CHAPTER

03

Electrochemistry

A Quick Recapitulation of the Chapter

1. **Electrochemistry** is the study of the chemical reactions which produce electrical energy or those chemical reactions which take place on passing electricity.
2. **Electrochemical cell**

Electrolytic cell	Galvanic or voltaic cell
Non-spontaneous redox reaction is made to occur.	Spontaneous redox reaction produces electric current.
Occur in same electrolytic solution.	Occur in different electrolytic solutions.
No salt bridge is required.	Salt bridge is essential to maintain electro neutrality.
Cathode is negative and anode is positive electrode.	Cathode is positive and anode is negative electrode.

3. **Representation of a Cell**
A cell can be represented as

$$\xrightarrow{e^-}$$

$$\underbrace{\text{Zn}(s)\,|\,\text{Zn}^{2+}(aq)}_{\text{Anode half-cell}}\,||\,\underbrace{\text{Cu}^{2+}(aq)\,|\,\text{Cu}(s)}_{\text{Cathode half-cell}}$$

$$\uparrow$$

Salt bridge

$$\xleftarrow[\text{Current}]{}$$

4. Potential difference between the metal and the metal ion, in which electrode is dipped is called **electrode potential**. At 1 atm pressure and 1M concentration of electrolyte, electrode potential is also known as **standard electrode potential** ($E°$).
5. The difference in potentials of the two half-cells of a cell is known as **electromotive force** (emf) or **cell potential**.

$$E°_{\text{cell}} = E°_{\text{ox}} + E°_{\text{red}}$$
$$= E°_{\text{cathode}} - E°_{\text{anode}} = E°_{\text{right}} - E°_{\text{left}}$$

6. The arrangement of elements in the order of increasing reduction potential values is called **electrochemical series.**

The standard electrode potential for flourine is highest which indicates that fluorine gas (F_2) has maximum tendency to get reduced to fluoride ion (F^-). In case of lithium ion is weakest oxidising agent while lithium metal is the most powerful reducing agent in an aqueous solution.

In this series Li is placed on the top and F at the bottom. Their positions signify that Li is the most power reducing agent which F is the most powerful oxidising agent known.

7. **Applications of Electrochemical Series**
   (i) Lower the value of reduction potential, greater is its reducing power.
   (ii) Metal having lesser value of standard potential can displace the other having higher value from its aqueous salt solution.
8. **Nernst Equation**

$$E_{\text{cell}} = E°_{\text{cell}} - \frac{0.0591}{n}\log\frac{[\text{Anode}]}{[\text{Cathode}]}$$

For concentration cell, $E = \frac{0.0591}{n}\log\frac{C_1}{C_2}$

Standard Gibb's energy of the reaction,

$$\Delta_r G° = -nFE°_{\text{cell}} = -2.303RT\log K$$

9. If two half reactions having potentials $E°_1$ and $E°_2$ are combined to give a third half reaction having a potential $E°_3$, then $\Delta G_3^\circ = \Delta G_1^\circ + \Delta G_2^\circ$

or $$E°_3 = \frac{n_1E_1^\circ + n_2E_2^\circ}{n_3}$$

10. **Relation between E_{cell} and pH**

$$E_{H^+/H_2} = E°_{H^+/H_2} - 0.0591 \log \frac{1}{[H^+]}$$
$$= 0 - 0.0591\,(-\log[H^+]) = -0.0591\,\text{pH}$$
$$[\because \text{pH} = -\log[H^+]]$$

Relation between E_{cell}° and K_C

$$E_{cell}^\circ = \frac{0.0591\,\text{V}}{n} \log K_C$$

11. **Conductance** (G) is the reciprocal of resistance and may be defined as the ease with which the electric current flows through the conductor, $G = \frac{1}{R}$

Here, R = Resistance [SI unit = Siemen (S) or ohm^{-1}]

12. **Conductivity** or **specific conductance (κ),** of an electrolytic solution may be defined as the conductance of a solution of 1 cm length with its area of cross section equal to 1 cm^2.

$$\kappa = \frac{1}{R} \times \text{cell constant}$$

Its unit is Sm^{-1} or $\text{ohm}^{-1}\ \text{cm}^{-1}$

13. **Molar conductivity** (Λ_m) is defined as the conducting power of all the ions produced by dissolving one gram mole of an electrolyte placed between two electrodes placed at a distance of 1 cm.

It is related to specific conductance as, $\Lambda_m = \frac{\kappa \times 1000}{M}$

where, M = molarity of solution

Its unit is $\Omega^{-1}\ \text{cm}^2\ \text{mol}^{-1}$

14. On dilution, equivalent and molar conductivities increase and specific conductivity decreases.

15. **Kohlrausch's law** states that limiting molar conductivity of an electrolyte is the sum of individual contributions of conductivity of cation and anion of the electrolyte.

$$\Lambda_m^\circ = x\lambda_m^\circ + y\lambda_m^\circ$$

(x and y = number of cations and anions)

16. Degree of dissociation, $\alpha = \frac{\Lambda_m}{\Lambda_m^\circ}$

[Λ_m = molar conductivity at a given concentration, Λ_m° = limiting molar conductivity].

17. Dissociation constant, $K_a = \frac{C\alpha^2}{(1-\alpha)}$ (for weak electrolyte)

[C = concentration, α = degree of dissociation]

18. **Faraday's first law of electrolysis** During electrolysis, amount of substance (w) liberated at electrode $= Z \times i \times t$.

19. **Faraday's second law of electrolysis:**

$$\frac{w_1}{w_2} = \frac{E_1}{E_2}$$

[w_1 and w_2 are the amounts of substances produced at different electrodes, E_1 and E_2 = equivalent masses]

20. If a number of cells are connected in series, the arrangement is called a **battery**. e.g. primary cells, i.e. non-rechargeable cell (e.g. : dry cell, mercury cell), secondary cells i.e. rechargeable cells (e.g lead storage battery, nickel-cadmium cell) and fuel cells i.e. cells using hydrogen, methane as fuel (e.g. $H_2 - O_2$ fuel cell)

21. **Corrosion** is the oxidative deterioration of a metal such as conversion of iron to rust. It is basically an electrochemical phenomenon.

Objective Questions Based on NCERT Text

Topic 1

Electrochemical Cells, Electrodes and Electrode Potential

1. The incorrect statement about electrochemistry is
   (a) the study of generation of electricity by spontaneous chemical reactions
   (b) the study of use of electrical energy to bring out a non-spontaneous chemical reaction
   (c) reaction based on it is not so efficient and ecofriendly
   (d) Transmission of sensory signals through cells to brain are of electrochemical origin

2. $\text{Zn}(s) + \text{Cu}^{2+}(aq) \longrightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s)$

The above redox reaction is used in
   (a) Galvanic cell (b) Daniell cell
   (c) Voltaic cell (d) All of these

3. For the generation of electrical potential of 1.1 V, the ideal concentration of zinc and copper ion (in mol dm^{-3}) in the Galvanic cell is
   (a) 1 and 1 (b) 2 and 1 (c) 3 and 2 (d) 4 and 2

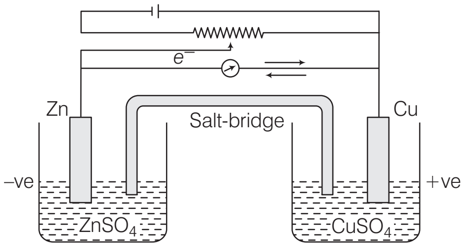
4. In general, in a Galvanic cell
(a) electrons flow from positive to negative electrode
(b) current flow from positive to negative electrode
(c) electrons and current flow in the same direction
(d) All of the above are correct

5. What happens when applied external opposite potential in a Daniell cell reaches to 1.1 V?
(a) Chemical reaction stops
(b) Current starts flowing in a direction opposite to that of the electrons
(c) Current starts flowing in the same direction as of electrons
(d) The cell behaves like an electrolytic cell

6. When applied external opposite potential in case of Daniell cell is less than 1.1V, then
(a) the cell behaves like an electrolytic cell
(b) electrons flow from zinc rod to copper rod
(c) zinc deposites at the zinc electrode and copper dissolves from copper rod
(d) Current flow from zinc rod to copper rod

7. Electrolytic cell is a device
(a) in which a non-spontaneous chemical reaction is carried out at the expense of electrical energy
(b) in which a spontaneous chemical reaction is carried out to generate electrical energy
(c) in which applied opposite potential is less than the cell potential
(d) Both(a) and (c)

8. Look at the Daniell cell,



The incorrect statement related to this cell is
(a) electron flow from zinc electrode to copper electrode, when applied external potential is 0.9 V
(b) current flow from zinc to copper electrode, when applied external potential is 1.5 V
(c) current flow from zinc to copper electrode when applied external potential is 0.9 V
(d) zinc electrode acts as cathode and copper as anode when applied external potential is 1.5 V

9. Select the correct statement(s) for the Galvanic cell.
(a) It is not an electrochemical cell
(b) It converts the electrical energy to chemical energy
(c) In this device the Gibbs energy of spontaneous redox reaction is used
(d) They are not used for running electrical gadgets

10. Select the false statement for Daniell cell.
(a) Reduction half-cell reaction,
$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$
(b) Oxidation half-cell reaction,
$$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$$
(c) Reduction half-cell portion is also called redox couples
(d) All of the above

11. The electrode potential is known as standard electrode potential
(a) when it is reduction potential (according to IUPAC)
(b) when concentrations of all the species involved is unity
(c) when it is oxidation potential (according to IUPAC)
(d) Both (a) and (b)

12. In a Galvanic cell
(a) anode is the electrode at which reduction takes place
(b) anode is the electrode which has a positive potential with respect to solution
(c) cathode is the site of reduction
(d) Both (b) and (c)

13. Select the incorrect statement
(a) Cell potential is the difference between electrode potentials of cathode and anode
(b) Cell potential is unitless quantity
(c) It is also called cell electromotive force (emf) when no current is drawn through cell
(d) In a cell anode is kept on left and cathode on right

14. In the cell representation,
(a) a single vertical line separates an electrode from the electrolytes
(b) two vertical lines shows salt bridge
(c) two vertical lines separate two electrolytes
(d) All of the above are true

15. Select the incorrect statement.
(a) Cell potential is given by the potential of half-cell on the right hand side minus the potential of half-cell on the left hand side
(b) Cell potential is given by
$$E^\circ_{cell} = E^\circ_{red} + E^\circ_{oxi}$$
(c) Cell potential is given by
$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$$
(d) None of the above

16. The correct cell representation of the following cell is
$$Zn + 2Ag^+ \longrightarrow Zn^{2+} + 2Ag$$
(a) $Zn(s)|Zn^{2+}(aq)||Ag^+(aq) + Ag(s)$
(b) $Zn^{2+}|Zn||Ag|Ag^+$
(c) $Zn(aq)|Zn^{2+}(s)||Ag^+(s)|Ag(aq)$
(d) $Zn(s)|Ag^+(aq)||Zn^{2+}(aq) + Ag(s)$

17. Select the correct statement(s) for the given reaction.

$$\mathrm{Cu}\,(s) + 2\mathrm{Ag}^+(aq) \longrightarrow \mathrm{Cu}^{2+}(aq) + 2\mathrm{Ag}\,(s)$$

(a) Reaction at cathode $\mathrm{Cu}\,(s) \longrightarrow \mathrm{Cu}^{2+}(aq) + 2e^-$
(b) $E_{\mathrm{cell}} = E^\circ_{\mathrm{Cu}^{2+}/\mathrm{Cu}} = E^\circ_{\mathrm{Ag}^+/\mathrm{Ag}}$
(c) $\mathrm{Ag}^+(aq)\,|\,\mathrm{Ag}(s)\,||\,\mathrm{Cu}^{2+}(aq)\,|\,\mathrm{Cu}\,(s)$
(d) None of the above

18. E°_{cell} of the reaction is

$$\mathrm{Zn}\,(s) + \mathrm{Cu}^{2+}(aq) \longrightarrow \mathrm{Zn}^{2+}(aq) + \mathrm{Cu}\,(s);$$

$[\text{If } E^\circ_{\mathrm{Zn}^{2+}/\mathrm{Zn}} = -0.76,\ E^\circ_{\mathrm{Cu}^{2+}/\mathrm{Cu}} = 0.34\ \mathrm{V}]$

(a) 0.34V (b) – 0.76 V (c) + 1.10 V (d) – 1.10 V

19. Calculate the standard cell potential for the following Galvanic cell, $\mathrm{Cr}|\mathrm{Cr}^{3+}||\mathrm{Cd}^{2+}|\mathrm{Cd}$

$[\text{Given, } E^\circ_{\mathrm{Cr}^{3+}/\mathrm{Cr}} = -0.74\mathrm{V} \text{ and } E^\circ_{\mathrm{Cd}^{2+}/\mathrm{Cd}} = -0.40\mathrm{V}]$

(a) 0.74 V (b) −0.34 V (c) + 0.34 V (d) 1.14 V

20. The reduction potential of hydrogen half-cell will be negative if **(AIEEE 2011)**

(a) $p(\mathrm{H}_2) = 1\,\mathrm{atm}$ and $[\mathrm{H}^+] = 2.0\,\mathrm{M}$
(b) $p(\mathrm{H}_2) = 1\,\mathrm{atm}$ and $[\mathrm{H}^+] = 1.0\,\mathrm{M}$
(c) $p(\mathrm{H}_2) = 2\,\mathrm{atm}$ and $[\mathrm{H}^+] = 1.0\,\mathrm{M}$
(d) $p(\mathrm{H}_2) = 2\,\mathrm{atm}$ and $[\mathrm{H}^+] = 2.0\,\mathrm{M}$

21. Select the incorrect statement about standard hydrogen electrode.

(a) $\mathrm{H}^+(aq)\,|\,\mathrm{H}_2(g)|\,\mathrm{Pt}(s)$
(b) $\frac{1}{2}\mathrm{H}_2(g) \longrightarrow \mathrm{H}^+(aq) + e^-$
(c) The electrode is dipped in basic solution and pure hydrogen gas is bubbled through it
(d) Concentration of both the oxidised and reduced forms of hydrogen is maintained at unity

22. Standard electrode potential (reduction potential) becomes equal to cell potential, when

(a) either anode or cathode is the hydrogen electrode
(b) anode is the hydrogen electrode
(c) cathode is the hydrogen electrode
(d) None of the above

23. $\mathrm{Pt}(s)|\,\mathrm{H}_2(g, 1\,\mathrm{bar})|\,\mathrm{H}^+(aq, 1\mathrm{M})\,||\mathrm{Cu}^{2+}(aq, 1\,\mathrm{M})|\mathrm{Cu}$

The measured emf of the cell is 0.34 V means

(a) Cu^{2+} ions get reduced easily than H^+ ion
(b) hydrogen ions cannot oxidise Cu
(c) copper does not dissolve in HCl
(d) All of the above

24. Select the correct statement for the cell

$\mathrm{Pt}(s)\,|\,\mathrm{H}_2(g, 1\,\mathrm{bar})|\,\mathrm{H}^+(aq, 1\mathrm{M})\,||\,\mathrm{Zn}^{2+}(aq, 1\mathrm{M})|\,\mathrm{Zn}$

Given, $\mathrm{Zn}^{2+}/\mathrm{Zn} = -0.76\ \mathrm{V}$

(a) emf of the cell = + 0.76 V
(b) $\mathrm{Zn}^{2+}(aq, 1\,\mathrm{M}) + 2e^- \longrightarrow \mathrm{Zn}(s)$
(c) hydrogen ion can reduce zinc
(d) All of the above

25. In the reaction of copper with nitric acid,

(a) Cu is oxidised to Cu^{2+} ion by hydrogen ion
(b) Cu is oxidised to Cu^{2+} ion by nitrate ion
(c) Cu is reduced by hydrogen ion
(d) Both (b) and (c)

Topic 2

Electrochemical Series and Nernst Equation

(Including Relation between E_{cell} and K or ΔG)

26. If the standard electrode potential of an electrode is greater than zero then

(a) its oxidised form is more stable as compared to hydrogen gas
(b) its reduced form is more stable as compared to hydrogen gas
(c) its reduced form is less stable as compared to hydrogen gas
(d) All of the above

27. Select the incorrect statement

(a) Fluorine gas is the strongest oxidising agent
(b) Lithium ion is weakest oxidising agent
(c) If the standard reduction potential is positive then hydrogen gas is more stable than reduced form
(d) Lithium metal is powerful reducing agent in an aqueous solution

28. Compound *XA*, *XB*, *XC* are added into separate test tubes containing *A, B, C* solutions. *XB* react with *A* and *C*. *XA* does not react with any of these. *XC* reacts with *A*.

Arrange the anion in the decreasing order of their oxidation.

(a) A^-, B^-, C^-
(b) B^-, C^-, A^-
(c) C^-, A^-, B^-
(d) B^-, A^-, C^-

29. Select the strongest reducing agent from the given electrode potential.

(a) $E^\circ_{\mathrm{Zn}^{2+}/\mathrm{Zn}}$ (b) $E^\circ_{\mathrm{Cr}^{3+}/\mathrm{Cr}}$
(c) $E^\circ_{\mathrm{H}^+/\mathrm{H}_2}$ (d) $E^\circ_{\mathrm{Fe}^{3+}/\mathrm{Fe}^{2+}}$

30. Standard electrode potential of three metals X, Y and Z are -1.2V, $+0.\ 5\text{V}$ and -3.0V respectively. The reducing power of these metals will be

(CBSE AIPMT 2011)

(a) $X > Y > Z$
(b) $Y > Z > X$
(c) $Y > X > Z$
(d) $Z > X > Y$

31. Given, $E^\circ_{\text{I}_2/\text{I}^-}$ $(1\text{M}) = +\ 0.54\text{V}$,

and $E^\circ_{\text{Br}_2/\text{Br}^-}$ $(1\text{M}) = 1.09\text{V}$

On this basis the feasible reaction is

(a) $2\text{Br}^- + \text{I}_2 \longrightarrow \text{Br}_2 + 2\text{I}^-$
(b) $\text{I}_2 + \text{Br}_2 \longrightarrow 2\text{I}^- + 2\text{Br}^-$
(c) $2\text{I}^- + \text{Br}_2 \longrightarrow \text{I}_2 + 2\text{Br}^-$
(d) $2\text{I}^- + 2\text{Br}^- \longrightarrow \text{I}_2 + \text{Br}_2$

32. For the reaction, $M^{n+}(aq) + ne^- \longrightarrow M(s)$, select the best suitable representation of Nernst equation, when the solid M is taken.

