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# NOTES CLUB

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Conductor :- Matter which can pass electricity are called conductor.

Non-conductor :- Matter which can not pass electricity are called non-conductor.

On the basis of flow of electricity conductors are divided into 3 parts :-

Metallic conductor :- These substances conduct electricity due to flow of electrons from high potential to low potential.

Electrolysis :- These are substances which gives ions in solution each conductor under go chemical changes by the flow of electricity.

Example :- Acid, base and salt solution

Electrolytes are two types :-

- ① weak electrolytes
- ② strong electrolytes

weak electrolytes	strong electrolytes
① These ionised partially in aqueous solution.	① These ionised completely in aqueous solution.
② These have low electrical conductivity.	② These have high electrical conductivity.

③ Dilution effect on their conductivity.

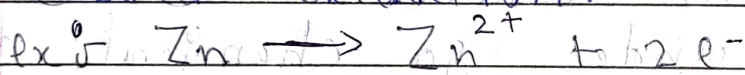
Example:  $\text{NH}_4\text{OH}$ ,  $\text{CH}_3\text{COOH}$  etc.

③ Dilution does not affect their conductivity.

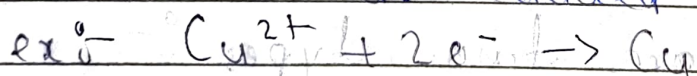
Example:  $\text{CuSO}_4$ ,  $\text{NaCl}$ ,  $\text{KCl}$ ,  $\text{NaOH}$ ,  $\text{H}_2\text{SO}_4$  etc.

Non-electrolytes: These are substances which cannot conduct electricity either in solid state or in solution.  
ex: ether, glucose.

Oxidation: De-electronation of a substance in a process is called oxidation.



Reduction: Electronation of a substance is called reduction.



Redox reaction: Reaction in which oxidation and reduction take place are called redox reaction.  
example:  $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$

Ohm's Law: Potential difference of conduction is proportional to passing through the conductor.

$$V \propto I$$
$$V = IR$$
$$R = \frac{V}{I}$$

Resistance :- The obstruction produced by conductor for flow of electric current passing through the conductor is called resistance of conductor.

$$R = \frac{V}{I}$$

Its unit is (ohm)  $\Omega$

Specific resistance :- Resistance is directly proportional to the length of conductor inversely proportional to the area of cross section.

$$R \propto \frac{l}{a}$$

$$R = \rho \frac{l}{a}$$

(row)  $\leftarrow \rho = \frac{Ra}{l}$

$\rho$  is called specific resistance its units is  $\Omega \text{ cm}$ .

electrolytic conductivity is electrical

Conductivity is reciprocal of resistance is called electrolytic conductivity.

$$C = \frac{1}{R}$$

Its unit is  $\Omega^{-1}$

Specific Conductivity :- The reciprocal of

specific resistance is called specific conductivity. It is denoted by (k) kappa.

$$k = \frac{1}{P}$$

its unit is  $\Omega^{-1} \text{ cm}^{-1}$

$$k = \frac{l}{Ra}$$

Factors affecting electrolytic conductivity

① Concentration of ions in solution :- In an electrolytic solution as the number of ions increases the conductivity of a solution also increases.

② Ionic mobilities :- Conductivity of electrolytic solution increases as

the mobility of ion increases. Its depend upon the following factors:

- (i) Inter-ionic attraction:- greater the attraction between the ions, lower is the conductivity of a solution.
- (ii) Solvation of ions:- when solute ions are attached by polar molecules of solvent solvation takes place. greater the solvation of solute ions, lower is the conductivity of solution.
- (iii) Viscosity of solvent:- Solution having high viscosity, solute have low conductivity.
- (iv) Temperature:- The conductivity of the solution increases as the temperature increases.

Variations of conductivity with concentration.

Effect of conductivity or specific conductivity:- As the solution is diluted, the number of ions carrying the electric current per meter cubic decreases, therefore conductivity or specific conductivity decreases with increase in dilution of solution.

