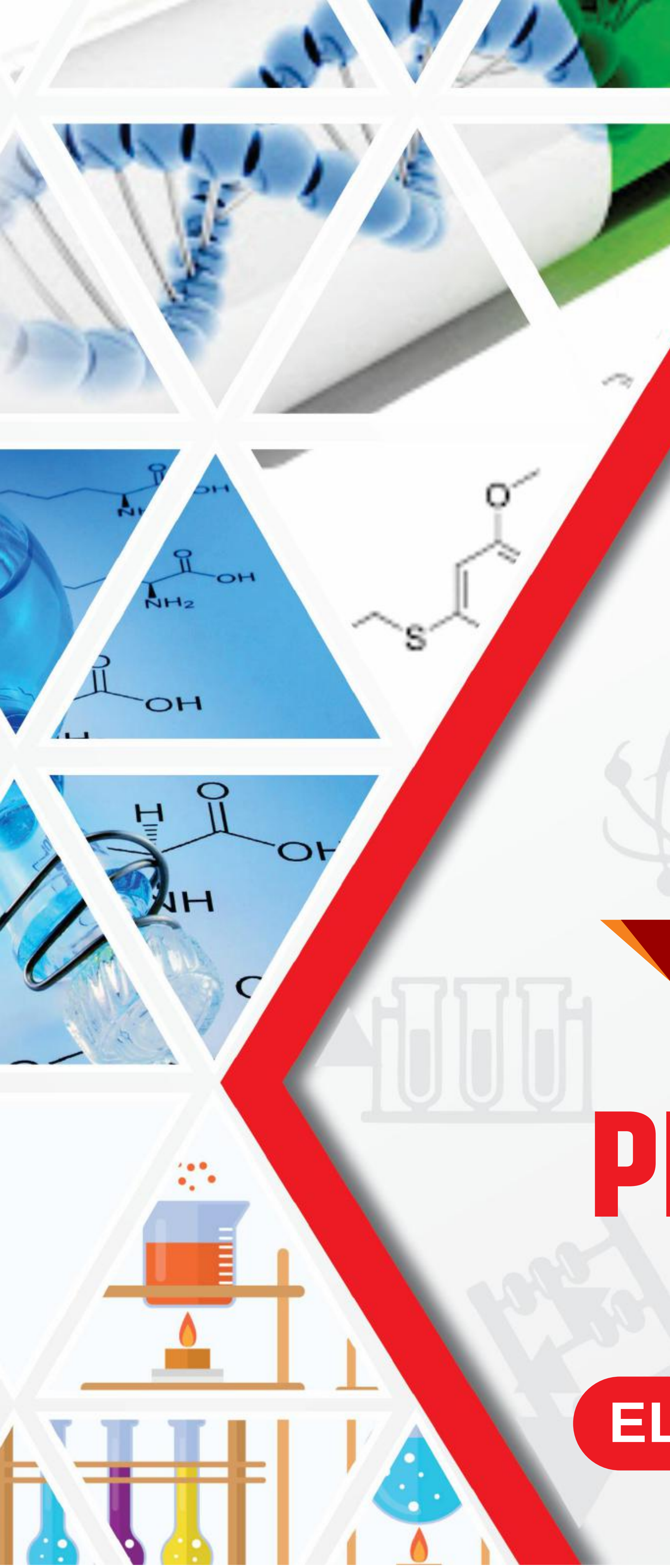


Ranjan Singh
Chemistry Classes
JEE (Main & Advanced)• NEET• XI• XII

Study Package

Ranjan Singh
M.Sc. Bio-Chemistry(P.U)
Ex-Faculty : Narayana & Goal

PHYSICAL
CHEMISTRY

ELECTROCHEMISTRY

Director's Message

Ranjan Singh
M.Sc.(Biochemistry), P.U.

Chemistry plays a central and important role in all competitive examinations as well as in day to day life. For last so many years, I have constantly been in touch with students, guiding them in Chemistry and looking into their difficulties for them to succeed in their board as well as competitive examinations JEE(Mains & Advance) | NEET.

I have felt a need for a good coaching centre to fulfil the requirements of students. Students need a highly experienced and qualified faculty in chemistry, who can guide them well, clear their doubts, provide them the effective & tricky notes, and make them do much needed practice. More importantly they should also be provided Classroom Monitoring, Periodical & Surprise Tests to guide them in the proper direction. I realize that, it is very important to diagnose the basic weaknesses and problems of students not succeeding in JEE(Mains & Advance) | NEET and Board exams. In fact, as question patterns are changing, now they need to have a different approach for these Examinations.

At RANJAN SINGH CHEMISTRY CLASSES, we have our own way to prepare students for Competitive Examinations as well as Board Examination at a time so they can crack the entrance exam like JEE(Mains & Advance) and NEET as well as 12th Board simultaneously. We act as a medium to provide the simplest, easiest and a comfortable way to make students achieve their target. At RANJAN SINGH CHEMISTRY CLASSES(RSCC), we guide our students with the best motivational classes so weak students are also able to believe that, They can do it.

When you join RANJAN SINGH CHEMISTRY CLASSES you become a part of the powerful force which propels you towards your goal and if you get a position among the rankers with my excellent guidance, I will think that our efforts have borne fruits.

M.Sc(Biochemistry), P.U.
Ex-faculty : Narayana IIT Academy
& Goal Institute

Ranjan Singh

Electrochemical cell :

It is device for converting chemical energy into electrical energy.

Electrochemical cell are of two types

Galvanic cells or Voltaic cell	Electrolytic cell.
A spontaneous chemical reaction generates an electric current.	An electric current drives a nonspontaneous reaction.

The two types of cells are therefore reverse of each other

I. **Anode :**

(i) Is where oxidation occurs

(ii) Is where electrons are produced

(iii) Has a negative sign

II. **Cathode :**

(i) Is where reduction occurs

(ii) Is where electrons are consumed

(iii) Has a positive sign

Construction of Cell :

- It has two **half–cells**,each having a beaker containing a metal strip that dips in its aqueous solution.
- The metal strips are called **electrodes** and are connected by a conducting wire.
- Two solutions are connected by a **salt bridge.**
- The oxidation and reduction half reactions a occur at separate electrodes and electric current flows through the wire.

Functions of Salt Bridge

- A **salt bridge** is a U–shaped inverted tube that contains a gel permeated with an inert electrolyte.
- It connects the solution of two half cell to complete the circuit.
- It minimise the liquid junction potential. The potential difference between the junction of two liquids.
- It maintains the electrical neutrality of the solution in order to give continous flow or generation of current.

Electrode Potential :

- The driving force that pushes the negative charged electrons away from the anode and pulls them towards the cathode is an electrical potential called **electromotive force** also known as **cell potential** or the **cell voltage.** Its unit is volt
- The potential difference developed between metal electrode and its ions in solution in known as electrode potential.
- The potential difference developed between metal electrodes and the solution of its ions at 1 M concentration at 1 bar pressure and 298 K is known as standard electrode potential.

Oxidation Potential (O.P.) :

- The electrode potential for oxidation half reaction
- Tendency to get oxidised.
- **Greater the O.P. greater will be tendency to get oxidised.**

Reduction Potential (R.P.)

