Amount of substance deposit or liberate at an electrode is directly proportional to the quantity of electricity

* $Al^{+3} + 1e\bar{s} \longrightarrow$ Less Al
* $Al^{+3} + 2e\bar{s} \longrightarrow$ 0
* $Al^{+3} + 3e\bar{s} \longrightarrow$ more Al

⇓

amount of substance deposit or loss on electrode due to current pass.

mass ✗ ← * $w \propto$ Current passed

* $w \propto Q$ * $w \propto I$
* $w \propto T$ (Time) * $w \propto It$

$$w = ZIt$$

Z = electrochemical equivalent. ← chemical equivalent

$$Z = \frac{\text{Equivalent weight (e)}}{F}$$

charge / Farad

$$Z = \frac{e}{F}$$

amount of substance liberated / deposit when one mole / 1F or 1C charge of current is passed from a substance

Chemical equivalent:

23g Na^+ + 1F ⟶ Na = 23g

24g Mg^{+2} + 1F ⟶ Mg^{+1} = 12g

For atom / Ion:

$$e = \frac{\text{Molar mass}}{\text{valency}}$$

* For Al = $\frac{27\,g}{3}$ = 9g

when one F current is passed from Al then 9g Al is deposit.

* For Mg^{+2}/Mg

$e = \frac{24}{2} \Rightarrow e = 12g$

For Acid:

$$e = \frac{\text{Molar mass}}{\text{No\# ionizable Hydrogen}}$$

H_2SO_4 $e = \frac{98}{2} \Rightarrow e = 49g$

For Base:

$$e = \frac{\text{Molar mass}}{\text{No\# ionizable OH}}$$

NaOH $e = \frac{40}{1}$ $e = 40g$

For Salt:

$$e = \frac{\text{Molar mass}}{\text{Total +ve charge / -ve charge}}$$

* $NaCl \longrightarrow \frac{58.5}{1} = 58.5\,g$

* $Mg^{+2}SO_4^{-2} \longrightarrow \frac{120}{2} = 60\,g$

(*) For Oxidizing & Reducing agents

$$e = \frac{\text{Molar mass}}{\text{change in O.S}}$$

$MnO_4^{-1} \longrightarrow MnO_2$

+7 (3) +4

$e = \frac{25}{3}$ $\quad e = 8.3\,g$

$w = ZIt$ (Z = equivalent weight)

$w = \frac{e}{F}It \Rightarrow w = \frac{M \cdot M}{nF} \times It$ (mass ← e; I in Faraday)

Answer in Calc

$1F = 96500\,C$ (change)

$1C = 6.25 \times 10^{18}\,e^-$

$W = \frac{}{n \times 96500} \times It$ (mass can be calculated)

$n = \frac{}{n \times 96500} \times It$ (w = amount of substance deposit)

$n = \frac{V}{V_m}$ $\quad e =$

$w =$

$V = \frac{It}{96500 \times n} \times V_m$

$w =$

Q# How much mass of Magnesium in $Mg(OH)_2$ is deposit when 2F current is pass for 1 min?

$$\boxed{w = \frac{M \cdot M}{n \times F} \times I \times t} \Rightarrow w = \frac{24}{2 \times F} \times 2 \times 60\ sec$$

$$\boxed{w = 1440\ g}$$

Q# How many moles of Na^{+1} in NaOH is deposit when 2C of current is passed for 0.5 min?

$$n = \frac{2 \times 30}{1 \times 96500} \qquad n = ?$$

Faraday's 2nd Law

* $Na^{+} \longrightarrow 1F = 23g \longrightarrow e = \frac{23}{1} = 23$
* $Mg^{+2} \longrightarrow 1F = 12g \longrightarrow e = 12g$
* $Al^{+3} \longrightarrow 1F = 9g \longrightarrow e = 9g$

When same amount of I is passed from different substance then amount of substance deposit depend on chemical equivalent.