- ② Molar conductivity :- It increases with the dilution of solution. In strong electrolyte, it increases less than weak electrolytes.

### Change in molar conductivity on Dilution:

- ① For strong electrolytes :- On Dilution there is ionisation increases due to increase in the number of ion the molar conductivity also increase. It can be calculated by Kohlrausch's law.

Kohlrausch's law :- At definite dilution when the dissociation of electrolyte is complete each ion makes a definite contribution towards the molar conductivity of electrolyte.

Irrespective of the nature of other ion with which it is associated.

$$\Delta m^\alpha = x \Delta a^\alpha + y \Delta c^\alpha$$

### Application:

- ① Calculation of molar conductivity for weak electrolyte.
- ② Calculation of degree of ionisation of weak electrolyte.
- ③ Calculation of the solubility of salt.
- ④ Calculation of ionic product of water.

**Galvanic cell :-** It is a electrochemical cell which converts chemical energy into electric energy. The working of this cell is based on redox reaction but oxidation and reduction half reactions indirectly take place in two separate containers. The working of galvanic cell may be explained by taking the example of Daniel cell or voltaic cell.

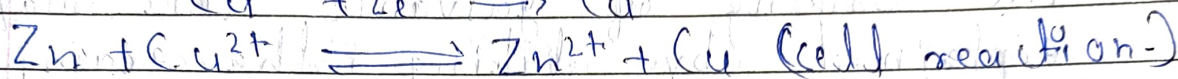
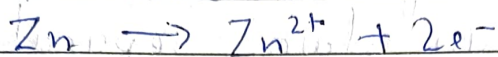
**Daniel cell :-** In this cell a rod of zinc metal is dipped in the solution of zinc sulphate and a rod of copper metal is dipped in the solution of copper sulphate. The solutions are connected together by an inverted U-tube filled with potassium chloride solutions. U-tube also called salt-bridge.

**working of the cell :-** zinc sulphate  $ZnSO_4$  solutions - contains  $Zn^{2+}$  ions when zinc (Zn) rod - comes in contact then electrode dissolving in the form of  $Zn^{2+}$ . This change is called oxidation.  
 $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$  (oxidation)  $\rightarrow A$

The zinc electrons pass through the wire to electrode unite with  $Cu^{2+}$  ions thus reducing them into free Cu atoms.  
 $Cu^{2+} + 2e^- \rightarrow Cu$  (reduction)  $\rightarrow B$



Electrode where oxidation takes place is called anode and where reduction takes place is called cathode. Cell reaction. The sum of the two half cell reaction is called redox reaction or cell reaction.

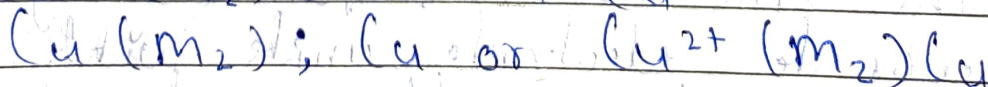
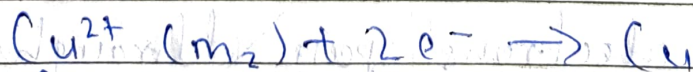
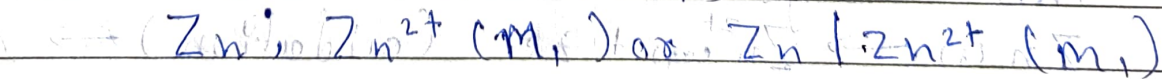
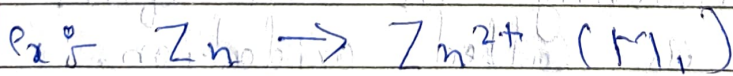
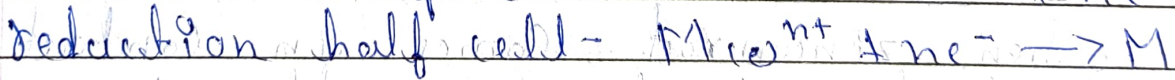
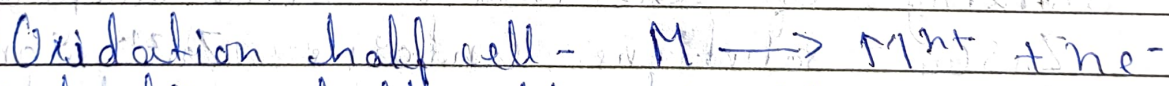


Function of salt bridge :-

① It allows the flow of current by completing the circuit.

