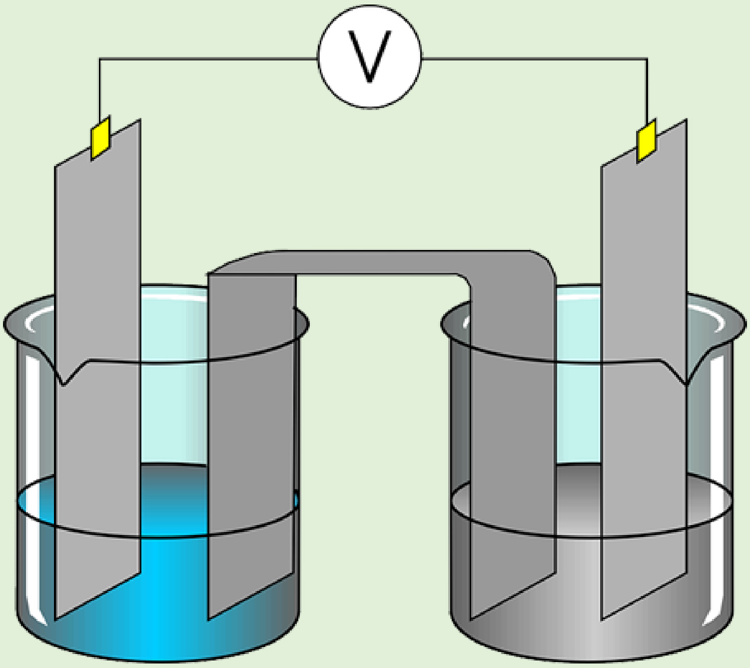
3. ELECTROCHEMISTRY



Chemistry Smart Booklet

Theory + NCERT MCQs + Topic Wise Practice MCQs + NEET PYQs

Electrochemistry

Electrode Potential

Potential difference between electrode and electrolyte.

$E_{cell} = E^0_{Righ} - E^0_{left}$

Also known as Deniel Cell:

Cathode: Copper
Anode: Zinc
Salt bridge: Agar-Agar
Electrolyte: $ZnSO_4$, $CuSO_4$
Cell reaction:
$Zn + CuSO_4 \longrightarrow ZnSO_4 + Cu$
Cell representation:
$Zn(s) | ZnSO_4(Sol) || CuSO_4(Sol) | Cu(s)$

Salt Bridge: U shaped inverted tube connecting two electrolyte solution

Anode: + vely charged Oxidation takes place

Cathode: + vely charged reduction takes place

Galvanic/Voltic Cell

Faraday Laws

Quantitative aspects

Faraday 1st Law

- Amount of chemical reactions which occurs at any electrode during electrolysis by a current is proportional to the quantity of electricity passed through electrolyte → w = Zit

Faraday 2nd Law

- Amount of substance deposited at electrodes during electrolysis is proportional to their chemical equivalents weights → $\frac{w_1}{E_1} = \frac{w_2}{E_2} = \frac{w_3}{E_3}$

Electrolytic Cell

Qualitative aspect depend on

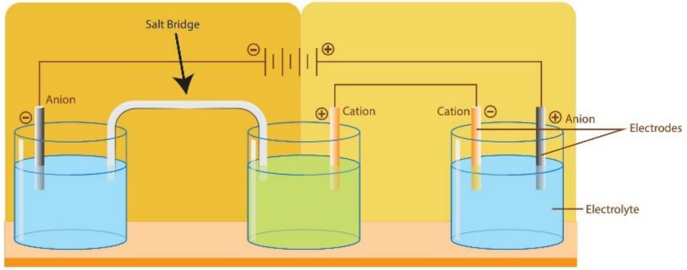
- Nature of Electrode
- Medium of Electrolyte

Device converting electrical energy into chemical energy.

Anode → + vely charged: oxidation occurs
Cathode → - vely charged: Reduction occurs

Cell representation

Oxidation half / Reduction half



Corrosion

Electro chemical phenomenon in which metal oxide or metal forms coating on metal surface.

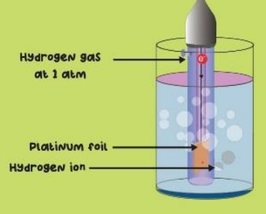
Prevention

- Painting, barrier protection, rust solution.

Example

- Rusting of iron.
- Furnishing of silver.

Standard Hydrogen Electrode (SHE)



For SHE, $E^\circ_{cell} = 0$

	Reduction Half-Reaction		E°(V)	
Stronger oxidizing agent	$F_2(g) + 2e^-$	$\longrightarrow 2F^-(aq)$	2.87	weaker reducing agent
	$H_2O_2(aq) + 2H^+(aq) + 2e^-$	$\longrightarrow 2\ H_2O\ (l)$	1.78	
	$MnO_4^-\ (aq) + 8H^+(aq) + 5e^-$	$\longrightarrow Mn^{2+}(aq) + 4H_2O(l)$	1.51	
	$Cl_2(g) + 2e^-$	$\longrightarrow 2Cl^-(aq)$	1.36	
	$Cr_2O_7^{2-}\ (aq) + 14H^+(aq) + 6e^-$	$\longrightarrow 2Cr^{3+}\ (aq) + 7H_2O\ (l)$	1.33	
	$O_2(g) + 4H^+(aq) + 4e^-$	$\longrightarrow 2H_2O(l)$	1.23	
	$Br_2(aq) + 2e^-$	$\longrightarrow 2Br^-(aq)$	1.09	
	$Ag^+(aq) + e^-$	$\longrightarrow Ag(s)$	0.80	
	$Fe^{3+}(aq) + e^-$	$\longrightarrow 2Fe^{2+}\ (aq)$	0.77	
	$O_2(g) + 2H^+(aq) + 2e^-$	$\longrightarrow H_2O_2\ (aq)$	0.70	
	$I_2(s) + 2e^-$	$\longrightarrow 2I^-(aq)$	0.54	
	$O_2(g) + 2H_2O(l) + 4e^-$	$\longrightarrow 4OH^-(aq)$	0.40	
	$Cu^{2+}(aq) + 2e^-$	$\longrightarrow Cu(s)$	0.34	
	$Sn^{4+}(aq) + 2e^-$	$\longrightarrow Sn^{2+}(aq)$	0.15	
	$2H^+(aq) + 2e^-$	$\longrightarrow H_2(g)$	0	
	$Pb_2+(aq) + 2e$	$\longrightarrow Pb(s)$	0.13	
	$Ni^{2+}(aq) + 2e^-$	$\longrightarrow Ni(s)$	-0.26	
	$Cd^{2+}(aq) + 2e^-$	$\longrightarrow Cd(s)$	-0.40	
	$Fe^{2+}(aq) + 2e^-$	$\longrightarrow Fe(s)$	0.45	
	$Zn^{2+}(aq) + 2^{e-}$	$\longrightarrow Zn(s)$	-0.76	
	$2H_2O(l) + 2e^-$	$\longrightarrow H2(g) + 2OH^-(aq)$	-0.83	
	$Al^{3+}(aq) + 3e^-$	$\longrightarrow Al(s)$	1.66	
weaker oxidizing agent	$Mg^{2+}(aq) + 2e^-$	$\longrightarrow Mg(s)$	-2.37	Stronger reducing agent
	$Na^+(aq) + e^-$	$\longrightarrow Na(s)$	-2.71	
	$Li^+(aq) + e^-$	$\longrightarrow Li(s)$	-3.04	

The arrangement of various electrodes in the increasing order of standard reduction potentials.

Nernst Equation

- For reaction: $M^{n+} + ne^- \longrightarrow M(s)$

$$E = E^\circ - \frac{2.303\ RT}{nF} \log \frac{1}{[M^{n+}]}$$

- For reaction: $aA + bB \longrightarrow cC + dD$

$$E_{cell} = E^\circ_{cell} - \frac{2.303\ RT}{nF} \log \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

- A Equilibrium $E_{cell} = 0$

$$E^\circ_{cell} = \frac{2.303\ RT}{nF} \log K_c$$

- $\Delta nG = nFE_{cell}$ or $\Delta nG = -NFE_{cell}$

Electrical Properties

- Conductance (G) = $\frac{1}{\text{Resistance}}$
  Unit: Ohm^{-1} or Siemens
- Specific conductivity (K) = $G\frac{l}{a}\left(\frac{l}{a} = \text{cell constant}\right)$
  Unit: Ohm^{-1} cm-1 or S cm^{-1}
- Molar conductance (Λ_m) = $\frac{1000 \times K}{M}$
  Unit: Scm-1mol^{-1}
- Equivalent conductance (Λ_{eq}) = $\frac{1000 \times K}{N}$
  Unit: $Cm^2\ ohm^{-1}\ g\text{-}eq^{-1}$

Kohlrausch's law of Independent Migration of Ions

At infinite dilution the molar conductivity of electrolytic is given by sum of ionic conductivities of cation & anions.

$$\Lambda^0_m = \nu_+\lambda^0_+ + \nu_-\lambda^0_-$$

APPLICATION

1. To determine Λ_m & Λ_{eq} of weak electrolytes at infinite dilution.
2. To calculate degree of dissociation: (α) $\alpha = \frac{\Lambda^c_m}{\Lambda^0_m}$
3. To calculate dissociation constant of weak Electrolyte: $k_c = \frac{c\alpha^2}{1-\alpha}$

Battery

(Primary)
(reaction occurs only once cannot be reuse)

Mercury
Anode: Zn-Hg
Cathode: Paste HgO & C
Electrolyte: Paste of KOH + ZnO

Leclanche cell (dry cell)
Zn
Graphite
Powdered MnO_2 + C +
Paste of NH4Cl + $ZnCl_2$

(Secondary)
(Can be reacharged by parring current f in opposite direction.)

Lead Storage battery Ni-cd cell
Pb
Pb + PBO
H_2SO_4 (38% by loss)

ELECTROCHEMISTRY

Conductivity of Solutions

What is Conductivity?

- Electrical resistance is represented by the symbol 'R', and it is measured in ohm (Ω).
- The electrical resistance of any object is directly proportional to its length (l) and inversely proportional to its area of cross-section (A), i.e.

$$R \propto \frac{l}{A} \qquad \text{OR} \qquad R = \rho \frac{l}{A}$$

where the constant of proportionality ρ is called resistivity (specific resistance).

- The inverse of resistance, R, is called conductance, G, and we have the relation

$$G = \frac{1}{R} = \frac{A}{\rho l} = \kappa \frac{A}{l}$$

where the constant □ is called conductivity (specific conductance).

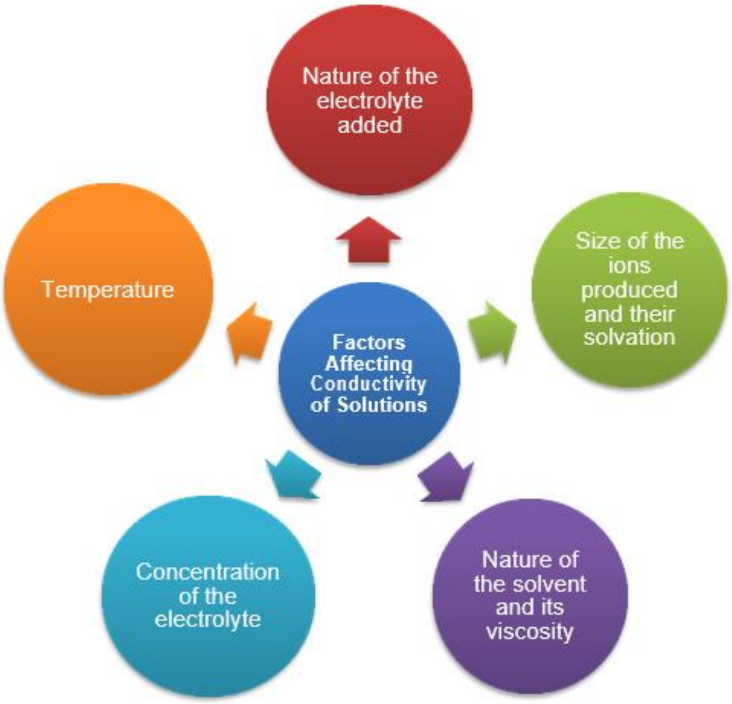
- The SI unit of conductance is Siemens, represented by the symbol 'S', and it is equal to ohm^{-1} (also known as mho) or Ω^{-1}. The SI unit of conductivity (□) is $S\ m^{-1}$.