- The electrode potential for reduction half reaction
- Tendency to get reduced.
- **Greater the R.P. greater will be tendency to get reduced.**

	Type of Electrode	**Electrode reaction in standard condition**	**Representation**
1.	Metal electrode	Reduction : $Zn^{2+} + 2e^- \rightarrow Zn(s)$	$E^0_{Zn^{2+}/Zn(s)}$ (SRP)
	(Zn electrode,	Oxidation : $Zn(s) \rightarrow Zn^{2+} + 2e^-$	$E^0_{Zn(s)/Zn^{2+}}$ (SOP)
	Cu electrode etc.)		
2.	Hydrogen peroxide	Reduction : $2e^- + 2H^+ + H_2O_2 \rightarrow 2H_2O$	$E^0_{H_2O_2/H_2O}$
	electrode	Oxidation : $H_2O_2 \rightarrow O_2 + 2H^+ + 2e^-$	$E^0_{H_2O_2/O_2}$
3.	Redox electrode	Reduction : $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	$E^0_{MnO_4^-/Mn^{2+}}$
4.	Metal Metal	Reduction : $AgCl(s) + e^- \rightarrow Ag(s) + Cl^-$	$E^0_{AgCl(s)/Ag(s)/Cl^-}$
	salt insoluble	Oxidation : $Ag(s) + Cl^- \rightarrow AgCl(s) + e^-$	$E^0_{Ag(s)/AgCl(s)/Cl^-}$
	electrode		

Reference electrode :

- The potential of a single electode cannot be determined, whereas the potential difference between two electrodes can be accurately measured using a reference electrode.
- An electrode is chosen as a reference with respect to which all other electrodes are valued.
- Standard Hydrogen Electrode **(SHE)** is taken as standard reference electrode. Its electrode potential is arbitrarily assumed to be 0.00 volt.
- Standard Hydrogen Electrode **(SHE)** consists of a platinum electrode in contact with H_2 gas and aqueous H^+ ions at standard state conditions (1 atm H_2 gas, 1 M H^+ (aq), 25°C).

$$2H^+(aq, 1M) + 2e^- \rightarrow H_2(g, 1\ atm) \qquad E° = 0V$$

$$H_2(g, 1atm) \rightarrow 2H^+(aq, 1M) + 2e^- \qquad E° = 0V$$

Cell potential :

- The difference in electrode potentials of the two half cell reactions (oxidation half cell and reduction half cell) is known as emf of the cell or cell potential. The emf of the cell or cell potential can be calculated from the values of electrode potential of the two half cell constituting the cell. The following three methods are in use :

When oxidation potential of anode and reduction potential of cathode are taken into account

E°_{cell} = oxidation potential of anode + reduction potential of cathode

$E^\circ_{cell} = E^\circ_{ox}$ (anode) + E°_{red}(cathode)

When reduction potential of both electrodes are taken into account :

E°_{cell} = Reduction potential of cathode – Reduction potential of anode

$E^\circ_{cell} = E^\circ_{red}$(cathode) – E°_{red}(anode)

When oxidation potential of both electrodes are taken into account :

E°_{cell} = oxidation potential of anode – Oxidation potential of cathode

$E^\circ_{cell} = E^\circ_{ox}$ (anode) – E°_{ox} (cathode)

The standard cell potential E° is the cell potential when both reactants and products are in their standard states – solutes at 1 M concentration, gases at a potential pressure of 1 atm, solids and liquids in pure from, with all at a specified temperature, usually 25° C.

Free energy changes for cell reaction :

The values of ΔG and E are directly proportional and are related by the equation.

$$\Delta G = -nFE$$

where

n = Number of moles of electron transfered in the reaction.

F = Faraday constant = 96485 C/mole e^- $\approx$ 96500 C/mole e^-

Shorthand Notation for Galvanic Cells

We require two half cells to produce an electrochemical cell, which can be represented by following few rules

The anode half-cell is always written on the left followed on the right by cathode half cell.

The separation of two phases (state of matter) is shown by a vertical line.

The various materials present in the same phase are shown together using commas.

The salt bridge is represented by a double slash (||).

The significant features of the substance viz. pressure of a gas, concentration of ions etc. are indicated in brackets immediately after writing the substance.

For a gas electrode, the gas is indicated after the electrode for anode and before the electrode in case of cathode. (i.e Pt H_2 / H^+ or H^+ /H_2 Pt)

$$\xrightarrow{\text{direction of e-flow}}$$

Anode half-cell   Cathode half-cell

$Cu(s) \mid Cu^{2+}(aq, 1M) \quad || \quad Cl_2(g, 1\ atm) \mid Cl^-(aq, 1M) \mid C(s)$

Electrochemical Series

Electrode	Reaction	SRP (at 298 K)
* Li	$Li^+ + e^- \rightarrow Li(s)$	– 3.05 V
K	$K^+ + e^- \rightarrow K(s)$	– 2.93 V
Ba		
Ca	$Ca^{+2} + 2e^- \rightarrow Ca(s)$	– 2.87 V
Na	$Na^+ + e^- \rightarrow Na(s)$	– 2.71 V
Mg	$Mg^{+2} + 2e^- \rightarrow Mg(s)$	– 2.37 V
Al		
* Zn	$Zn^{+2} + 2e^- \rightarrow Zn(s)$	– 0.76 V
Cr	$Cr^{+3} + 3e^- \rightarrow Cr(s)$	– 0.74 V
* Fe	$Fe^{2+} + 2e^- \rightarrow Fe$	– 0.44 V

Increasing strength of reducing agent

Increasing strength of oxidising agent

* Electrolytes (H_2O)	$H_2O(l) + e^- \rightarrow \frac{1}{2}H_2 + OH^-$	– 0.41 V
Cd	$Cd^{+2} + 2e^- \rightarrow Cd(s)$	– 0.40 V
Co		
Ni	$Ni^{+2} + 2e^- \rightarrow Ni(s)$	– 0.24 V
Sn	$Sn^{+2} + 2e^- \rightarrow Sn(s)$	– 0.14 V
Pb	$Pb^{+2} + 2e^- \rightarrow Pb(s)$	– 0.13 V
* H_2	$2H^+ + 2e^- \rightarrow H_2(g)$	0.00 V
Cu	$Cu^{2+} + 2e^- \rightarrow Cu(s)$	0.34 V
I_2		
Fe	$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77 V
Hg	$Hg_2^{2+} + 2e^- \rightarrow Hg(l)$	0.79 V
Ag	$Ag^+ + e^- \rightarrow Ag$	0.80 V
Hg	$Hg^{2+} \rightarrow Hg(l)$	0.85 V
Br_2	$Br_2 + 2e^- \rightarrow 2Br^-$	1.06 V
* Electrolytes	$\frac{1}{2}O_2 + 2H^+ + 2e^- \rightarrow H_2O(\ell)$	1.23 V
*	$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$	1.33 V
*	$Cl_2 \rightarrow Cl_2 + 2e^- \rightarrow 2Cl^-$	1.36 V
*	$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51 V
*	$F_2 + 2e^- \rightarrow 2F^-$	2.87 V

- **Calculation of Electrode Potential of unknown electrode with the help of given (two) electrode.**
  - Obtain the reaction of the 3rd electrode with the help of some algebraic operations on reactions of the given electrodes.
  - Then calculate ΔG of the 3rd reaction with the help of some algebaric operations of ΔG^0 of 1st and 2nd reactions.
  - Use $\Delta G^0 = -nF E^0_{elec.}$ to calculate unknown Electrode Potential
  - E^0_{cell} is intensive property so if we multiply/divide electrode reaction by any number the E^0_{cell} value would not changed

    i.e. $Zn^{2+} + 2e^- \rightarrow Zn(s)$ $E° = -0.76$ V

    Multiply by 2

    $2Zn^{2+} + 4e^- \rightarrow 2Zn(s)$ $E° = -0.76$ V (remain same)

Nernst Equation :

- Cell potentials depend on temperature and on the composition of the reaction mixtures.
- It depends upon the concentration of the solute and the partial pressure of the gas, if any.
- The dependence upon the concentration can be derived from thermodynamics.

From thermodynamics

$$\Delta G = \Delta G° + RT \ln Q$$

$$-nFE = -nFE° + 2.303\,R\,T \log Q$$

$$E = E° - \frac{2.303RT}{nF} \log Q$$

Take T = 298 K

R = 8.314 J/mol K

$F = 96500\ C$

Now we get

$$E = E° - \frac{0.059}{n}\log Q$$

Where n = number of transferred electron

Q = reaction quotient

- Nerst equation can be used to calculate cell potentials for non standard conditions also.
- Nerst equations can be applied to half cell reactions also.

Nernst Equation for cell Potential :

$$aA + bB \rightleftharpoons CC + dD$$

$$E_{cell} = E^0_{cell} - \frac{RT}{nF}\ln Q$$

n = no. of electrons which gets cancelled out while making cell reaction.