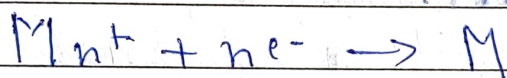
② Maintain electrical neutrality.

\* Voltaic cell of galvanic cell is if M metal and  $M^{n+}$  are the ions in the solution are  $C$  is the concentration of ions then

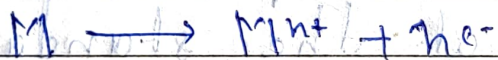


Electrode potential :- when a strip of metal (M) is brought in contact with the solution containing its own ions ( $M^{n+}$ ) then either of the following three possible process can take place.

1. The metal ion  $M^{n+}$  make colloid with metallic strip and bounce bay without any change.
2. The metal ion colloid with the strip gain  $n$  electron and get converted into metal atom the ion is reduced.



3. Metal atom on the strip may lose  $n$  electrons and enter the solution as  $M^{n+}$  ion metal is oxidised.



known if (the metal has) relatively high tendency to get oxidised its atom would start losing electrons change into positive ions and pass into solution. The electron lost develop the negative charge on the metal strip. similarly if the metal ion greater tendency to get reduced they will except electrons at the strip from the metal atom a net positive

charge develop on the metal strip the separation of charges at the equilibrium state result the electrical potential difference between the metal and the solution of its ions and is known as electrode potential.

Electro Motive Force or EMF of the cell  
It may be defined as the potential difference between two electrode of the cell either no current is drawn from it. It is measured with the help of potentiometer, vacuum tube, voltmeter.

The EMF of the cell depends upon the nature of reactants, concentration of the solution in the two half cells at the temperature. The EMF of the cell at the standard state condition is called standard EMF.

and can be calculated from the standard electrode potential of the two half cells.  
( $E_{\text{cell}} = E_{\text{cathode}} - E_{\text{anode}}$ )

SHE :- Standard Hydrogen Electrode  
Consist of a platinum wire sealed into a glass tube and carrying a platinum foil at one end the

Platinum foil is coated with finely divided platinum. The electrode is placed in a weaker containing aqueous solution of some acid having one molar concentration of  $H^+$  ions. Hydrogen gas at one bar pressure is continuously bubbled through at a temperature of 298 Kelvin. The oxidation or reduction in the SHE take place at platinum foil hence it can act as anode as well as cathode. If SHE acts as anode then oxidation will take place.  $H_2(g) \rightleftharpoons 2H^+ + 2e^-$   
If SHE acts as cathode then reduction will take place.  $2H^+ + 2e^- \rightarrow H_2(g)$ .

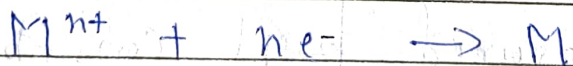
	EMF	Potential difference
①	EMF is the potential difference between the two electrodes of the cell when no current is flowing in the circuit.	It is the difference between the electrode potential of the two electrodes under any condition.
②	It can be measured by potentiometer.	It can be measured by simple volt meter.
③	It is the maximum voltage that can be obtained from the cell.	It is less than maximum voltage.
④	Work done by cell is maximum when calculated by EMF.	Work done is less maximum when calculated by potential difference.

Electrochemical series: The arrangement of various elements in order of increasing value of standard reduction potential is called electrochemical series.

Application of electrochemical series:

1. To determine the relative strength of different oxidizing agent.
2. Standard EMF of the cell can be calculated.
3. Comparison of reactivities of metal.

NERST Equation: - For single electrode the tendency of metal to lose electron or tendency of its ions to gain electrons depends upon the concentration of ions in solution; at the same time the tendency to lose or gain electrons is expressed in term of electrode potential.