Conductivity of Electrolytic (Ionic) Solutions

- Very pure water has small amounts of hydrogen and hydroxyl ions (~10^{-7}M) which lend it very lowconductivity ($3.5 \times 10^{-5}\ S\ m^{-1}$).
- When electrolytes are dissolved in water, they dissociate to give their own ions in the solution; hence, its conductivity also increases.

Electrolytic or ionic conductance:

Conductance of electricity by ions present in solutions



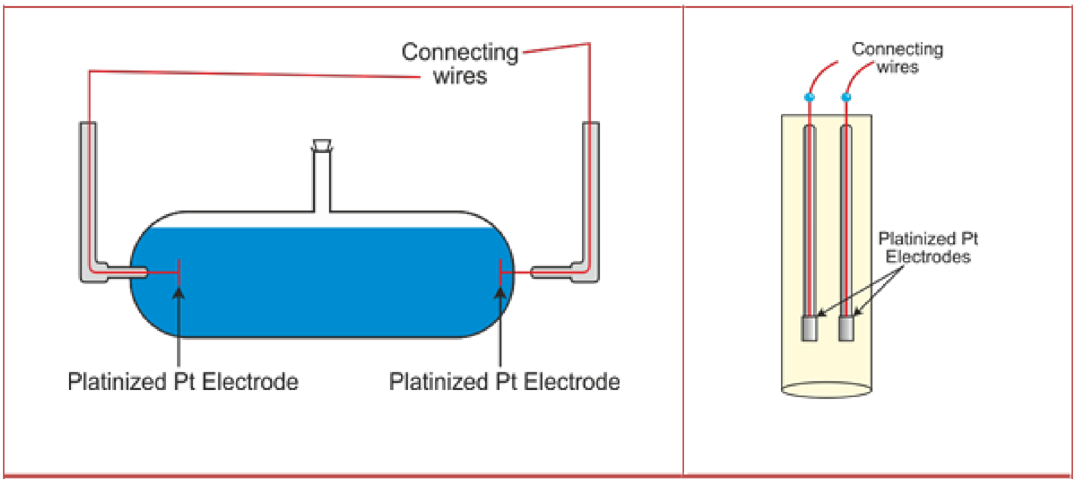
Measurement of Conductivity of Ionic Solutions

- We first need to find the resistance of an ionic solution to measure conductivity.
- We face two problems when measuring the resistance of an ionic solution:

1. Passing direct current (DC) changes the composition of the solution
2. A solution cannot be connected to the bridge like a metallic wire or other solid conductor

- The first difficulty is resolved by using an alternating current (AC) source of power. The second problem is solved by using a specially designed vessel called **conductivity cell**.

➢ Conductivity Cell:



Two different types of conductivity cells

- It consists of two platinum electrodes coated with platinum black (finely divided metallic Pt is deposited on the electrodes electrochemically).
- These have area of cross-section equal to 'A' and are separated by distance 'l'. The resistance of such a column of solution is then given by the equation:

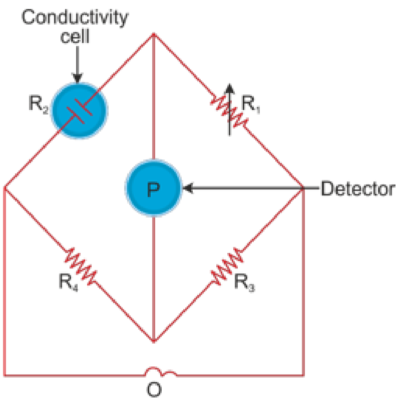
$$R = \rho \frac{l}{A} = \frac{l}{\kappa A}$$

- The quantity 'l/A' is called cell constant and is denoted by the symbol G*. It depends on the distance between the electrodes and their area of cross-section and has the dimension $[L^{-1}]$.
- Hence, the cell constant G* is given by the equation:

$$G^* = \frac{l}{A} = R\kappa$$

- When the cell constant is determined, we can use it for measuring the resistance or conductivity of any solution.
- The set up for the measurement of the resistance is nothing but the well-known

Wheatstone bridge.



Arrangement for measurement of resistance of a solution of an electrolyte

- It consists of two resistances R_3 and R_4, a variable resistance R_1 and the conductivity cell having the unknown resistance R_2.
- The Wheatstone bridge is fed by an oscillator O (a source of AC power in the audio frequency range 550−5000 cycles per second).
- P is a suitable detector (a headphone or other electronic device).
- The bridge is balanced when no current passes through the detector.
- Under these conditions,

$$\text{Unknown Resistance, } R_2 = \frac{R_1 R_4}{R_3}$$

- When the cell constant and the resistance of the solution in the cell are determined, the conductivity of the solution is given by the equation:

$$\kappa = \frac{\text{Cell Constant}}{R} = \frac{G^*}{R}$$

Molar Conductivity

- Conductivity of solutions of different electrolytes in the same solvent and at a given temperature differs due to

1. Charge and size of ions in which they dissociate
2. Concentration of ions or ease with which the ions move under a potential gradient

- Therefore, it becomes necessary to define a quantity called molar conductivity denoted by the symbol (λ_m). It is related to the conductivity of the solution by the equation:

$$\text{Molar Conductivity} = \lambda_m = \frac{\kappa}{c}$$

- Unit of λm is in S m^2 mol^{-1}.
- Hence, molar conductivity can be given by the formula

$$\lambda_m\left(S m^2 mol^{-1}\right)=\frac{\kappa\left(S m^{-1}\right)}{1000\,L\,m^{-3}\times Molarity\left(mol\,L^{-1}\right)}$$

Variation of Conductivity and Molar Conductivity with Concentration

- Both conductivity and molar conductivity change with the concentration of the electrolyte.
- Conductivity always decreases with a decrease in concentration for both weak and strong electrolytes. It is because the number of ions per unit volume which carry the current in a solution decreases on dilution.
- Molar conductivity increases with a decrease in concentration. This is because the total volume (V) of solution containing one mole of electrolyte also increases.
- The decrease in κ on dilution of a solution is more than compensated by an increase in its volume.

Molar conductivity (λ_m): Conductance of the electrolytic solution kept between the electrodes of a conductivity cell at unit distance but having area of cross section large enough to accommodate sufficient volume of solution which contains one mole of the electrolyte.

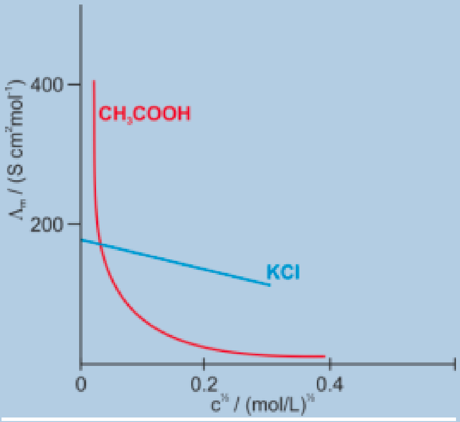
- When concentration approaches zero, the molar conductivity is known as limiting molar conductivity and is represented by the symbol λ^o_m.
- The variation in λ_m with concentration is different for strong and weak electrolytes.

➢ Strong Electrolytes:

- λ increases slowly with dilution and can be represented by the equation:

$$\lambda_m=\lambda^o_m-Ac^{1/2}$$

- It can be seen that if we plot λm against $c^{1/2}$, we obtain a straight line with intercept equal to λ^o_m and slope equal to '−A'.



Molar conductivity λ_m versus $c^{1/2}$ for acetic acid (weak electrolyte) and potassium chloride (strong electrolyte) in aqueous solutions

- The value of the constant 'A' for a given solvent and temperature depends on the type of electrolyte.

➢ Weak Electrolytes:

- Weak electrolytes such as acetic acid have a lower degree of dissociation at higher concentrations. Hence, for such electrolytes, the change in λ_m with dilution is due to

1. Increase in the degree of dissociation
2. The number of ions in total volume of solution which contains 1 mol of electrolyte

- In such cases, λ^o_m increases steeply on dilution, especially near lower concentrations. Therefore, it cannot be obtained by extrapolating λ_m to zero concentration.
- At infinite dilution, electrolyte dissociates completely ($\alpha = 1$), but at such low concentration, the conductivity of the solution is so low that it cannot be measured accurately. Therefore, λ^o_m for weak electrolytes is obtained by using Kohlrausch's law of independent migration of ions.
- Thus, at any concentration c, if α is the degree of dissociation, then it can be approximated to the ratio of molar conductivity, λ_m, at the concentration c to limiting molar conductivity λ^o_m.

$$\therefore \alpha = \frac{\lambda_m}{\lambda^0_m}\;{}_{m}$$

- But we know that for a weak electrolyte,

$$K_a = \frac{C\alpha^2}{(1-\alpha)}$$

- By putting the value of α in the above equation, we get the equation:

$$K_a = \frac{c\lambda^2_m}{\lambda^0_m\left(1-\frac{\lambda_m}{\lambda^0_m}\right)} = \frac{c\lambda^2_m}{\lambda^0_m\left(\lambda^0_m-\lambda_m\right)}$$

Numerical

1) The resistance of a conductivity cell containing 0.001M KCl solution at 298 K is 1500Ω. What is the cell constant if conductivity of 0.001M KCl solution at 298K is 0.146×10^{-3} Scm^{-1}?

Answer :

Given Data: R =1500Ω ; κ=0.146 × 10^{-3} Scm^{-1}

G^{*} = ?

Solution :

The cell cons tant is givenby the equation,

Cell constant=G^{*}=Rκ = 1500× 0.146 × 10^{-3}= 0.219 m^{-1}

Hence, cell constant of given conductivity cell is 0.219 m^{-1} .

2) Conductivity of 0.00241 M acetic acid is 7.896 × 10^{-5} Scm^{-1}. Calculate its molar conductivity and if λ_m^o for acetic acid is 390.5 Scm^2 mol^{-1}, what is its dissociation constant?

Answer:

Given Data:

$\kappa = 7.896 \times 10^{-5}\ Scm^{-1}$

$c = 0.00241\ M$

$\lambda_m^o = 390.5\ Scm^2\,mol^{-1}$

$\lambda_m = ?$

$K_a =$

Solution:

The molar conductivity can be calculated by formula,

$$\lambda_m = \frac{\kappa}{c}$$

$$= \frac{7.896 \times 10^{-5}}{0.00241}$$

$$\therefore \lambda_m = 327.63 \times 10^{-4}\ Sm^2\,mol^{-1}$$

The dissociation constant of acetic acid can be given by,

$$K_a = \frac{c\lambda_m^2}{\lambda_m^0\left(\lambda_m^0 - \lambda_m\right)}$$

$$= \frac{0.00241 \times \left(327.63 \times 10^{-4}\right)^2}{390.5\left(390.5 - 327.63 \times 10^{-4}\right)}$$

$$\therefore K_a = 1.696 \times 10^{-11}\ molL^{-1}$$

Hence, the molar conductivity of given solution is $327.63\square 10^{\square 4} Sm^2 mol^{-1}$ and dissociation cons tant is 1.696 x10^{-11} $molL^{-1}$.