(a) $E_{M^{n+}/M} = E^\circ_{(M^{n+}/M)} - \dfrac{RT}{nF} \ln \dfrac{[M]}{[M^{n+}]}$

(b) $E_{M^{n+}/M} = E^\circ_{M^{n+}/M} - \dfrac{RT}{nF} \ln \dfrac{1}{[M^{n+}]}$

(c) $E_{M^{n+}/M} = E^\circ_{M^{n+}/M} - \dfrac{RT}{nF} \ln \dfrac{[M^{n+}]}{[M]}$

(d) $E_{M^{n+}/M} = E^\circ_{M^{n+}/M} - \dfrac{RT}{nF} \ln\, [M^{n+}]$

33. Select the most appropriate Nernst equation for Daniell cell containing Cu^{2+} and Zn^{2+} ions.

(a) $E_{(\text{Cu}^{2+}/\text{Cu})} = E^\circ_{(\text{Cu}^{2+}/\text{Cu})} - \dfrac{RT}{F} \ln \dfrac{1}{[\text{Cu}^{2+}(aq)]}$

(b) $E_{(\text{Zn}^{2+}/\text{Zn})} = E^\circ_{(\text{Zn}^{2+}/\text{Zn})} - \dfrac{RT}{2F} \ln \left[\dfrac{1}{\text{Zn}^{2+}(aq)}\right]$

(c) $E_{\text{cell}} = E^\circ_{\text{cell}} - \dfrac{RT}{2F} \ln \dfrac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$

(d) All of the above

34. Select the appropriate statement for Daniell cell containing Cu^{2+} and Zn^{2+} ions.

(a) E_{cell} does not depend on the concentration of Cu^{2+} and Zn^{2+} ions
(b) E_{cell} decreases with increase in the concentration of Cu^{2+} ions
(c) E_{cell} increases with decrease in the concentration of Zn^{2+} ions
(d) All of the above

35. The electrode potentials for

$\text{Cu}^{2+}(aq) + e^- \longrightarrow \text{Cu}^+(aq)$ and

$\text{Cu}^+(aq) + e^- \longrightarrow \text{Cu}\,(s)$ are $+0.\ 15\text{V}$ and $+0.50\text{V}$ respectively. The value of $E^\circ_{\text{Cu}^{2+}/\text{Cu}}$ will be

(CBSE AIPMT 2011)

(a) 0.150 V (b) 0.500 V
(c) 0.325 V (d) 0. 650 V

36. For the cell,

$$\text{Cu}\,|\,\text{Cu}^{2+}\,||\,\text{Ag}^+\,|\ \text{Ag},\ E^\circ_{\text{cell}} = +0.46\ \text{V}$$

If concentration of Cu^{2+} ions is doubled, then $E^\circ_{\textbf{cell}}$ will be

(a) halved (b) doubled
(c) four times (d) remains the same

37. Standard electrode potential for $\text{Sn}^{4+}/\text{Sn}^{2+}$ couple is $+0.15\text{V}$ and that for the Cr^{3+}/Cr couple is -0.74V. These two couples in their standard state are connected to make a cell. The cell potential will be

(CBSE AIPMT 2011)

(a) + 1.83 V (b) +1.19 V
(c) + 0.89 V (d) + 0.18 V

38. Consider the following cell reaction

$$2\text{Fe}(s) + \text{O}_2(g) + 4\text{H}^+(aq) \longrightarrow 2\text{Fe}^{2+}(aq) + 2\text{H}_2\text{O}(l),\ E^\circ = 1.67\ \text{V}$$

At $[\text{Fe}^{2+}] = 10^{-3}$ M, $p_{\text{O}_2} = 0.1$ atm and pH = 3, the cell potential at 25°C is **(JEE Main 2011)**

(a) 1.47 V (b) 1.77 V
(c) 1.87 V (d) 1.57 V

39. A button cell used in watches, functions as following

$$\text{Zn}(s) + \text{Ag}_2\text{O}\,(s) + \text{H}_2\text{O}(l) \rightleftharpoons 2\text{Ag}(s) + \text{Zn}^{2+}(aq) + 2\text{OH}^-(aq)$$

If half-cell potentials are

$$\text{Zn}^{2+}(aq) + 2e^- \longrightarrow \text{Zn}(s);\ E^\circ = -0.76\ \text{V}$$

$\text{Ag}_2\text{O}\,(s) + \text{H}_2\text{O}(l) + 2e^- \longrightarrow 2\text{Ag}\,(s) + 2\text{OH}^-;$

$E^\circ = 0.34\text{V}$. The cell potential will be **(NEET 2013)**

(a) 1.1 V (b) 0.42 V
(c) 0.84 V (d) 1. 34 V

40. A hydrogen gas electrode is made by dipping platinum wire in a solution of HCl and $\text{pH} = 10$ and by passing hydrogen gas around the platinum wire at 1 atm pressure. The oxidation potential of electrode would be **(NEET 2013)**

(a) 0.059 V (b) 0.59 V
(c) 0.118 V (d) 0.18 V

41. Calculate the equilibrium constant for the reaction,

$$\mathrm{Zn}(s)+\mathrm{Cu}^{2+}(aq) \longrightarrow \mathrm{Zn}^{2+}(aq)+\mathrm{Cu}(s)$$

[Given, $E^{\circ}_{\text{cell}} = 1.1$ V]

(a) 2×10^{32} (b) 2×10^{34}
(c) 2×10^{37} (d) 2×10^{39}

42. Equilibrium constant (K) is related with standard cell potential (E°_{cell}) but not with E_{cell}. This is because

(a) E°_{cell} is easier to calculate as compared to E_{cell}
(b) E_{cell} becomes zero at equilibrium but E°_{cell} does not
(c) At a given temperature, E_{cell} changes hence, values of K cannot be measured
(d) K also relates with E_{cell}

43. In the given reaction,

$$2\mathrm{Cu}^{+}(aq) \rightleftharpoons \mathrm{Cu}^{2+}(aq)+\mathrm{Cu}(s)$$

$E^{\circ}_{\mathrm{Cu}^{+}/\mathrm{Cu}} = 0.6$ V and $E^{\circ}_{\mathrm{Cu}^{2+}/\mathrm{Cu}} = 0.41$ V

Find out the equilibrium constant.

(a) 2.76×10^{2}
(b) 2.76×10^{4}
(c) 2.76×10^{6}
(d) 2.76×10^{8}

44. Which of the following statements is/are incorrect?

(a) Electrical work done in one second is equal to electrical potential multiplied by total charge passed
(b) Charge has to be passed reversibly in order to obtain the maximum work from Galvanic cell
(c) The reversible work done by Galvanic cell is equal to decrease in its Gibbs energy
(d) $\Delta G = -nE_{\text{cell}}$

45. If the E°_{cell} for a given reaction has a negative value, then which of the following gives the correct relationships for the values of ΔG° and K_{eq}? **(CBSE AIPMT 2011)**

(a) $\Delta G^{\circ} > 0;\ K_{\text{eq}} < 1$ (b) $\Delta G^{\circ} > 0;\ K_{\text{eq}} > 1$
(c) $\Delta G^{\circ} < 0;\ K_{\text{eq}} > 1$ (d) $\Delta G^{\circ} < 0;\ K_{\text{eq}} < 1$

46. Calculate the standard Gibbs free energy of

$$\mathrm{Zn}(s)+\mathrm{Cu}^{2+}(aq) \longrightarrow \mathrm{Zn}^{2+}(aq)+\mathrm{Cu}(s)$$

(a) $-91\,FE_{\text{cell}}$ (b) $-2\,FE_{\text{cell}}$ (c) $-3\,FE_{\text{cell}}$ (d) $-4\,FE_{\text{cell}}$

47. The Gibbs energy for the decomposition of $\mathrm{Al_2O_3}$ at 500°C is as follows

$$\frac{2}{3}\mathrm{Al_2O_3} \longrightarrow \frac{4}{3}\mathrm{Al}+\mathrm{O_2};\ \Delta_r G = +960\ \text{kJ mol}^{-1}$$

The potential difference needed for the electrolytic reduction of aluminium oxide ($\mathrm{Al_2O_3}$) at 500°C is atleast **(CBSE AIPMT 2012)**

(a) 4.5V (b) 3.0V (c) 2.5V (d) 5. 0V

Topic 3
Conductance, Conductivity and Kohlrausch's Law

48. Select the incorrect statement.

(a) SI units of resistance is $\text{kg m}^2/\text{S}^3\text{A}^2$
(b) Wheat stone bridge is used to measure resistance as well as resistivity
(c) Resistance = resistivity, when $l = 1\text{m}$ and $A = 1\text{m}^2$
(d) Inverse of resistance is called specific conductance

49. The correct statement about conductivity is

(a) it is the inverse of resistivity
(b) its SI unit is Sm^{-1}
(c) it is equal to conductance, when $\frac{l}{\text{A}} = 1\text{m}^{-1}$
(d) All of the above

50. Organic conducting polymers are

(a) polyaniline (b) polypyrrole
(c) polythiophene (d) All of these

51. What happens to the composition of metallic conductor when electron enters at one end and go out through other end?

(a) There will be change in metallic conductor
(b) Metallic conductor remain unchanged
(c) There will be change on metallic conductor only in the presence of applied pressure
(d) Metallic conductor will behave as insulator

52. Select the appropriate statement(s) about ionic conductance.

(a) It is the conductance of electricity by ions present in the solutions
(b) It depends on nature of solvent, its viscosity and of electrolyte
(c) It depends upon the size of ions produced and solvation
(d) All of the above

53. Which of the following statement is correct?

(a) Electrolytic conductance decreases but electrical conductance increases with increase in temperature
(b) Electrolytic conductance increases but electrical conductance decreases with increase in temperature
(c) Electrolytic as well as electrical conductance decrease with rise in temperature
(d) Electrolytic as well as electrical conductance increase with rise in temperature

54. Alternating current (AC) is used instead of direct current (DC) for the measurement of conductivity of ionic solutions. This is because

(a) alternating current changes the composition of the solutions
(b) direct current changes the composition of the solution
(c) direct current is more powerful
(d) both (b) and (c)

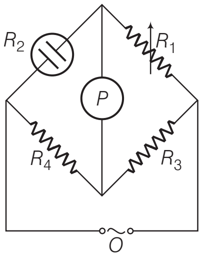
55. Which of the following statement(s) is/are true regarding to the cell constant?

(a) It is the ratio of distance between the electrodes and their area of cross-section
(b) Its SI unit is cm^{-1}
(c) It is the measured directly from l and A
(d) All of the above are correct

56. KCl solution is generally used to determine the cell constant because

(a) it is highly ionic in nature
(b) its conductivity is known accurately at various concentration and different temperatures
(c) size of cations and anion are comparable
(d) All of the above

57. Which of the following information(s) is/are false for this figure?



(a) This assembly is used for measuring conductivity of solution
(b) O is an oscillator, i.e, a source of AC power
(c) P is the conductivity cell
(d) unknown resistance is measured by using the formula, $R_2 = \frac{R_1 R_4}{R_3}$

58. The resistance of the cell containing KCl solution at 23°C was found to be 55 Ω. Its cell constant is 0.616 cm^{-1}. The conductivity of KCl solution $(\Omega^{-1} \text{ cm}^{-1})$ is

(a) 1.21×10^{-3}
(b) 1.12×10^{-2}
(c) 1.12×10^{-3}
(d) 1.21×10^{-2}

59. Which of the following statement is true for molar conductivity?

(a) It differs due to charge and size of ions in which they dissociate
(b) $\Lambda_m (\text{S cm}^2 \text{ mol}^{-1}) = \frac{\kappa \text{ (S cm}^{-1})}{1000 \text{ L cm}^{-3} \times \text{Molarity (mol L}^{-1})}$
(c) $\Lambda_m (\text{S cm}^2 \text{ mol}^{-1}) = \frac{\kappa \text{(S cm}^{-1}) \times 1000 \text{ (cm}^3 \text{ / L)}}{\text{Molarity (mol L}^{-1})}$
(d) All of the above

60. Resistance of 0.2 M solution of an electrolyte is 50Ω. The specific conductance of the solution is 1.3 Sm^{-1}. If resistance of the 0.4 M solution of the same electrolyte is 260 Ω, its molar conductivity is **(AIEEE 2012)**

(a) $6250 \text{ S m}^2 \text{ mol}^{-1}$
(b) $6.25\times10^{-4} \text{ S m}^2 \text{ mol}^{-1}$
(c) $625\times10^{-4} \text{ S m}^2 \text{ mol}^{-1}$
(d) $62.5 \text{ S m}^2 \text{ mol}^{-1}$

61. Select the correct statement.

(a) Conductivity decreases but molar conductivity increases with increase in dilution
(b) Conductivity increases but molar conductivity decreases with increase in dilution
(c) Conductivity decreases but molar conductivity increases with increase in concentration
(d) Conductivity as well as molar conductivity decrease with decrease in concentration

62. Conductivity decreases with decrease in concentration for weak as well as strong electrolytes.

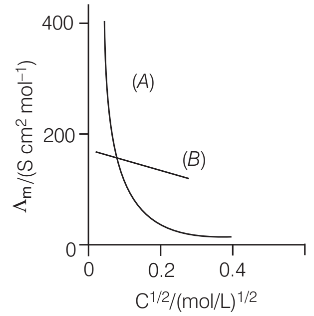
This is because

(a) number of current carrying particles per unit volume decreases
(b) volume of solution increases
(c) number of current carrying particles per unit volume increases
(d) None of the above

63. The conductance of electrolytic solution kept between the electrodes of conductivity cell at unit distance but having area of cross-section large enough to accommodate sufficient volume of solution is called

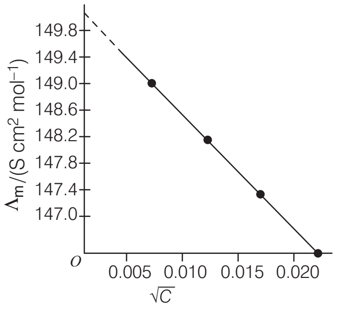
(a) limiting molar conductivity
(b) molar conductivity
(c) conductivity
(d) All of the above

64. The graph of molar conductivity *versus* $C^{1/2}$ is plotted. Which type of electrolyte are used in *A* and *B* ?



	A	B
(a)	Weak electrolyte	Strong electrolyte
(b)	Strong electrolyte	Weak electrolyte
(c)	Weak electrolyte	Weak electrolyte
(d)	Strong electrolyte	Strong electrolyte

65. Consider the following graph.



Here the limiting molar conductivity is

(a) 148.6 (b) 150 (c) 87.46 (d) 147

66. Select the correct statement for $\Lambda_m = \Lambda^\circ_m - AC^{1/2}$.

(a) This equation is for weak electrolyte

(b) Intercept is equal to Λ_m

(c) Slope is '*A*'

(d) Value of *A* depends on the charges of cation and anion

67. Two electrolytes *X* and *Y* are diluted. Λ_m of *Y* increases 1.5 times and for *X* it increases 25 times. Predict the strong electrolyte among *X* and *Y*.

(a) *X* (b) *Y*

(c) Both (a) and (b)

(d) There is no effect of dilution on nature of electrolyte

68. "Limiting molar conductivity of an electrolyte can be represented as sum of the individual contributions of anion and cation of the electrolyte".

Which law states the above statement?

(a) Henry's law

(b) Debye Onsager's law

(c) Kohlrausch law of independent migration of ions

(d) All of the above

69. Which of the following equation(s) is/are incorrect?

(a) $\Lambda^\circ_{m(KCl)} - \Lambda^\circ_{m(NaCl)} = \Lambda^\circ_{m(KBr)} - \Lambda^\circ_{m(NaBr)}$

(b) $\Lambda^\circ_{m(KBr)} - \Lambda^\circ_{m(NaBr)} = \Lambda^\circ_{m(KI)} - \Lambda^\circ_{m(NaI)}$

(c) $\Lambda^\circ_{m(NaBr)} - \Lambda^\circ_{m(NaCl)} = \Lambda^\circ_{m(KBr)} - \Lambda^\circ_{m(KCl)}$

(d) None of the above

70.

Ion	Λ° (S cm^2 mol^{-1})	**Ion**	Λ° (S cm^2 mol^{-1})
H^+	349.6	OH^-	199.1
Na^+	50.1	Cl^-	76.3
K^+	73.5	Br^-	78.1
Ca^{2+}	119.0	CH_3COO^-	40.9
Mg^{2+}	106.0	SO_4^{2-}	160.0

Calculate the limiting molar conductivity of NaCl and KBr from the above table.

	NaCl	KBr
(a)	126.4	150.6
(b)	126.4	151.6
(c)	150.6	126.4
(d)	151.6	126.4

71. The limiting molar conductivity of calcium acetate is (use the data given in Q. 70)

(a) 159.9 (b) 200.8 (c) 78.1 (d) 156.2

72. Molar conductivities (Λ°_m) at infinite dilution of NaCl, HCl and CH_3COONa are 126.4,425.9 and 91.0 S cm^{2}mol^{-1} respectively. Λ°_m for CH_3COOH will be **(NEET 2013)**

(a) 425.5 S cm^2 mol^{-1} (b) 180.5 S cm^2 mol^{-1}

(c) 290.85 S cm^2 mol^{-1} (d) 390.5 S cm^2 mol^{-1}

73. Select the correct statement(s) for the weak electrolyte.

(a) They have higher degree of dissociation at higher concentration

(b) The change in Λ_m with dilution is due to decrease in the degree of dissociation and number of ions in the solution that contain electrolyte

(c) Λ_m increases steeply on dilution near high concentration

(d) At infinite dilution, electrolyte dissociate completely but at such low concentration the conductivity of solution is low and cannot be measured accurately.