* 1st Law → Same substance
* 2nd Law → different substance

$W \propto e$ $\quad W \propto \frac{M \cdot M}{\text{Valency}}$

$W \propto M \cdot M$ (if Valency = Same)

$Li^+ < Na^+ < K^+ < Rb^+ < Cs^+$

$W \propto \frac{1}{\text{valency}}$ $\quad \frac{W_1}{W_2} = \frac{e_1}{e_2}$

Q# When Same Current is passed from the following which one will deposit more?

a) HCl $\quad$ b) HNO_3 $\quad$ c) H_2SO_4 $\quad$ d) Same

$e = \frac{36.5}{1}$ $\quad e = \frac{63}{1}$

$e = 36.5$ $\quad e = 63g$ $\quad e = 49g$

Q# How many moles of e^- are required to deposit 18g of Al^{+3} in $AlCl_3$:

electrons:

1 mol Al^{+3} = 3F

or 27g of Al^{+3} = 3F

18g : x

x = 2F → 2 moles of e^-

Q# How many e^- are required to deposit 48g of Mg^{+2} ion:

24g Mg^{+2} = 2F

48 : x $\quad 24x = 96$ $\quad x = \frac{96}{24}$

x = 4 mole $\quad$ x = 4NA

⊛ Cell Representation

1) Anode ⟷ Cathode

2) Salt brigde = ||

3) phase boundry = /

4) State ↳ Concentration mention

5) salt product ↳ Common ion ⟶ Ignore

Q# =

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

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

$$\boxed{Zn_{(s)}/Zn^{+2}_{(aq)} \;||\; Cu^{+2}_{(aq)}/Cu_{(s)}}$$

Q# $E_{cell} = ?$

e^- (gain)

Cu^{+2}/Cu → 0.3 volt (R.P) Anode

Zn^{+2}/Zn → 0.4 volt (R.P) cathode

(e^- gain)

$E_{cell} = 0.4 - 0.3$

$\boxed{E_{cell} = 0.1\ V}$

Q#

e^- gain

Ag^{+1}/Ag = 0.4 (R.P) cathode

Mg/Mg^{+2} = 0.3 (O.P) Anode

(e^- loss)

$E_{cell} = 0.4 - 0.3$

$\boxed{E_{cell} = 0.1\ V}$

Q# Cell represent

* Ag $E_{oxid} = -0.4$ Cathode
* Mg $E_{oxid} = -0.2$ Anode

$Mg \longrightarrow Mg^{+2} + 2e^-$ (Anode)

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

$Mg_{(s)} | Mg^{+2}_{(aq)} \;//\; Ag^{+1}_{(aq)} | Ag_{(s)}$

Types of Cell

Electrolytic	electrochemical cell ; Galvanic / Voltaic cell
* Electrical energy → chemical energy	* chemical energy → electrical energy
* Non-spontaneous	* Spontaneous.
* Endothermic	* Exothermic
* Anode = +ve, Cathode = -ve	* Anode = -ve, cathode = +ve
* No salt bridge	* Contain salt bridge
* one Container	* more than one container.
* Limited to lab process	* Commercial cell.

in this cell electrolysis of conc. aqueous NaCl takes place

① Nelson's Cell

* Electrolytic cell
* produce $\boxed{NaOH}$
* Anode → Graphite (inert)
* Cathode → perforated steel (Inert)
* electrolyte → $\boxed{\text{Conc aq } NaCl}$

Anode:

$NaCl \longrightarrow Na^+ + Cl^-$ (more)

$H_2O \longrightarrow H^+ + OH^-$ (less)

Cl^- & OH^- — Cl^- → excess

$2Cl^- \longrightarrow \boxed{Cl_2} + 2e^-$ → gas evolve.

Cathode:

Na^+ & H^+

$2H^+ + 2e^- \longrightarrow \boxed{H_2}$ ✓ gas evolve

At Basin:

$\boxed{Na^+ + OH^- \longrightarrow NaOH}$

② Daniel Cell

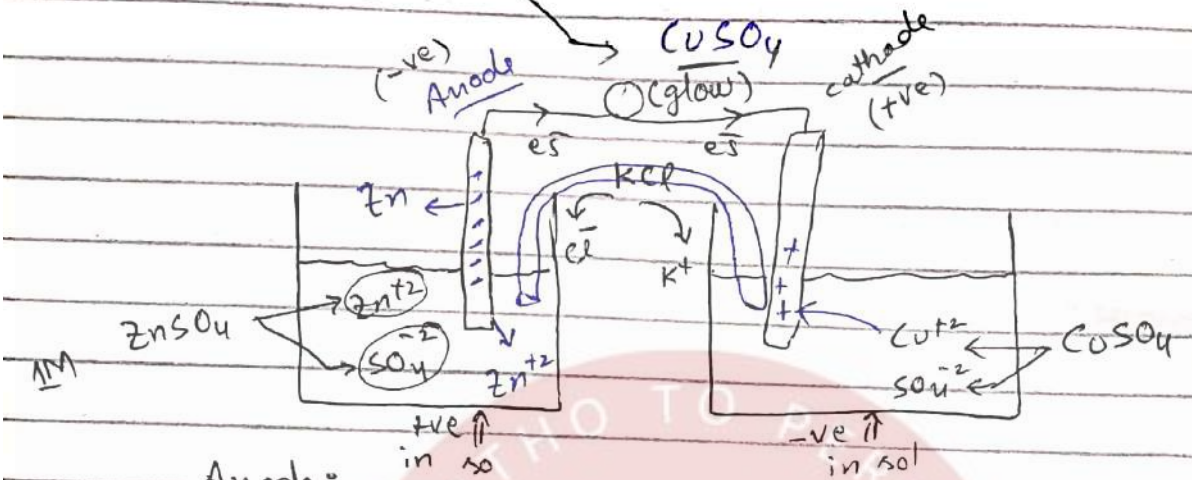
* Voltaic / Galvanic Cell
* Reduction potential $\boxed{Cu > Zn}$
  - Cu → Cathode
  - Zn → Anode
*

Anode ⟶ Zn (Active)

Cathode ⟶ Cu (Active)

electrolytes → $ZnSO_4$

→ $CuSO_4$



Anode:

$Zn \longrightarrow Zn^{+2} + 2es^-$

* es^- at anode ⇑
* in Solution +ve charge ⇑

Cathode:

solution ← $Cu^{+2} + 2es^- \longrightarrow Cu$

* at electrode ⟶ +ve charge ⇑
* in Solution ⟶ -ve charge ⇑

Flow of es:

From Anode toward Cathode

Salt bridge:

↓ gelly solution

KSO_4, $NaSO_4$

KCl, used

- maintain electrical neutrality
- formed electrical contact
- Those electrolytes will be present in which ionic mobility of ions same prevent mixing of subst

* Cation of salt bridge → move toward Cathode

* Anion of salt bridge → " " Anode

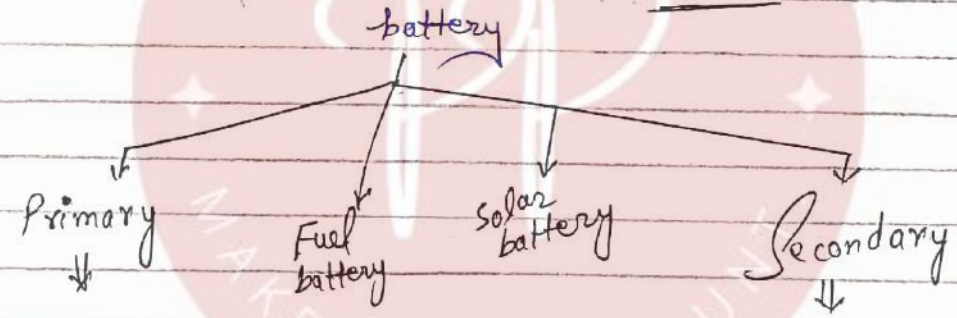
* if No salt bridge → No flow of es

* if s. bridge stop suddenly → then current stop gradually

× × ×

Battery → Voltaic cell

A container which contain one or more than one voltaic cell called battery



* Not Rechargible
* Not Reversible b/c electrolyte used in rxn.

* Rechargible
* Reversible
  ↓
  electrolyte not used in rxn.