NCERT LINE BY LINE QUESTIONS

(1.) In hydrogen-oxygen fuel cell, combustion of hydrogen occurs to **[QR code, Page: 90]**

(a.) Produce high purity water

(b.) Create potential difference between two electrodes

(c.) generate hear

(d.) remove adsorbed oxygen from electrode surface

(2.) Fluorine is a best oxidising agent because it has

(a.) highest electron affinity

(b.) highest E°_{red}

(c.) highest E^o_{oxid}

(d.) lowest electron affinity

(3.) (180.) E° for the half cell Zn^{2+}/Zn is -0.76 emf of the cell $Zn/Zn^{2+}(1M) \| 2H^+(1M)|H_2(1\text{ atm})$ is

(a.) $-0.76V$

(b.) $+0.76$

(c.) $-0.38V$

(d.) $+0.38V$

(4.) The electrolyte used in Leclanche cell is **[QR code, Page: 88]**

(a.) Paste of KOH and ZnO

(b.) 38% solution of H_2SO_4

(c.) Moist paste of NH_4Cl and $ZnCl_2$

(d.) Moist $NaOH$

(5.) **Assertion:** $E_{Ag^+/Ag}$ increases with increase in concentration of Ag^+ ions.

Reason: $E_{Ag^+/Ag}$ has a positive value. **[Page: 71]**

(a.) Both and R are true and R is correct explanation of A.

(b.) Both A and R are true and R is not correct explanation of A.

(c.) A is true but R is false.

(d.) Both A and R are false.

(6.) Hydrogen electrode is placed in the solution where pH is 10. The potential of this electrode will be **[Page: 73]**

(a.) +0.591 V

(b.) −0.591 V

(c.) 0 V

(d.) none of these

(7.) For a reaction **[Page: 73]**

$A(s)+2B^+ \rightarrow A^{2+}+2B(s)$

K_C has been found to be 10^{12}. The E^o_{cell} is

(a.) 0.354 V

(b.) 0.708 V

(c.) 0.0098 V

(d.) 1.36 V

(8.) The order of equivalent conductance at infinite dilution of $LiCl$, $NaCl$ and KCl?

(a.) $LiCl > NaCl > KCl$

(b.) $KCl > NaCl > LiCl$

(c.) $NaCl > KCl > LiCl$

(d.) $LiCl > KCl > NaCl$

(9.) The reference electrode is made by using **[Page: 73]**

(a.) $ZnCl_2$ (b.) $CuSO_4$

(c.) $HgCl_2$ (d.) Hg_2Cl_2

(10.) Which one of the following is a strong electrolyte? **[Page: 81]**

(a.) $Ca(NO_3)_2$ (b.) HCN

(c.) H_2SO_3 (d.) NH_4OH

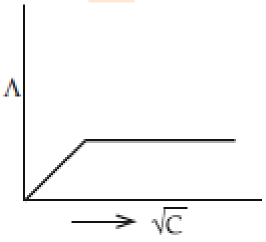
(11.) Which of the following decrease on dilution of electrolyte solution? **[Page: 81]**

(a.) equivalent conductance (b.) molar conductance

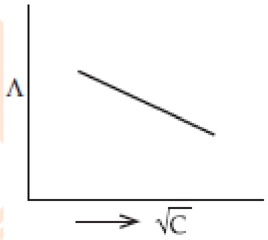
(c.) specific conductance (d.) conductance

(12.) The variation of equivalent conductance of strong electrolyte with concentration is correctly shown in which figure: **[QR code, Page: 82]**

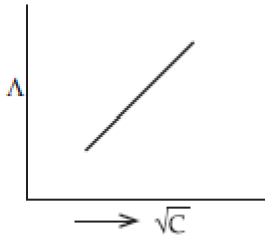
(a.)



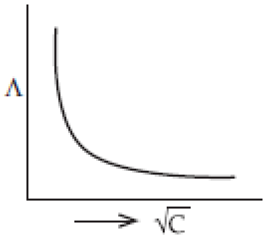
(b.)



(c.)



(d.)



(13.) The quantity of charge required to obtain one mole of aluminum from Al_2O_3 is **[NCERT Exemplar, Page: 86]**

(a.) 1 P (b.) 6 F

(c.) 3 F (d.) 2 F

(14.) Electrolysis of H_2SO_4 (conc.) gives the following at anode? **[Page: 87, HOTS]**

(A) H_2 **(B)** O_2 **(C)** $H_2S_2O_3$ **(D)** $H_2S_2O_8$

(a.) $\Lambda = \frac{X \times 1000}{0.1}$ (b.) $\Lambda = \frac{2 \times 1000}{X \times 0.1}$

(c.) $\Lambda = \frac{X \times 1000}{0.5}$ (d.) $\Lambda = \frac{0.5}{1000X}$

(15.) Using the data given below find out the strong oxidising agent **[NCERT Exemplar, Page: 71]**

$E^{\circ}_{Cl_2O_7^{2}/Cr^{3+}} = 1.33V$ $E^{o}_{Cl_2/Cl^-} = 1.36V$

$E^{o}_{MnO_4/Mn^{2+}} = 1.51V$ $E^{o}_{Cr^{3+}/Cr} = -0.74V$

(a.) Cl^-

(b.) Mn^{2+}

(c.) MnO_4^-

(d.) Cr^{3+}

(16.) E^{o} for the cell $Zn\,|\,Zn^{2+}(aq)\,\|\,Cu^{2+}(aq)/Cu$isl.10V at $25^{\circ}C$ is 1. 10V at$25^{o}C$. The equilibrium constant for the cell reaction **[Page: 73]**

$Zn+Cu^{2+}(aq) \rightleftharpoons Cu+Zn^{2+}(aq)$ is of the order of

(a.) 10^{-37} (b.) 10^{37}

(c.) 10^{-17} (d.) 10^{17}

(17.) When dil.HNO_3 is electrolysed **[Page: 87]**

(a.) $H_2(g)$ is formed at anode (b.) $O_2(g)$ is formed at anode

(c.) NO_2 is formed at anode (d.) N_2 is formed at anode

(18.) The equivalent conductivity of a solution containing 2.54 g of $CuSO_4$ per litre is $91\Omega^{-1}cm^{2}eq^{-1}$. Its conductivity would be

(a.) $2.9\times10^{-3}\Omega^{-1}cm^{-1}$ (b.) $1.8\times10^{-2}\Omega^{-1}cm^{-1}$

(c.) $2.4\times10^{-4}\Omega^{-1}cm^{-1}$ (d.) $3.6\times10^{-3}\Omega^{-1}cm^{-1}$

(19.) Prevention of corrosion of iron by zinc coating is called **[Page: 91]**

(a.) electrolysis (b.) Photoelectrolysis

(c.) cathodic protection (d.) galvanization

(20.) Which of the following is not an anodic reaction?

(a.) $Ag^{+}\rightarrow Ag-e^{-}$ (b.) $Cu\rightarrow Cu^{2+}+2e^{-}$

(c.) $Fe^{2+}\rightarrow Fe^{3+}+e^{-}$ (d.) $4OH^{-}\rightarrow 2H_2O+O_2+4e^{-}$

(21.) When 0.1 mol MnO_4^{2-} is oxidised the quantity of electricity required to completely oxidise MnO_4^{2-} to MnO_4^{-} is **[NEET-2014, Page: 8O]**

(a.) 96500 C (b.) $2\times96500C$

(c.) 9650 C (d.) 96.50 C

(22.) The oxidation potential of a hydrogen electrode at $pH=1$ is $(T=298K)$ **[Page: 69]**

(a.) 0.059 V (b.) O

(c.) –0.059 V (d.) 0.59 V

(23.) On the basis of E° values, the strongest oxidizing agent is **[Page: 71]**

$\left[Fe(CN)_6\right]^{4}\rightarrow\left[Fe(CN)_6\right]^{3-}+e^{-};E^{\circ}=-0.35V$

$Fe^{2+}\rightarrow Fe^{3+}+e^{-};E^{\circ}=-0.77V$

(a.) Fe^{3+}

(b.) $\left[Fe(CN)_6\right]^{3-}$

(c.) $\left[Fe(CN)_6\right]^{4-}$

(d.) Fe^{2+}

(24.) Molar conductance $\left(\Lambda_m^o\right)$ at infinite dilution of $NaCl, HCl$ and CH_3COONa are 126.4, 425.9 and 91.0 Scm^2mol^{-1} resp. **[Page: 83]**
Λ_m^o for CH_3COOH will be

(a.) $425.5 Scm^2 mol^{-1}$

(b.) $180.5 Scm^2 mol^{-1}$

(c.) $290.8 Scm^2 mol^{-1}$

(d.) $390.5 Scm^2 mol^{-1}$

(25.) The passage of electricity in the Daniel cell when Zn and Cu electrodes are connected is: **[Page: 66]**

(a.) From Cu to Zn in the cell

(b.) From Cu to Zn outside the cell

(c.) From Zn to Cu outside the cell

(d.) In any direction of cell

(26.) KCl is used in salt bridge because **[Page: 66]**

(a.) it forms a good jelly with agar-agar

(b.) it is a strong electrolyte

(c.) it is a good conductor of electricity

(d.) the transference number of K^+ and Cl^- ions are almost equal

(27.) The cell constant of a conductivity cell — **[NCERT Exemplar, Page: 78]**

(a.) changes with change of electrolyte

(b.) changes with change of concentration of electrolyte

(c.) change with temperature of electrolyte

(d.) remains constant for a cell

(28.) The number of electrons delivered at the cathode during the electrolysis by a current of 1 A in 60 seconds is (charge one electron $= 1.6\times10^{-19}$C) **[Page: 86]**

(a.) 6×10^{23}

(b.) 6×10^{20}

(c.) 3.75×10^{20}

(d.) 7.48×10^{23}

(29.) The specific conductance in $ohm^{-1}cm^{-1}$ of four electrolytes P, Q, R and S are given below:
$P = 5\times10^{-5}$
$Q = 7\times10^{-8}$
$R = 1\times10^{-10}$
$S = 9.2\times10^{-3}$
The one that offers highest resistance to the passage of current is

(a.) P

(b.) S

(c.) R

(d.) Q

(30.) Which of the following is a cathode $R\times N$?

(a.) $Fe^{2+} \to Fe^{3+}$

(b.) $4OH^- \to 2H_2O + O_2$

(c.) $2H_2O \to 2OH^- + H_2$

(d.) $2SO_4^{2-} \to S_2O_8^{2-}$

(31.) If the E_{cell}^o for a given reaction has negative value, which of the following gives the correct relationship for the values of ΔG° and K_{eq}? **[NEET 2016, Page: 74]**

(a.) $\Delta G^\circ >, K_{eq} < 1$

(b.) $\Delta G^o > 0; K_{eq} > 1$

(c.) $\Delta G^\circ < 0; K_{eq} > 1$

(d.) $\Delta G^o < 0; K_{eq} < 1$

(32.) Cell reaction is spontaneous when

(a.) ΔG° is negative
(b.) ΔG^o is positive
(c.) ΔE^o_{red} is positive
(d.) ΔE^o_{red} is negative

(33.) For a given half cell: $Al^{3+} + 3e^- \rightarrow Al,\cdot$ on increasing $\left[Al^{3+}\right]$, the electrode potential **[Page: 72]**

(a.) increase
(b.) decrease
(c.) No change
(d.) first decrease then increase

(34.) **Assertion:** Mercury cell does not give steady potential.
Reason: In the cell reaction, ions are not involved in solution.
[NCERT Exemplar, Page: 88]

(a.) Both A and R are true but R is correct explanation of A.
(b.) Both A and R are true and R is not correct explanation of A.
(c.) A is true bur R is false.
(d.) A is false but R is true.