Equilibrium in electrochemical cell

$$\Delta G^0 = -nFE^0_{cell}$$

$$\mathbf{\Delta G = -nF}E^0_{cell}$$

From thermodynamics

$\Delta G = \Delta G^0 + RT\ell nQ$

at chemical eqlibrium $\Delta G = 0$, $Q = K_{eq}$

$E_{cell} = 0 \rightarrow$ cell will be of no use

so, $\Delta G^0 = -RT\ell n\ K_{eq}$

at equilibrium $-nFE^0_{cell} = -2.303\ RT\log(K_{eq})$

$$\log K_{eq} = \frac{nF}{2.303RT}E^0_{cell}$$

at 298 K and R = 8.314 J/mol K

$$\log Keq = \frac{n}{0.059}E^0_{cell}$$

Concentration cells :

- A concentration cell consists of two electrodes of the same material, each electrode dipping in a solution of its own ions and the solution being at different concentrations.
- The two solutions are separated by a salt bridge.

e.g.

$Ag(s)\ |\ Ag^+(a_1)\ ||\ Ag^+(a_2)\ |\ Ag(s)$ $(a_1 < a_2)$ a_1, a_2 are concentrations of each half cell

At LHS electrode Anode : $Ag(s) \longrightarrow Ag^+(a_1) + e^-$

At RHS electrode Cathode : $Ag^+(a_2) + e^- \longrightarrow Ag(s)$

The net cell reaction is : $Ag^+(a_2) \longrightarrow Ag^+(a_1)$

The nernst eq. is

$$E_{cell} = -\frac{0.059}{n}\log\frac{a_1}{a_2} \qquad \text{(Here n = 1, Temp, 298 K)}$$

Electrolysis & Electrolytic cell :

Electrolysis :

- Electrolyte is a combination of cations and anions which in fused state can conduct electricity.
- This is possible due to the movement of ions from which it is made of and electrolyte.
- The process of using an electric current to bring about chemical change is called electrolysis.
- Electrolysis is a process of oxidation and readuction due to current in the electrolyte.
- The product obtained during electrolysis depends on following factors.
  - The nature of the electrolyte
  - The concentration of electrolyte
  - The charge density flowing during electrolysis.
  - The nature of the electrode

Faraday's Law of Electrolysis :

- **1st Law :** The mass deposited/released/produced of any substance during electrolysis is proportional to the amount of charge passed into the electrolyte.

  $W \propto Q$

  $W = ZQ$

  Z – electrochemical equivalent of the substance.

  $$\text{Unit of } Z = \frac{\text{mass}}{\text{coulomb}} = \text{Kg/C or Gm/C}$$

  Z = Mass deposited when 1 C of charge is passsed into the solution.

Equivalent mass (E) : mass of any substance produced when 1 mole of e^- are passed through the solution during electrolysis.

$$E = \frac{\text{Molar mass}}{\text{no.of } e^- \text{ involved in oxidation / reduction}}$$

$$W = \frac{i \times t}{96500} \times \frac{\text{Molar mass}}{(\text{no. of } e^- \text{ involved})}$$

- **2nd Law :** When equal charge is passed through 2 electrolytic cells and these cells are connected in series then mass deposited at electrode will be in the ratio of their electrochemical equivalents or in the ratio of their equivalent masses.

  $$W = ZQ = \frac{EQ}{96500}$$

  $$\frac{W_1}{W_2} = \frac{z_1}{z_2} = \frac{E_1}{E_2} \text{ (Q = same)}$$

Current Efficiency :

$$\text{current efficiency} = \frac{\text{charge actually used in electricity}}{\text{charge passed}} \times 100$$

$$\text{current efficiency} = \frac{\text{mass actually produced}}{\text{mass that should have been produced}} \times 100$$

Electrolytic conduction

Conductors	**Insulators**	**Semiconductors**
Metals	Genrally non–metals	Si, Ge

Conductors :

(1) Electronic conductors (2) Electrolytic conductors

Electronic Conductors: No decomposition of any substance current flow because of e^-(charge carries are e^-

On increasing temprature (T), resistance (R) increases.

Electrolytic Conductors : Solutions of electrolytes (WA, SA, WB, SB, salt solution)

– When current is passed there will be decomposition of electrolyte (electrolysis)

– Charge carriers are ions.

– On increasing temprature (T), resistance (R) decrease. (because more charge carriers are produced)

– Factors effecting electrolytic conductance :

(1) Nature of the electrolyte,

(2) Nature of the solvent : polar more conductance
non–polar less conductance

(3) Concentration : Generally on increasing concentration conductivity decreases.

(4) Temperature : On increasing temprature (T), conductivity increases

Resistance

$$R = \frac{V}{I} \text{ (Ohm's law } (\Omega))$$

$$R = \frac{\rho \ell}{A}$$

ρ – resistivity/specific resistance

– resistance of unit length wire of unit area of cross section = constant = (Ω m)

$$\rho = \frac{RA}{\ell}$$

Resistivity of a solution is defined as the resistance of the solution between two electrodes of 1 cm^2 area of cross section and 1 cm apart.

or

Resistance of 1 cm^3 of solution will be it's resistivity.

Onductance (C OR G) → Inverse of R :

The conductance offered by an electrolytic conductor is expressed in terms of reciprocal of resistance as its is the ease provided by solution in the passage of current.

$$C = \frac{1}{R} = \Omega^{-1} \text{ (mho)}$$

[C expressed in ohm^{-1} or in Ω^{-1} or mho or Siemens (S)]

Moh scale → on this scale hardness of substance is measured (diamond = 10)

Conductivity (K or σ) or specific conductance (specific conductivity) κ_{kappa}

$$K = \frac{1}{\rho} = \frac{\ell}{aR} = C.\frac{\ell}{a}$$

$$\therefore \quad K = C.\frac{\ell}{a}$$

= $\Omega^{-1}\ m^{-1}$ (SI unit)
= $\Omega^{-1}\ cm^{-1}$ [In this chapter we use this unit]
or ($\mathbf{K}$ expressed in $S\ Cm^{-1}$)

Note :

- Conductivity of a conductor is reciprocal to its resistivity.
- $\left(\frac{\ell}{a}\right)$ is known as cell constant expressed in cm^{-1} and in SI unit as m^{-1}.
- Conductance is an additive property. For example, an aqueous solution containing several electrolytes, the total conductance is given by $C_T = \sum C_i + C_{water}$, where $\sum C_i$ is total conductances of all the ions and C_{water} is conductance of water utilized for making the solution.
- C_{water} is often negligible in comparion to $\sum C_i$ as repeatedly distilled water (known as conductivity water) of very low conductivity $3.5 \times 10^{-5}\ S\ m^{-1}$ is used for making solution.
- If electrodes are of different area of cross–section then common area of electrodes is taken as effective area of cross–section.

Electrolytic conduction :

Specific conductivity (K) : Conductance of a unit cube ($1cm^3$) of solution is called its specific conductivity.

$$K = C \times \frac{\ell}{a}$$

Molar conductivity : – (λ_m)

The conductance of solution kept between the electrodes at unit distance apart and having area of cross-section large enough to accomodate sufficient volume of the solution that contains 1 mole of electrolyte.

Relationship between conductivity (κ) and molar conductivity (λ_m) :

Let V cm^3 of the solution containing 1 mole of electrolyte be palced betwen two large electrode 1 cm apart.

Let measured conductance of the solution = x.

The solution contains 1 gram/mole of electrolyte so the measured conductance also represent the molar conductivity of solution.

$$\lambda_M = X.$$

But we know that conductivity (K) is the conductance of 1 centimeter cube of solution.

$$\kappa = \frac{X}{V} \qquad \text{Molarity (M)} = \frac{1 \times 1000}{V_{ml}}$$

or $$\kappa = \frac{\lambda_M}{V}, \qquad M = \frac{1000}{V}$$

$$\lambda_M = \kappa \times V. \qquad V = \frac{1000}{M}$$

$$\lambda_M = \frac{K \times 1000}{M} \qquad \text{Unit of } \lambda_M = \Omega^{-1}\ cm^2\ mole^{-1}$$

Equivalent conductivity : Equivalent conductivity of a solution is defined as conductance of all the ions produced by 1gm equivalent of the electrolyte dissolved in V cm^3 of the solution when the distance between the electrodes is 1cm and area of the electrodes is so large that whole of the solution is contained between them.

Let V cm^3 of the solution containing 1 gm equivalent of electrolyte be palced between two large electrode 1 cm apart.

Let measured condctance of the solution = x.