74. At 25°C, molar conductance of 0.1 molar aqueous solution of ammonium hydroxide is $9.54\Omega^{-1}$cm^{2}mol^{-1} and at infinite dilution, its molar conductance is 238 ohm^{-1}cm^{2}mol^{-1}. The degree of ionisation of ammonium hydroxide at the same concentration and temperature is **(NEET 2013)**

(a) 2.080% (b) 20.800% (c) 4 .008 % (d) 40 .80 %

Topic 4

Electrolytic Cell and Electrolysis

75. Which of the following statements is/are incorrect for electrolytic cell?
(a) External source of voltage is applied to carry the chemical reaction
(b) These cells are mainly used in laboratory and chemical industry
(c) Simplest electrolytic cell consists of two copper strips dipped in an aqueous solution of $CuSO_4$
(d) None of the above

76. Select the incorrect statement.
(a) Highly reactive metals are produced by electrochemical reduction of their respective cations
(b) Fused chloride of sodium is used to produce sodium by electrolysis
(c) Oxides of Mg or Al are used to produce these metals by electrolysis
(d) Electrolysis of Al is done in the presence of cryolite

77. In order to obtain pure copper from impure copper by electrolysis
(a) anode is made from impure copper
(b) copper is oxidised at anode and deposited at cathode
(c) a solution of copper sulphate is used as electrolyte
(d) All of the above

78. Which of the following statement is incorrect?
(a) Mass of any substance deposited or liberated during electrolysis is proportional to the charge passed through the electrolyte
(b) Amounts of different substances liberated by same quantity of electricity are proportional to their chemical equivalent weights
(c) Equivalents of a substance are obtained by dividing its atomic mass by *n*-factor
(d) Coulometer is an instrument which determine the quantity of electricity passed from the amount of metal deposited or consumed

79. What will be the weight of silver deposited, if 96.5 A of current is passed into aqueous solution of $AgNO_3$ for 100 s?
(a) 1.08 g (b) 10.8 g
(c) 108 g (d) 1080 g

80. When 0.1 mole of MnO_4^{2-} is oxidised, the quantity of electricity required to completely oxidise MnO_4^{2-} to MnO_4^- is **(CBSE AIPMT 2014)**
(a) 96500 C
(b) $2 \times 96500\,C$
(c) 9650C
(d) 96.50 C

81. An electrolytic cell contains alumina. If we have to obtain 50 g Al by using 105 A of current, the time required is
(a) 1.54 hr (b) 2.15 hr
(c) 1.42 hr (d) 1.32 hr

82. 1.5 A current is flowing through a metallic wire. If it flows for 3 hrs, how many electrons would flow through the wire?
(a) 2.05×10^{22} electrons (b) 1.0×10^{23} electrons
(c) 10^{24} electrons (d) 4.5×10^{23} electrons

83. How many coulombs are required in order to reduce 12.3 g of nitrobenzene to aniline?
(a) 579 C (b) 5790 C
(c) 57900 C (d) 579000 C

84. Calculate the time required to deposite 2×10^{-3} cm thick layer of silver (density 1.05 g cm^{-3}) on the surface of area $100\,cm^{-3}$ by passing a current of 10 A through silver nitrate solution?
(a) 125 s (b) 115 s
(c) 18.7 s (d) 27.25 s

85. 6 A current with 75% efficiency is passed through a cell for 6 h? ($z = 4 \times 10^{-4}$). The amount of metal deposited will be
(a) 22.4 g (b) 36.0 g
(c) 32.4 g (d) 38.8 g

86. Same amount of electricity is passed through the solutions of HCl and $CuSO_4$. If 6.35 g of copper is deposited from $CuSO_4$ solution, the amount of hydrogen liberated at STP will be
(a) 0.01 (b) 0.1
(c) 0.001 (d) 1

87. 10 F of electricity is passed through the solutions of silver nitrate, copper sulphate and ferric chloride. The amount of metal deposited at cathode in each case respectively are
(a) Ag = 1080 g Cu = 3175 g Fe = 186.7 g
(b) Ag = 108 g Cu = 31.75 g Fe = 18.67 g
(c) Ag = 1080 g Cu = 317.5 g Fe = 186.7 g
(d) Ag = 10.8 g Cu = 317.5 g Fe = 186.7 g

88. When aqueous sodium chloride solution is electrolysed
(a) at cathode H^+ is reduced into H_2 instead of Na^+
(b) at cathode Na^+ is reduced to Na
(c) Cl^- is oxidised into Cl_2 at cathode
(d) Both (b) and (c)

89. What will happen to pH of brine solution if electrolysed?
(a) pH will decrease (b) pH will increase (c) pH remains the same (d) pH = 7

90. Which of the following reaction is preferred at anode during the electrolysis of H_2SO_4 at higher concentration?
(a) $2H_2O(l) \longrightarrow O_2(g) + 4H^+(aq) + 4e^-$
(b) $2SO_4^{2-}(aq) \longrightarrow S_2O_8^{2-}(aq) + 2e^-$
(c) Both (a) and (b)
(d) None of these

91. Which of the following products are formed at cathode and anode during the electrolysis of an aqueous solution of $MgSO_4$ between inert electrodes?
(a) $O_2(g)$ and $SO_2(g)$ (b) $O_2(g)$ and $Mg(s)$ (c) $O_2(g)$ and $H_2(g)$ (d) $H_2(g)$ and $O_2(g)$

Topic 5

Batteries, Fuel Cells and Corrosion

92. The incorrect statement related to batteries is
(a) these are the assemblies having one or more than one cell connected parallely
(b) these convert chemical energy of redox reaction into electrical energy
(c) these should be reasonably light and compact
(d) their voltage should not vary during the use

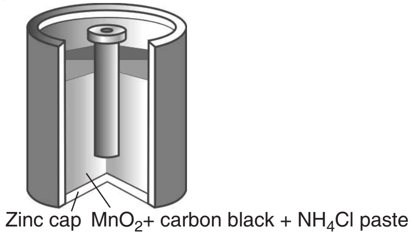
93. Select the incorrect statement.
(a) Primary batteries are non-rechargeable, so can be used only once
(b) Secondary batteries are rechargeable
(c) Primary batteries may act as electrolytic cells
(d) Secondary batteries may act as electrolytic cells

94. Which of the following statement is incorrect for dry cell?
(a) Reaction at cathode

$$MnO_2 + NH_4^+ + e^- \longrightarrow MnO(OH) + NH_3$$

(b) Manganese is reduced from +5 to +4 state
(c) Cell potential is 1.5 V
(d) Ammonia form complex with Zn^{2+} to give $[Zn(NH_3)_4]^{2+}$

95. What is the anode and cathode in the cell given in the figure?

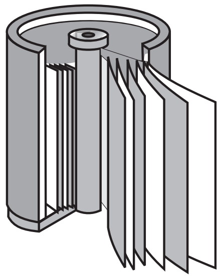


	Anode	Cathode
(a)	Zinc cup	Carbon rod
(b)	Carbon cup	Zinc rod
(c)	Zinc rod	Carbon cup
(d)	Zinc cup	Zinc rod

96. The true statement about mercury cell is
(a) it is used in high current devices
(b) here Zn − Hg acts as cathode and a paste of C + HgO as anode
(c) the cell potential remains constant due to non-involvement of such a ion, concentration of which can change during its whole life
(d) All of the above

97. The incorrect statement about lead storage battery is
(a) it is used in high current devices like invertors
(b) here lead acts as anode and lead packed with lead dioxide acts as cathode
(c) during discharging, Pb and PbO_2, both form $PbSO_4$
(d) during charging PbO_2 is formed at anode

98. Which of the following statements is/are false for the given figure?



(a) It has jelly roll arrangement
(b) It is separated by layer soaked in moist nickel solution
(c) Overall reaction

$$Cd(s) + 2Ni(OH)_3(s) \longrightarrow CdO(s) + 2Ni(OH)_2(s) + H_2O(l)$$

(d) There is a separator between positive and negative plate

99. A device that converts energy of combustion of fuels like hydrogen and methane, directly into electrical energy is known as

(a) fuel cell
(b) electrolytic cell
(c) dynamo
(d) Ni-Cd cell

100. Mark the incorrect statement(s) for fuel cells.

(a) It is used in Apollo space programme and for drinking water supply for astronauts
(b) Hydrogen and oxygen are bubbled through porous carbon electrodes
(c) Catalyst like finely divided Pt or Pd are incorporated into electrodes
(d) Anode reaction is $O_2 + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$

101. Mark the incorrect statement about corrosion.

(a) It slowly coats the surfaces of metallic objects
(b) In it metal is oxidised by giving its electrons to oxygen
(c) It damages buildings, bridges, ships and almost all the objects made of (metals especially that of iron)
(d) All of the above are correct

102. During the rusting of iron

(a) metal acts as cathode
(b) hydrogen ion acts as anode
(c) formula of rust is $Fe_3O_4.xH_2O$
(d) the overall reaction is

$$2Fe + O_2 + 4H^+ \longrightarrow 2Fe^{2+} + 2H_2O$$

103. Given the standard electrode potentials.

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

and $E^\circ_{H^+/O_2/H_2O} = 1.23$ V

Calculate the E°_{cell} of the corrosion.

(a) – 0.79 V
(b) – 1.67 V
(c) 1.67 V
(d) + 0.79 V

104. Galvanisation is

(a) zinc plating on aluminium sheet
(b) zinc plating on iron sheet
(c) iron plating on zinc sheet
(d) aluminium plating on zinc sheet

Special Format Questions

I. More Than One Correct Option

105. When an external opposite potential greater than 1.1 V is applied to the Galvanic cell

(a) electrons flow from copper electrode to zinc electrode
(b) zinc gets deposited at the zinc electrode
(c) current flows from zinc to copper electrode
(d) copper gets deposited at the copper electrode

106. Select the correct statements.

(a) When a metallic electrode is dipped into the solution of its electrolyte, a half-cell is formed
(b) A metallic wire is used to connect the two half-cells
(c) A salt bridge is required when both the electrodes are dipped in the solution of same electrolyte
(d) Salt bridge connects the electrolytes of two half-cells

107. In which of the following cells $E_{cell} = E^\circ_{cell}$?

(a) $Cu\,(s)|Cu^{2+}(0.01\text{ M})||Ag^+(0.1\text{ M})|Ag(s)$
(b) $Pt(H_2)|pH = 1||Zn^{2+}(0.01\text{ M})|Zn\,(s)$
(c) $Pt(H_2)|pH = 1||Zn^{2+}(1\text{ M})|Zn\,(s)$
(d) $Pt(H_2)|H^+ = 0.01\text{ M}|\ |Zn^{2+}(0.01\text{ M})|Zn\,(s)$

108. Given, $E^\circ_{Ni^{2+}/Ni = 0.25\text{ V}}$, $E^\circ_{Cu^{2+}/Cu = 0.34\text{ V}}$ $E^\circ_{Ag^+/Ag = 0.80\text{V}}$ and $E^\circ_{Zn^{2+}/Zn = -0.76\text{V}}$

Which of the following reactions under standard condition will not take place in the specified direction?

(a) $Ni^{2+}(aq) + Cu\,(s) \longrightarrow Ni(s) + Cu^{2+}(aq)$
(b) $Cu\,(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag\,(s)$
(c) $Cu\,(s) + 2H^+(aq) \longrightarrow Cu^{2+}(aq) + H_2(g)$
(d) $Zn\,(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + 3H_2(g)$

109. Select the correct statements.

(a) 1F is the charge carried by one mole of electrons
(b) If same quantity of electricity is passed through the solutions of 0.1 M $AgNO_3$ and 0.1 M $CuSO_4$, same weight of silver and copper will be deposited
(c) Units of electrochemical equivalent are grams per coulomb.
(d) Passage of 1F of electricity producers one gram equivalent of the substance at the electrode

110. The correct statements among the following is

(a) charge required for the oxidation or reduction of a substance depends upon the stoichiometry of the electrode reaction
(b) 2 moles of electrons are required for the reduction of 1mole of calcium ion
(c) 50,000 A current is equal to 0.518 F per second
(d) Charge required for reducing 1 mole of MnO_4^- into Mn^{2+} is 4.28×10^5 C

111. When very dilute solution of sodium hydroxide is electrolysed using platinum electrodes, then

(a) hydrogen is evolved at cathode
(b) hydrogen is evolved at anode
(c) sodium is obtained at cathode
(d) oxygen is produced at anode

112. Fuel cell involve following reactions

(a) $O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$ (at cathode)
(b) $O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$ (at anode)
(c) $2H_2(g) + 4OH^-(aq) \longrightarrow 4H_2O(l) + 4e^-$ (at anode)
(d) $2H_2(g) + 4OH^-(aq) \longrightarrow 4H_2O(l) + 4e^-$ (at cathode)

113. The reactions involved during rusting of iron are

(a) $Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$ (at anode)
(b) $O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$ (at cathode)
(c) $4Fe^{2+}(aq) + O_2(g) + 4H_2O(l) \longrightarrow 2Fe_2O_3(s) + 8H^+$
(d) $Fe_2O_3(s) + xH_2O(l) \longrightarrow Fe_2O_3 \cdot xH_2O$

114. Select the incorrect statements.

(a) Zn-Cu cell is called Daniell cell
(b) Rust in Fe_2O_3
(c) Saline water slow down rusting
(d) Pure metals undergo corrosion faster than impure metal

II. Statement Based Questions

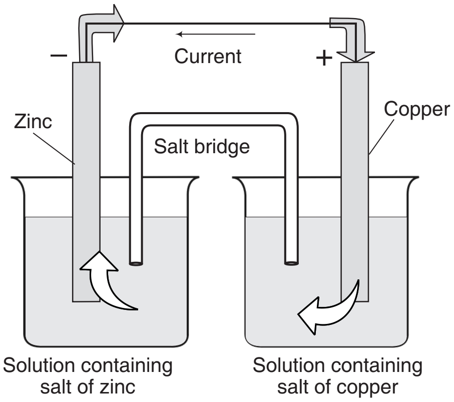
115. What happens when opposing voltage applied to the Galvanic cell reaches 1.1 V?

I. Reactions stop together.
II. No current flows through the cell.
III. Rate of reaction increases.

Which of the above statements is/are incorrect?

(a) I and II
(b) Only I
(c) Only III
(d) Only II

116. I. Metallic electrodes are dipped into electrolyte.
II. Half-cells are connected by metallic wire through voltmeter and switch.
III. There is no need of salt bridge if the electrodes are dip in the same electrolyte.



Which of the following statements is/are true for the above diagram? Choose the correct option.

(a) I and II
(b) II and III
(c) III and I
(d) I, II and III

117. I. The potential of individual half-cell can be measured.
II. Difference between the potentials of two half-cells can only be measured.
III. $Pt(s)|H_2(g)|H^+(aq)$ half-cell is called standard hydrogen electrode.

Select the correct statement(s) and choose the appropriate option.

(a) I and II
(b) II and III
(c) III and I
(d) I, II and III

118. Consider the following statements.

I. Lower the value of $E°$, more will be the reducing power.
II. If $E° = -$ve, hydrogen gas is more stable than the reduced form of the species.
III. Lithium has the lowest electrode potential.
IV. Lithium ion is the weakest oxidising agent.

The correct statements are

(a) I and III
(b) I, II and III
(c) II, III and IV
(d) I, II, III and IV

119. Electrochemical cells are used

I. in determining pH of solutions.
II. in potentiometric titrations.
III. to predict solubility product.
IV. to calculate equilibrium constant.

The correct option is

(a) I, II and III
(b) II and IV
(c) Only IV
(d) I, II, III and IV

120. Consider the following statements about the application of electrochemical series.

I. Comparison of relative oxidising and reducing powers of substances.
II. Prediction of evolution of hydrogen during reaction of a metal with acid.
III. Prediction of spontaneity of a redox reaction.
IV. Calculation of amount of metal deposited at cathode.

The true statement(s) is/are

(a) I and II only
(b) I, II and III
(c) IV only
(d) II, III and IV

121. Some organic substances are conducting polymers. The true statement(s) regarding them is/are

I. These being lighter are used to make light weight batteries.
II. Being flexible these are used to make electronic devices like transistors

(a) Only I
(b) Only II
(c) Both I and II
(d) Neither I nor II

122. I. $Q = It$
II. Charge required for oxidation or reduction depends on the stoichiometry of electrode reaction.
III. Charge on 1 electron $= 1.6021 \times 10^{-19}$ C
IV. Charge on 1 mole of electron $= 1.6021 \times 10^{-19}$ C
V. Unit of current is Coulomb (C)
VI. $1\,\text{F} = 96500\,\text{C mol}^{-1}$

Which of the following statements are incorrect?

(a) I and II
(b) II and III
(c) IV and V
(d) VI and I

123. Consider the following statements.

I. In dry cell, a moist paste of NH_4Cl and $ZnCl_2$ is present in between the electrodes.
II. Ammonia gas is produced in Leclanche cell.
III. A part of KOH and ZnO is used as electrolyte in mercury cell.

True statements are

(a) I and II
(b) I and III
(c) II and III
(d) I, II and III

124. I. Rusting of iron is an example of corrosion but turnishing of sodium metal is not.
II. Covering of *bis*-phenol protect iron from rusting.
III. Sn and Zn are used to protect iron from rusting.

True statements are

(a) I and II
(b) I and III
(c) II and III
(d) I, II and III

III. Assertion-Reason Type Questions

■ **Directions** (Q.Nos. 125-143) *In the following questions a statement of Assertion* (A) *followed by a statement of Reason* (R) *is given. Choose the correct answer out of the following choices.*

(a) Assertion and Reason both are correct and Reason is the correct explanation of Assertion.
(b) Assertion and Reason both are correct but Reason is not the correct explanation of Assertion.
(c) Assertion is correct but Reason is incorrect.
(d) Assertion is incorrect but Reason is correct.

125. Assertion (A) The electrode may be negatively charged or positively charged, with respect to solution.

Reason (R) Metal atom of electrode has tendency to go into solution as ion and leave behind the electron at the electrode.

126. Assertion (A) Platinum and gold are used as inert electrodes.

Reason (R) Pt and Au do not participate in the reaction and provide the surface for oxidation or reduction.

127. Assertion (A) Copper sulphate solution is stored in zinc pot.

Reason (R) Zinc is more reactive than copper so it displaces copper from copper sulphate solution.

128. Assertion (A) Electrical conductance through metals is called electronic conductance.

Reason (R) This is due to movement of electrons.

129. Assertion (A) Conductivity always increases with decrease in concentration for strong and weak electrolytes.

Reason (R) Number of ions per unit volume decreases on dilution.

130. Assertion (A) Conductivity of pure water is $3.5 \times 10^{-5}\ \text{S m}^{-1}$.

Reason (R) High amounts of hydrogen and hydroxyl ions are present in water.

131. Assertion (A) Conductivity of electrolytes decreases when dissolved in water.

Reason (R) They furnish their own ions.

132. Assertion (A) Anhyd. HCl is bad conductor of electricity, while *aq.* HCl is good conductor.

Reason (R) Aqueous HCl is not fully ionised but anhyd. HCl is ionised to produce H^+ and Cl^-.

133. **Assertion** (A) DC can not be used during measurement of resistance of an ionic solution.

Reason (R) DC changes the composition of solution and solution cannot be connected to the bridge like metallic wire.

134. **Assertion** (A) Pure water conduct electricity.

Reason (R) It is unionised.

135. **Assertion** (A) Impure copper is converted into pure copper by electrolysis.

Reason (R) Copper is dissolved at cathode and deposited at anode.

136. **Assertion** (A) Solid NaCl does not conduct electricity.

Reason (R) Solid NaCl has no free ions.

137. **Assertion** (A) Electrolysis of NaCl solution gives chlorine at anode instead of O_2.

Reason (R) Formation of oxygen at anode requires overvoltage.

138. **Assertion** (A) A solution of $Ni(NO_3)_2$ is electrolysed between platinum electrodes using current of 5 A for 20 minute. The weight of Ni deposited is 1.825 g.

Reason (R) The mass of substance deposited during the electrolysis of an electrolyte is inversely proportional to the quantity of electricity passed through the electrolyte.

139. **Assertion** (A) Mercury cells give a constant voltage throughout its life.

Reason (R) Electrolyte KOH is not involved in the reaction.

140. **Assertion** (A) H_2-O_2 fuel cell is not used over the ordinary cell.

Reason (R) They do not cause any pollution and its efficiency is 60–70%.

141. **Assertion** (A) Thermal plants produce electricity more efficiently than Galvanic cells.

Reason (R) In thermal plants, chemical energy is used for converting water into high pressure steam and used to run a turbine to produce electricity.

142. **Assertion** (A) Magnesium blocks are fixed to the bottom of ship.

Reason (R) Magnesium acts as sacrificial electrode.

143. **Assertion** (A) Stainless steel undergoes rusting.

Reason (R) Chromium forms an oxide layer over the stainless steel.

IV. Matching Type Questions

144. Match the following Column I with units given in Column II and choose the correct option from the codes given below.