(35.) Which of the following statements is not correct about an inert electrode in a cell **[NCERT Exemplar, Page: 69]**

(a.) It does not participate in the cell reaction.
(b.) It provides the surface either for oxidation or reduction reaction.
(c.) It provides surface for the conduction of electrons.
(d.) It provides surface for redox reaction.

(36.) The standard oxidation potentials for the half cell reactions are

$Zn \rightarrow Zn^{2+} + 2e^-; E^o = +0.76V$

$Pe \rightarrow Fe^{2+} + 2e^-; E^o = +0.41V$

The emf of the cell reaction:

$Fe^{2+} + Zn \rightarrow Zn^{2+} + Fe$ is

(a.) $-0.35V$
(b.) $+0.35V$
(c.) $+1.17V$
(d.) $-1.17V$

(37.) (186.) Zn cannot displace following ions from their aqueous solution:

(a.) Ag^+
(b.) Cu^{2+}
(c.) Fe^{2+}
(d.) Na^+

(38.) Matrix match. **[QR code, NCERT Exemplar, Page: 79]**

Column 1	**Column 2**
(A) Λ_m	(p) Scm^{-1}
(B) E_{cell}	(q) m^{-1}
(C) K	(r) Scm^2mol^{-1}
(D) G°	(s) V

(a.) $A \rightarrow r, B \rightarrow p, C \rightarrow q, D \rightarrow s$
(b.) $A \rightarrow r, B \rightarrow s, C \rightarrow p, D \rightarrow q$
(c.) $A \rightarrow r, B \rightarrow s, C \rightarrow p, D \rightarrow q$
(d.) $A \rightarrow r, B \rightarrow q, C \rightarrow p, D \rightarrow s$

(39.) What will be the emf for the given cell **[QR code, Page: 72]**

$Pt|H_2(P_1)|H^+(aq) \parallel H_2(P_2)|Pt$

(a.) $\frac{RT}{F}\log e\frac{P_1}{P_2}$
(b.) $\frac{RT}{2F}\log e\frac{P_1}{P_2}$
(c.) $\frac{RT}{F}\log e\frac{P_2}{P_1}$
(d.) none of the above

(40.) How much electric charge is required to oxidize 1 mole of FeO to Fe_2O_3?

(a.) 49600 C
(b.) $2\times 96500C$
(c.) 96500 C
(d.) $4\times 96500C$

(41.) **Assertion:** Λ_m for weak electrolytes show a sharp increase when the electrolytic solution is diluted. Reason: For weak electrolytes degree of dissociation increases with dilution of solution. **[NCERT Exemplar, Page: 83]**

(a.) Both A and R are true and R is correct explanation of A.
(b.) Both A and R are true and R is nor correct explanation of A.
(c.) A is true but R is false.
(d.) Both A and R are false.

(42.) Which one is not called a anode reaction from the following? **[Page: 68]**

(a.) $Cl^{!} \rightarrow \frac{1}{2}Cl_2 + e^-$
(b.) $Cu \rightarrow Cu^{2+} + 2e^-$
(c.) $Hg^+ \rightarrow Hg^{2+} + e^-$
(d.) $Zn^{2+} + 2e^- \rightarrow Zn$

(43.) The unit of equivalent conductivity is **[Page: 75]**

(a.) ohm cm
(b.) $ohm^{-1}cm^2(g\ equivalent)^{-1}$
(c.) $ohm\ cm^2(g\ equivalent)$
(d.) $S\ cm^{-2}$

(44.) The molar conductivity of a $0.5\ mol/dm^3$ solution of $AgNO_3$ with electrolytic conductivity of $5.76\times 10^{-3} Scm^{-1}$ at $298\ K$ is **[NEET-2O16 Page: 84]**

(a.) $2.88\ S\ cm^2/mol$
(b.) $11.52\ S\ cm^2/mol$
(c.) $0.086\ S\ cm^2/mol$
(d.) $28.8\ S\ cm^2/mol$

(45.) **Assertion:** Cu is less reactive than hydrogen.
Reason: $E^\circ_{Cu^{2+}/Cu}$ is negative. **[NCERT Exemplar, Page: 71]**

(a.) Both A and R are true R is correct explanation of A.
(b.) Both A and R are true and R is not the correct explanation of A.
(c.) A is true but reason is false.
(d.) Both A and R are false.

(46.) The emf of the cell $Zn|Zn^{2+}(a = 0.1/1M) \parallel Fe^{2+}(a = 0.01M)/Fe$ is$0.2905\ V$. The equilibrium constant for the cell reaction is **[Page: 73]**

(a.) $10^{0.32/0.591}$
(b.) $10^{032/0.0295}$
(c.) $10^{0.26/0.0295}$
(d.) $10^{032/0295}$

(47.) How many moles of electrons are needed for the reduction of 20 ml of 0.5 M solution of $KMnO_4$ in acidic medium **[Page: 86]**

(a.) 5×10^{-2} (b.) 5×10^{-3}
(c.) 10^{-2} (d.) 10^{-5}

(48.) Standard reduction potentials at $25^\circ C, Li^+/Li, Ba^{2+}/Ba, Na^+/Na$ and Mg^{2+}/Mg are $-3.05, -2.90, -2.71$ and $-2.37V$ respectively. Which one of the following is strongest oxidising agent? **[Page: 71]**

(a.) Ba^{2+} (b.) Mg^{2+}
(c.) Na^+ (d.) Li^+

(49.) How many gram of cobalt metal will be deposited when a solution of cobalt (II) chloride is electrolyzed with a current of 10 A for 109 minutes $(1F = 96500C$, atomic mass of $Co = 59u)$ **[Page: 86]**

(a.) 4.0 (b.) 20.0
(c.) 40.0 (d.) 0.66

(50.) In SHE, the pH of acid solution should be **[Page: 75]**

(a.) 7 (b.) 14
(c.) 0 (d.) 4

TOPIC WISE PRACTICE QUESTIONS

TOPIC 1: Cells and Electrode Potential

1. The function of a salt bridge is to
1) maintain electrical neutrality of both half cells.
2) increase the cell potential at the positive electrode.
3) decrease the cell potential at the negative electrode.
4) eliminated the impurities present in the electrolyte.

2. Standard electrode potential of SHE at 298 K is:
1) 0.05 V 2) 0.10 V 3) 0.50 V 4) 0.00 V

3. The standard EMF of quinhydrone electrode is 0.699 V. The EMF of the quinhydrone electrode dipped in a solution with pH = 10 is:
1) 0.109 V 2) – 0.109 V 3) 1.289 V 4) – 1.289 V

4. If $E^o_{Fe^{2+}/Fe} = -0.441V$ and $E^o_{Fe^{3+}/Fe^{2+}} = +0.771V$ the standard EMF of the reaction $Fe + 2Fe^{3+} \rightarrow 3Fe^{2+}$ will be
1) 1.653 V 2) 1.212 V 3) 0.111 V 4) 0.330 V

5. In the electrolytic cell, flow of electrons is from
1) cathode to anode in solution. 2) cathode to anode through external supply.
3) cathode to anode through internal supply. 4) anode to cathode through internal supply.

6. Which device converts chemical energy of a spontaneous redox reaction into electrical energy?
1) Galvanic cell 2) Electrolytic cell 3) Daniell cell 4) both 1) and 3)

7. Which of the following reaction is possible at anode?
1) $2Cr^{3+} + 7H_2O \rightarrow Cr_2O_7^{2-} + 14H^+$ 2) $F_2 \rightarrow 2F^-$
3) $(1/2)O_2 + 2H^+ \rightarrow H_2O$ 4) None of these

8. Which one is not called a anode reaction from the following?
1) $Cl^- \rightarrow \frac{1}{2}Cl_2 + e^-$ 2) $Cu \rightarrow Cu^{++} + 2e^-$ 3) $Hg^+ \rightarrow Hg^{++} + e^-$ 4) $Zn^{2+} + 2e^- \rightarrow Zn$

9. Which one of the following statement is true for a electrochemical cell?
1) H_2 is cathode and Cu is anode. 2) H_2 is anode and Cu is cathode.

3) Reduction occurs at H_2 electrode. 4) Oxidation occurs at Cu electrode.

10. $\underset{\text{(anode)}}{Zn(s)\,|\,Zn^{2+}(aq)}\,||\,\underset{\text{(cathode)}}{Cu^{2+}(aq)\,|\,Cu(s)}$ is

1) Weston cell 2) Daniel cell 3) Calomel cell 4) Faraday cell

11. If the following half cells have E° values as

$A^{3+} + e^- \rightarrow A^{2+}, E^\circ = y_2 V$

$A^{2+} + 2e^- \rightarrow A, E^\circ = -y_1 V$

The E° of the half-cell $A^{3+} + 3e \rightarrow A$ will be

1) $\frac{2y_1 - y_2}{3}$ 2) $\frac{y_2 - 2y_1}{3}$ 3) $2y_1 - 3y_2$ 4) $y_2 - 2y_1$

12. A concentration cell is a galvanic cell in which

1) Decrease in free energy in a spontaneous chemical process appears as electrical energy.
2) Decrease in free energy in a spontaneous physical process appears as electrical energy.
3) Decrease in free energy in a spontaneous physical or chemical process appears as electrical energy.
4) A non-spontaneous physical or chemical process produces electrical energy.

13. Based on the cell notation for a spontaneous reaction, at the anode :

Ag (s) | AgCl (s) | Cl^- (aq) || Br^- (aq) | Br_2 (*l*) | C (s)

1) AgCl gets reduced 2) Ag gets oxidized 3) Br– gets oxidized 4) Br_2 gets reduced

14. Zn can displace:

1) Mg from its aqueous solution. 2) Cu from its aqueous solution.
3) Na from its aqueous solution. 4) Al from its aqueous solution.

15. The oxidation potential of a hydrogen electrode at pH = 10 and H_2 p = 1 atm is

1) – 0.59 V 2) 0.00 V 3) + 0.59 V 4) 0.059 V

16. The tendency of an electrode to lose electrons is known as

1) Electrode potential 2) reduction potential 3) Oxidation potential 4) e.m.f.

17. A smuggler could not carry gold by depositing iron on the gold surface since

1) Gold is denser 2) iron rusts
3) Gold has higher reduction potential than iron 4) Gold has lower reduction potential than iron

18. The correct order of $E^\circ_{M^{2+}/M}$ values with negative sign for the four successive elements Cr, Mn, Fe and Co is

1) Mn > Cr > Fe > Co 2) Cr < Fe > Mn > Co 3) Fe > Mn > Cr > Co 4) Cr > Mn > Fe > Co

TOPIC 2: Nernst Equation

19. The E° at 25 °C for the following reaction is 0.22 V. Calculate the equilibrium constant at 25° C :

$H_2(g) + 2AgCl(s) \rightarrow 2Ag(s) + 2HCl(aq)$

1) 2.8×10^7 2) 5.2×10^8 3) 5.2×10^6 4) 5.2×10^3

20. For the given Nernst equation

$$E_{cell} = E^\circ_{cell} - \frac{RT}{2F}\ln\frac{[Mg^{2+}]}{[Ag^+]^2}$$

Which of the following representation is correct?