The solution contains 1 gram/eq. of electrolyte so the measured conductance also represent the eq. conductivity of solution.

$$\lambda_N = X.$$

But we know that conductivity (K) is the conductance of 1 centimeter cube of solution.

$$\kappa = \frac{X}{V} \qquad \text{Normalty (N)} = \frac{1 \times 1000}{V_{ml}}$$

or $\kappa = \frac{\lambda_N}{V}$, $\qquad N = \frac{1000}{V}$

$$\lambda_N = \kappa \times V. \qquad V = \frac{1000}{N}$$

$$\lambda_N = \frac{K \times 1000}{N}$$

$\lambda_{eq} = \kappa \times$ volume of solution containing one 1gm eq. of electrolyte

$$= \kappa \times \frac{1000}{N} \; (\Omega^{-1}\,cm^{-1}) \left(\frac{cm^3}{eq}\right)$$

$$= \Omega^{-1}\,cm^2\,\text{equivalent}^{-1}$$

Variation of conductivities with dilution :

- **Varition of κ**

  On dilution (or on decreasing concentration) κ (specific conductance of solution decreases)

- **Varition of λ_{eq} / λ_M**

  λ_{eq} or λ_M increases on dilution

  Because effect of volume increament is greater than effect of decreament in κ.

Variation of λ_{eq} / λ_M of a solution with concentration :

- **Strong electrolyte**

  These solution are found to follow debye huckle onsagar eqation at low concentrations.

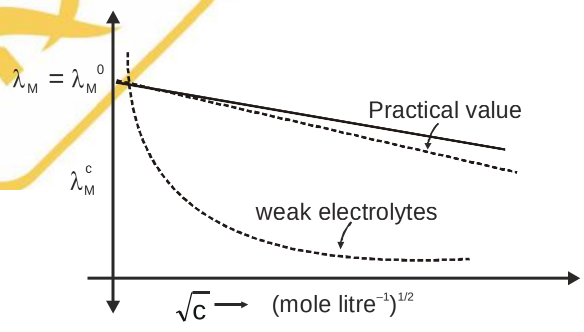
$$\lambda_M^c = \lambda_M^\infty - b\sqrt{c}$$

$$= \lambda_M^\infty - b\sqrt{c} \qquad \text{[b constant]}$$

λ_M^∞ or λ_M^0 = molar conductivity at infinite dilution

= molar conductivity at zero concentration

b = constant



- **Weak electrolytes :**

  For the same concentration of weak electrolytes as of strong electrolytes the λ_{eq} of λ_M values are very low in comparison to those of strong electrolytes.

  at ∞ dilution WE (Weak electrolytes) is also 100% dissociated so have high values of λ_{eq} of λ_M .

  In case of SE (Strong electrolytes) we can calculated λ_M^∞ or λ_M^0 by extrapolation of their straight line variation of λ_M with c

  But in case of weak electrolytes we can not be calculated λ_M^∞ or λ_M^0 by extraplotation or practical method.

Kohlrausch law :

At infinite dilution each ion makes a definite contribution towards equivalent conductance of electrolyte irrespective of the nature of ion with which it is associated and the value of equivalent conductance at infinite dilution for any electrolyte is the sum of contributions from its constituents ions.

	λ_m^0	**diff.**	**common ions**	**diff.**
KCl	235		NaCl	
NaCl	200	35	NaBr	55
KBr	195		KCl	
NaBr	160	35	KBr	55
KI	155		RbCl	
NaI	120	35	RbBr	55

$$\lambda_M^\infty\ (KCl) = \lambda_{K^+}^\infty + \lambda_{Cl^-}^\infty$$

♦ **Equivalent Conductivity :**

$$\lambda^\infty_{eq}(\text{electrolyte}) = \lambda^\infty_{C^+} + \lambda^\infty_{a^-}$$
$$\downarrow$$

Limting ionic conductivity of cation

$$\lambda^\infty_{eq(NaCl)} = \lambda^\infty_{Na^+} + \lambda^\infty_{Cl^-}$$

$$\lambda^\infty_{eq(BaCl)} = \lambda^\infty_{Ba^{+2}} + \lambda^\infty_{Cl^-}$$

$$\lambda^\infty_{eq(AlCl_3)} = \lambda^\infty_{Al^{+3}} + \lambda^\infty_{Cl^-}$$

$$\lambda^\infty_{eq}\ (Al_2(SO_4)_3) = \lambda^\infty_{Al^{+3}} + \lambda^\infty_{SO_4{}^{2-}}$$

♦ **Molar conductivity :**

$$\lambda_M^0 = m\lambda_c^0 + n\lambda_a^0$$

where m & n are no of cations and anions present in one formula unit of the electrolyte

$$\lambda^0_{M\ (NaCl)} = \lambda^0_{Na^+} + \lambda^0_{Cl^-}$$
$$\lambda^0_{M\ (BaCl_2)} = \lambda^0_{Ba^{2+}} + 2\lambda^0_{Cl^-}$$
$$\lambda^0_{M\ (AlCl_3)} = \lambda^0_{Al^{3+}} + 3\lambda^0_{Cl^-}$$
$$\lambda^0_{M\ (Al_2(SO_4)_3)} = 2\lambda^0_{Al^{3+}} + 3\lambda^0_{SO_4{}^{2-}}$$

Application of kohlrausch law :

♦ **Calculation of λ^0_M of weak electrolytes :**

$$\lambda^0_{M\ (CH_3COOH)} = \lambda^0_{CH_3COO^-} + \lambda^0_{H^+}$$
$$\lambda^0_{M\ (CH_3COONa)} = \lambda^0_{CH_3COO^-} + \lambda^0_{Na^+}$$
$$\lambda^0_{M\ (HCl)} = \lambda^0_{H^+} + \lambda^0_{Cl^-}$$
$$\lambda^0_{M\ (NaCl)} = \lambda^0_{Na^+} + \lambda^0_{Cl^-}$$
$$\lambda^0_{M\ (CH_3COOHl)} = \lambda^0_{M(CH_3COONa)} + \lambda^0_{M(HCl)} - \lambda^0_{M(NaCl)}$$

♦ **Calculation of degree of dissociation and K_{eq} of weak electrolytes (dissociation constants) :**

Let at concentration C if molar conductivity = λ_m^C
and at ∞ dilution the molar conductivity = λ^∞_m

$$\text{then} \quad \alpha = \frac{\lambda_m^c}{\lambda_m^0}$$

$$K_{eq} = \frac{c\alpha^2}{(1-\alpha)}$$

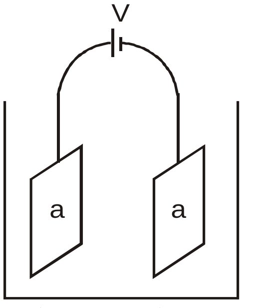
♦ **To evaluate absolute ionic mobilities :**

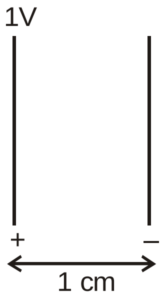
Ionic Mobility :

It is the distance travelled by the ion per second under the potential gradient of 1 volts per cm. It's unit is $cm^2\ s^{-1}\ v^{-1}$.

Potential gradient :

The ratio of potential difference across the electrode to the distance between. Them,





$$\text{Potential gradient} = \frac{V}{\ell}$$

Absolute ionic mobility :

Ionic mobility at inifinite dilution is called absolute ionic mobility and represented by μ^0_c or μ^0_a.

or

Speed of the ion at infinite dilution under unit potential gradient (in $cm^2\ sec^{-1}\ vol^{-1}$).

$$\lambda^0_c \propto \mu_c \qquad ; \qquad \lambda^0_a \propto \mu_a$$

$$\lambda^0_c = F\mu^0_c \qquad ; \qquad \lambda^0_a = F \times \mu^0_a.$$

$$\text{Ionic Mobility } \mu = \frac{\nu}{(V/\ell)} \begin{matrix} \longrightarrow \text{speed} \\ \longrightarrow \text{potential gradient} \end{matrix}$$

$$= \left(\frac{cm/sec}{volt/cm}\right) = cm^2\ volt^{-1}\ sec^{-1}$$

Transport Number :

It is the fraction of total current carried by each ion of the electrolyte. Now greater the mobility of the ion larger will be magnitude of current carried by the ion and consequently larger will be its transport number.

$$t_c = \left[\frac{\mu_c}{\mu_c + \mu_a}\right], \qquad t_a = \left[\frac{\mu_a}{\mu_a + \mu_c}\right].$$