Dry Cell

* Leclanch cell (S. name)
* Types of primary battery.
* Voltaic cell
* ~~Anode~~ cathode → Graphite (inert)

Anode ~~Cathode~~ ⟶ Zn (Active)

electrolyte → paste of $ZnCl_2$, MnO_2, NH_4Cl

→ in moist form

Anode:

inactive Zn ⟶ $Zn^{+2} + 2e^-$

↖ Cathode: (complex Rxn occurs)

electrolyte used

* $Zn^{+2} + NH_4Cl \longrightarrow ZnCl_2 + NH_4^+$
* $MnO_2 + NH_4^+ + 2e^- \longrightarrow Mn_2O_3 + NH_3 + H_2O$ (gas)
* $Zn^{+2} + NH_3 \longrightarrow [Zn(NH_3)_4]Cl_2$

Note:

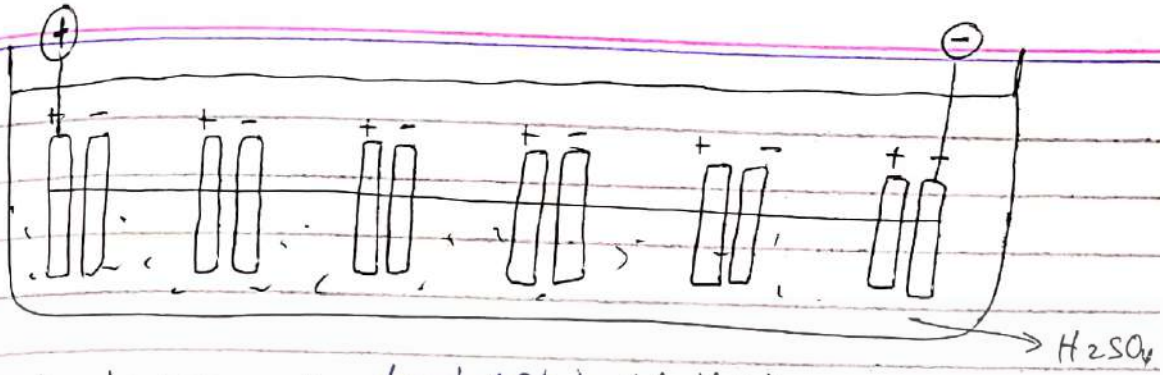
* Conce of H_2O ⇑
* $MnO_2 \longrightarrow Mn_2O_3$

Reduced ← +4 ⟶ +3 (Mn) ⇓

* NH_3 form tetrahedral Complex with zinc.
* Conce of $ZnCl_2$ ⇑
* $NH_4^+ \longrightarrow NH_3$

Lead storage Battery

* Lead acid battery
* Secondary battery
* 6 voltaic cell → Connect in Series



* Anode ⟶ lead (Pb) (Active)

* Cathode ⟶ PbO_2 (Lead oxide (Active)

* electrolyte ⟶ H_2SO_4 (Acidic)

Rxn:

$$H_2SO_4 \longrightarrow SO_4^{-2} + 2H^+ + 2e^-$$

Discharging Case:

Anode

$$Pb + H_2SO_4 \longrightarrow PbSO_4 + 2H^+ + 2e^-$$

$$PbO + H_2SO_4 \xrightarrow{+2H^+ + 2e^-} PbSO_4 + 2H_2O$$

$$PbO + PbO + 2H_2SO_4 \longrightarrow 2PbSO_4 + 2H_2O$$

* (molar) Concent of $PbSO_4$ ↑ ⟶ $C = \frac{n}{V}$ ↓

* $PbSO_4$ ⟶ diluted

* Concent of acid ⟶ decrease

* PH of acid ⟶ increase

$$PH \propto \frac{1}{[H^+]}$$

* Acidic character ⟶ decrease

* Basic character ⟶ increase

* POH ⟶ decrease

Charging Case:

* Current reverse
* Cathode ⟶ Anode
* Anode ⟶ Cathode.

Cathode ⟶ Anode:

$PbSO_4 + 2H_2O \longrightarrow PbO + 2H_2SO_4 + 2H^+ + 2e^-$

Anode ⟶ Cathode:

$PbSO_4 + 2H^+ + 2e^- \longrightarrow Pb + H_2SO_4$

$2PbSO_4 + 2H_2O \longrightarrow PbO + Pb + H_2SO_4$.