(a) $Ag^+|Ag||Mg^{2+}|Mg$ (b) $Mg^{2+}|Mg||Ag|Ag^+$
(c) $Mg|Mg^{2+}||Ag^+|Ag$ (d) $Mg|Mg^{2+}||Ag|Ag^+$

21. For a given reaction: M(x+n) + ne^- $\rightarrow M^{x+}$, E^- red is known along with $M^{(x+n)}$ and M^{x+} ion concentrations. Then

1) n can be evaluated. 2) x can be evaluated.
3) (x + n) can be evaluated. 4) n, x, (x + n) can be evaluated.

22. Without losing its concentration $ZnCl_2$ solution cannot be kept in contact with

1) Au 2) Al 3) Pb 4) Ag

23. The oxidation potentials of A and B are +2.37 and +1.66 V respectively. In chemical reactions

1) A will be replaced by B 2) A will replace B
3) A will not replace B 4) A and B will not replace each other

24. The emf of the cell

Ni | Ni^{2+} (1.0 M)| | Au^{3+} (1.0M)| Au is
[Given $E^0 Ni^{2+} / Ni = -0.25$ V and $E^0 Au^{3+} / Au = +1.5$ V]
1) 2.00 V 2) 1.25 V 3) – 1.25 V 4) 1.75 V

25. In the cell reaction
$Cu(s) + 2Ag^+ (aq) \rightarrow Cu^{2+} (aq) + 2Ag(s)$,
$E^o_{cell} = 0.46$ V. By doubling the concentration of Cu^{2+}, E^o_{cell} will become
1) Doubled 2) halved
3) Increases but less than double 4) decreases by a small fraction

26. Following cell has EMF 0.7995V.
Pt | H_2 (1 atm) | HNO_3 (1M) || $AgNO_3$ (1M) | Ag
If we add enough KCl to the Ag cell so that the final Cl^- is 1M. Now the measured emf of the cell is 0.222 V. The K_{sp} of AgCl would be –
1) $1 \times 10^{-9.8}$ 2) $1 \times 10^{-19.6}$ 3) 2×10^{-10} 4) 2.64×10^{-14}

27. In a cell that utilises the reaction
$Zn(s) + 2H^- (aq) \rightarrow Zn^{2+} (aq) + H_2 (g)$ addition of H_2SO_4 to cathode compartment, will
1) Increase the E and shift equilibrium to the right
2) Lower the E and shift equilibrium to the right
3) Lower the E and shift equilibrium to the left
4) Increase the E and shift equilibrium to the left

28. At equilibrium:
1) $E^0_{cell} = 0, \Delta G^\circ = 0$ 2) $E_{cell} = 0, \Delta G = 0$ 3) Both are correct 4) None is correct

29. Equivalent conductivity can be expressed in terms of specific conductance (κ) and concentration (N) in gram equivalent dm^{-3} as :
1) $\kappa \times N$ 2) $\dfrac{\kappa \times 1000}{N}$ 3) $\dfrac{\kappa \times N}{1000}$ 4) $\kappa \times N \times 1000$

30. The ionic conductivity of Ba^{2+} and Cl^- at infinite dilution are 127 and 76 $ohm^{-1}\ cm^2\ eq^{-1}$ respectively. The equivalent conductivity of $BaCl_2$ at infinity dilution (in $ohm^{-1} cm^2\ eq^{-1}$)
would be :
1) 203 2) 279 3) 101.5 4) 139.5

31. The increase in equivalent conductivity of a weak electrolyte with dilution is due to
1) increase in degree of dissociation and decrease in ionic mobility.
2) decrease in degree of dissociation and decrease in ionic mobility.
3) increase in degree of dissociation and increase in ionic mobility.
4) decrease in degree of dissociation and increase in ionic mobility.

32. For a relation
$\Delta_r G = -nFE_{cell}$
$E_{cell} = E^0_{cell}$ in which of the following condition?
1) Concentration of any one of the reacting species should be unity
2) Concentration of all the product species should be unity.
3) Concentration of all the reacting species should be unity.
4) Concentration of all reacting and product species should be unity.

33. 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
1) 1000 x/N 2) 1000/Nx 3)1000N/x 4)Nx/1000

34. For an electrolyte solution of 0.05 mol L^{-1}, the conductivity has been found to be 0.0110 S cm– 1. The molar conductivity is
1) 0.055 S $cm^2\ mol^{-1}$ 2) 550 S $cm^2\ mol^{-1}$ 3) 0.22 S $cm^2\ mol^{-1}$ 4) 220 S $cm^2\ mol^{-1}$

35. On which of the following magnitude of conductivity does not depends?
1) Nature of material 2) Temperature 3) Pressure 4) Mass of the material

36. The resistance of 0.01 N solution of an electrolyte was found to be 220 ohm at 298 K using a conductivity cell with a cell constant of $0.88 cm^{-1}$. The value of equivalent conductance of solution is –
1) 400 mho cm^2 g eq^{-1} 2) 295 mho cm^2 g eq^{-1}
3) 419 mho cm^2 g eq^{-1} 4) 425 mho cm^2 g eq^{-1}

37. The correct order of equivalent conductance at infinite dilution of LiCl, NaCl and KCl is

1) LiCl > NaCl > KCl 2) KCl > NaCl > LiCl
3) NaCl > KCl > LiCl 4) LiCl > KCl > NaCl

38. Which of the following expression correctly represents molar conductivity?

1) $\wedge_m = \frac{K}{C}$ 2) $\wedge_m = \frac{KA}{1}$ 3) $\wedge_m = KV$ 4)all of these

39. At 25°C, the molar conductance at infinite dilution for the strong electrolytes NaOH, NaCl and $BaCl_2$ are 248×10^{-4}, 126×10^{-4} and 280×10^{-4} Sm^2mol^{-1} respectively.

$\wedge_m^0 Ba(OH)_2$ in S m^2mol^{-1} is

1) 52.4×10^{-4} 2) 524×10^{-4} 3) 402×10^{-4} 4) 262×10^{-4}

40. The ion of least limiting molar conductivity among the following is

1) SO_4^{2-} 2) H^+ 3) Ca^{2+} 4) CH_3COO^-

41. Specific conductance of a 0.1 N KCl solution at 23ºC is 0.012 ohm^{-1} cm^{-1}. Resistance of cell containing the solution at same temperature was found to be 55 ohm. The cell constant is

1) 0.0616 cm–1 2) 0.66 cm–1 3) 6.60 cm–1 4) 660 cm–1

42. The equivalent conductances of two strong electrolytes at infinite dilution in H_2O (where ions move freely through a solution) at 25°C are given below:

$\Lambda^0_{CH_3COONa} = 91.0S \ \ cm^2/equiv.$

$\Lambda^0_{HCl} = 426.2S \ \ cm^2/equiv.$

What additional information/ quantity one needs to calculate Λ^0 of an aqueous solution of acetic acid?

1) Λ^0 of chloroacetic acid ($ClCH_2COOH$) 2) Λ^0 of NaCl
3) Λ^0 of CH_3COOK 4) the limiting equivalent coductance of $H^+\left(\lambda^0_{H^+}\right)$

TOPIC 4: Electrolysis and Types of Electrolysis

43. The electric charge required for electrode deposition of one gram - equivalent of a substance is :

1) one ampere per second. 2) 96500 coulombs per second.
3) one ampere for one hour. 4) charge on one mole of electrons.

44. How many minutes will it take to plate out 5.2 g of Cr from a $Cr_2(SO_4)_3$ solution using a current of 9.65 A?

(Atomic mass: Cr = 52.0)

1) 200 2) 50 3) 100 4) 103

45. Which of the following is the use of electrolysis?

1) Electrorefining 2) Electroplating 3) Both 1) & 2) 4) None of these

46. By virtue of Faraday's second law of electrolysis, the electrochemical equivalent of the two metals liberated at the electrodes has the same ratio as that of their

1) atomic masses 2) molecular masses 3) equivalent masses 4) any of three

47. In electrolysis of very dilute NaOH using platinum electrodes

1) H_2 is evolved at cathode and O_2 at anode 2) NH_3 is produced at anode
3) Cl_2 is obtained at cathode
4) O_2 is produced at cathode and H_2 at anode

48. Reaction that takes place at graphite anode in dry cell is

1) $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$ 2) $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

3) $Mn^{2+}(aq) + 2e^- \rightarrow Ms(s)$ 4) $Mn(s) \rightarrow Mn^+(aq) + e^- + 1.5V$

49. Electrolysis of fused NaCl will give

1) Na 2) NaOH 3) NaClO 4) None of these

50. An unknown metal M displaces nickel from nickel (II) sulphate solution but does not displace manganese from manganese sulphate solution. Which order represents the correct order of reducing power?

1) Mn > Ni > M 2) Ni > Mn > M 3) Mn > M > Ni 4) M > Ni > Mn

51. A solution of sodium sulphate in water is electrolysed using inert electrodes. The products at the cathode and anode are respectively

1) H_2, O_2 2) O_2, H_2 3) O_2, Na 4) O_2, SO_2

52. If 0.5 amp. current is passed through acidified silver nitrate solution for 100 minutes. The mass of silver deposited on cathode, is (eq.wt.of silver = 108)
1) 2.3523 g 2) 3.3575 g 3) 5.3578 g 4) 6.3575 g
53. Find the charge in coulombs required to convert 0.2 mole VO_3^{2-} into VO_4^{3-}
1) 1.93×10^4 2) 9.65×10^4 3) 1.93×10^5 4) 9.65×10^5
54. The amount of electricity that can deposit 108 g of Ag from $AgNO_3$ solution is:
1) 1 F 2) 2 A 3) 1 C 4) 1 A
55. Three faradays electricity was passed through an aqueous solution of iron (II) bromide. The weight of iron metal
(at. wt = 65) deposited at the cathode (in gm) is
1) 56 2) 84 3) 112 4) 168

TOPIC 5: Commercial Cell and Corrosion

56. Which of the following statement is false for fuel cells?
1) They are more efficient 2) They are free from pollution
3) They run till reactants are active 4) Fuel burned with O_2
57. Among the following cells:
(i) Leclanche cell (ii) Nickel-Cadmium cell (iii) Lead storage battery (iv) Mercury cell
primary cells are
1) (i) and (ii) 2) (i) and (iii) 3) (ii) and (iii) 4) (i) and (iv)
58. Identify the correct statement :
1) Corrosion of iron can be minimized by forming a contact with another metal with a higher reduction potential.
2) Iron corrodes in oxygen free water
3) Corrosion of iron can be minimized by forming an impermeable barrier at its surface.
4) Iron corrodes more rapidly in salt water because its electrochemical potential is higher.
59. Several blocks of magnesium are fixed to the bottom of a ship
1) make the ship lighter. 2) prevent action of water and salt.
3) prevent puncturing by under-sea rocks. 4) keep away the sharks.
60. Which of the following batteries cannot be reused?
1) Lead storage battery 2) Ni-Cd cell 3) Mercury cell 4) Both (2) and (3)

NEET PREVIOUS YEARS QUESTIONS

1. In the electrochemical cell :
Zn | $ZnSO_4$ (0.01M) | | $CuSO_4$ (1.0 M) | Cu, the emf of this Daniel cell is E_1. When the concentration of $ZnSO_4$ is changed to 1.0M and that of $CuSO_4$ changed to 0.01M, the emf changes to E_2. From the followings, which one is the relationship between E_1 and E_4? (Given,RT/F= 0.059) **[2017]**
1) $E_1 < E_2$ 2) $E_1 > E_2$ 3) $E_2 = 0 \neq E_1$ 4) E1 = E2
2. The pressure of H_2 required to make the potential of H_2-electrode zero in pure water at 298 K is : **[2016]**
1) 10^{-14} atm 2) 10^{-12} atm 3) 10^{-10} atm 4) 10^{-4} atm
3. A device that converts energy of combustion of fuels like hydrogen and methane, directly into electrical energy is known as : **[2015]**
1) Electrolytic cell 2) Dynamo 3) Ni-Cd cell 4) Fuel Cell
4. When 0.1 mol MnO_4^{2-} is oxidised the quantity of electricity required to completely oxidise MnO_4^{2-} to MnO_4^- is: **[2014]**
1) 96500 C 2) 2×96500 C 3) 9650 C 4) 96.50 C
5. The weight of silver (at wt. = 108) displaced by a quantity of electricity which displaces 5600 mL of O_2 at STP will be :- **[2014]**
1) 5.4 g 2) 10.8 g 3) 54.9 g 4) 108.0 g
6. For a cell involving one electron $E^{\ominus}_{cell} = 0.59V$ at 298 K, the equilibrium constant for the cell reaction is :-