Where t_c = Transport Number of cation & t_a = Transport Number of cation

◆

To calculate solubility of sparingly soluble salt & their K_{sp} values :

The saturated solutions of sparingly soluble salts can be taken to be infinitely dilute

$$\lambda^o_M(AgCl) = \lambda^0_{M(Ag^+)} + \lambda^0_{M(Cl^-)}$$

$$= \kappa \times \frac{1000}{\text{molarity}} = \kappa \times \frac{1000}{\text{solubility}}$$

TOPIC WISE MCQS

Electrolysis

01. In electrolysis of a fused salt, the weight deposited on an electrode will not depend on-
(A) Temperature (B) Current intensity
(C) Electrochemical equivalent of ions
(D) Time for electrolysis

02. The electrolysis of a solution resulted in the formation of H_2 at the cathode and Cl_2 at the anode. The liquid is-
(A) Pure water (B) H_2SO_4 solution
(C) NaCl solution in water
(D) $CuCl_2$ solution in water

03. Which loses charge at cathode-
(A) Ions (B) Cations
(C) Anions
(D) Both anions and cations

04. In the electrolysis of $CuSO_4$, the reaction : $Cu^{2+} + 2e^{-} \rightarrow Cu$, takes place at :
(A) Anode (B) Cathode
(C) In solution (D) None

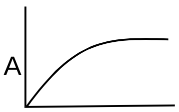
05. If mercury is used as cathode in the electrolysis of aqueous NaCl solution, the ions discharged at cathode are-
(A) H^+ (B) Na^+ (C) OH^- (D) Cl^-

Electrolytic conductance

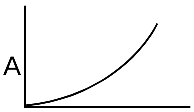
06. The specific conductance of a solution is 0.3568 ohm^{-1}. When placed in a cell the conductance is 0.0268 ohm^{-1}. The cell constant is-
(A) 1.331 cm^{-1} (B) 13.31 cm^{-1}
(C) 0.665 cm^{-1} (D) 6.65 cm^{-1}

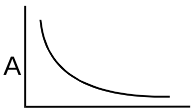
07. A conductance cell was filled with a 0.02 M KCl solution which has a specific conductance of 2.768×10^{-3} ohm^{-1} cm^{-1}. If its resistance is 82.4 ohm at 25ºC, the cell constant is-
(A) 0.2182 cm^{-1} (B) 0.2281 cm^{-1}
(C) 0.2821 cm^{-1} (D) 0.2381 cm^{-1}

08. The variation of equivalent conductance vs decrease in concentration of a strong electrolyte is correctly given in the plot -

(A) 

(B) 

(C) 

(D) 

09. Which of the following solutions has the highest equivalent conductance ?
(A) 0.01M NaCl (B) 0.050 M NaCl
(C) 0.005M NaCl (D) 0.02M NaCl

10. The resistance of 0.01N solution of an electrolyte AB at 328K is 100 ohm. The specific conductance of solution is (cell constant = $1cm^{-1}$)-
(A) 100ohm (B) 10^{-2} ohm^{-1}
(C) 10^{-2} ohm^{-1} cm^{-1} (D) 10^{2} ohm–cm

11. For an electrolytic solution of 0.05 mol L^{-1}, the conductivity has been found to be 0.0110 Scm^{-1}.The molar conductivity is-
(A) 0.055 S cm^2 mol^{-1} (B) 550 S cm^2 mol^{-1}
(C) 0.22 S cm^2 mol^{-1} (D) 220 S cm^2 mol^{-1}

12. Two electrodes are fitted in conductance cell 1.5 cm apart while the area of cross section of each electrode is 0.75 cm^2. The cell constant is-
(A) 1.125 (B) 0.5 cm
(C) 2.0 cm^{-1} (D) 0.2 cm^{-1}

13. The best conductor of electricity is in 1M solution of-
(A) CH_3COOH (B) H_2SO_4
(C) H_3PO_4 (D) Boric acid

Faraday's Law of Electrolysis

14. A certain current liberates 0.504 g of H_2 in 2 hours. How many grams of copper can be liberated by the same current flowing for the same time in $CuSO_4$ solution -
(A) 31.8 g (B) 16.0 g
(C) 12.7 g (D) 63.5 g

15. A currnet of 2.6 ampere is passed through $CuSO_4$ solution for 6 minutes 20 seconds. The amount of Cu deposited is (At. wt. of Cu = 63.5, Faraday = 96500 C)-
(A) 6.35 g (B) 0.635 g
(C) 0.325 g (D) 3.175 g

16. Three Faradays of electricity are passed through molten Al_2O_3, aqueous solution of $CuSO_4$ and molten NaCl taken in three different electrolytic cells. The amount of Al, Cu and Na deposited at the cathodes will be in the ratio of-
(A) 1 mole : 2 mole : 3 mole
(B) 1 mole : 1.5 mole : 3 mole
(C) 3 mole : 2 mole : 1 mole
(D) 1 mole : 1.5 mole : 2 mole

17. The quantity of electricity required to liberate 0.01g equivalent of an element at the electrode is -

(A) 9650C (B) 96500C
(C) 965C (D) 96.5C

18. The unit of electrochemical equivalent is-

(A) gm ampere^{-1} (B) gm/coulomb
(C) gm-ampere (D) coulomb/gram

19. One faraday of electricity will liberate one mole of metal from a solution of-

(A) $AuCl_3$ (B) $CuSO_4$(C) $BaCl_2$ (D) KCl

20. The number of faraday required to generate 1 mole of Mg from $MgCl_2$ is-

(A) 1 (B) 2 (C) 3 (D) 4

21. What weight of copper (At.mass 63.5) deposits when 2Faraday of electricity is passed through cupric salt-

(A) 63.5g (B) 31.75g (C) 127g (D) 2.0g

22. How many coulombs of electricity are required for the oxidation of 1mole of H_2O to O_2.

(A) 9.65×10^4C (B) 4.825×10^5C
(C) 1.93×10^5C (D) 1.93×10^4C

23. How long 2.5amp of current is passed to supply 54000 C of charge-

(A) 1hr (B) 2.5hr (C) 6hr (D) 9hr

24. 1 Faraday of electricity will liberate 1 mole of the metal from the solution of

(A) Copper sulphate (B) Calcium chloride
(C) Gold (III) Chloride (D) Silver (I) Chloride

25. Charge in coulambs is equal to-

(A) $\dfrac{\text{Faraday}}{\text{av. number}}$
(B) Faraday × av. number
(C) $\dfrac{\text{av.number}}{\text{Faraday}}$ (D) None of these

26. A current of 2 ampere was passed through solutions of $CuSO_4$ and $AgNO_3$ in series. 0.635 g of copper was deposited. Then the weight of silver deposited will be-

(A) 0.59 g (B) 3.24 g (C) 1.08 g(D) 2.16 g

27. On passing 3ampere of electricity for 50 minutes, 1.8g of metal deposits. The equivalent mass of metal is-

(A) 20.5 (B) 25.8 (C) 19.3 (D) 30.7

28. The number of coulombs required to deposit 5.4g of Al when the electrode reaction is-

$Al^{3+} + 3e^- \rightarrow Al$

(A) 1.83×10^5C (B) 57900C
(C) 5.86×10^5C (D) None of the above

Electro chemical series and Electrode potential

29. The reaction 1/2 H_2 (g) + AgCl (s) = H^+(aq) + Cl^- (aq) + Ag (s) can be represented in the galvanic cell as-

(A) Ag |AgCl (s)| KCl (sol) | | $AgNO_3$ (sol)| Ag
(B) PtH_2 (g) | HCl (sol) | | $AgNO_3$ (sol)
(C) Pt|H_2 (g) | HCl (sol) | | AgCl (s) | Ag
(D) Pt|H_2 (g) | KCl (sol) | | AgCl (s)| Ag

30. At 298 K, the standard reduction potentials for the following half reactions are given as

Zn^{2+} (aq) + $2e^- \rightarrow$ Zn (s); – 0.762
Cr^{3+} (aq) + $3e^- \rightarrow$ Cr (s); – 0.740
$2H^+$ (aq) + $2e^- \rightarrow H_2$ (g); 0.00
Fe^{3+} (aq) + $e^- \rightarrow Fe^{2+}$ (aq); + 0.770

The strongest reducing agent is -

(A) Zn (s) (B) H_2(g)
(C) Cr (s) (D) Fe^{2+}(aq)

31. The standard electrode potential of Zn, Ag and Cu are –0.76, 0.80 and 0.34 volt respectively; then -

(A) Ag can oxidise Zn and Cu
(B) Ag can reduce Zn^{2+} and Cu^{2+}
(C) Zn can reduce Ag^+ and Cu^{2+}
(D) Cu can oxidise Zn and Ag

32. An aqueous solution of $CuSO_4$ is stirred with a silver spoon. The following will happen -

(A) Cu^+ will be formed
(B) Ag^+ will be formed
(C) Cu will be precipitated
(D) Nothing will happen

33. Consider following half-cell reaction-

I. $A + e^- \rightarrow A^-$ E° = 0.96 V
II. $B^- + e^- \rightarrow B^{2-}$ E° = – 0.12 V
III. $C^+ + e^- \rightarrow C$ E° = + 0.18V
IV. $D^{2+} + 2e^- \rightarrow D$ E° = –1.12 V

What combination of two half-cells would result in a cell with the largest potential?