	Column I		Column II
A.	Resistivity	1.	ohm^{-1}
B.	Conductance	2.	ohm m
C.	Conductivity	3.	$S\ cm^{-1}$

Codes

	A	B	C
(a)	2	3	1
(b)	1	2	3
(c)	2	1	3
(d)	1	3	2

145. Match the following Column I (expressions) with Column II and choose the correct option from the codes given below.

	Column I		Column II
A.	$\rho l / A$	1.	Conductivity
B.	$1 / R$	2.	Conductance
C.	RA / l	3.	Resistance
D.	$1/\rho$	4.	Resistivity

Codes

	A	B	C	D
(a)	3	4	2	1
(b)	3	2	4	1
(c)	4	3	1	2
(d)	4	1	3	2

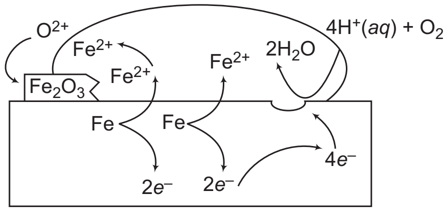
146. Match the following Column I with type of electrolyte in Column II and choose the correct option from the codes given below.

	Column I		Column II
A.	NaCl	1.	2-1 electrolyte
B.	$CaCl_2$	2.	2-2 electrolyte
C.	$MgSO_4$	3.	1-1 electrolyte

Codes

	A	B	C
(a)	1	3	2
(b)	3	2	1
(c)	3	1	2
(d)	1	2	3

147. Match Column I with Column II related to the figure given below and then select the appropriate option from the codes given below.



Column I		Column II
A. Figure represents	1.	Reduction
B. $O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$	2.	Corrosion of iron of atmosphere
C. $Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$	3.	Oxidation
D. $2Fe^{2+}(aq) + 2H_2O(l) + \frac{1}{2}O_2 \longrightarrow Fe_2O_3 + 4H^+(aq)$		

Codes

	A	B	C	D
(a)	2	1	3	2
(b)	2	3	1	2
(c)	1	2	1	3
(d)	1	2	3	1

V. Matrix Matching Type Questions

148. Match the following columns.

Column I		Column II
A. $Zn + Cu^{2+} \longrightarrow Cu + Zn^{2+}$	p.	Cathode half reaction is $Cu^{2+} + 2e^- \longrightarrow Cu$
B. $Mg + Cu^{2+} \longrightarrow Cu + Mg^{2+}$	q.	Electrons flows from metal to Cu side.
C. $2Ag^+ + Cu \longrightarrow Cu^{2+} + 2Ag$	r.	$E^\circ_{anode} < E^\circ_{cathode}$
D. $H_2 + Cu^{2+} \longrightarrow H^+ + Cu$	s.	Hydrogen electrode is used

Codes

	A	B	C	D
(a)	p,q,r	p,q,r	r	p,r,s
(b)	p,q,r	r	q	p
(c)	p,q	r,q	q	p,r
(d)	r,q	p,q	p,q,s	p,r

149. Match the following columns.

Column I		Column II
A. Fe^{3+} and I^-	p.	$E^\circ_{cell} = +ve$
B. Fe^{3+} and Br^-	q.	$\Delta G^\circ = +ve$
C. Zn and H^+	r.	$E^\circ_{cell} = -E^\circ_{electrode}$
D. Br and Fe^{2+}	s.	$\Delta G^\circ = -ve$

[Given, $E^\circ_{Fe^{3+}/Fe^{2+}} = 0.77$, $E^\circ_{\left[\frac{1}{2}I_2/I^-\right]} = +0.54$,

$E^\circ_{½ Br_2/Br^-} = 1.09$, $E^\circ_{Zn^{2+}/Zn} = -0.76$]

Codes

	A	B	C	D
(a)	p	q	p	q,r
(b)	p,s	q	p,r,s	p,s
(c)	p,r	p,s	p,r	p,r,s
(d)	q	r,s	p,s	p,q,r

150. Match the salts given in Column I with their use given in Column II.

Column I		Column II
A. Hg_2Cl_2	p.	Salt bridge
B. Agar-Agar	q.	Calomel electrode
C. 0.1 N KCl	r.	Ice cream
D. Quinhydron	s.	Redox electrode

Codes

	A	B	C	D
(a)	p	p,s	q,r	q,s
(b)	q	p,r	p	s
(c)	q	p,s	p,r	q,s
(d)	q,r	q,s	r,s	p,q,s

151. Match the physical quantities given in Column I with their units given in Column II

Column I		Column II
A. Resistance	p.	Ω
B. Resistivity	q.	volt/A^{-1}
C. Conductivity	r.	$\Omega\, m$
D. Specific conductance	s.	$\Omega^{-1} m^{-1}$

Codes

	A	B	C	D
(a)	r,s	q	q	s
(b)	p,q,r	p,s	p,r	s
(c)	p,q	r	s	s
(d)	r	r,q	r,p	r,s

152. Match the following columns.

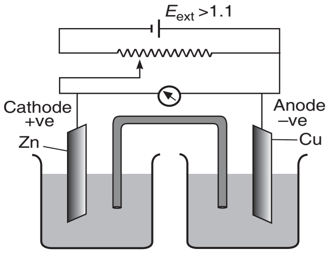
Column I (Substance)		Column II (Product after electrolysis)
A. Aqueous solution of $AgNO_3$ using Ag electrodes	p.	Oxygen is produced at anode
B. Aqueous solution of $AgNO_3$ using Pt electrodes	q.	Hydrogen is produced at cathode
C. Dilute solution of H_2SO_4 using Pt electrodes	r.	Silver is deposited at cathode
D. Aqueous solution of $CuCl_2$ using Pt electrodes	s.	Neither O_2 nor H_2 is produced

Codes

	A	B	C	D
(a)	r,q	p,q	q,s	p,s
(b)	r,q	p,q,s	q,s	q,r,s
(c)	p	q	r	s
(d)	r,s	r,p	p,q	s

VI. Passage Based Questions

■ **Directions** (Q. Nos. 153-155) *Answer the following questions on the basis of the diagram given below.*



153. Directions of current is from

(a) Cu to Zn
(b) Zn to Cu
(c) salt bridge
(d) no current will flow

154. Flow of electron is from

(a) Cu to Zn
(b) Zn to Cu
(c) salt bridge
(d) no current will flow

155. Find out the correct statement.

(a) Zn dissolves at anode and copper deposits at cathode
(b) Zn dissolves at cathode and copper deposits at anode
(c) Zn deposits at cathode and copper dissolves at anode
(d) Zn deposits at anode and copper dissolves at cathode

■ **Directions** (Q. Nos. 156-159) *Read the following paragraph and answer the following questions given below.*

The reaction which occur in the Galvanic cell is

$$MnO_4^- + 8H^+ + 5Fe^{2+} \longrightarrow Mn^{2+} + 5Fe^{3+} + 4H_2O$$

$$E^\circ(MnO_4^-, Mn^{2+}, H^+ \mid Pt) = 1.51 \text{ V } \textit{and}$$

$$E^\circ(Fe^{3+}, Fe^{2+} \mid Pt) = 0.77 \text{ V}$$

156. Calculate the standard emf of cell.

(a) 0.74 V
(b) – 0.74 V
(c) 2.28
(d) – 2.28

157. How would the emf of the cell be increased above the standard emf?

(a) By decreasing pH
(b) By increasing $[Fe^{3+}]$
(c) By decreasing $[MnO_4^-]$
(d) By increasing $[Mn^{2+}]$

158. If $[H_3O^+] = 2.0$ M, $[Fe^{3+}] = 2.0$ M, $[Fe^{2+}] = 0.5$ M, $[Mn^{2+}] = 0.5$ M, $[MnO_4^-] = 1.0$ M. Calculate the emf of the cell.

(a) – 0.747 V
(b) + 0.747 V
(c) – 0.733 V
(d) + 0.736 V

159. The other concentrations are kept at unity except reducing $[Mn^{2+}]$ to 0.50 M, the emf of the cell is increased by

(a) 0.012 V
(b) 3.55×10^{-3} V
(c) 0.059 V
(d) 0.018 V

■ **Directions** (Q. Nos. 160-161) *Read the following paragraph and answer the following questions given below.*

Resistance of a conductivity cell filled with 2 mol L^{-1} KCl solution is 100 Ω. The resistance of the same cell when filled with 0.2 mol L^{-1} KCl solution is 520 Ω. The conductivity of 1 mol L^{-1} KCl solution is 1.29 S/m.

160. Calculate the conductivity of 0.2 mol L^{-1} KCl solution.

(a) 0.248 S cm^{-1}
(b) 0.248 S m^{-1}
(c) 2.48 S cm^{-1}
(d) 2.48 S m^{-1}

161. Calculate the molar conductivity of 0.2 mol L^{-1} KCl solution.

(a) 124×10^{-2} S m^2 mol^{-1}
(b) 124×10^{-3} S m^2 mol^{-1}
(c) 124×10^{-4} S m^2 mol^{-1}
(d) 124×10^{-5} S m^2 mol^{-1}

■ **Directions** (Q. Nos. 162-164) *Read the following paragraph and answer the following questions given below.*

In a fuel cell H_2 and O_2 react to produce electricity. In the process H_2 gas is oxidised at anode and oxygen is reduced at cathode. 67.2 L of H_2 at STP react in 15 min. Entire current is used for electro deposition of Cu from Cu^{2+}.

162. Find out the average current produced.

(a) 640.33
(b) 642.33
(c) 643.33
(d) None of these

163. How many grams of Cu is deposited?

(a) 170
(b) 180
(c) 190
(d) 200

164. How many equivalents of H_2 is oxidised?

(a) 2
(b) 3
(c) 5
(d) 6

■ **Directions** (Q. Nos. 165-167) *Read the following paragraph and answer the following questions given below.*

A 4.0 M aqueous solution of NaCl is prepared and 500 mL of this solution is electrolysed. This leads to evolution of chlorine gas at one of the electrodes.

165. The total charge required for the complete electrolysis will be

(a) 96500 C
(b) 24125 C
(c) 48250 C
(d) 193000 C

166. Find out the maximum weight of amalgam formed from the solution [if cathode is Hg].

(a) 200
(b) 400
(c) 225
(d) 446

167. Calculate the number of moles of chlorine gas evolved during the electrolysis.

(a) 0.5
(b) 1.0
(c) 2.0
(d) 3.0

■ **Directions** (Q. Nos. 168-171) *Read the following paragraph and answer the following questions given below.*

NaCl aqueous solution on electrolysis gives $H_2(g)$, $Cl_2(g)$ and NaOH.

$$2Cl^-(aq) + 2H_2O \longrightarrow 2OH^-(aq) + H_2(g) + Cl_2(g)$$

In 20 L of NaCl solution (20% by weight), current of 25. A with 62% of current efficiency is passed.

168. Which of the following reaction will occur at anode?

(a) $2Cl^- \rightarrow Cl_2 + 2e^-$
(b) $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$
(c) Both (a) and (b)
(d) None of these

169. Which of the following reaction will occur at cathode?

(a) $2Cl^- \longrightarrow Cl_2 + 2e^-$
(b) $2H_2O + 2e^- \longrightarrow H_2 + 2OH^-$
(c) Both (a) and (b)
(d) None of the above

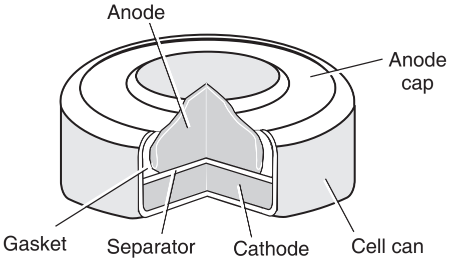
170. How long will it take to produce 1 kg of Cl_2?

(a) 48.72 hr
(b) 4.872 hr
(c) 48.72 s
(d) 48.72 min

171. What is the amount of electricity required (experimentally) in terms of Faraday?

(a) 28.17 F
(b) 2.817 F
(c) 45.44 F
(d) 454.4 F

■ **Directions** (Q. 172-176) *From the following figure, answer the following questions given below.*



172. Name the cell given in the figure.

(a) Dry cell
(b) Mercury cell
(c) Leclanche cell
(d) All of the above

173. What is anode and cathode used in the above cell?

	Anode	Cathode
(a)	Zn(Hg)	HgO
(b)	Zn	HgO
(c)	Zn	HgO and carbon
(d)	Zn(Hg) amalgam	HgO and carbon

174. What is cathodic reaction occurring in the above cell?

(a) $Zn(Hg) + 2OH^- \longrightarrow ZnO(s) + H_2O + 2e^-$
(b) $HgO + H_2O + 2e^- \longrightarrow Hg(l) + 2OH^-$
(c) $Zn + 2OH^- \longrightarrow ZnO(s) + H_2O + 2e^-$
(d) $Zn(Hg) + HgO(s) \longrightarrow ZnO(s) + Hg(l)$

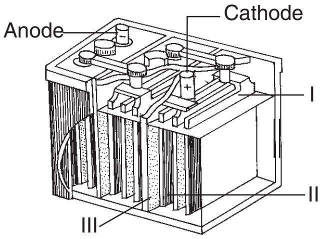
175. Cell potential of the above cell is

(a) 1.35 V
(b) 1.96 V
(c) 1.5 V
(d) 1.23 V

176. What is the reducing agent in the cell?

(a) Zinc
(b) Mercury oxide
(c) Carbon
(d) KOH

■ **Directions** (Q. Nos. 177-179) *Answer the following questions from the figure.*



177. Name the figure given above.

(a) Mercury battery
(b) Lead storage battery
(c) Nickel-cadmium battery
(d) Dry battery

178. What is I, II and electrolyte used in the above figure?

	I	II	Electrolyte
(a)	Spongy lead	PbO_2	50% H_2SO_4 solution
(b)	Spongy lead	PbO_2	38% H_2SO_4 solution
(c)	PbO_2	Spongy lead	50% H_2SO_4 solution
(d)	PbO_2	Spongy lead	38% H_2SO_4 solution

179. Which of the following statements is/are true for the above figure?

(a) Cathode reaction

$$Pb(s) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s) + 2e^-$$

(b) Anode reaction

$$PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \longrightarrow PbSO_4(s) + 2H_2O(l)$$

(c) On charging the battery $PbSO_4$ give Pb at anode and PbO_2. at cathode.
(d) All of the above

NCERT & NCERT Exemplar Questions

NCERT

180. An electrochemical cell can behave like an electrolytic cell when ....

(a) $E_{\text{cell}} = 0$ (b) $E_{\text{cell}} > E_{\text{ext}}$ (c) $E_{\text{ext}} > E_{\text{cell}}$ (d) $E_{\text{cell}} = E_{\text{ext}}$

181. Which cell will measure standard electrode potential of copper electrode?

(a) $\text{Pt}\,(s)\,|\,\text{H}_2\,(g,\,0.1\text{ bar})\,\big|\,\text{H}^+(aq.,1\text{ M})\,\big\|\,\text{Cu}^{2+}(aq,1\text{M})\,\big|\,\text{Cu}$

(b) $\text{Pt}\,(s)\,|\,\text{H}_2\,(g,\,1\text{ bar})\,\big|\,\text{H}^+(aq,1\text{ M})\,\big\|\,\text{Cu}^{2+}(aq,2\text{ M})\,\big|\,\text{Cu}$

(c) $\text{Pt}\,(s)\,|\,\text{H}_2\,(g,1\text{ bar})\,\big|\,\text{H}^+(aq,1\text{ M})\,\big\|\,\text{Cu}^{2+}(aq,1\text{ M})\,\big|\,\text{Cu}$

(d) $\text{Pt}\,(s)\,|\,\text{H}_2\,(g,0.1\text{bar})\,\big|\,\text{H}^+(aq,0.1\text{M})\,\big\|\,\text{Cu}^{2+}(aq,1\text{M})\,\big|\,\text{Cu}$

182. The difference between the electrode potentials of two electrodes when no current is drawn through the cell is called......... .

(a) cell potential (b) cell emf

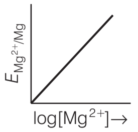
(c) potential difference (d) cell voltage

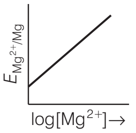
183. Which of the following statements is not correct about an inert electrode in a cell?

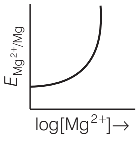
(a) It does not participate in the cell reaction
(b) It provides surface either for oxidation or for reduction reaction
(c) It provides surface for conduction of electrons
(d) It provides surface for redox reaction

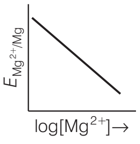
184. Electrode potential for Mg electrode varies according to the equation

$$E_{\text{Mg}^{2+}/\text{Mg}} = E^{s}_{\text{Mg}^{2+}/\text{Mg}} - \frac{0.059}{2}\log\frac{1}{[\text{Mg}^{2+}]}.$$ The graph of $E_{\text{Mg}^{2+}/\text{Mg}}$ *vs* $\log[\text{Mg}^{2+}]$ is

(a)


(b)


(c)


(d)


185. Which of the following statement is correct?

(a) E_{cell} and $\Delta_r G$ of cell reaction both are extensive properties
(b) E_{cell} and $\Delta_r G$ of cell reaction both are intensive properties
(c) E_{cell} is an intensive property while $\Delta_r G$ of cell reaction is an extensive property
(d) E_{cell} is an extensive property while $\Delta_r G$ of cell reaction is an intensive property

186. Using the data given below find out the strongest reducing agent.

$$E^{s}_{\text{Cr}_2\text{O}_7^{2-}/\text{Cr}^{3+}} = 1.33\text{ V};\; E^{s}_{\text{Cl}_2/\text{Cl}^-} = 1.36\text{V}$$
$$E^{s}_{\text{MnO}_4^-/\text{Mn}^{2+}} = 1.51\text{V};\; E^{s}_{\text{Cr}^{3+}/\text{Cr}} = -0.74\text{V}$$

(a) Cl^- (b) Cr

(c) Cr^{3+} (d) Mn^{2+}

187. Use the data given in Q. 186 and find out which of the following is the strongest oxidising agent?

(a) Cl^- (b) Mn^{2+}

(c) MnO_4^- (d) Cr^{3+}

188. Using the data given in Q.186 find out in which option the order of reducing power is correct?

(a) $\text{Cr}^{3+} < \text{Cl}^- < \text{Mn}^{2+} < \text{Cr}$
(b) $\text{Mn}^{2+} < \text{Cl}^- < \text{Cr}^{3+} < \text{Cr}$
(c) $\text{Cr}^{3+} < \text{Cl}^- < \text{Cr}_2\text{O}_7^{2-} < \text{MnO}_4^-$
(d) $\text{Mn}^{2+} < \text{Cr}^{3+} < \text{Cl}^- < \text{Cr}$

189. Use the data given in Q. 191 and find out the most stable ion in its reduced form.

(a) Cl^- (b) Cr^{3+}

(c) Cr (d) Mn^{2+}

190. Use the data of Q. 186 and find out the most stable oxidised species.

(a) Cr^{3+} (b) MnO_4^- (c) $\text{Cr}_2\text{O}_7^{2-}$ (d) Mn^{2+}

191. The cell constant of a conductivity cell ....

(a) changes with change of electrolyte
(b) changes with change of concentration of electrolyte
(c) changes with temperature of electrolyte
(d) remains constant for a cell

NCERT Exemplar

192. Which of the following statements about solution of electrolytes is not correct?

(a) Conductivity of solution depends upon size of ions
(b) Conductivity depends upon viscosity of solution
(c) Conductivity does not depend upon solvation of ions present in solution
(d) Conductivity of solution increases with temperature

193. $\Lambda^{\circ}_{\text{m}\,(\text{NH}_4\text{OH})}$ is equal to ....