[Given that $\frac{2.303RT}{F} = 0.059V$ at $T = 298K$] [2019]

(1) 1.0×10^2 (2) 1.0×10^5 (3) 1.0×10^{10} (4) 1.0×10^{30}

7. For the cell reaction $2Fe^{3+}(aq) + 2I^-(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$

$E^{\ominus}_{cell} = 0.24V$ at 298 K. The standard Gibbs energy $(\Delta_r^0 G^{\ominus})$ of the cell reaction is : [2019]

[Given that Faraday constant F = 96500 C mol^{-1}]

(1) –46.32 kJ mol^{-1} (2) –23.16 kJ mol^{-1} (3) 46.32 kJ mol^{-1} (4) 23.16 kJ mol^{-1}

8. Following limiting molar conductivities are given as **[2019-ODISSA]**

$\lambda^0_{m(H_2SO_4)} = x\,Scm^2\,mol^{-1}$

$\lambda^0_{m(K_2SO_4)} = y\,Scm^2\,mol^{-1}$

$\lambda^0_{m(CH_3COOK)} = z\,Scm^2\,mol^{-1}$

$\lambda^0_m\,(in\,S\,cm^2\,mol^{-1})$ for CH_3COOH will be

(1) $x - y + 2z$ (2) $x + y - z$ (3) $x - y + z$ (4) $\frac{(x-y)}{2} + z$

9. The standard electrode potential (E.) values of Al^{3+}/Al, Ag^+/Ag, K^+/K and Cr^{3+}/Cr are .1.66 V, 0.80V, – –2.93 V and –0.74 V, respectively. The correct decreasing order of reducing power of the metal is : **[2019-ODISSA]**

(1) Ag > Cr > Al > K (2) K > Al > Cr > Ag (3) K > Al > Ag > Cr (4) Al > K > Ag > Cr

10. Identify the reaction from following having top position in EMF series (Std.red. potential) according to their electrode potential at 298 K. **[2020-COVID-19]**

(1) $Mg^{2+} + 2e^- \rightarrow Mg_{(s)}$ (2) $Fe^{2+} + 2e^- \rightarrow Fe_{(s)}$ (3) $Au^{3+} + 3e^- \rightarrow Au_{(s)}$ (4) $K^+ + 1e^- \rightarrow K_{(s)}$

11. In a typical fuel cell, the reactants (R) and product (P) are :- **[2020-COVID-19]**

(1) R = $H_{2(g)}$, $O_{2(g)}$; P = $H_2O_{2(l)}$ (2) R = $H_{2(g)}$, $O_{2(g)}$; P = $H_2O_{(l)}$

(3) R = $H_{2(g)}$, $O_{2(g)}$, $Cl_{2(g)}$; P = $HClO_{4(aq)}$ (4) R = $H_{2(g)}$, $N_{2(g)}$; P = $NH_{3(aq)}$

12. On electrolysis of dil sulphuric acid using platinum (Pt) electrode, the product obtained at anode will be **[2020]**

1. SO_2 gas 2. Hydrogen gas 3. Oxygen gas 4. H_2S gas

13. The number of Faradays (F) required to produce 20g of calcium from molten $CaCl_2$ (Atomic mass of Ca = 40 g mol^{-1}) is **[2020]**

1) 4 2) 1 3) 2 4) 3

14. The molar conductance of NaCl, HCl and CH_3COONa at infinite dilution are 126.45,426.16 and 91.0 S cm^2mol^{-1} respectively. The molar conductance of CH_3COOH at infinite dilution is. Choose the right option for your answer **[NEET-2021]**

1) 390.71 S cm^2mol^{-1} 2) 698.28 S cm^2mol^{-1} 3) 540.48 S cm^2mol^{-1} 4) 201.28 S cm^2mol^{-1}

15. The molar conductivity of 0.007 M acetic acid is $20\,S\,cm^2mol^{-1}$. What is the dissociation constant of acetic acid? Choose the correct option. **[NEET-2021]**

$$\begin{bmatrix} \Lambda^\circ_{H^+} = 350\,S\,cm^2mol^{-1} \\ \Lambda^\circ_{CH_3COO^-} = 50\,S\,cm^2mol^{-1} \end{bmatrix}$$

1) $2.50 \times 10^{-4}\,mol\,L^{-1}$ 2) $1.75 \times 10^{-5}\,mol\,L^{-1}$ 3) $2.50 \times 10^{-5}\,mol\,L^{-1}$ 4) $1.75 \times 10^{-4}\,mol\,L^{-1}$

16. Given below are half-cell reactions: **[NEET-2022]**

$$\mathbf{MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O, E^o_{Mn^{2+}/MnO_4^-} = -1.510V}$$

$$\mathbf{\frac{1}{2}O_2 + 2H^+ + 2e^- \rightarrow H_2O, E^o_{O_2/H_2O} = +1.223V}$$

Will the permanganate ion, MnO_4^- liberate O_2 from water in the presence of an acid?

1) Yes, because $E^o_{cell} = +0.287V$
2) No, because $E^o_{cell} = -0.287V$
3) Yes, because $E^o_{cell} = +2.733V$
4) No, because $E^o_{cell} = -2.733V$

17. At 298 K, the standard electrode potentials of $Cu^{2+}/Cu, Zn^{2+}/Zn, Fe^{2+}/Fe$ and Ag^+/Ag are $0.34V, -0.76V, -0.44V$ and $0.80V$ respectively. **[NEET-2022]**

On the basis of standard electrode potential, predict which of the following reaction cannot occur?

1) $CuSO_{4(aq)} + Zn_{(s)} \rightarrow ZnSO_{4(aq)} + Cu_{(s)}$
2) $CuSO_{4(aq)} + Fe_{(s)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$
3) $FeSO_{4(aq)} + Zn_{(s)} \rightarrow ZnSO_{4(aq)} + Fe_{(s)}$
4) $2CuSO_{4(aq)} + 2Ag_{(s)} \rightarrow 2Cu_{(s)} + Ag_2SO_{4(aq)}$

18. Find the emf of the cell in which the following reaction takes place at 298 K **[NEET-2022]**

$$\mathbf{Ni(s) + 2Ag^+(0.001M) \rightarrow Ni^{2+}(0.001M) + 2Ag(s)}$$

$$\mathbf{(Given\ that\ \ E^o_{cell} = 1.05\ V, \frac{2.303\,RT}{F} = 0.059\ at\ 298\,K)}$$

1)1.385 V
2) 1.385 V
3) 0.9615 V
4) 1.05 V

NCERT LINE BY LINE QUESTIONS – ANSWERS

(1.)	b	**(2.)**	b	**(3.)**	b	**(4.)**	c	**(5.)**	b
(6.)	b	**(7.)**	a	**(8.)**	c	**(9.)**	d	**(10.)**	a
(11.)	c	**(12.)**	b	**(13.)**	c	**(14.)**	d	**(15.)**	c

(16.)	b	**(17.)**	b	**(18.)**	a	**(19.)**	d	**(20.)**	a
(21.)	c	**(22.)**	a	**(23.)**	a	**(24.)**	d	**(25.)**	b
(26.)	d	**(27.)**	d	**(28.)**	c	**(29.)**	c	**(30.)**	c
(31.)	a	**(32.)**	a	**(33.)**	a	**(34.)**	d	**(35.)**	d
(36.)	b	**(37.)**	d	**(38.)**	c	**(39.)**	b	**(40.)**	c
(41.)	a	**(42.)**	d	**(43.)**	b	**(44.)**	b	**(45.)**	c
(46.)	b	**(47.)**	a	**(48.)**	b	**(49.)**	b	**(50.)**	c

TOPIC WISE PRACTICE QUESTIONS - ANSWERS

1) 1	2) 4	3) 1	4) 2	5) 4	6) 4	7) 1	8) 4	9) 2	10) 2
11) 2	12) 2	13) 2	14) 2	15) 3	16) 3	17) 3	18) 1	19) 1	20) 3
21) 1	22) 2	23) 2	24) 4	25) 4	26) 1	27) 1	28) 2	29) 2	30) 2
31) 3	32) 3	33) 2	34) 4	35) 4	36) 1	37) 2	38) 4	39) 2	40) 4
41) 2	42) 2	43) 4	44) 2	45) 3	46) 3	47) 1	48) 2	49) 1	50) 3
51) 1	52) 2	53) 1	54) 1	55) 2	56) 4	57) 4	58) 3	59) 2	60) 3

NEET PREVIOUS YEARS QUESTIONS-ANSWERS

1) 2	2) 1	3) 4	4) 3	5) 4	6) 3	7) 1	8) 4	9) 2	10) 3
11) 2	12) 3	13) 2	14) 1	15) 2	16) 1	17) 4	18) 3		

NCERT LINE BY LINE QUESTIONS – SOLUTIONS

(1.) (b) In H_2-O_2 fuel cell, the combustion of H_2 occurs to create potential difference between the two electrodes.

(4.) (c) The electrolyte used in Leclanche cell is moist paste of NH_4Cl and $ZnCl_2$.

(5.) (b) $Ag^+ + e^- \rightarrow Ag$

$$E_{Ag+/Ag} = E^o_{Ag^+/Ag} - \frac{0.059}{1}\log\frac{1}{\left[Ag^+\right]}$$

$$= E^o_{Ag^+/Ag} + 0.059\left[Ag^+\right]$$

On increasing $\left[Ag^+\right], E^o_{Ag^+/Ag}$ will increase and it has a positive value.

(6.) (b) $-0.591V$

(9.) (d) Calomel electrode is used as reference electrode.

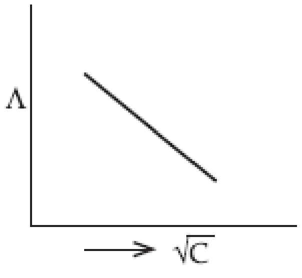
(10.) (a) Weak acid

$$\left.\begin{array}{lcl} HCN & \rightarrow & \textit{Weak acid} \\ NH_4OH & \rightarrow & \textit{Weak base} \\ Ca\left(NO_3\right)_2 & \rightarrow & \textit{Weak acid} \end{array}\right\}$$

Weak electrolytes $\rightarrow Ca\left(NO_3\right)_2$ → Strong Base →

Strong Acidic Salt → Strong electrolyte

(12.) (b)



i.e., on ↑ concentration dissociation of electrolytes decrease Λ_m or Λ_n^o decrease.

(13.) (c) $Al_2O_3 \rightarrow 2Al^{3+} + 3O^{2-}$

$Al^{3+} + 3e^- \rightarrow Al(\text{For 1 mole})$

3 F charge require to obtain 1 mole Al from Al_2O_3

(14.) (d) Electrolysis of Conc.H_2SO_4

$H_2SO_4 \rightarrow H^+ + HSO_4^-$

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

at anode $2HSO_4^- \rightarrow H_2S_2O_8(\text{Marshall'sacid}) + 2e^-$

(15.) (c)

(16.) (b) $E^o_{cell} = \frac{0.059}{2} \log K_C$ or $\frac{1.10 \times 2}{0.059} = \log K_c$

$K_c = 1.9 \times 10^{37}$

(17.) (b) O_2 is formed at anode at anode and H_2 is formed at cathode.