(A) I and II (B) I and III
(C) I and IV (D) II and IV

34. Standard E^0 of the half cell Fe | Fe^{2+} is + 0.44V and standard E^0 of half cell Cu | Cu^{2+} is – 0.32V, then-

(A) Cu oxidises Fe^{2+} ion
(B) Cu^{2+} oxidises Fe
(C) Cu reduces Fe^{2+} ion
(D) Cu^{2+} reduces Fe

Emf of the cell

35. Which of the following will increase the voltage of the cell

Sn (s) + 2Ag$^+$ (aq) ® Sn^{2+} (aq) + 2 Ag(s)-

(A) Increase in the concentration of Sn^{2+} ions

(B) Increase in the concentration of Ag$^+$ ions

(C) Increase in the size of silver rod

(D) None

36. The standard oxidation potentials, E^o for the half reactions are as Zn = Zn^{2+} + 2e, E^o = 0.76 V, Fe = Fe^{2+} + 2e, E^o = 0.41 V. The emf for the cell reaction Fe^{2+} + Zn $\rightarrow$ Zn^{2+} + Fe, is -

(A) – 0.35 V (B) + 0.35 V

(C) + 1.17 V (D) – 1.17 V

37. The single electrode potential E of 0.1 M solution of M$^+$ ions [E^o_R = – 2.36 V] is -

(A) + 2.41 (B) – 2.41

(C) – 4.82 (D) + 4.82

38. E^o values of Mg^{2+} |Mg, Fe^{2+}| Fe and Zn^{2+} | Zn are – 2.37 V, – 0.44 V and – 0.76 V respectively. The correct statement is -

(A) Mg oxidises Fe (B) Zn oxidises Fe

(C) Zn reduces Mg^{2+} (D) Zn reduces Fe^{2+}

39. For the reactions

$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$, E° = 1.51 V

$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$, E° = 1.23V

then for the reaction

$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$, E° is -

(A) 1.70 V (B) 5.09 V

(C) 0.28 V (D) 0.84 V

40. Consider the following equations for a cell reaction

$A + B \rightleftharpoons C + D$ E° = x volt, $K_{eq} = K_1$

$2A + 2B \rightleftharpoons 2C + 2D$E° = y volt,$K_{eq} = K_2$

then :

(A) $x = y,\ K_1 = K_2$ (B) $x = 2y,\ K_1 = 2K_2$

(C) $x = y,\ K_1^2 = K_2$ (D) $x^2 = y,\ K_1^2 = K_2$

41. Standard electrode potentials of

Fe^{2+} + 2e $\rightarrow$ Fe and Fe^{3+} + 3e $\rightarrow$ Fe

are – 0.440 V and – 0.036V respectively. The standard electrode potential (E^0) for Fe^{3+} + e $\rightarrow$ Fe^{2+} is -

(A) – 0.476 V (B) – 0.404 V

(C) + 0.404 V (D) + 0.772V

42. At equilibrium -

(A) $E^\circ_{cell} = 0,\ \Delta G^\circ = 0$

(B) $E_{cell} = 0,\ \Delta G = 0$

(C) both are correct

(D) none is correct

43. emf of cell Ni | Ni^{2+} (1.0M) || Au^{3+} (1.0M)| Au is ....., if E^0 for Ni^{2+} | Ni is – 0.25 V, E^0 for Au^{3+} | Au is 1.50V -

(A) + 1.25 V (B) – 1.75V

(C) + 1.75V (D) + 4.0V

44. Normal Al - AlCl$_3$ coupled with standard hydrogen electrode gives an emf of 1.66V. The standrad oxidation electrode potential of aluminium is-

(A) – 1.66V (B) + 1.66V

(C) – 0.83V (D) + 0.83V

45. The cell reaction Zn + Cu^{2+} $\rightarrow$ Zn^{2+} + Cu, is best represented by -

(A) Cu / Cu^{2+} || Zn^{2+} / Zn

(B) Zn / Zn^{2+} || Cu^{2+} / Cu

(C) Cu^{2+} / Cu || Zn / Zn^{2+}

(D) Pt / Zn^{2+} || Pt / Cu^{2+}

46. Consider a voltaic cell based on these half-cells

$Ag^+ (aq) + e^- \longrightarrow Ag(s)$; E^o = +0.80 V

$Cd^{2+} (aq) + 2e^- \longrightarrow Cd(s)$; E^o = –0.40 V

Identify the anode and give the voltage of the cell under standard conditions -

(A) Ag; E_{cell} = 0.40 V (B) Ag; E_{cell} = 2.00V

(C) Cd; E_{cell} = 1.20 V (D) Cd; E_{cell} = 2.00V

47. The value of the reaction quotient, Q, for the cell

Zn(s)|Zn^{2+}(0.01 M)||Ag$^+$(1.25 M)|Ag(s) is -

(A) 156 (B) 125

(C) 1.25×10^{-2} (D) 6.40×10^{-3}

Types of cell & corrosion

48. The Zn acts as sacrificial or cathodic protection to prevent rusting of iron because -

(A) E^0_{OP} of Zn < E^0_{OP} of Fe

(B) E^0_{OP} of Zn > E^0_{OP} of Fe

(C) E^0_{OP} of Zn = E^0_{OP} of Fe

(D) Zn is cheaper than iron

49. In electrochemical corrosion of metals, the metal undergoing corrosion -

(A) Acts as anode (B) Acts as cathode

(C) Undergoes reduction (D) None

50. When a lead storage battery is charged it acts as -

(A) A fuel cell (B) An electrolytic cell

(C) A galvanic cell

(D) A concentra tion cell

MISCELLANEOUS QUESTIONS

Q.1 On passing electricity through dilute H_2SO_4 solution the amount of substance liberated at the cathode and anode are in the ratio-

(A) 1 : 8 (B) 8 : 1 (C) 16 : 1 (D) 1 : 16

Q.2 When sodium chloride solution is electrolysed, the gas that is liberated at the cathode is-

(A) Oxygen (B) Hydrogen
(C) Chlorine (D) Air

Q.3 During electrolysis of an aqueous solution of Cu^{2+} sulphate, 0.635g of copper was deposited at cathode. The amount of electricity consumed in coulomb is-

(A) 1930 (B) 3860
(C) 9650 (D) 4825

Q.4 When electric current is passed through a cell having an electrolyte, the positive ions move towards the cathode and the negative ions towards the anode. If the cathode is pulled out of the solution -

(A) The positive and the negative ions both will move towards the anode
(B) The positive ions will start moving towards the anode: the negative ions will stop moving
(C) The negative ions will continue to move towards the anode: the positive ions will stop moving
(D) The positive ions and the negative ions will start moving randomly

Q.5 What products are formed during the electrolysis of a concentrated aqueous solution of sodium chloride?