(a) $\Lambda^{\circ}_{\text{m}\,(\text{NH}_4\text{OH})} + \Lambda^{\circ}_{\text{m}\,(\text{NH}_4\text{Cl})} - \Lambda^{\circ}_{\text{m}\,(\text{HCl})}$
(b) $\Lambda^{\circ}_{\text{m}\,(\text{NH}_4\text{Cl})} + \Lambda^{\circ}_{\text{m}\,(\text{NaOH})} - \Lambda^{\circ}_{\text{m}\,(\text{NaCl})}$
(c) $\Lambda^{\circ}_{\text{m}\,(\text{NH}_4\text{Cl})} + \Lambda^{\circ}_{\text{m}\,(\text{NaCl})} - \Lambda^{\circ}_{\text{m}\,(\text{NaOH})}$
(d) $\Lambda^{\circ}_{\text{m}\,(\text{NaOH})} + \Lambda^{\circ}_{\text{m}\,(\text{NaCl})} - \Lambda^{\circ}_{\text{m}\,(\text{NH}_4\text{Cl})}$

194. In the electrolysis of aqueous sodium chloride solution which of the half-cell reaction will occur at anode?

(a) $Na^+(aq) + e^- \longrightarrow Na\,(s);\ E^{\circ}_{cell} = -2.71\,V$

(b) $2H_2O\,(l) \longrightarrow O_2\,(g) + 4H^+\,(aq) + 4e^-;\ E^{\circ}_{cell} - 1.23\,V$

(c) $H^+\,(aq) + e^- \longrightarrow \frac{1}{2}H_2\,(g);\ E^{\circ}_{cell} = 0.00\,V$

(d) $Cl^-\,(aq) \longrightarrow \frac{1}{2}Cl_2\,(g) + e^-;\ E^{\circ}_{cell} = 1.36\,V$

195. The quantity of charge required to obtain one mole of aluminium from Al_2O_3 is

(a) 1 F
(b) 6 F
(c) 3 F
(d) 2 F

196. While charging the lead storage battery.

(a) $PbSO_4$ anode is reduced to Pb
(b) $PbSO_4$ cathode is reduced to Pb
(c) $PbSO_4$ cathode is oxidised to Pb
(d) $PbSO_4$ anode is oxidised to PbO_2

197. For the given cell, $Mg\,|\,Mg^{2+}\,||\,Cu^{2+}\,|\,Cu$

(a) Mg is cathode
(b) Cu is cathode
(c) The cell reaction is $Mg + Cu^{2+} \longrightarrow Mg^{2+} + Cu$
(d) Cu is the oxidising agent

198. The positive value of the standard electrode potential of Cu^{2+}/Cu indicates that

(a) this redox couple is a stronger reducing agent than the H^+/H_2 couple
(b) this redox couple is a stronger oxidising agent than H^+/H_2 couple
(c) Cu can displace H_2 from acid
(d) Cu cannot displace H_2 from acid

199. E°_{cell} for some half-cell reactions are given below.

(i) $H^+(aq) + e^- \longrightarrow \frac{1}{2}H_2(g);\ E^{\circ}_{cell} = 0.00\ V$

(ii) $2H_2O(l) \longrightarrow O_2(g) + 4H^+(aq) + 4e^-;\ E^{\circ}_{cell} = 1.23V$

(iii) $2SO_4^{2-}\,(aq) \longrightarrow S_2O_8^{2-}(aq) + 2e^-;\quad E^{\circ}_{cell} = 1.96V$

On the basis of these mark the correct answer.

(a) In dilute sulphuric acid solution, hydrogen will be reduced at cathode
(b) In concentrated sulphuric acid solution, water will be oxidised at anode
(c) In dilute sulphuric acid solution, water will be oxidised at anode
(d) In dilute sulphuric acid solution, SO_4^{2-} ion will be oxidised to tetrathionate ion at anode

200. $E^{\circ}_{cell} = 1.1\ V$ for Daniell cell. Which of the following expressions are correct description of state of equilibrium in this cell?

(a) $1.1 = K_C$
(b) $\frac{2.303RT}{2F}\log K_C = 1.1$
(c) $\log K_C = \frac{2.2}{0.059}$
(d) $\log K_C = 1.1$

201. Conductivity κ, is equal to......... .

(a) $\frac{1}{R}\frac{l}{A}$
(b) $\frac{G^*}{R}$
(c) Λ_m
(d) $\frac{l}{A}$

202. Conductivity of an electrolytic solution depends on

(a) nature of electrolyte
(b) concentration of electrolyte
(c) power of AC source
(d) distance between the electrodes

203. Molar conductivity of ionic solution depends on ....

(a) temperature
(b) distance between electrodes
(c) concentration of electrolytes in solution
(d) surface area of electrodes

204. What will happen during the electrolysis of aqueous solution of $CuSO_4$ by using platinum electrodes?

(a) Copper will deposit at cathode
(b) Copper will deposit at anode
(c) Oxygen will be released at anode
(d) Copper will dissolve at anode

205. What will happen during the electrolysis of aqueous solution of $CuSO_4$ in the presence of copper electrodes?

(a) Copper will deposit at cathode
(b) Copper will dissolve at anode
(c) Oxygen will be released at anode
(d) Copper will deposit at anode

■ **Directions** (Q.Nos. 206-215) *In the following questions a statement of Assertion* (A) *followed by a statement of reason* (R) *is given. Choose the correct answer out of the following choices.*

(a) Both Assertion and Reason are correct and the Reason is the correct explanation of Assertion.
(b) Both Assertion and Reason are correct and Reason is not the correct explanation of Assertion.
(c) Assertion is correct but the Reason is incorrect.
(d) Both Assertion and Reason are incorrect.
(e) Assertion is incorrect but Reason is correct.

206. **Assertion** (A) E_{cell} should have a positive value for the cell to function.

Reason (R) $E_{cathode} < E_{anode}$

207. Assertion (A) Mercury cell does not give steady potential.

Reason (R) In the cell reaction, ions are not involved in solution.

208. Assertion (A) Current stops flowing when $E_{cell} = 0$.

Reason (R) Equilibrium of the cell reaction is attained.

209. Assertion (A) Cu is less reactive than hydrogen.

Reason (R) $E^\circ_{Cu^{2+}/Cu}$ is negative.

210. Assertion (A) $E_{Ag^+/Ag}$ increases with increase in concentration of Ag^+ ions.

Reason (R) $E_{Ag^+/Ag}$ has a positive value.

211. Assertion (A) Copper sulphate can be stored in zinc vessel.

Reason (R) Zinc is less reactive than copper.

212. Assertion (A) Conductivity of all electrolytes decreases on dilution.

Reason (R) On dilution number of ions per unit volume decreases.

213. Assertion (A) Λ_m for weak electrolytes shows a sharp increase, when the electrolytic solution is diluted.

Reason (R) For weak electrolytes degree of dissociation increases with dilution of solution.

214. Assertion (A) For measuring resistance of an ionic solution, an AC source is used.

Reason (R) Concentration of ionic solution will change if DC source is used.

215. Assertion (A) Electrolysis of NaCl solution gives chlorine at anode instead of O_2.

Reason (R) Formation of oxygen at anode requires over voltage.

216. Match the items of Column I and Column II on the basis of data given below.

$$E^\circ_{F_2/F^-} = 2.87\,V,\ E^\circ_{Li^+/Li} = -3.5\,V,$$
$$E^\circ_{Au^{3+}/Au} = 1.4\,V,\ E^\circ_{Br_2/Br^-} = 1.09\,V$$

Column I	Column II
A. F_2	1. Metal is the strongest reducing agent.
B. Li	2. Metal ion which is the weakest oxidising agent.
C. Au^{3+}	3. Non-metal which is the best oxidising agent.
D. Br^-	4. Unreactive metal.
E. Au	5. Anion that can be oxidised by Au^{3+}.
F. Li^+	6. Anion which is the weakest reducing agent.
G. F^-	7. Metal ion which is an oxidising agent.

	A	B	C	D	E	F	G
(a)	3	1	5	7	2	4	6
(b)	3	1	7	5	4	2	6
(c)	3	1	5	7	4	6	2
(d)	1	3	5	7	4	2	6

217. Match the terms given in Column I with the items given in Column II.

Column I	Column II
A. Λ_m	1. Intensive property
B. E°_{cell}	2. Depends on number of ions/volume
C. κ	3. Extensive property
D. $\Delta_r G_{cell}$	4. Increases with dilution

	A	B	C	D
(a)	1	4	3	2
(b)	1	4	2	3
(c)	4	1	3	2
(d)	4	1	2	3

218. Match the terms given in Column I with the units given in Column II.

Column I	Column II
A. Λ_m	1. $S\ cm^{-1}$
B. E_{cell}	2. m^{-1}
C. κ	3. $S\ cm^2\ mol^{-1}$
D. G^*	4. V

	A	B	C	D
(a)	3	4	1	2
(b)	4	3	2	1
(c)	3	4	2	1
(d)	4	3	1	2

219. Match the items of Column I and Column II.

Column I	Column II
A. κ	1. $I \times t$
B. Λ_m	2. $\Lambda_m / \Lambda^\circ_m$
C. α	3. $\frac{\kappa}{C}$
D. Q	4. $\frac{G^*}{R}$

	A	B	C	D
(a)	3	4	1	2
(b)	4	3	2	1
(c)	3	4	2	1
(d)	4	3	1	2

220. Match the items of Column I and Column II.

Column I	Column II
A. Lead storage battery	1. Maximum efficiency
B. Mercury cell	2. Prevented by galvanisation
C. Fuel cell	3. Gives steady potential
D. Rusting	4. Pb is anode, PbO_2 is cathode

	A	B	C	D		A	B	C	D
(a)	3	4	1	2	(b)	4	3	2	1
(c)	3	4	2	1	(d)	4	3	1	2

221. Match the items of Column I and Column II.

Column I	Column II
A. Leclanche cell	1. Cell reaction, $2H_2 + O_2 \longrightarrow 2H_2O$
B. Ni-Cd cell	2. Does not involve any ion in solution and is used in hearing aids.
C. Fuel cell	3. Rechargeable
D. Mercury cell	4. Reaction at anode, $Zn \longrightarrow Zn^{2+} + 2e^-$
	5. Converts energy of combustion into electrical energy

	A	B	C	D
(a)	4	3	5	2
(b)	1	4	3	2
(c)	3	4	5	1
(d)	4	3	5	1

222. Two Faraday of electricity is passed through a solution of $CuSO_4$. The mass of copper deposited at the cathode is (atomic mass of Cu = 63.5 u)

(a) 0 g (b) 63.5 g (c) 2 g (d) 127 g

223. The equivalent conductance of NaCl at concentration C and at infinite dilution are λ_C and λ_∞, respectively. The correct relationship between λ_C and λ_∞ is given as (where, the constant B is positive)

(a) $\lambda_C = \lambda_\infty + (B)C$ (b) $\lambda_C = \lambda_\infty - (B)C$

(c) $\lambda_C = \lambda_\infty - (B)\sqrt{C}$ (d) $\lambda_C = \lambda_\infty + (B)\sqrt{C}$

224. Given below are the half-cell reactions,

$Mn^{2+} + 2e^- \longrightarrow Mn; E^\circ = -1.18$ eV

$2(Mn^{3+} + e^- \longrightarrow Mn^{2+});\ E^\circ = +1\ .51$ eV

The E° for $3\,Mn^{2+} \longrightarrow Mn + 2Mn^{3+}$ will be

(a) −2.69 V, the reaction will not occur

(b) −2.69 V, the reaction will occur

(c) −0.33 V, the reaction will not occur

(d) −0.33 V, the reaction will occur

225. Given, $E^\circ_{Cr^{3+}/Cr} = 0.74$ V; $E^\circ_{MnO_4^-/Mn^{2+}} = 1.51$ V;

$E_{Cr_2O_7^{2-}/Cr^{3+}} = 1.33$ V; $E^\circ_{Cl/Cl^-} = 1.36$ V;

Based on the data given above strongest oxidising agent will be

(a) Cl (b) Cr^{3+}

(c) Mn^{2+} (d) MnO_4^-

226. The electrochemical cell shown below is a concentration cell.

M/M^{2+} (saturated solution of a sparingly soluble salt, MX_2) || M^{2+} (0.00001 mol dm^{-3} / M).

The emf of the cell depends on the difference in concentration of M^{2+} ions at the two electrodes. The emf of the cell at 298 is 0.059 V.

The value of ΔG (kJ mol^{-1}) for the given cell is (take 1F = 96500C mol^{-1})

(a) −5.7 (b) 5.7

(c) 11.4 (d) −11.4

227. The standard reduction potential for Zn^{2+}/Zn, Ni^{2+}/Ni and Fe^{2+}/Fe are −0.76, −0.23 and −0.44 V respectively. The reaction $X + Y^{2+} \longrightarrow X^{2+} + Y$ will be spontaneous if

(a) X = Ni, Y = Fe (b) X = Ni, Y = Zn

(c) X = Fe, Y = Zn (d) X = Zn, Y = Ni

Answers

1.	(c)	**2.**	(d)	**3.**	(a)	**4.**	(b)	**5.**	(a)	**6.**	(b)	**7.**	(a)	**8.**	(c)	**9.**	(c)	**10.**	(c)	**11.**	(d)	**12.**	(c)	**13.**	(b)	**14.**	(d)	**15.**	(d)
16.	(a)	**17.**	(d)	**18.**	(c)	**19.**	(c)	**20.**	(c)	**21.**	(c)	**22.**	(b)	**23.**	(d)	**24.**	(b)	**25.**	(b)	**26.**	(b)	**27.**	(c)	**28.**	(b)	**29.**	(a)	**30.**	(d)
31.	(c)	**32.**	(b)	**33.**	(c)	**34.**	(c)	**35.**	(c)	**36.**	(d)	**37.**	(c)	**38.**	(d)	**39.**	(a)	**40.**	(b)	**41.**	(c)	**42.**	(b)	**43.**	(c)	**44.**	(d)	**45.**	(a)
46.	(b)	**47.**	(c)	**48.**	(d)	**49.**	(d)	**50.**	(d)	**51.**	(b)	**52.**	(d)	**53.**	(b)	**54.**	(b)	**55.**	(a)	**56.**	(b)	**57.**	(c)	**58.**	(b)	**59.**	(d)	**60.**	(b)
61.	(a)	**62.**	(a)	**63.**	(b)	**64.**	(a)	**65.**	(b)	**66.**	(d)	**67.**	(b)	**68.**	(c)	**69.**	(d)	**70.**	(b)	**71.**	(b)	**72.**	(d)	**73.**	(d)	**74.**	(c)	**75.**	(d)
76.	(c)	**77.**	(d)	**78.**	(c)	**79.**	(b)	**80.**	(c)	**81.**	(c)	**82.**	(b)	**83.**	(c)	**84.**	(c)	**85.**	(d)	**86.**	(b)	**87.**	(c)	**88.**	(a)	**89.**	(b)	**90.**	(b)
91.	(d)	**92.**	(a)	**93.**	(c)	**94.**	(b)	**95.**	(a)	**96.**	(c)	**97.**	(d)	**98.**	(b)	**99.**	(a)	**100.**	(d)	**101.**	(d)	**102.**	(d)	**103.**	(c)	**104.**	(b)	**105.**	(abc)
106.	(abd)	**107.**	(ab)	**108.**	(ab)	**109.**	(acd)	**110.**	(abc)	**111.**	(ad)	**112.**	(ac)	**113.**	(abcd)	**114.**	(bcd)	**115.**	(c)	**116.**	(d)	**117.**	(b)	**118.**	(d)	**119.**	(d)	**120.**	(b)
121.	(c)	**122.**	(c)	**123.**	(d)	**124.**	(c)	**125.**	(b)	**126.**	(a)	**127.**	(d)	**128.**	(a)	**129.**	(d)	**130.**	(c)	**131.**	(d)	**132.**	(c)	**133.**	(a)	**134.**	(d)	**135.**	(c)
136.	(a)	**137.**	(a)	**138.**	(c)	**139.**	(c)	**140.**	(d)	**141.**	(d)	**142.**	(a)	**143.**	(d)	**144.**	(c)	**145.**	(b)	**146.**	(c)	**147.**	(a)	**148.**	(a)	**149.**	(b)	**150.**	(b)
151.	(c)	**152.**	(d)	**153.**	(d)	**154.**	(a)	**155.**	(c)	**156.**	(a)	**157.**	(a)	**158.**	(d)	**159.**	(b)	**160.**	(b)	**161.**	(d)	**162.**	(c)	**163.**	(c)	**164.**	(d)	**165.**	(d)
166.	(d)	**167.**	(b)	**168.**	(a)	**169.**	(b)	**170.**	(a)	**171.**	(c)	**172.**	(b)	**173.**	(d)	**174.**	(b)	**175.**	(a)	**176.**	(a)	**177.**	(b)	**178.**	(b)	**179.**	(c)	**180.**	(c)
181.	(c)	**182.**	(b)	**183.**	(d)	**184.**	(b)	**185.**	(c)	**186.**	(b)	**187.**	(c)	**188.**	(b)	**189.**	(d)	**190.**	(a)	**191.**	(d)	**192.**	(c)	**193.**	(d)	**194.**	(d)	**195.**	(c)
196.	(a)	**197.**	(bc)	**198.**	(bd)	**199.**	(ac)	**200.**	(bc)	**201.**	(ab)	**202.**	(ab)	**203.**	(ac)	**204.**	(ac)	**205.**	(ab)	**206.**	(c)	**207.**	(e)	**208.**	(a)	**209.**	(c)	**210.**	(b)
211.	(d)	**212.**	(a)	**213.**	(a)	**214.**	(a)	**215.**	(a)	**216.**	(b)	**217.**	(d)	**218.**	(d)	**219.**	(b)	**220.**	(d)	**221.**	(a)	**222.**	(b)	**223.**	(c)	**224.**	(a)	**225.**	(d)

Hints & Explanations

1. (*c*) The reaction carried out electrochemically can be energy efficient and less polluting, i.e. ecofriendly. All other given statements are true.

2. (*d*) Galvanic cells are also known as voltaic cells, e.g. Daniell cell. The reaction occuring in this cell is

$$\text{Zn}(s) + \text{Cu}^{2+}(aq) \longrightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s)$$

4. (*b*) In a Galvanic cell, electrons flow from negative electrode to positive electrode. The direction of flow of current is opposite to that of flow of electron, i.e. from positive electrode to negative electrode.

5. (*a*) When $E_{\text{ext}} = 1.1\text{V}$ in a Daniell cell and chemical reactions stops and no current flows through the cell.

6. (*b*) When $E_{\text{ext}} < 1.1\text{V}$

(i) Electrons flow from zinc electrode to copper electrode, hence, current flows from copper to zinc electrode.

(ii) Zinc dissolves at anode and copper deposites at cathode.