(19.) (d) Prevention of corrosion by zinc coating is called galvanization.

(21.) (c)

$$\overset{+6}{\text{Mn}}O_4^{2-} \rightarrow \overset{+7}{\text{Mn}}O_4^- + e^-$$

$0.1\,mol \qquad 0.1\,mol$

$Q = 0.1 \times F = 0.1 \times 96500\ C = 9650\ C$

(22.) (a) $0.059\,V$

(23.) (a) $[Fe(CN)_6]^3 \rightarrow [Fe(CN)_6]^{4-}; B^o = +035V$

$Fe^{3+} \rightarrow Fe^{2+}; F:^o = +0.77V$

Higher the positive reduction potential, stronger will be the oxidising agent. Oxidising agent oxidises other compounds and get itself reduced easily.

(24.) (d) $\Lambda^\circ_{NaCl} = 126.4S\ cm^2mol^{-1}$

$\Lambda^\circ_{HCl} = 425.9\ S\ cm^2mol^{-1}$

$\Lambda^\circ_{CH_3COONa} = 91.0\ S\ cm^2mol^{-1}$

$\Lambda^\circ_{CH_3COOH} = \Lambda^\circ_{CH_3COONa} + \Lambda^\circ_{HCl} = -\Lambda^\circ_{NaCl}$

$= 91 + 425.9 - 126.4 = 390.5\ S\ cm^2mol^{-1}$

(25.) (b) From Cu to Zn outside the cell.

(27.) (d) The cell constant of a conductivity cell remain constant for a cell.

(28.) (c) $Q = 1 \times t$

$Q = 1 \times 60 = 60C$

Now, $1.60 \times 10^{19} C = 1$ electron

$\therefore 60C = \frac{60}{1.6 \times 10^{19}} = 37.5 \times 10^{19} = 3.75 \times 10^{20}$ electrons

(31.) (a) $\Delta G^\circ = -nFE^o_{cell}$

If $E^o_{cell} = -ve$ then $\Delta G^o = +Ve$

i.e. $\Delta G^o > 0$

$\Delta G^\circ = -nRT \log k_{eq}$

For $\Delta G^\circ = +ve, K_{eq} = -ve$ i.e., $K_{eq} < 1$

(33.) (a)

$$E = E^\circ + \frac{0.059}{3} \log \frac{[A1^{3+}]}{[A1]}$$

so, if $\left[Al^{3+}\right]\uparrow = E\uparrow$

(34.) (d) A is false but R is true.

(35.) (d)

(38.) (c) $A \rightarrow r, B \rightarrow s, C \rightarrow p_2 D \rightarrow q$

(39.) (b) RHS $2H^+ + 2e^- \square \quad H_2(P_2)$

LHS $\quad H_2(P_1) \square \quad 2H^+ 2e^-$

Overall Rxn: $H_2(P_1) \square \quad (P_2)$

$$E = E^\circ - \frac{RT}{nF}\ln\frac{P_2}{P_1} = 0 - \frac{RT}{nF}\ln\frac{P_2}{P_1} = \frac{RT}{nF}\ln\frac{P_1}{P_2}$$

(41.) (a) For weak electrolyte, Λ_m increase steeply on dilution due to increase in the no. of ions (or degree of dissociation).

(42.) (d) $Zn^{2+} + 2e^- \rightarrow Zn$

(43.) (b) $ohm^{-1} cm^2 (geq)^{-1}$

(44.) (b) $\Lambda_m = \frac{K \times 100}{Molarity(M)}$

$$= 5.76 \times \frac{10^{-3} Scm^{-1} \times 1000}{0.5\ mol\ cm^{-3}} = 11.52\ S\ cm^2 mol^{-}{}'$$

(45.) (c) Cu is less reactive than hydrogen because SRP $E^o_{Cu^{2+}/Cu} = 0.34V$ and $E^o_{H^+/H} = 0.0V$

(46.) (b) $E = E^\circ - \frac{0.0591}{n}\log Q$

$$0.2905 = E^o - \frac{0.0591}{2}, \log\frac{[0.1]}{[0.01]}$$

$$E^\circ = 0.295 + 0.0295 = +0.32V$$

$$0.32 = \frac{0.0591}{2}\log K$$

$$K = 10^{0.32/0.265}$$

(47.) (a) $n_{KMnO_4} = 0.5 \times 20 \times 10^{-8} = 10^{-2}$

$Mn^{7+} + 5e^- \rightarrow Mn^{2+}$

So, 1 mole $KMnO_4 \rightarrow 5$ mole of e^-

10^{-2} mol $KMnO_4 \rightarrow 5 \times 10^{-2}$ mole e^-

(48.) (b) A cation with maximum value of SRP has the strongest oxidising power.

(49.) (b) $W = \frac{I \times T \times E}{96500} = \frac{10 \times 109 \times 60 \times 59}{96500 \times 2} = 20$

(50.) (c) In SHE concentration of solution is $1M\,HCl$

$1M\,HCl \rightarrow H^+ + Cl^-$

$pH = -\log\left[H^+\right] = -\log 1 = 0$

TOPIC WISE PRACTICE QUESTIONS – SOLUTIONS

1. (1) Maintain electrical neutrality of both half cells.
2. 4) $E^\circ_{H_2|H^+} = E^\circ_{H_2|H^+} = 0V$
3. 1) E = E° – 0.059 Ph = 0.699 – 0.059 × 10 = 0.109 V
4. 2)

(b) $Fe \longrightarrow Fe^{2+} + 2e$ [Anode] $E = -0.441V$

$[Fe^{3+} + e \longrightarrow Fe^{2+}] \times 2$ [Cathode] $E = +0.771V$

$Fe + 2Fe^{3+} \longrightarrow 3Fe^{2+}$

$$E^o_{cell} = E^o_{cathode} - E^o_{Anode}$$
$$= .771 - (-.441) = .771 + .441 = 1.212\,V$$

5. (4) In electrolytic cell the flow of electrons is from anode to cathode through internal supply.
6. (4)
7. (1) $2Cr^{3+} + 7H_2O \to Cr_2O_7^{2-} + 14H^+$

O.S. of Cr changes from +3 to +6 by loss of electrons. At anode oxidation takes place.

8. 4) $\overset{\text{Reduction}}{Zn^{2+} + 2e^- \longrightarrow Zn}$ It shows reduction reaction.
9. (2) In electrochemical cell H2 has greater tendency to release electrons.

$\therefore$ H_2 is liberated at anode and Cu is deposited at the cathode.

10. (2) The cell in which Cu and Zn rods are dipped in its solution is called Daniel cell.
11. 2)

(i) $A^{3+} + e^- \longrightarrow A^{2+}$, $\Delta G_1 = -1\,F\,y_2$

(ii) $A^{2+} + 2e^- \longrightarrow A$, $\Delta G_2 = -2F(-y_1) = 2Fy_1$

Add, (i) and (ii) we get

$A^{3+} + 3e^- \longrightarrow A$;

$\Delta G_3 = \Delta G_1 + \Delta G_2$

$-3FE^o = -Fy_2 + 2Fy_1$

$-3FE^o = -F\,(y_2 - 2y_1)$

$$E^o = \frac{y_2 - 2y_1}{3}$$

12. 2) In concentration cell the spontaneous process is physical in nature involving transfer of matter from higher concentration to lower concentration in indirect manner.
13. (2) Ag becomes oxidized and Br– becomes reduced.
14. 2)

$E^o_{Zn^{2+}|Zn} < E^o_{Cu^{2+}|Cu}$;

$E^o_{Cu^{2+}|Cu} + E^o_{Zn|Zn^{2+}} = +ve$

15. 3) $E_{oxidation} = 0.059\ pH = 0.059 \times 10 = 0.59V$
16. 3) The magnitude of the electrode potential of a metal is a measure of its relative tendency to loose or gain electrons. i.e., it is a measure of the relative tendency to undergo oxidation (loss of electrons) or reduction (gain of electrons).

$M \to M^{n+} + ne^-$ (oxidation potential)

$M^{n+} + ne^- \to M$ (reduction potential)

17. 3) Gold having higher E^o_{Red} oxidises $Fe \to Fe^{2+}$
18. 1)

The value of $E^\circ_{M^{2+}/M}$ for given metal ions are

$E^\circ_{Mn^{2+}/Mn} = -1.18\ V,$

$E^\circ_{Cr^{2+}/Cr} = -0.9\ V,$

$E^\circ_{Fe^{2+}/Fe} = -0.44\ V$ and

$E^\circ_{Co^{2+}/Co} = -0.28\ V.$

The correct order of $E^\circ_{M^{2+}/M}$ values without considering negative sign would be

$Mn^{2+} > Cr^{2+} > Fe^{2+} > Co^{2+}.$

19. 1) $E^\circ = \frac{0.0591}{2}\log K_{eq}; \log K_{eq} = \frac{2\times0.22}{0.0591} \Rightarrow 7.44$ or $K_{eq} \approx 2.8\times10^7$

20. 3)

21. 1) $E_{red} = E^\circ_{red} + \frac{0.059}{n}\log\frac{\left[M^{(x+n)}\right]}{\left[M^{x+}\right]}$

22. 2) Without losing its concentration $ZnCl_2$ solution cannot kept in contact with Al because Al is more reactive than Zn due to its highly negative electrode reduction potential.

23. (2) A will replace B.

24. (4) Here Ni is anode and Au is cathode

Given $E^\circ_{Ni^{2+}/Ni} = 0.25V$ and $E^\circ_{Au^{3+}/Au} = +1.5V$

$E^\circ_{cell} = E^\circ_C - E^\circ_A = 1.5-(-0.25) = 1.5+0.25 = 1.75V$

25. 4) emf will decrease

26. 1)

$2Ag^+ + H_2 \longrightarrow 2H^+ + 2Ag$

$E = E^\circ - \frac{0.0591}{2}\log\frac{[H^+]^2}{P_{H_2}\times[Ag^+]^2}$

$0.222 = 0.7995 - \frac{0.0591}{2}\log\frac{1}{[Ag^+]^2}$

$[Ag^+] = 10^{-9.8}$

$K_{sp} = [Ag^+][Cl^-] = (10^{-9.8})\times(1) = 10^{-9.8}$

27. 1)

$Zn(s) + 2H^+(aq) \rightleftharpoons Zn^{2+}(aq) + H_2(g)$

$E_{cell} = E^\circ_{cell} - \frac{0.059}{2}\log\frac{[Zn^{2+}][H_2]}{[H^+]^2}$

Addition of H_2SO_4 will increase [H+] and E_{cell} will also increase and the equilibrium will shift towards RHS.

28. 2) At equilibrium, $\Delta G = 0, E_{cell} = 0$

29. 2) $\Lambda_{eq} = \frac{\kappa\times1000}{N}$

30. 2) $\Lambda^\infty_m(BaCl_2) = \lambda^\infty_m(Ba^{2+}) + 2\lambda^\infty_m(Cl^-) = 127 + 2\times76 = 279\ ohm^{-1}cm^2eq^{-1}$

31. (3) On dilution degree of dissociation of a weak electrolyte increase, hence increase in ionic mobility.