The nearest equation can be written in two ways:

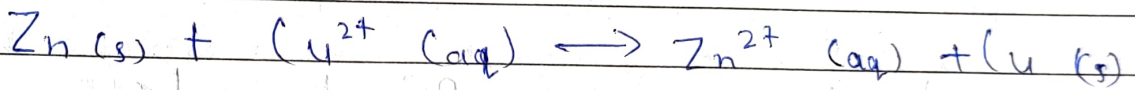
$$E^+(M^{n+}/M) \rightarrow E^-(M^{n+}/M) + \frac{RT}{nF} \ln \frac{M^{n+}}{M}$$

$$\rightarrow E^{\circ} (m^{n+}/m) + 2.303 \frac{RT}{nF}$$

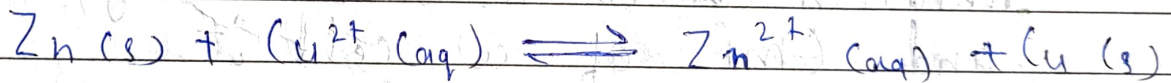
$$E^{\circ} (m^{n+}/m) = E^{\circ} + \frac{0.059}{n} \log_{10} \frac{M_{n+}}{m}$$

Application of Nernst equation

- ① work done by the cell can be calculated
- ② Total work done by cell can be calculated.
- ③ Total change in free energy of cell can be calculated.



At equilibrium



$$E = E^{\circ} - \frac{RT}{nF} \ln \frac{Zn^{2+}}{Cu^{2+}}$$

$$0 = E^{\circ} - \frac{RT}{nF} \ln K_c$$

$$E^{\circ} = 2.303 \frac{RT}{nF} \log_{10} K_c$$

work done by the cell  
= Decrease in

$$= nF E_{\text{cell}}$$

$$-\Delta_r G^\circ = nF E^\circ_{\text{cell}}$$

$$E^\circ = 2.303 \frac{RT}{nF} \log_{10} K_c$$

$$\frac{-\Delta_r G^\circ}{nF} = 2.303 \frac{RT}{nF} \log_{10} K_c$$

$$= nF \cdot 2.303 \frac{RT}{nF} \log_{10} K_c$$

$$\Delta_r G^\circ = -2.303 RT \log_{10} K_c$$

Relation between Gibbs energy change  
and EMF of a cell is

Gibbs energy: The available energy  
used up in the chemical reaction  
is Gibbs free energy

work done by the cell = Amount of  
electricity  $\times$  EMF

Corrosion: It may be defined as the  
process of slow conversion  
of metals into their undesirable

Compounds (usually oxide) by reaction with moisture and other gases present in atmosphere.

### Classification

- ① Atmospheric Corrosion.
- ② Immersed Corrosion
- ③ Chemical Corrosion.
- ④ Underwater Corrosion.
- ⑤ Underground Corrosion.

### Factors which affect Corrosion:-

- ① Reactivity of the metal:- The more active metal are more affect to Corrosion.
- ② Presence of Impurities:- Presence of impurities helps in Corrosion to occur rapidly.
- ③ Air and moisture:- Air and moisture are helpful in Corrosion no rusting is caused if iron is kept in vacuum.

Strains in metal:- Rusting is more on the area having bends - dents, scratches, nicks, cuts.

Present of electrolytes:- The presence of electrolytes also make the Corrosion process faster.



Electrochemical theory of rusting:-

The impure iron surface behave like small electrochemical cell in presence of water containing dissolved oxygen of  $\text{CO}_2$  such a cell is also called Corrosion cells. In these Corrosion cells pure iron act as anode and impure surface act as cathode moisture having dissolved oxygen or  $\text{CO}_2$  in it.

Prevention of Corrosion:-

1. Barrier protection:- A barrier film is introduced between iron and atmospheric oxygen and moisture.

- ① By painting the surface.
- ② By coating the surface with a thin film of oil and grease.
- ③ By electroplating with some non-corrosion metal.