7. (*a*) When $E_{\text{ext}} > E$ generated, the cell behaves like an electrolytic cell. In this cell a non-spontaneous reaction is carried out at the expense of electrical energy.

8. (*c*) In Daniell cell potential is 1.1 V. Since $E_{\text{ext}}(0.9) < 1.1$, so in this case, current flows from copper to zinc electrode.

10. (*c*) The reduction half-reaction and oxidation half-reaction together form redox couple.

12. (*c*) In a Galvanic cell, anode is always the site of oxidation (i.e. its potential is negative with respect to solution) and cathode is always the site of reduction (i.e. its potential is positive with respect to solution).

13. (*b*) Unit of cell potential is volt.

15. (*d*) All the statements are true.

17. (*d*) The cell representation of the given reaction is $\text{Cu}(s)|\text{Cu}^{2+}(aq)||\text{Ag}^{+}(aq)|\text{Ag}(s)$

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

Anode reaction is $\text{Cu}(s) \longrightarrow \text{Cu}^{2+}(aq) + 2e^{-}$ and cathode reaction is $2\text{Ag}^{+}(aq) + 2e^{-} \longrightarrow 2\text{Ag}(s)$.

18. (*c*) For the given reaction,

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{Cu}^{2+}/\text{Cu}} - E^{\circ}_{\text{Zn}^{2+}/\text{Zn}}$$
$$= 0.34 - (-0.76) = +1.10\ \text{V}$$

19. (*c*) $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{Cd}^{2+}/\text{Cd}} - E^{\circ}_{\text{Cr}^{2+}/\text{Cr}}$

$$= -0.40 - (-0.74) = +0.34$$

20. (*c*) Reduction hydrogen half-cell is

$\text{H}^{+}|(z\text{M})|\text{Pt}(\text{H}_2)$; Pressure, p_{H_2}

Half-cell reaction is

$$2\text{H}^{+}(aq) + 2e^{-} \longrightarrow \text{H}_2(g)$$

Reaction quotient = $Q = p_{\text{H}_2}/[\text{H}^{+}]^2$ and $n = 2$

$$E^\circ_{\text{cell}} = E^\circ_{\text{red}} - \frac{0.0591}{n}\log Q = 0 - \frac{0.0591}{2}\log Q$$

p_{H_2}	$[\mathrm{H}^+]$	Q	E_{red}
1atm	2.0M	0.25	+ve
1atm	1.0M	1.0	0
2atm	1.0M	2.0	−ve
2atm	2.0M	0.50	+ve

$E^\circ_{\text{red}} = 0.00$ V for standard hydrogen electrode

If $Q > 1$, then $E^\circ_{\text{red}} = -ve$

Thus, correct answer is (c).

21. (*c*) In standard hydrogen electrode, the electrode is dipped in acidic solution (HCl solution) and pure hydrogen gas is bubbled through it.

22. (*b*) When anode is the hydrogen electrode,

$$E^\circ_{\mathrm{H}_2} = 0$$

$$E^\circ = E^\circ_R - E^\circ_L = E^\circ_R - 0 = E^\circ_R$$

24. (*b*) $E^\circ_{\text{cell}} = E_{\mathrm{Zn}^{2+}/\mathrm{Zn}} - E^\circ_{\mathrm{H}_2/2\mathrm{H}^+} = -(0.76 - 0) = -0.76$ V

Thus, hydrogen ions can only oxidise zinc and get reduced by them.

25. (*b*) In the reaction of Cu with $\mathrm{HNO_3}$, it is nitrate ion not the hydrogen ion, which oxidises Cu.

27. (*c*) If standard electrode potential is negative then hydrogen gas is more stable than reduced form of the species.
The standard electrode potential for fluorine is highest which indicates that fluorine gas ($\mathrm{F_2}$) has maximum tendency to get reduced to fluoride ions ($\mathrm{F^-}$) and fluorine gas is the strongest oxidising agent but fluoride ion is the weakest reducing agent.

Lithium has lowest electrode potential indicating that Li ion is weakest oxidising agent while Li metal is the most powerful reducing agent.

28. (*b*) Oxidising tendency $\propto \dfrac{1}{\text{electrode potential}}$

$$XA \longrightarrow \text{No reaction}$$
$$XB \longrightarrow A, C$$
$$XC \longrightarrow A$$

Order of electrode potential is $XB < XC < XA$.

Order of oxidation of anion $B^- > C^- > A^-$.

29. (*a*) More negative the E°, higher is the reducing power. So $\mathrm{Zn}(s)$ is the strongest reducing agent among the given.

30. (*d*) $Z > X > Y$; higher the reduction potential, lesser the reduction power.

31. (*c*) $2\mathrm{I}^- \longrightarrow \mathrm{I}_2 + 2e^-$ (oxidation)

$$\mathrm{Br}_2 + 2e^- \longrightarrow 2\mathrm{Br}^- \text{(reduction)}$$

$$2\mathrm{I}^- + \mathrm{Br}_2 \longrightarrow \mathrm{I}_2 + 2\mathrm{Br}^- \text{(net reaction)}$$

$$E^\circ_{\text{cell}} = 1.09 - 0.54 = +0.55 \text{ V}$$

Positive E°_{cell} means reaction is feasible.

32. (*b*) Electrode potential of an electrode is measured with respect to standard hydrogen electrode by the expression

$$E_{M^{n+}/M} = E^\circ_{M^{n+}/M} - \frac{RT}{nF}\ln\frac{[M]}{[M^{n+}]}$$

but concentration of solid M is taken as unity

So, $$E_{M^{n+}/M} = E^\circ_{M^{n+}/M} - \frac{RT}{nF}\ln\frac{1}{[M^{n+}]}$$

33. (*c*) In Daniell cell, the electrode potential for concentration of $\mathrm{Cu^{2+}}$ and $\mathrm{Zn^{2+}}$ ions can be written as

For cathode

$$E_{\mathrm{Cu}^{2+}/\mathrm{Cu}} = E^\circ_{\mathrm{Cu}^{2+}/\mathrm{Cu}} - \frac{RT}{2F}\ln\frac{1}{[\mathrm{Cu}^{2+}(aq)]}$$

For anode

$$E_{\mathrm{Zn}^{2+}/\mathrm{Zn}} = E^\circ_{\mathrm{Zn}^{2+}/\mathrm{Zn}} - \frac{RT}{2F}\ln[\mathrm{Zn}^{2+}(aq)]$$

$$E_{\text{cell}} = E_{\mathrm{Cu}^{2+}/\mathrm{Cu}} - E_{\mathrm{Zn}^{2+}/\mathrm{Zn}}$$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{2F}\ln\frac{[\mathrm{Zn}^{2+}]}{[\mathrm{Cu}^{2+}]}$$

34. (*c*) E_{cell} depends on the concentration of both $\mathrm{Cu^{2+}}$ and $\mathrm{Zn^{2+}}$ ions. It increases with increase in the concentration of $\mathrm{Cu^{2+}}$ ions and decrease in the concentration of $\mathrm{Zn^{2+}}$ ions.

35. (*c*) $\Delta G_3 = \Delta G_1 + \Delta G_2$

$$-2FE^\circ = -1\,\mathrm{F} \times 0.15 + (-1\mathrm{F} \times 0.50)$$
$$-2FE^\circ = -0.15\mathrm{F} - 0.50\mathrm{F}$$
$$= -F(0.15 + 0.50)$$
$$E^\circ = -\frac{0.65}{2} = 0.325 \text{ V}$$

36. (*d*) $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$

i.e. it does not depend upon the concentration of ions , so remain the same.

37. (*c*) $E^\circ_{\text{cell}} = E^\circ_{\text{cathode(RP)}} - E^\circ_{\text{anode(RP)}}$

$$= 0.15 - (-0.74) = +0.89 \text{ V}$$

38. (*d*) The half reactions are

$$[\mathrm{Fe}(s) \longrightarrow \mathrm{Fe}^{2+}(aq) + 2e^-] \times 2$$
$$\mathrm{O_2}(g) + 4\mathrm{H}^+ + 4e^- \longrightarrow 2\mathrm{H_2O}$$
$$2\mathrm{Fe}(s) + \mathrm{O_2}(g) + 4\mathrm{H}^+ \longrightarrow 2\mathrm{Fe}^{2+}(aq) + 2\mathrm{H_2O}(l)$$

$$E = 1.67 - \frac{0.059}{4}\log\frac{(10^{-3})^2}{(10^{-3})^4\,(0.1)} = 1.57 \text{ V}$$

39. (*a*) Anode is always the site of oxidation and cathode is the site of reduction, Thus, anode half-cell is

$$\mathrm{Zn}^{2+}(aq) + 2e^- \longrightarrow \mathrm{Zn}(s);\ E^\circ = -0.76 \text{ V}$$

Cathode half-cell is

$$\mathrm{Ag_2O}(s) + \mathrm{H_2O}(l) + 2e^- \longrightarrow 2\mathrm{Ag}(s) + 2\mathrm{OH}^-(aq);$$

$$E^\circ = 0.34 \text{ V}$$

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$$
$$= 0.34 - (-0.76)$$
$$= 1.1 \text{ V}$$

40. (*b*) For hydrogen electrode, oxidation half-reaction is

$$\underset{(1\text{ atm})}{H_2} \longrightarrow \underset{(\text{at pH } 10)}{2H^+} + 2e^-$$

If $\text{pH} = 10, H^+ = 1\times 10^{-\text{pH}} = 1\times 10^{-10}$

From Nernst equation, $E = E° - \frac{0.0591}{2}\log\frac{[H^+]^2}{p_{H_2}}$

For hydrogen electrode, $E° = 0$

$$E = -\frac{0.0591}{2}\log\frac{(10^{-10})^2}{1} = -\frac{0.0591\times 2}{2}\log(10)^{-10}$$

$$= 0.0591\times 10\times \log 10 = 0.59\text{ V}$$

41. (*c*) $\text{Zn}(s) + \text{Cu}^{2+}(aq) \longrightarrow \text{Zn}^{2+}(aq) + \text{Cu}(s)$

$$0 = E_{\text{cell}} = E°_{\text{cell}} - \frac{2.303RT}{2F}\log\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

or $E°_{\text{cell}} = \frac{2.303RT}{2F}\log\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$

But at equilibrium

$$\frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]} = K_C$$

$$E°_{\text{cell}} = \frac{0.059\text{ V}}{2}\log K_C = 1.1\text{ V}$$

$$\log K_C = \frac{1.1\text{ V}\times 2}{0.059\text{ V}} = 37.288$$

$$K_C = 2\times 10^{37}\text{ at 298 K}$$

43. (*c*) Right hand cell reaction,

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

Left hand cell reaction,

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

Cell reactions, $2\text{Cu}^+ \longrightarrow \text{Cu} + \text{Cu}^{2+}$

Cell potential $E°_{\text{cell}} = E°_{\text{Cu}^+/\text{Cu}} - E°_{\text{Cu}^{2+}/\text{Cu}^+}$

$$= 0.60 - 0.41 = 0.19\text{ V}$$

$$-nFE° = -RT\ln K_{\text{eq}}$$

$$\log K_{\text{eq}} = \frac{nE°}{(2.303RT/F)} = \frac{2\times 0.19\text{ V}}{0.059\text{ V}} = 6.44$$

$$K_{\text{eq}} = 10^{6.44} = 2.76\times 10^6$$

44. (*d*) $\Delta_r G° = -nFE°_{\text{cell}}$

ΔG is directly proportional to E_{cell} but in equality Faraday constant must be included.

All other statements are true.

45. (*a*) The standard free energy ($\Delta G°$) is related to standard electrode potential ($E°$) and equilibrium constant as

$$\Delta G° = -nFE°$$

and $\Delta G° = -RT\log_e K_{\text{eq}}$

Thus, if $E°_{\text{cell}}$ is −ve, then $\Delta G°$ is +ve, i.e., $\Delta G° > 0$.

Hence, K_{eq} would be negative and less than 1, $K_{\text{eq}} < 1$

46. (*b*) $\Delta_r G = -nFE_{\text{cell}}$

$\because$ $n = 2$, so $\Delta_r G = -2FE_{\text{cell}}$

47. (*c*)

$$4\text{Al}^{3+} \longrightarrow 4\text{Al} + 12e^-$$

$$12e^- + 6\text{O}^{2-} \longrightarrow 3\text{ O}_2$$

or

$$\frac{4}{3}\text{Al}^{3+} \longrightarrow \frac{4}{3}\text{Al} + 4e^-$$

$$4e^- + 2\text{O}^{2-} \longrightarrow \text{O}_2;$$

$$\Delta G° = -nFE°$$

$$\Delta G = +\ 960\text{ kJ mol}^{-1} = 960\times 1000\text{ J mol}^{-1} +$$

$$n = 4F = 4\times 96500\text{ C}$$

$\therefore$ $960\times 1000 = -4\times 96500\times E°$

$$E° = -\frac{960000}{4\times 96500} = -2.48\text{ V}$$

Potential difference ≈ 2.5 V

48. (*d*) Reciprocal or inverse of resistance is called conductance, while specific conductance or conductivity is the reciprocal of resistivity. All other statements are true.

51. (*b*) It remains unchanged because only electron moves but no ion left its place.

52. (*d*) The conductance of electricity by ions present in the solutions is called electrolytic or ionic conductance. It depends on the nature of electrolyte added, size of ions produced and their solvation, nature of solvent and its viscosity, concentration of electrolyte and temperature (it increases with increase of temperature).

53. (*b*) Electrolytic conductance increases with increase in temperature due to increase in the speed of ions at higher temperature.

Electrical conductance, however, decreases with increase in temperature because of the disturbance of kernel at high temperature.

55. (*a*) Cell constant $= \frac{l}{A}$. Its SI unit is m^{-1}

It cannot be measured directly from l/A.

It is measured from resistance and conductivity.

56. (*b*) Conductivity of KCl solution is known accurately at various concentrations and different temperatures, so it is generally used in conductivity cell to measure cell constant.

57. (*c*) Wheatstone bridge consists of two resistance, R_3 and R_4 and a variable resistance R_1 and conductivity cell having unknown resistance R_2. It is fed by an oscillator O (a source of AC power in the range of 550 to 5000 cycles per second). P is suitable detector and bridge is balanced, when no current passes through the detector.

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

58. (*b*) Conductivity, $\kappa = \frac{\text{Cell constant}}{\text{Resistance}} = \frac{0.616}{55} = 1.12\times 10^{-2}\Omega^{-1}\text{ cm}^{-1}$

60. (*b*) Specific conductance = conductance × cell constant

$$1.3\text{Sm}^{-1} = \frac{1}{50}\text{S}\times\text{cell constant}$$

$\therefore$ Cell constant $= 1.3\times 50\text{ m}^{-1} = 65\text{ m}^{-1} = (65/100)\text{ cm}^{-1}$

$$\text{Molar conductivity} = \frac{1000\times\text{conductance }\times\text{cell constant}}{\text{molarity}}$$

$$= \frac{1000}{0.4}\times\frac{1}{260}\times\frac{65}{100} = 6.25\text{ Scm}^2\text{mol}^{-1}$$

$$= 6.25\times 10^{-4}\text{Sm}^2\text{ mol}^{-1}$$

61. (*a*) Conductivity always decreases with decrease in concentration because the number of ions (current carrying species) per unit volume decreases on dilution. Molar conductivity, however, increases on dilution because total volume of solution containing one mole of electrolyte also increases.

65. (*b*) When concentration approaches zero, the molar conductivity is known as limiting molar conductivity. So here limiting molar conductivity is near about 150.

66. (*d*) For strong electrolytes $\Lambda_m = \Lambda_m^\circ - AC^{1/2}$

If we plot Λ_m against $C^{1/2}$, we obtain straight line with intercept equal to Λ_m° and slope is equal to $-A$. The constant A depends on the type of electrolyte, i.e. charges on cation and anion produced on dissociation of electrolyte.

67. (*b*) Y is strong electrolyte because on dilution number of ions remains almost the same in case of strong electrolyte. Only interionic attraction decreases and increase in Λ_m is small.

70. (*b*) $\Lambda^\circ_{NaCl} = 50.1 + 76.3 = 126.4 \text{ S cm}^2 \text{ mol}^{-1}$

$$\Lambda^\circ_{KBr} = 73.5 + 78.1 = 151.6 \text{ S cm}^2 \text{ mol}^{-1}$$

71. (*b*) $\Lambda^\circ_{(CH_3COO)_2Ca} = 2\times\Lambda_{CH_3COO^-} + \Lambda_{Ca^{2+}}$

$$= 2\times 40.9 + 119.0 = 81.8 + 119.0$$
$$= 200.8$$

72. (*d*) $CH_3COONa + HCl \longrightarrow NaCl + CH_3COOH$

$$\overset{\circ}{\Lambda}_{m(CH_3COOH)} = \overset{\circ}{\Lambda}_{m(CH_3COONa)} + \overset{\circ}{\Lambda}_{m(HCl)} - \overset{\circ}{\Lambda}_{m(NaCl)}$$
$$\Lambda^\circ_{m(CH_3COOH)} = (91.0 + 425.9 - 126.4) \text{ S cm}^2 \text{ mol}^{-1}$$
$$\overset{\circ}{\Lambda}_{m(CH_3COOH)} = 390.5 \text{ S cm}^2 \text{ mol}^{-1}$$

73. (*d*) Weak electrolytes have lower degree of dissociation at higher concentrations. Change in Λ_m with dilution is due to increase in the degree of dissociation and consequently the number of ions in total volume of solution that contain 1 mole of electrolyte. Λ_m increases steeply on dilution, specially near lower concentrations.

At infinite dilution electrolyte dissociate completely but at such low concentration the conductivity of the solution is so low that it cannot be measured accurately.

74. (*c*) Given, molar conductance at 0.1 M concentration,

$$\lambda_c = 9.54\ \Omega^{-1}\text{cm}^2\text{mol}^{-1}$$

Molar conductance at infinite dilution,

$$\lambda_c^\infty = 238\ \Omega^{-1}\ \text{cm}^2\text{mol}^{-1}$$

We know that, degree of ionisation,

$$\alpha = \frac{\Lambda^C}{\lambda^\infty}\times 100 = \frac{9.54}{238}\times 100 = 4.008\%$$

75. (*d*) All the statements are true.

76. (*c*) Oxides of Al and chloride of Mg are electrolysed to produce these metals. Other statements are true.

78. (*c*) Equivalent $= \dfrac{\text{Atomic mass}}{\text{Equivalent weight}}$

79. (*b*) $Q = it = 96.5\times 100 = 9650$ C

96500 C charge deposited weight of Ag $= 108$ g

9650 C deposited weight of Ag $= \dfrac{108 \times 9650}{96500} = 10.8$ g

80. (*c*) $\underset{+6}{MnO_4^{2-}} \xrightleftharpoons{1\text{mol } e^-[=1F]} \underset{(+7)}{MnO_4^-}$

As per the equation, for 1 mole of MnO_4^{2-}, 1 F of electricity is required. Thus, for 0.1 mole of MnO_4^{2-}, 0.1 F of electricity is required.

Since; $1\,F = 96500\,C$

$\therefore \quad 0.1\,F = 0.1\times 96500\,C = 9650\,C$

Hence, 9650 C of electricity is required to completely oxidise MnO_4^{2-} to MnO_4^-.