32. (3) When the concentration of all reacting species kept unity, then $E_{cell} = E^\circ_{cell}$ and the given relation will become

$\Delta_r G = -nFE^\circ_{cell}$

33. (2) Specific resistance = x

$\therefore$ Specific conductance (or conductivity) = $\kappa = \frac{1}{x}$

$$\therefore \Lambda_{eq} = \frac{\kappa \times 1000}{N} = \frac{1000}{xN}$$

34. 4) $\Lambda_m = \frac{\kappa \times 1000}{M} = \frac{0.0110 \times 1000}{0.05} = 220 \,\text{S cm}^2\,\text{mol}^{-1}$

35. 4) Conductivity does not depend upon mass or weight of material.

36. 1) $\Lambda_{eq} = \kappa \times \frac{1000}{N} = \frac{1}{R} \times \frac{\ell}{a} \times \frac{1000}{N} = \frac{1}{R} \times \text{cell constant} \times \frac{1000}{N} = \frac{1}{220} \times 0.88 \times \frac{1000}{0.01} = 400 \text{mho cm}^2 \text{ g eq}^{-1}$

37. (2) As we go down the group 1 (i.e. from Li^+ to K^+), the ionic radius increases, degree of solvation decreases and hence effective size decreases resulting in increase in ionic mobility. Hence equivalent conductance at infinite dilution increases in the same order.

38. 4)

39. 2)

$$\Lambda^\circ_{Na^+} + \Lambda^\circ_{OH^-} = 248 \times 10^{-4} \text{Sm}^2\text{mol}^{-1}$$

$$\Lambda^\circ_{Na^+} + \Lambda^\circ_{Cl^-} = 126 \times 10^{-4} \text{Sm}^2\text{mol}^{-1}$$

$$\Lambda^\circ_{Ba^{2+}} + \Lambda^\circ_{2Cl^-} = 280 \times 10^{-4} \text{ S m}^2\text{mol}^{-1}$$

Now, $\Lambda^\circ_{Ba(OH)_2} = \Lambda^\circ_{BaCl_2} + 2\Lambda^\circ_{NaOH} - 2\Lambda^\circ_{NaCl}$

$$\Lambda^\circ_{Ba(OH)_2} = 280 \times 10^{-4} + 2 \times 248 \times 10^{-4} - 2 \times 126 \times 10^{-4}$$

$$\Lambda^\circ_{Ba(OH)_2} = 524 \times 10^{-4} \text{Sm}^2\text{mol}^{-1}.$$

40. (4) Larger the size, lower the speed.

41. (2) Specific conductance of the solution (κ) = 0.012 ohm^{-1} cm^{-1} and resistance (R) = 55 ohm.
Cell constant = Specific conductance × Observed resistance = 0.012 × 55= 0.66 cm^{-1}.

42. (b) According to Kohlrausch's law, molar conductivity of weak electrolyte acetic acid (CH_3COOH) can be calculated as follows:

$$\Lambda^\circ_{CH_3COOH} = \left(\Lambda^\circ_{CH_3COONa} + \Lambda^\circ_{HCl}\right) - \Lambda^\circ_{NaCl}$$

$\therefore$ Value of Λ°_{NaCl} should also be known for calculating value of $\Lambda^\circ_{CH_3COOH}$.

43. 4) $\underset{1\text{mole}}{A^+} + e^- \rightarrow \underset{1\text{mole}}{A}$

$\therefore$ Charge of 1 mole electron is required for deposition of 1 equivalent of substance.

44. 2) $5.2 = \frac{52}{3} \times \frac{9.65}{96{,}500} \times t(\text{sec})$

$$t(\text{sec}) = 3000$$

$$t(\text{min}) = 50$$

45. (3) Electrorefining and electroplating are done by electrolysis.

46. (3)

47. (1) Reduction and oxidation of H_2O occurs to give H_2 at cathode and O_2 at anode.

48. (2)

49. (1) When molten or fused NaCl is electrolysed, it yields metallic sodium and gaseous chlorine. Reactions involved are as follows:

$$NaCl\,(s) \longrightarrow Na^+\,(l) + Cl^-\,(l)$$
$$Na^+ + e^- \longrightarrow Na\,(l) \quad \text{(at cathode)}$$
$$Cl^-\,(l) \longrightarrow Cl\,(g) + e^- \quad \text{(at anode)}$$
$$Cl\,(g) + Cl(g) \longrightarrow Cl_2\,(g)$$

50. (3) Oxidation potential of M is more than Ni and less than Mn. Hence reducing power Mn > M > Ni

51. (1) Water is reduced at the cathode and oxidized at the anode instead of Na^+ and SO_4^{2-}

Cathode : $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

Anode : $H_2O \rightarrow 2H^+ + \frac{1}{2}O_2 + 2e^-$.

52. (b) Given current (i) = 0.5 amp;
Time (t) = 100 minutes × 60 = 6000 sec
Equivalent weight of silver (E) = 108.
According to Faraday's first law of electrolysis

$$W = \frac{Eit}{96500} = \frac{108 \times 0.5 \times 6000}{96500} = 3.3575g$$

53. (1) Charge = 0.2 × 1 Faraday = 0.2 × 96500 coulombs = 19300 = 1.93×10^4 coulombs

54. (1) According to Faraday law's of electrolysis, amount of electricity required to deposit 1 mole of metal = 96500 C = 1 F i.e., for deposition of 108 g Ag electricity required = 1 F

55. (2) $Fe^{2+} + 2e^- \rightarrow Fe; E_{Fe} = \frac{56}{2} = 28$
1 Faraday liberates = 28 g of Fe
3 Faraday liberates = 3 × 28 = 84 g

56. (4) Combustion takes place.

57. (4) Primary cells are those cells, in which the reaction occurs only once and after use over a period of time, it becomes dead and cannot be reused again. e.g., Leclanche cell and mercury cell.

58. (3) Corrosion of iron can be minimized by forming an impermeable barrier at its surface.

59. (2) Magnesium provides cathodic protection and prevent rusting or corrosion.

60. (3) Mercury cell being primary in nature can be used only once.

NEET PREVIOUS YEARS QUESTIONS-EXPLANATIONS

1. 2)
For cell,
$Zn|ZnSO_4(0.01\ M)\,||\,CuSO_4(1.0\ M)|Cu$

$$E_{cell} = E^\circ_{cell} - \frac{2.303RT}{nF}\log\frac{[Zn^{2+}]}{[Cu^{2+}]^2}$$

$$\therefore\ E_1 = E^\circ_{cell} - \frac{2.303RT}{2 \times F} \times \log\frac{(0.01)}{(1)^2}$$

When concentrations are changed for $ZnSO_4$ and $CuSO_4$, we can write

$$E_2 = E^\circ_{cell} - \frac{2.303RT}{2F} \times \log\frac{1}{(0.01)^2}\ \therefore\ E_1 > E_2$$

2. 1) $2H^+(aq) + 2e^- \rightarrow H_2(g)$

$$\therefore E = E^0 - \frac{0.0591}{2}\log\frac{P_{H_2}}{[H^+]^2}$$

$$0 = 0 - 0.0295\log\frac{P_{H_2}}{(10^{-7})^2}$$

$$\frac{P_{H_2}}{(10^{-7})^2} = 1$$

$$P_{H_2} = 10^{-14}\text{atm}$$

3. 4) A device that converts energy of combustion of fuels, directly into electrical energy is known as fuel cell.

4. 3)

$$\underset{0.1\,\text{mole}}{\overset{(+6)}{MnO_4^{2-}}} \longrightarrow \overset{(+7)}{MnO_4^{-}} + e^{-}$$

Quantity of electricity required = 0.1F = 0.1 × 96500 = 9650 C

5. 4) $w_{O_2} = n_{O_2} \times 32$

$$w_{O_2} = \frac{5600}{22400} \times 32 = 8g = 1 \text{ equivalent of } O_2 = 1 \text{ equivalent of } Ag = 108$$

6. 3)

$$E_{cell} = E^{o}_{cell} - \frac{2.303\,RT}{nF}\log_{10} Q$$

at equlibrium $E_{cell} = 0,\ Q = K_{eq.}$

$$0 = E^{o}_{cell} - \frac{0.0591}{1}\log_{10} K_{eq.}$$

$E^{o}_{cell} = +0.0591 \log_{10} K_{eq.}$

$0.59 = +\ 0.0591 \log_{10} K_{eq.}$

$+10 = \log_{10} K_{eq.}$

$K_{eq.} = 10^{+10}$

7. 1) $2Fe^{3+}(aq) + 2I^{-}(aq) \rightarrow 2Fe^{2+}(aq) + I_2(aq)$

$n = 2\ ;\ \Delta G^{0} = -nFE^{0}$

$= -2 \times 96500 \times (0.24) = -46320J$

$= -46.32\,kJ\,mol^{-1}$

8. 4)

$CH_3COOH \rightarrow CH_3COO^- + H^+$ ...(1)

$H_2SO_4 \rightarrow 2H^+ + SO_4^{-2}$ ...(2)

$K_2SO_4 \rightarrow 2K^+ + SO_4^{-2}$ ...(3)

$CH_3COOK \rightarrow CH_3COO^- + K^+$ ...(4)

According to Kohlrausch's law-

$$\lambda^{o}_{CH_3COOH} = \lambda^{o}_{CH_3COO^-} + \lambda^{o}_{H^+}$$

$$\text{eq. (1)} = \text{eq. (4)} + \text{eq.}\frac{(2)}{2} - \text{eq.}\frac{(3)}{2}$$

$$\therefore\ \lambda^{o}_{CH_3COOH} = z + \frac{x}{2} - \frac{y}{2}$$

$$\boldsymbol{\lambda^{o}_{CH_3COOH} = \frac{(x-y)}{2} + z(S \times cm^2 mol^{-1})}$$

Reducing power of metal $\propto \dfrac{1}{SRP}$

9. 2) **K > Al > Cr > Ag.**

10. (3) According to electrode potential series,

$Au^{+3} + 3e^{-} \rightarrow Au_{(s)}$ (has topmost position ⇒ max. SRP)

11. (2) In typical fuel cell

Reactants = H_2, O_2

Products = H_2O

12. 3)On electrolysis of dil $H_2\,SO_4$

As anodes $H_2O \rightarrow 2H^+ + \frac{1}{2}O_2 + 2e^-$ oxygen gas is liberated

13. 2)E of Ca = 40/2 = 20 gr
For the deposition of 1gram equivalent wt (20gr) 1F is required

14. 1) $\wedge_m^0 CH_3OOH = \wedge_{HCl}^0 + \wedge_{CH_3COONa}^0 - \wedge_{NaCl}^0 = (426.16+91)-126.45 = 390.71\,S\,cm^2 mol^{-1}$

15. 2) $\lambda_0 = \Lambda^{\circ}_{H^+} + \lambda_0\left(CH_3COO^-\right)$ = 350 + 50= 400 ; λ =20 ; C = 0.007 m

$$K_a = \frac{C\lambda^2}{\lambda_0\left(\lambda_0-\lambda\right)} = \frac{7\times10^{-3}\times(20)^2}{400(400-20)};\ K_c = 1.8\times10^{-5}$$

16. In this reaction $CunO_4^-$ reduce at cathode and H_2O oxidise at anode. So given $E^o_{MnO_4^-/Mn^{+2}} = +1.510V$,

$E^o_{O_2/H_2O} = +1.223V$

$E^o_{cell} = 1.510 - 1.223 = +0.287$ so, the reaction is spontaneous

17. $2CuSO_4 + 2Ag \rightarrow 2Cu + Ag_2SO_4$ Ag cannot displace copper from $CuSO_4$ solute

18.: $E = 1.05 - 0.0295\log\frac{0.001}{(0.001)^2\,10^{-3}} = 1.05 - 0.0295\times3 = 10.5 - 0.0885 = 0.9615v$