2. Sacrificial protection:- In this method

the surface of iron is covered with a layer of more active metal like - zinc, zinc gets oxidised and Corrosion of iron is prevented.

Electrical protection:- In this method

Iron is protected attached with

Some more active metal such as magnesium or zinc the more active metal act as anode and loss its electrons in presence of iron the iron surface act as cathode.

Use of anticorrosion solutions :- The alkaline solutions of some phosphate and chromate salts act as anticorrosion solutions.

Laws of electrolysis :-

① - Faraday's first law :- The weight of substance (deposited) liberated during electrolysis is directly proportional to the quantity of electricity passed.

$$w \propto Q$$
$$w = ZIt$$

② Faraday's second law :- when the same quantity of electricity is passed through different electrolytes the masses or ions deposited at the electrodes are directly proportional to their chemical equivalents.

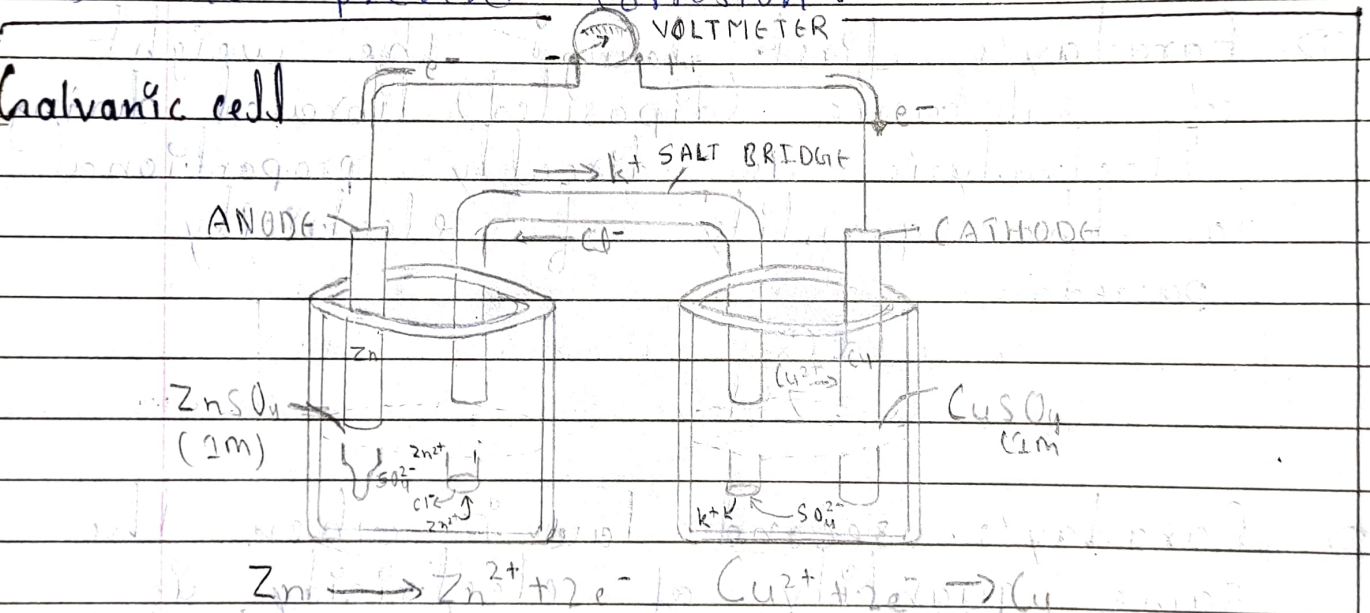
$$w \propto E$$
$$\left( \frac{w_1}{w_2} = \frac{E_1}{E_2} \right)$$