81. (*c*) $Al^{3+} + 3e^- \longrightarrow Al$

Equivalent weight of Al $= \dfrac{27}{3} = 9$

From Faraday's first law, $w = zIt = \dfrac{E}{96500}It$

$$\therefore \quad t = \frac{w\times 96500}{E\times I} = \frac{50\times 96500}{9\times 105} = 5105.82s = 1.42 \text{ hr}$$

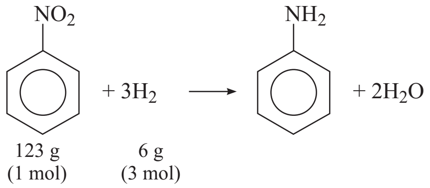
82. (*b*) We know that

Charge, $q = It = 1.5\times 3\times 60\times 60 = 16200C$

$\because$ Charge on one electron $= 1.6\times 10^{-19}$ C

$\therefore$ 16200 C charge is on $\dfrac{1\times 16200}{1.6\times 10^{-19}} = 1.0\times 10^{23}$ electrons

83. (*c*)



Moles of nitrobenzene used in the reaction $= \dfrac{12.3}{123} = 0.1$

H_2 required for reduction $= 0.1\times 3 = 0.3$ mol

Amount of charge required to liberate 0.3 mole of hydrogen

$$= 2 \times 96500 \times 0.3 = 57900 \text{ C.}$$

84. (*c*) Mass of silver in coating $= V \times d$

$$= 2\times 10^{-3}\times 100\times 1.05 = 2\times 0.105\,\text{g}$$
$$w = \frac{I\times t\times E}{96500}$$
$$\text{Or} \qquad t = \frac{w\times 96500}{I\times E} = \frac{0.105\times 2\times 96500}{10\times 108} = 18.7s$$

85. (*d*) $w = zit = 4\times 10^{-4}\times 6\times \dfrac{75}{100}\times 6\times 3600 = 38.8$ g

86. (*b*) From Faraday's second law,

$$\frac{w_{H_2}}{w_{Cu}} = \frac{E_{H_2}}{E_{Cu}} \Rightarrow w_{H_2} = \frac{1}{63.5}\times 6.35 = 0.1 \text{ g}$$

87. (*c*) $Ag^+ + \underset{\substack{1\text{ mol}\\1\text{ F}}}{e^-} \longrightarrow Ag;\ Cu^{2+} + \underset{\substack{2\text{ mol}\\2\text{ F}}}{2e^-} \longrightarrow Cu$

$$Fe^{2+} + \underset{\substack{3\text{ mol}\\3\text{ F}}}{3e^-} \longrightarrow Fe$$

Gram of Ag liberated $= \dfrac{108\times 10}{1} = 1080$ g

Gram of Cu liberated $= \dfrac{63.5\times 10}{2} = 317.5\,\text{g}$

Gram of Fe liberated $= \dfrac{56\times 10}{3} = 186.7\ \text{g}$

88. (*a*) When aqueous solution of NaCl is electrolysed, there is a competition between the following reductions reactions at cathode.

$$Na^+ + e^- \longrightarrow Na\ ; E^\circ_{cell}\ =\ -2.71\ \ V$$

$$H^+ + e^- \longrightarrow \frac{1}{2}H_2 ; E^\circ_{cell}\ = 0.00\ V$$

The reaction with higher value of E° is preferred and therefore, the second reaction occurs at cathode, i.e. H^+ is reduced instead of Na^+.

89. (*b*) H_2 is liberated at cathode and Cl_2 at anode during the electrolysis of NaCl solution. NaOH is formed so pH will rise.

90. (*b*) During the electrolysis of H_2SO_4 (dilute)

$$2H_2O(l)\ \longrightarrow\ O_2(g) + 4H^+(aq) + 4\,e^- ; E^\circ_{cell} = 1.23\ V$$

During the electrolysis of conc. H_2SO_4

$$2SO_4^{2-}(aq)\ \longrightarrow\ S_2O_8^{2-}(aq) + 2\,e^- ; E^\circ_{cell} = 1.96\ V$$

91. (*d*) When $MgSO_4\ (aq)$ is electrolysed, H_2 and O_2 are evolved at cathode and anode respectively.

92. (*a*) In batteries, cells are connected in series, not parallely.

93. (*c*) Primary batteries act as only electrochemical cell but not as electrolytic cell.

94. (*b*) **Anode** $Zn(s)\ \longrightarrow\ Zn^{2+}(aq) + 2e^-$

Cathode $MnO_2 + NH_4^+ + e^-\ \longrightarrow\ MnO(OH) + NH_3$

At cathode manganese is reduced from +4 oxidation state to +3 oxidation state. Ammonia produced in reaction form complex with Zn^{2+} to give $[Zn(NH_3)_4]^{2+}$. The cell has potential of 1.5 V.

95. (*a*) The cell consists of zinc container that act as anode and cathode is a carbon (graphite) rod surrounded by powdered manganese dioxide and carbon. The space between electrodes is filled by moist paste of NH_4Cl and $ZnCl_2$.

96. (*c*) Mercury cell is a low current device. In it $Zn - Hg$ acts as anode and a paste of $C + HgO$ as cathode. The cell potential remains the same due to the non-involvement of such a ion, concentration of which can change during its whole life.

98. (*b*) A rechargeable nickel-cadmium cell is a jelly roll arrangement and separated by layer soaked in moist sodium or potassium hydroxide.

Overall reaction during discharge is

$$Cd(s) + 2Ni(OH)_3(s) \longrightarrow CdO(s) + 2Ni(OH)_2(s) + H_2O(l)$$

99. (*a*) In thermal power plants, electricity is produced by using chemical energy of fossil fuels like coal etc.

100. (*d*) In hydrogen-oxygen fuel cell,

anode reaction is $2H_2(g) + 4OH^-(aq) \longrightarrow 4H_2O(l) + 4e^-$

Given reaction occurs at cathode.

102. (*d*) During rusting of iron, metal acts as anode, hydrogen ion as cathode. The overall reaction is

$$2Fe + O_2 + 4H^+ \longrightarrow 2Fe^{2+} + 2\ H_2O$$

Formula of rust is $Fe_2O_3 \cdot xH_2O$.

103. (*c*) $E^\circ_{cell} = E^\circ_{H^+/O_2/H_2O} - E^\circ_{Fe^{2+}/Fe} = 1.23 - (-0\ .\ 44) = 1.67\ V$

104. (*b*) Galvinisation is the process of applying a protective zinc coating to iron or steel.

105. (*a, b,c*) When $E_{ext} > 1.1V$ (in Daniell cell)

(i) electrons flow from Cu to Zn and current flows in opposite direction.

Zn is deposited at the zinc electrode and Cu dissolves from Cu electrode.

106. (*a, b, d*) No salt bridge is required when both the electrodes are dipped in the solution of same electrolyte.

107. (*a, b*) From Nernst equation, $E_{cell} = E^\circ_{cell} - \dfrac{0.0591}{nF}\log Q$.

If $Q = 1$, $\log Q = 0$ so, $E_{cell} = E^\circ_{cell}$

This condition is found in (a) and (b).

108. (*a, b*) Because E°_{cell} is not positive only for these reaction.

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

$$= 0.25 - 0.34 = -0.09\ V$$

Similarly, $E^\circ_{cell} = E^\circ_{H^+/H_2} - E^\circ_{Cu^{2+}/Cu}$

$$= -0.34\ \ V$$

109. (*a, c, d*) When same quantity of electricity is passed, same gram equivalents are obtained but masses are different.

110. (*a, b, c*) $\overset{+7}{Mn}O_4^- + 5e^- \longrightarrow Mn^{2+}$

$$5\ e^- \equiv 5F = 5\times 96500 = 4.82\times 10^5 C$$

111. (*a, d*) $NaOH \longrightarrow Na^+ + OH^-$

At anode

$$2H_2O \longrightarrow O_2 + 4H^+ + 4\,e^-$$

At cathode

$$H^+ + e^- \longrightarrow \frac{1}{2}H_2$$

113. (*a, b, c, d*) All the reactions are involved.

114. (*b, c, d*) Rust is $Fe_2O_3 \cdot xH_2O$

Saline water speeds up the rusting process and impure metal corrodes more faster as compared to pure metal.

115. (*c*) No current will flow through the cell and the reaction stops.

117. (*b*) The potential of individual half-cell cannot be measured. The difference between potentials of two half-cells give the emf of cell. $Pt(s)|H_2(g)|H^+(aq)$ half-cell is called standard hydrogen electrode.

120. (*b*) Amount of metal deposited can not be obtained with the help of electrochemical series.

125. (*b*) At electrode-electrolyte interface, there is a tendency of metal ion to deposit on metal electrode and make it positively charged. At the same time metal atom have tendency to go into solution as ions and leave behind the electrons at the electrode and make it negatively charged.

126. (*a*) Metals like platinum or gold are used as inert electrodes. They do not participate in the reaction but provide their surface for oxidation or reduction reactions and for conduction of electrons.

127. (*d*) $CuSO_4$ solution can not be stored in zinc pot because Zn is more reactive than Cu so it displaces Cu from $CuSO_4$ solution.

128. (*a*) Electrical conductance through metals is called metallic or electronic conductance because it is due to movement of electrons.

129. (*d*) Conductivity always decreases with decrease in concentration for weak and strong electrolytes. This is explained by the decrease in the number of ions per unit volume that carry current in solution, on dilution.

130. (*c*) Pure water has small amounts of hydrogen and hydroxyl ions ($\sim 10^{-7}$ M) due to which it has very low conductivity (3.5×10^{-5} S m^{-1}).

131. (*d*) When electrolytes are dissolved in water, they furnish their own ions in the solution, hence the conductivity increases.

132. (*c*) Anhyd. HCl does not conduct the electricity because it does not produce the ions.

133. (*a*) DC can not be used for the measurement of resistance of an ionic solution. This is because on passing DC the composition of solution changes. Secondly such a solution cannot be connected to the bridge like metallic wire or solid conductor.

134. (*d*) Pure water do not conduct electricity because it is unionised.

135. (*c*) Cu gets discharged at cathode $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

Copper metal is deposited at cathode and at anode, copper is converted into Cu^{2+} ion

$$Cu\,(s) \longrightarrow Cu^{2+}(aq) + 2e^-$$

i.e. Copper is dissolved from anode. It is the industrial process by which impure copper is converted into copper of high purity.

136. (*a*) Solid NaCl does not conduct electricity due to the absence of free ions.

138. (*c*) Electricity used $= 5 \times 20 \times 60 = 6000$ C.

$$Ni(NO_3)_2 + 2H^+ + \underset{2\,F}{2e^-} \longrightarrow \underset{58.7\text{ g}}{Ni} + 2HNO_3$$

Charge required to deposit 1 mole

$$Ni = 2\text{ F} = 2 \times 96500 = 1.93 \times 10^5 \text{ C}$$

1.93×10^5 C of charge produce Ni = 58.7 g

6000 C of charge will produce $Ni = \dfrac{58.7 \times 6000 \text{ C}}{1.93 \times 10^5} = 1.825$ g

Assertion is true but reason is false as mass of substance deposited is directly proportional to the quantity of electricity passed.

139. (*c*) Mercury cell gives the constant voltage throughout its life because the overall reaction does not involve any ion in the solution, concentration of which charges during its whole life.

140. (*d*) $H_2 - O_2$ fuel cell is used over ordinary cell because they do not cause any pollution and have high efficiency of 60-70%.

141. (*d*) Production of electricity by thermal plants is not very efficient and cause environmental pollution.

142. (*a*) Blocks of magnesium metal provide cathodic protection in order to protect sheet from oxidation.

143. (*d*) Chromium is more electropositive metal than iron. In stainless steel, chromium forms an oxide layer and protects it from corrosion.

144. (*c*) A – 2; B – 1; C – 3

145. (*b*) The inverse of resistance *R* is called conductance *G*.

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

Inverse of resistivity is called conductivity and is represented by κ.

147. (*a*) Given figure represents the corrosion of iron in the atmosphere. Oxidation occurs at anode and reduction occurs at cathode

Cathode $O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l);$

$$E^\circ_{H^+|O_2|H_2O} = 1.23 \text{ V}$$

Anode $2Fe(s) \longrightarrow 2Fe^{2+}(aq) + 4e^-; E^\circ_{Fe^{2+}/Fe} = -0.44$ V

Overall reaction

$$2Fe(s) + O_2(g) + 4H^+(aq) \longrightarrow 2Fe^{2+}(aq) + 2H_2O(l);\ E^\circ_{cell} = 1.67 \text{ V}$$

148. (*a*) A→p, q, r; B→p, q, r; C–r; D→p, r, s

Hydrogen is not a metal.

149. (*b*) A→p, s; B→q; C→p, r, s; D→p, s

A reaction is feasible, if its E°_{cell} value is positive.

A. $Fe^{3+}(aq) + I^-(aq) \longrightarrow Fe^{2+}(aq) + \frac{1}{2}I_2(g)$

$E^\circ_{cell} = E^\circ_{(Fe^{3+}/Fe^{2+})} - E^\circ_{1/2\ I_2/I^-} = (0.77 - 0.54) = 0.23$

The reaction is feasible.

i.e. $\Delta G^\circ = -ve$

B. $Fe^{3+}(aq) + Br^-(aq) \longrightarrow Fe^{2+}(aq) + \frac{1}{2}Br_2(g)$

$E^\circ_{cell} = E^\circ_{(Fe^{3+}/Fe^{2+})} - E^\circ_{1/2\ Br_2/Br^-} = 0.77 - 1.09 = -0.32$ V

The reaction is not feasible.

i.e. $\Delta G^\circ = +ve$

C. $Zn + 2H^+ \longrightarrow Zn^{2+} + H_2$

$E^\circ_{cell} = E^\circ_{H^+/H_2} - E^\circ_{Zn^{2+}/Zn} = 0 - (-0.76) = +\ 0.76$

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

D. $\frac{1}{2}Br_2(aq) + Fe^{2+}(aq) \longrightarrow Br^-(aq) + Fe^{3+}(aq)$

$E^\circ_{cell} = E^\circ_{(1/2\ Br_2/Br^-)} - E^\circ_{(Fe^{3+}/Fe^{2+})} = 1.09 - 0.77 = 0.32$ V

The reaction is feasible.

$\Delta G^\circ = -ve$

150. (*b*) A→q; B→p, r; C→p; D→s

151. (*c*) A→p, q; B→r; C→s; D→s

152. (*d*) A→r, s; B→r, p; C→p, q; D→s

A. An aqueous solution of $AgNO_3$ with silver electrodes.

In aqueous solution, ionisation of $AgNO_3$ and H_2O takes place.

$$AgNO_3(s) \xrightleftharpoons{(aq)} Ag^+(aq) + NO_3^-(aq)$$

$$H_2O\,(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

At cathode Ag^+ ions has less discharge potential than H^+ ions so silver will be deposited at cathode.

$$Ag^+(aq) + e^- \longrightarrow Ag(s)$$

At anode An equivalent amount of silver will be oxidised to Ag^+ ions by releasing electrons.

$$Ag(s) \longrightarrow Ag^+(aq) + e^-$$

Ag anode is attacked by NO_3^- ions, so it will also produce Ag^+ in the solution.

B. An aqueous solution of $AgNO_3$ with platinum electrodes. In aqueous solution, ionisation of $AgNO_3$ and H_2O both occur.

$$AgNO_3(s) \xrightleftharpoons{(aq)} Ag^+(aq) + NO_3^-(aq)$$
$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

As platinum electrodes are non-attackable electrodes, they will not be reacted upon by NO_3^- ions.

At cathode Ag will be deposited at cathode.

$$Ag^+(aq) + e^- \longrightarrow Ag(s)$$

At anode Out of NO_3^- and OH^- ions, only OH^- ions will be oxidised (due to less discharge potential) preferentially and NO_3^- ions will remain in the solution.

$$OH^-(aq) \longrightarrow OH + e^-$$
$$4\,OH \longrightarrow 2\,H_2O(l) + O_2(g)$$

So, oxygen gas is produced at anode. The solution remains acidic due to the presence of HNO_3.

$$H^+(aq) + NO_3^-(aq) \rightleftharpoons HNO_3(aq)$$

C. A dilute solution of H_2SO_4 with platinum electrodes. Both H_2SO_4 and water ionise in the solution.

$$H_2SO_4(aq) \rightleftharpoons 2H^+(aq) + SO_4^{2-}(aq)$$
$$H_2O(l) \rightleftharpoons 2H^+(aq) + OH^-(aq)$$

At cathode H^+ ions will be reduced and hydrogen gas is produced at cathode.

$$H^+(aq) + e^- \longrightarrow H(g)$$
$$H(g) + H(g) \longrightarrow H_2(g)$$

At anode OH^- ions will be released preferentially and not SO_4^- ions due to less discharge potential.

$$OH^-(aq) \longrightarrow OH + e^-$$
$$4\,OH \longrightarrow 2H_2O(l) + O_2(g)$$

Oxygen gas is produced at anode.

Solution will be acidic and will contain H_2SO_4.

D. An aqueous solution of $CuCl_2$ with platinum electrodes. Both $CuCl_2$ and water ionise as usual.

$$CuCl_2 \xrightleftharpoons{(aq)} Cu^{2+}(aq) + 2Cl^-(aq)$$
$$H_2O(l) \rightleftharpoons 2H^+(aq) + OH^-(aq)$$

At cathode Cu^{2+} ions will be reduced preferentially due to less discharge potential than H^+ ions.

$$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$$

Copper metal is deposited at cathode.

At anode Cl^- ions will be discharged in preference to OH^- ions and chlorine gas is produced at anode.

$$Cl^-(aq) \longrightarrow Cl(g) + e^-$$
$$Cl(g) + Cl(g) \longrightarrow Cl_2(g)$$

153. (*b*) Increase in the external potential (>1.1) again starts the reaction but in opposite direction, so direction of current is from zinc to copper.

156. (*a*) $E^\circ_{cell} = E^\circ_{H^+,\,Mn^{2+},\,MnO_4^-\,|\,Pt} - E^\circ_{Fe^{3+},\,Fe^{2+}\,|\,Pt}$

$$= 1.51 - (0.77) = +0.74\ V$$

157. (*a*) $E = E^\circ - \dfrac{RT}{5\,F} \ln\left(\dfrac{[Mn^{2+}][Fe^{3+}]^5}{[MnO_4^-][H^+]^8[Fe^{2+}]^5}\right)$

Lowering pH means increasing $[H^+]$ and it will make term less positive and E_{cell} becomes positive.