* Acid ⟶ concentrated (H_2O ⇓)
* PH ⟶ decrease
* Acidic char ⟶ increase
* Concen of $PbSO_4$ ⟶ Decrease

Fuel Cells

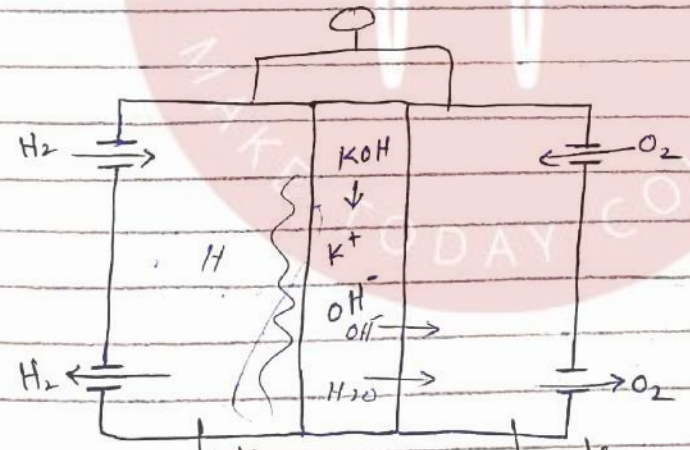
* Any cell which produce energy due through Combustion Rxn
  ↓
  Rxn of any Compound with O_2.
* Type of voltaic cell
  ⇓
  (Chemical energy ⟶ electrical energy)

* Last longing or life span ⇑
* efficient cells.
* Need continuous supply of fuel / energy

⊛ Alkaline Fuel Cell

=> Hydrogen & O_2 fuel cell.

* R.P of $O_2 > H_2$ — H_2 will come ⇑
* Anode (-ve) → porous graphite with Pt, Pd, Ni
* Cathode (+ve) → porous graphite or carbon coated with Ni & Nickel oxide with Pt, Pd, Ni — O_2 will come
* electrolyte → strong Basic Solution e.g. KOH, NaOH.



Anode:

$2H_2 \longrightarrow 4H^+ + 4e^-$

$4H^+ + 4OH^- \xrightarrow{+4e^-} 4H_2O$

$\boxed{2H_2 + 4OH^- \longrightarrow 4H_2O + 4e^-}$

Cathode:

$2H_2O + O_2 + 4e^- \longrightarrow 4OH^-$

Net Rxn:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Note:

* Net Rxn product ⟶ Formation of H_2O
* H_2O concentration ⟶ increase
* Concent of Base (OH^-) ⟶ decrease
* POH ⟶ increase
* Basic strength ⟶ Decrease.
* PH strength ⟶ decrease
* Acidic strength ⟶ increase

New

Corrosion

* The slow & continuous eating away of metal.

Rusting

* The corrosion of iron called Rusting
  => property of metal Not non-metal.
  * Redox Rxn.
  * electrochemical process.
  * form (likely) electrolytic cell
  * exothermic
  * Spontaneous

* slow process
* Brownish mass form called Rust.

⊛ Condition:

=> iron + H_2O + O_2 (Air)

in dry air ⟶ No Rusting

* Anode => Fe
* Cathode => Water Film (layer) (H_2O, O_2)
* Catalyst ↓ H_2O

⊛ Rust: ⟶ class of compound

* Iron (iii) oxide (Fe_2O_3)
* Iron (iii) Hydroxide $Fe(OH)_3$ ✓
* Iron (iii) oxide hydroxide FeO(OH)

Rxn:

Anode: (active) ⟶ oxidized through O_2

oxidation: $2Fe \longrightarrow 2Fe^{+2} + 4e^-$

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

Fe ⟶ Reducing agent.

Cathode: ⟶ (Active)

$H_2O + O_2 + 4e^- \longrightarrow 4OH^-$

O_2 ⟶ oxidizing agent.

Rust:

$Fe^{+3} + 3OH^- \longrightarrow Fe(OH)_3$

(iron hydroxide)