## Applications

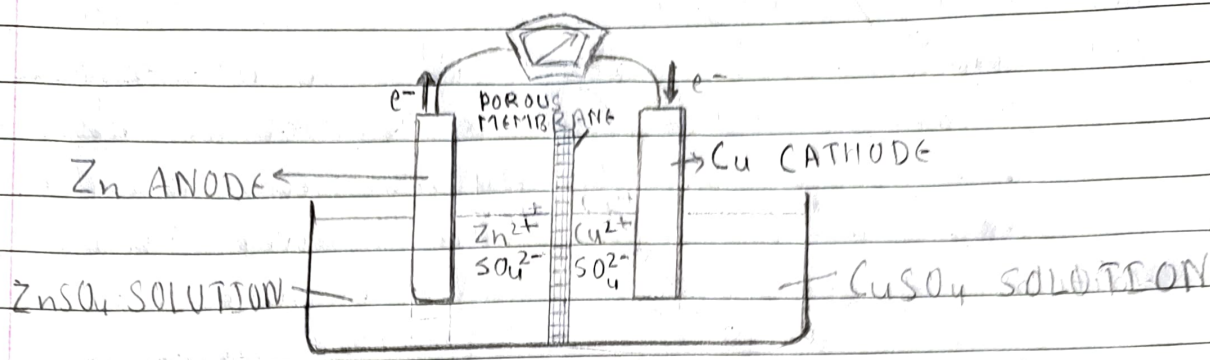
- ① Electrochemical equivalents may be calculated.
- ② Equivalent weight of any element may be determined.
- ③ Amount of substance deposited on electrode may be determined.

## Application of electrolysis

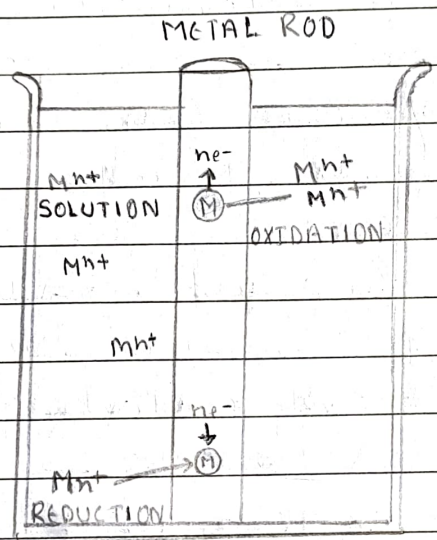
- ① In coating of metal.
- ② In purification of metal.
- ③ To prevent corrosion.



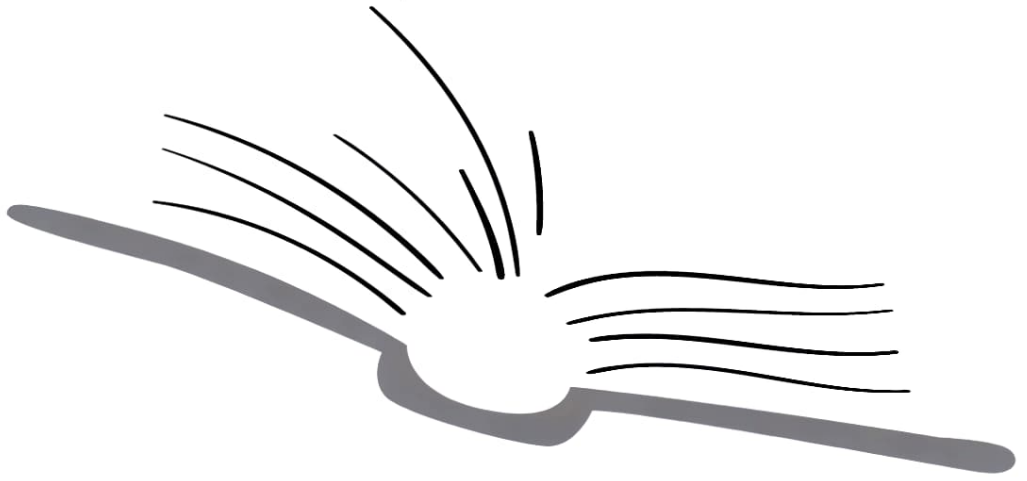
## Electrochemical cell or Galvanic cell



A galvanic cell (with porous membrane)



Electrode Equilibrium/Potential



**THANKYOU**  
**FOR**  
**READIN**