I : $Cl_2(g)$ II : NaOH(aq) III : $H_2(g)$

(A) I only (B) I and II
(C) I and III (D) I, II and III

Q.6 When an electric current is passed through an aqueous solution of sodium chloride-

(A) H_2 is evolved at the anode
(B) Oxygen is evolved at the cathode
(C) Its pH progresively decreases
(D) Its pH progressively increases

Q.7 When aqueous solution of KCl is electrolysed, resultant solution -

(A) has pH > 7
(B) turns red litmus blue
(C) is basic in nature
(D) all of these

Q.8 In the electrolysis of $CuCl_2$ solution using Cu electrodes the mass of cathode increases by 3.18g. What happened at the other electrode-

(A) 0.05 mole of Cu^{2+} ions passed into solution
(B) 0.112 litre of Cl_2 was liberated
(C) 0.56 litre O_2 was liberated
(D) 0.1 mole of Cu^{2+} ions passed into the solution

Q.9 The molar conductances of HCl, NaCl and CH_3COONa are 426, 126 and 91 Ω^{-1} cm^2 mol^{-1} respectively. The molar conductance for CH_3COOH is-

(A) $561\Omega^{-1}$ cm^2 mol^{-1} (B) $391\Omega^{-1}$ cm^2 mol^{-1}
(C) $261\Omega^{-1}$ cm^2 mol^{-1} (D) $612\Omega^{-1}$ cm^2 mol^{-1}

Q.10 If x is the specific resistance of the solution and N is the normality of the solution. The equivalent conductivity of the solution is given by-

(A) $\frac{1000\,x}{N}$ (B) $\frac{1000}{N\,x}$ (C) $\frac{1000\,N}{x}$ (D) $\frac{N\,x}{1000}$

Q.11 According to Kohlrausch law, the limiting value of molar conductivity of an electrolyte A_2 B is

(A) $\lambda^{\infty}_{(A^+)}+\lambda^{\infty}_{(B^-)}$ (B) $\lambda^{\infty}_{(A^+)}-\lambda^{\infty}_{(B^-)}$
(C) $2\lambda^{\infty}_{(A^+)}+\frac{1}{2}\lambda^{\infty}_{(B^-)}$ (D) $2\lambda^{\infty}_{(A^+)}+\lambda^{\infty}_{(B^-)}$

Q.12 The same amount of electricity was passed through two cells containing molten Al_2O_3 and molten NaCl. If 1.8g of Al were liberated in one cell, the amount of Na liberated in the other cell is-

(A) 4.6 g (B) 2.3 g (C) 6.4 g (D) 3.2 g

Q.13 In which one of the following one Faraday of electricity will liberate 1/2 mole of the metal-

(A) $AuCl_3$ (B) $FeCl_3$ (C) $CuSO_4$ (D) NaCl

Q.14 How many coulombs of electricity are consumed when a 100 mA current is passed through a solution of $AgNO_3$ for half an hour during an electrolysis experiment-

(A) 1080 (B) 18000 (C) 180 (D) 2000

Q.15 Number of faraday required to liberate 8g of H_2 is-

(A) 8 (B) 16 (C) 4 (D) 2

Q.16 A certain current liberates 0.500 g of H_2 in 2.00 h. How many gram of oxygen can be liberated by the same current in the same time?

(A) 0.500 g (B) 8.00 g
(C) 4.00 g (D) 16.00 g

Q.17 The weight ratio of Al and Ag deposited using the same quantity of current is-
(A) 9 : 108 (B) 2 : 12
(C) 108 : 9 (D) 3 : 8

Q.18 20g of chlorine are evolved in 6 hour from sodium chloride solution by the current of-
(A) 5 ampere (B) 10 ampere
(C) 2.5 ampere (D) 50 ampere

Q.19 The reduction potential of the two half cell reactions (occuring in an electrochemical cell) are
$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$ (E^0 = - 0.31V)
$Ag^+ (aq) + e^- \rightarrow Ag(s)$ (E^0 = + 0.80V)
The feasible reaction will be -
(A) $Pb + SO_4^{2-} + 2Ag^+ (aq) \rightarrow 2Ag(s) + PbSO_4$
(B) $PbSO_4 + 2Ag^+ (aq) \rightarrow Pb + SO_4^{2-} + 2Ag(s)$
(C) $Pb + SO_4^{2-} + Ag (s) \rightarrow Ag^+ (aq) + PbSO_4$
(D) $PbSO_4 + Ag (s) \rightarrow Ag^+ (aq) + Pb + SO_4^{2-}$

Q.20 Consider the following electrolytes -
1. $AgNO_3$ 2. $CuSO_4$
3. $AlCl_3$ 4. $Bi_2(SO_4)_3$
The quantity of electricity needed to electrolytes separately 1 M solutions of these electrolytes will be (F is faraday) -
(A) 1F, 2F, 3F, 6F (B) 2F, 3F, 6F, 1F
(C) 3F, 6F, 2F, 1F (D) 6F, 3F, 2F, 1F

Q.21 When a copper wire is immersed in a solution of $AgNO_3$, the colour of the, solution becomes blue because copper-
(A) Forms a soluble complex with $AgNO_3$
(B) Is oxidized to Cu^{++}
(C) Is reduced to Cu^{--}
(D) Splits up into atomic form and dissolves

Q.22 Which of the following solution can be safely stored in a copper vessel-
(A) $ZnSO_4$ (B) $AgNO_3$
(C) $AuCl_3$ (D) All of them

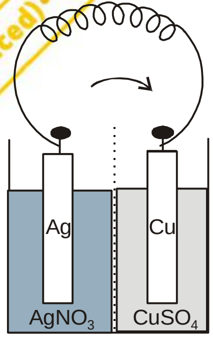
Q.23 Which one will liberate Br_2 from KBr-
(A) HI (B) I_2
(C) Cl_2 (D) SO_2

Q.24 Given
$Ag^+ / Ag \rightarrow + 0.80V$
$Co^{2+} / Co \rightarrow - 0.28V$
$Cu^{2+} / Cu \rightarrow + 0.34V$
$Zn^{2+} / Zn \rightarrow - 0.76 V$
The most reactive metal which displaces other metals from their salt solution is :
(A) Ag (B) Cu (C) Co (D) Zn

Q.25 A student made the following observations in the laboratory-
(i) Clean copper metal did not react with 1-molar $Pb(NO_3)_2$ solution
(ii) Clean lead metal dissolved in a 1-molar $AgNO_3$ solution and crystals of Ag metal appeared
(iii) Clean silver metal did not react with 1-molar $Cu(NO_3)_2$ solution
The order of decreasing reducing character of the three metals is-
(A) Cu, Pb, Ag
(B) Cu, Ag, Pb
(C) Pb, Cu, Ag
(D) Pb, Ag, Cu

Q.26 The electrode potential measures the -
(A) Tendency of the electrode to gain or lose electrons
(B) Tendency of a cell reaction to occur
(C) Difference in the ionisation potential of electrode and metal ion
(D) Current carried by an electrode

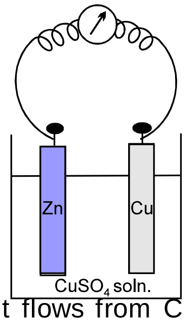
Q.27 In the arrangement shown, the arrow represents-



(A) Direction of flow of electrons
(B) Direction of flow of current
(C) No significant feature
(D) Movement of ions

Q.28 Red hot carbon will remove oxygen from the oxides XO and YO but not from ZO. Y will remove oxygen from XO. Use this evidence to deduce the order of activity of the three metals X, Y and Z putting the most reactive first-
(A) X, Y, Z
(B) Z, Y, X
(C) Y, X, Z
(D) Z, X, Y

Q.29 In the given arrangement, on pressing the key, the



(A) Current flows from Cu to Zn rod
(B) Current flows from Zn to Cu rod
(C) No flow of current occurs
(D) Electron flow occurs through wire from Zn to Cu.

Q.30 Point out the correct statement about Zn – $CuSO_4$ cell
(A) The flow of electrons occurs from copper to zinc
(B) The value of E^0_{Red} of copper electrode is less than that of zinc electrode.
(C) Zinc is anode while Cu is cathode electrode
(D) All the statements are correct

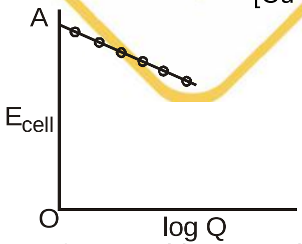
Q.31 The standard reduction potential of Pb and Zn electrodes are - 0.126 and - 0.763 volts respectively. The e.m.f. of the cell-
Zn / Zn^{2+} (0.1M) || Pb^{2+} (1M) / Pb is
(A) 0.637V (B) < 0.637V
(C) > 0.637V (D) 0.889

Q.32 The e.m.f. of the cell involving the reaction $2Ag^+(aq.) + H_2(g) \rightarrow 2Ag(s) + 2H^+(aq.)$ is 0.80 V. The standard oxidation potential of silver electrode is-
(A) 0.80V (B) – 0.80V (C) 0.40V (D) 0.20V

Q.33 For a reaction $A(s) + 2B^+ \rightarrow A^{2+} + 2B$
K_C has been found to be 10^{12}. The E^0_{cell} is-
(A) 0.354V (B) 0.708V
(C) 0.0098V (D) 1.36V

Q.34 $Zn + Cu^{2+} (aq) \rightleftharpoons Cu + Zn^{2+} (aq)$

Reaction quotient is $Q = \frac{[Zn^{2+}]}{[Cu^{2+}]}$.



Variation of E_{cell} with log Q is of the type with OA = 1.10 V. E_{cell} will be 1.1591 V when :

(A) $\frac{[Cu^{2+}]}{[Zn^{2+}]} = 0.01$ (B) $\frac{[Zn^{2+}]}{[Cu^{2+}]} = 0.01$

(C) $\frac{[Zn^{2+}]}{[Cu^{2+}]} = 0.1$ (D) $\frac{[Zn^{2+}]}{[Cu^{2+}]} = 1$

Q.35 For the cell reaction,

$Cu^{2+}(C_1) + Zn(s) \rightleftharpoons Zn^{2+}(C_2) + Cu(s)$

of the electrochemical cell, the change in free energy, at a given temperature is function of -
(A) ln (C_1) (B) ln (C_2/C_1)
(C) ln (C_2) (D) ln $(C_1 + C_2)$

Q.36 An electrochemical cell is set up as -
$Pt(H_2)|HCl(0.1M)||CH_3COO\ H(0.1M)|Pt(H_2)$
The e.m.f. of this cell will not be zero, because-
(A) the temperature is constant
(B) H^+ conc. of 0.1 M HCl and 0.1 M CH_3COOH is not same
(C) acids used in two compartments are different
(D) e.m.f depends on molarities of acids used

Q.37 The oxidation potential of hydrogen electrode will be greater than zero if-
(A) Conc. of H_3O^+ ions is 2M
(B) Conc. of H_3O^+ ions is 1M
(C) Partial pressure of H_2 is 2atm
(D) E_{oxi} can never be +ve

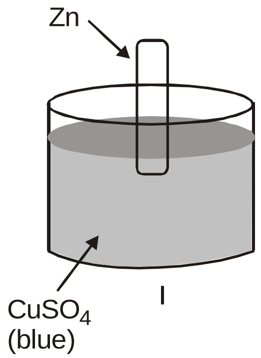
Q.38 The solution of nickel sulphate in which nickel rod is dipped is diluted to 10 times. The reduction potential of nickel-
(A) Decreases by 60mV
(B) Increases by 30mV
(C) Decreases by 30mV
(D) Decreases by 60V

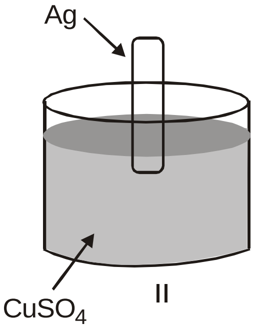
Q.39 The electrode potential of hydrogen electrode at the pH = 10 will be-
(A) 0 (B) +ve
(C) – ve (D) unpredictable

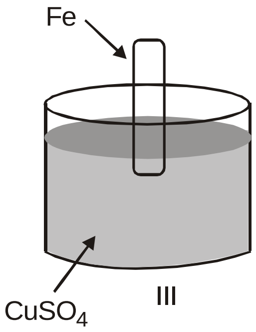
Q.40 $E^o_{Al^{3+}/Al} = -1.66$ V and K_{sp} of Al $(OH)_3$ $=1.0 \times 10^{-33}$. Reduction potential of the above couple at pH = 14 is :
(A) –2.31 V (B) + 2.31 V
(C) + 1.01 V (D) – 1.01 V

Q.41 Four colourless salt solutions are placed in separate test-tubes and a strip of copper is placed in each. Which solution finally turns blue-
(A) $Pb(NO_3)_2$ (B) $Zn(NO_3)_2$
(C) $AgNO_3$ (D) $Cd(NO_3)_2$

Q.42 Consider following sets



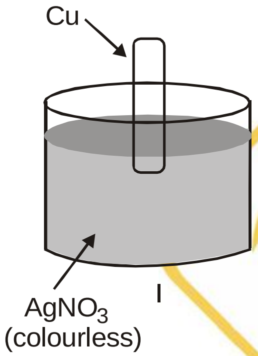


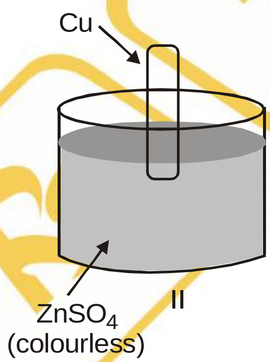


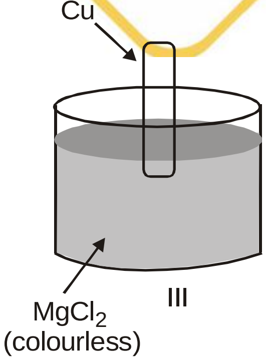
Blue colour solution changes to colourless (or fades) in -

(A) I, II, III (B) I, II (C) II, III (D) I, III

Q.43 Consider following sets :







Colourless solution changes to blue coloured solution in

(A) I (B) II (C) III (D) I, III

Q.44 The thermodynamic efficiency of fuel cell is given by -

(A) ΔH / ΔG (B) nFE / ΔG
(C) nFE / ΔH (D) nFzxE^0

Q.45 Which is correct representation for a cell at equilibrium.

(A) $\Delta G° = -2.303\ RT \log K_{eq.}$

(B) $E^0 = \dfrac{2.303RT}{nF} \log K_{eq.}$

(C) $-\Delta G° = RT \ln K_{eq.}$

(D) All

Q.46 In an electrochemical cell-

(A) Kinetic energy changes into potential energy
(B) Chemical energy changes into electrical energy
(C) Potential energy changes into kinetic energy
(D) Kinetic energy changes into chemical energy

Q.47 An example of a simple fuel cell is-

(A) Lead storage battery
(B) Laclanche cell
(C) $H_2 - O_2$ cell
(D) All

Q.48 The function of the salt bridge is to -

(A) Allow solutions of two half cells to intermix
(B) Does not allow the ions to move from anode to cahthode
(C) Keep the solution electrically neutral in two half cells
(D) None of the above

Q.49 Which one of the following does not hold good for S.H.E.-

(A) The pressure of hydrogen gas is 1.5 atmosphere
(B) The concentration of H$^+$ in solution is 1M
(C) The temperature is 298K
(D) The surface of platinum electrode is coated with platinum black

Q.50 The value of equilibrium constant for a feasible cell reaction is-

(A) < 1 (B) Zero
(C) = 1 (D) > 1

◆ ◆ ◆ ◆ ◆ ◆

ANSWER KEY

TOPIC WISE MCQS

Q.No.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	A	C	B	B	B	B	B	A	C	C	D	C	B	B	C
Q.No.	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
Ans.	B	C	B	D	B	A	C	C	D	A	D	C	B	C	A
Q.No.	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45
Ans.	C	D	C	B	B	A	B	D	A	C	D	B	C	B	B
Q.No.	46	47	48	49	50										
Ans.	C	D	B	A	B										

MISCELLANEOUS QUESTIONS

Ques.	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15
Ans.	A	B	A	D	D	D	C	A	B	B	D	A	C	C	A
Ques.	16	17	18	19	20	21	22	23	24	25	26	27	28	29	30
Ans.	C	A	C	A	A	B	A	C	D	C	A	B	B	C	C
Ques.	31	32	33	34	35	36	37	38	39	40	41	42	43	44	45
Ans.	C	B	A	B	B	B	C	C	C	A	C	D	A	C	D
Ques.	46	47	48	49	50										
Ans.	B	C	C	A	D										

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