158. (*d*) $E_{cell} = 0.74\ V - \dfrac{RT}{5\,F} \ln\left[\dfrac{(0.5)(2.0)^5}{(1.0)(2.0)^8(0.5)^5}\right]$

$$= 0.74\ V - \frac{0.059}{5} \log\left[\frac{1}{(2.0)^3(0.5)^4}\right]$$
$$= 0.74\ V - \left(\frac{0.059}{5}\right) \log 2$$
$$= 0.74\ V - \frac{0.059 \times 0.301}{5} = 0.736\ V$$

159. (*b*) $\Delta E = -\left(\dfrac{0.059}{5}\right) \log(0.5) = +\,3.55 \times 10^{-3}\ V$

160. (*b*) The cell constant is given by the equation:

Cell constant, $G^* = \text{conductivity} \times \text{resistance}$

$$= 1.29\ S/m \times 100\,\Omega$$
$$= 129\ m^{-1} = 1.29\ cm^{-1}$$

Conductivity of 0.2 mol L^{-1} KCl solution

$$= \text{cell constant/ resistance}$$
$$= \frac{G^*}{R} = \frac{129\ m^{-1}}{520\ \Omega} = 0.248\ S\ m^{-1}$$

161. (*d*) Concentration $= 0.2\ mol\ L^{-1}$

$$= 1000 \times 0.2\ mol\ m^{-3} = 200\ mol\ m^{-3}$$

Molar conductivity $= \Lambda_m = \dfrac{\kappa}{C}$

$$= \frac{248 \times 10^{-3}\ S\ m^{-1}}{200\ mol\ m^{-3}}$$
$$= 124 \times 10^{-5}\ S\ m^2\ mol^{-1}$$

162. (*c*) Number of gram-equivalents of hydrogen produced

$$= \frac{67.2}{11.2} = 6$$
$$I = \frac{Q}{t} = \frac{6 \times 96500}{15 \times 60} = 643.33\ A$$

163. (*c*) Weight of copper deposited $= \dfrac{6 \times 63.5}{2} = 190.5\ g$

164. (*d*) Six equivalents of H_2 is oxidised.

165. (*d*) $Na^+ + e^- \xrightarrow{Hg} Na(Hg)$

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

Moles of NaCl electrolysed is $4 \times \dfrac{500}{1000} = 2.0$

Two Faraday of electric charge is required for electrolysis of 2 moles of NaCl.
Total coulombs $= 2 \times 96500 = 193000$ C.

166. (*d*) 2 moles of Na formed produced two moles of Na(Hg) amalgam.
Mass of amalgam $= 2 \times (23 + 200) = 446$ g

167. (*b*) 1 mole of Cl_2 is produced.

168. (*a*) **Anode** $2Cl^-(aq) \longrightarrow Cl_2(g) + 2e^-$

169. (*b*) **Cathode** $2H_2O(aq) + 2e^- \longrightarrow H_2(g) + 2OH^-$

170. (*a*) $1\text{kg } Cl_2 = \dfrac{1000}{35.5}$ equivalent of $Cl_2 = 28.17$ equivalent of Cl_2

Theoretical electricity required $= 28.17$ F
Efficiency is 62%

Electricity required (experimental) $= \dfrac{28.17 \times 100\,F}{62} = 45.44$ F

$\Rightarrow 45.44 \times 96500 = 25t; \quad t = 48.72$ hr

171. (*c*) The amount of electricity required (experimentally) in terms of Faraday is 45.44 F.

173. (*d*) Zinc-mercury amalgam acts as anode and a paste of HgO and carbon as cathode in the mercury cell. The electrolyte is the paste of KOH and ZnO.

174. (*b*) **Anode** $Zn(Hg) + 2OH^-(aq) \longrightarrow ZnO(s) + H_2O + 2e^-$

Cathode $HgO(s) + H_2O + 2e^- \longrightarrow Hg(l) + 2OH^-$

The overall reaction is

$$Zn(Hg) + HgO(s) \longrightarrow ZnO(s) + Hg(l)$$

176. (*a*) The reducing agent is zinc and oxidising agent is mercury (II) oxide in this cell.

178. (*b*) Lead storage battery consists of spongy lead anode and a grid of lead packed with lead dioxide (PbO_2) as cathode. 38% solution of H_2SO_4 is used as an electrolyte.

179. (*c*) **Anode** $Pb(s) + SO_4^{2-}(aq) \longrightarrow PbSO_4(s) + 2\,e^-$

Cathode $PbO_2(s) + SO_4^{2-}(aq) + 4H^+ + 2\,e^- \longrightarrow PbSO_4(s) + 2H_2O(l)$

Overall cell reaction

$$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \longrightarrow 2PbSO_4(s) + 2H_2O(l)$$

At the time of charging, a reverse reaction takes place. In which $PbSO_4$ (*s*) at anode and cathode is converted into Pb and PbO_2 respectively.

180. (*c*) If an external opposite potential is applied on the Galvanic cell and increased, reaction continue to take place till the opposing voltage reaches the value of 1.1 V. At this stage no current flow through the cell and if there is any further increase in the external potential then reaction starts functioning in opposite direction.
Hence, this works as an electrolytic cell.

181. (*c*) Standard electrode potential of copper electrode can be calculated by constructing a concentration cell composed of two half-cell reactions in which concentration of species on left hand and right hand side are unity.

In such case cell potential is equal to standard electrode potential.

$$\underbrace{Pt(s)\,|\,H_2(g, 1\text{ bar})\,|\,|\,H^+(aq, 1\text{ M})\,|}_{\text{Oxidation half -cell reaction}}\;\underbrace{|\,Cu^{2+}(aq, 1\text{ M})\,|\;Cu}_{\text{Reduction half -cell reaction}}$$

183. (*d*) An inert electrode in a cell provides surface for either oxidation or for reduction reaction by conduction of electrons through its surface but does not participate in the cell reaction.
It does not provide surface for redox reaction.

184. (*b*) Electrode potential for Mg electrode varies according to the equation,

$$E_{Mg^{2+}/Mg} = E^\circ_{Mg^{2+}/Mg} - \frac{0.059}{2}\log\frac{1}{[Mg^{2+}]}$$

$$E_{Mg^{2+}/Mg} = E^\circ_{Mg^{2+}/Mg} + \frac{0.059}{2}\log[Mg^{2+}]$$

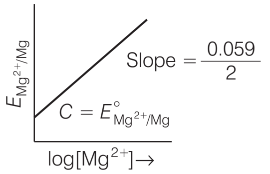
$$E_{Mg^{2+}/Mg} = \frac{0.059}{2}\log\,[Mg^{2+}] + E^\circ_{Mg^{2+}/Mg}$$

This equation represents equation of straight line. It can be correlated as

$$\begin{array}{ccccccc} E_{Mg^{2+}/Mg} & = & \left(\dfrac{0.059}{2}\right) & \log[Mg^{2+}] & + & E^\circ_{Mg^{2+}/Mg} \\ \uparrow & & \uparrow & \uparrow & & \uparrow \\ Y & & m & x & & +C \end{array}$$

So, intercept $(C) = E^\circ_{Mg^{2+}/Mg}$

This, equation can be diagramatically represented as



185. (*c*) E_{cell} is an intensive property as it does not depend upon mass of species (number of particles) but $\Delta_r G$ of the cell reaction is an extensive property because this depends upon mass of species (number of particles).

186. (*b*) Here, out of given four options standard reduction potential of chromium has highest negative value, hence most powerful reducing agent is chromium.

187. (*c*) Higher the positive value of standard reduction potential of metal ion, higher will be its oxidising capacity.

Since, $E^\circ_{MnO_4^-/Mn^{2+}}$ has value equal to 1.51 V, hence it is the strongest oxidising agent.

188. (*b*)

Species (ions)	**SRP values**
Mn^{2+}	1.51 V
Cl^-	1.36 V
Cr^{3+}	1.33 V
Cr	− 0.74 V

On moving top to bottom SRP value decreases from positive to negative which will increase the reducing capacity.

189. (*d*) $E^\circ_{MnO_4^-/Mn^{2+}}$ has positive value equal to 1.51 V which is highest among given four choices. So, Mn^{2+} is the most stable ion in its reduced form.

190. (*a*) $E^\circ_{Cr^{3+}/Cr}$ has most negative value of E°, i.e. – 0.74 lowest among given four choices. So, Cr^{3+} is the most stable oxidised species.

191. (*d*) Cell constant is defined as the ratio of distance between the electrodes and their area of cross section.

$$G = \frac{l}{A}$$

Since, l and A remain constant for any particular cell hence value of cell constant always remains constant for a cell.

192. (*c*) Solution consists of electrolytes is known as electrolytic solution and conductivity of electrolytic solution depends upon the following factors

(i) **Size of ions** As ion size increases, ion mobility decreases and conductivity decreases.

(ii) **Viscosity of solution** Greater the viscosity of the solution, lesser will be the conductivity of the solution.

(iii) **Solvation of ions** Greater the solvation of ions of an electrolyte, lesser will be the electrical conductivity of the solution.

(iv) **Temperature of medium** Conductivity of solution increases with increase in temperature.

193. (*d*)

$$\begin{aligned}
\Lambda^\circ_{m\,(NH_4Cl)} &= \Lambda^\circ_{m\,(NH_4^+)} + \Lambda^\circ_{m\,(Cl^-)}\\
\Lambda^\circ_{m\,(NaOH)} &= \Lambda^\circ_{m\,(Na^+)} + \Lambda^\circ_{m\,(OH^-)}\\
\Lambda^\circ_{m\,(NaCl)} &= \Lambda^\circ_{m\,(Na^+)} + \Lambda^\circ_{m\,(Cl^-)}\\
&\underline{-\qquad\qquad -\qquad\qquad\qquad -}
\end{aligned}$$

$$\Lambda^\circ_{m\,(NH_4Cl)} + \Lambda^\circ_{m\,(NaOH)} - \Lambda^\circ_{m\,(NaCl)} = \Lambda_{m\,(NH_4OH)}$$

194. (*d*) In case of electrolysis of *aq*. NaCl oxidation reaction occurs at anode as follows:

$$Cl^-(aq) \longrightarrow \frac{1}{2}Cl_2(g) + e^-;\ E^\circ = 1\cdot 36\ V$$

$$2H_2O(l) \longrightarrow O_2(g) + 4H^+(aq) + 4e^-;\quad E^\circ_{cell} = 1.23\ V$$

But due to lower E°_{cell} value, water should get oxidised in preference of Cl^- (*aq*).

However, the actual reaction taking place is the first one i.e. Cl_2 is produced.

This unexpected result is explained on the basis of the concept of 'overvoltage', i.e. water needs greater voltage for oxidation to O_2 (as it is kinetically slow process) than that needed for oxidation of Cl^- ions to Cl_2.

195. (*c*) The quantity of charge required to obtain one mole of aluminium from Al_2O_3 is equal to number of electrons required to convert Al_2O_3 to Al.

$$Al^{3+}(aq) \xrightarrow{+3e} Al(s)$$

Hence, total 3F is required.

196. (*a*) While charging the lead storage battery the reaction occurring in the cell is reversed and $PbSO_4$ (*s*) on anode and cathode is converted into Pb and PbO_2 respectively.

197. (*b, c*)

(i) Left side of cell represents oxidation half-cell i.e. oxidation of Mg and right side of cell represents reduction half-cell i.e. reduction of copper.

(ii) Cu is reduced and reduction occurs at cathode.

(iii) Mg is oxidised and oxidation occurs at anode.

(iv) Whole cell reaction can be written as

Oxidation

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

Reduction

198. (*b, d*) Lesser the E° value of redox couple, higher is the reducing power.

$$Cu^{2+} + 2e^- \longrightarrow Cu \qquad E^\circ = 0.34\ V$$

$$2H^+ + 2e^- \longrightarrow H_2 \qquad E^\circ = 0.00\ V$$

Since, $2H^+ / H_2$ has lesser SRP than Cu^{2+} /Cu redox couple.

Therefore,

(i) This redox couple is a stronger oxidising agent than H^+ / H_2.

(ii) Cu cannot displace H_2 from acids.

199. (*a, c*) During the electrolysis of dilute sulphuric acid solution all the given three reactions occur each of which represents particular reaction either oxidation half-cell reaction or reduction half-cell reaction.

Oxidation half-cell reactions occur at anode are as follows

$$2SO_4^{2-}(aq) \longrightarrow S_2O_8^{2-} + 2e^-\ E^\circ_{cell} = 1.96\ V$$

$$2H_2O(l) \longrightarrow O_2(g) + 4H^+(aq) + 4e^-;\ E^\circ_{cell} = 1.23\ V$$

Reaction having lower value of E°_{cell} will undergo faster oxidation. Hence, oxidation of water occur preferentially for dilute sulphuric acid but in case of concentrated H_2SO_4 former reaction (oxidation of SO_4^{2-}) is preferred.

Reduction half-cell reaction occurs at cathode is

$$H^+(aq) + e^- \longrightarrow \frac{1}{2}H_2(g) \qquad \text{At cathode}$$

200. (*b, c*) At state of equilibrium,

$$\begin{aligned}
\Delta G &= -RT\ln K\\
-nFE^\circ &= -RT\ 2.303\log K_C\\
E^\circ &= \frac{+RT\ 2.303\ \log K_C}{+2F} \qquad (n = 2 \text{ for Daniell cell})
\end{aligned}$$

$\because$ At equilibrium $E^\circ = 1.1$

$$\therefore\ \frac{2.303\,RT}{2F}\log K_C = 1.1 \Rightarrow \log K_C = \frac{2.2}{0.059} \qquad \text{[on solving]}$$

201. (*a, b*) As we know conductance is reciprocal of resistance and conductivity is the conductance of 1 cm^3 of substance. Also,conductivity is reciprocal of resistivity.

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

$$\rho = \frac{R\cdot A}{l} \Rightarrow \kappa = \frac{1}{\left(\frac{R\cdot A}{l}\right)}$$

$$\kappa = \frac{1}{R}\cdot\frac{l}{A} = \frac{1}{R}\times G^* = \frac{G^*}{R}$$

202. (*a, b*) Conductivity of electrolytic solution is due to the presence of mobile ions in the solution. This type of conductance is known as ionic conductance. Hence, conductivity of electrolytic solutions depend upon

(i) the nature of electrolyte added
(ii) size of the ion produced and their solvation
(iii) concentration of electrolyte
(iv) nature of solvent and its viscosity
(v) temperature

Power of source or distance between the electrodes has no effect on conductivity of electrolytic solution.

203. (*a, c*) Molar conductivity is the conductivity due to ions furnished by one mole of electrolyte in solution. It depends on

(i) **Temperature** It increases with increase in temperature.
(ii) **Concentration of electrolytes in solution** As concentration of electrolyte increases, molar conductivity decreases.

$$\therefore \qquad \Lambda_m = \frac{\kappa}{C}$$

204. (*a, c*) For electrolysis of aqueous solution of $CuSO_4$,

$$CuSO_4\,(aq) \longrightarrow Cu^{2+} + SO_4^{2-}$$

$$H_2O \longrightarrow 2H^+ + O^{2-}$$

$$\text{At anode} \qquad 2O^{2-} \longrightarrow O_2 + 2e^-$$

$$\text{At cathode} \qquad Cu^{2+} + 2e^- \longrightarrow Cu\,(s)$$

205. (*a, b*) Electrolysis of $CuSO_4$ can be represented by following two half-cell reactions

At cathode $Cu^{2+} + 2e^- \longrightarrow Cu(s)$
At anode $Cu(s) \longrightarrow Cu^{2+} + 2e^-$

Here, Cu will deposit at cathode while copper will dissolved at anode.

206. (*c*) Assertion is true but the reason is false. Feasibility of chemical reaction depends on Gibb's free energy which is related to E°_{cell} as, $\Delta G^\circ = -nFE^\circ_{cell}$

When value of E°_{cell} is positive then ΔG° becomes negative.
Hence, reaction becomes feasible.
Correct reason is $E_{cathode} > E_{anode}$.

207. (*e*) Assertion is false but reason is true. Correct assertion is mercury cell gives steady potential.

Reason is correct as ions are not involved in the cell reaction.

208. (*a*) Current stops flowing when $E_{cell} = 0$

As at $E_{cell} = 0$, reaction reaches the equilibrium.

209. (*c*) Electrode potential of Cu^{2+}/Cu is + 0.34 V and electrode potential of $2H^+/H_2$ is 0.00 V.

Hence, correct reason is due to positive value of Cu^{2+}/Cu it is less reactive than hydrogen.

210. (*b*) Both Assertion and Reason are correct but Reason is not the correct explanation of Assertion.

$$E = E^\circ - \frac{0.0591}{1}\log\frac{1}{[Ag^+]}$$

$$= E^\circ + 0.059 \log [Ag^+]$$

Thus, $E_{Ag^+/Ag}$ increases with increase in concentration of Ag^+.

211. (*d*) Copper sulphate cannot be stored in zinc vessel as zinc is more reactive than copper due to negative value of standard reduction potential of Zn.

212. (*a*) Since conductivity depends upon the number of ions per unit volume. Therefore, the conductivity of all electrolytes decreases on dilution due to decrease in number of ions per unit volume.

213. (*a*) Molar conductivity of weak electrolytic solution increases on dilution, because as we add excess water to increase the dilution degree of dissociation increases which leads to increase in the number of ions in the solution. Thus, Λ_m show a very sharp increase.

214. (*a*) Concentration of ionic solution changes on using DC current as a source of energy while on passing AC current concentration does not change. Hence, AC source is used for measuring resistance.

215. (*a*) Electrolysis of NaCl is represented by following chemical reactions

$$\textbf{At cathode} \qquad H^+(aq) + e^- \longrightarrow \frac{1}{2}H_2(g)$$

$$\textbf{At anode} \qquad Cl^-(aq) \longrightarrow \frac{1}{2}Cl_2 + e^-;\ E^\circ_{cell} = 1.36\text{ V}$$

$$2H_2O\,(aq) \longrightarrow O_2(g) + 4H^+\,(aq) + 4e^-;\ E^\circ_{cell} = 1.23\text{ V}$$

E°_{cell} for this reaction has lower value but formation of oxygen at anode requires over potential.

216. (*b*) A. → (3) B. → (1) C. → (7) D. → (5)
E. → (4) F. → (2) G. → (6)

A. F_2 is a non-metal and best oxidising agent because SRP of F_2 is + 2.87 V.
B. Li is a metal and strongest reducing agent because SRP of Li is − 3.05 V.
C. Au^{3+} is a metal ion which is an oxidising agent as SRP of Au^{3+} is +1.40 V.
D. Br^- is an anion that can be oxidised by Au^{3+} as Au^{3+} ($E^\circ = 1.40$) is greater than Br^-($E^\circ = 1.09$ V).
E. Au is an unreactive metal.
F. Li^+ is a metal ion having least value of SRP (−3.05 V), hence it is the weakest oxidising agent.
G. F^- is an anion which is the weakest reducing agent as F^-/F_2 has low oxidation potential (−2.87 V).

217. (*d*) A. → (4) B. → (1) C. → (2) D. → (3)

A. Λ_m (molar conductivity) is the conductivity due to number of ions furnished by one mole of electrolyte. As dilution increases, number of ions present in the solution increases hence, molar conductivity increases.
B. E°_{cell} of any atom/ion does not depend upon number of atom/ion, hence E°_{cell} of any atom/ion is an intensive property.
C. κ represents specific conductivity which depends upon number of ions present in per unit volume.
D. $\Delta_r G_{cell}$ is an extensive property as it depends upon number of particles(species).

218. (*d*) A. → (4) B. → (3) C. → (1) D. → (2)

219. (*b*) A. → (4) B. → (3) C. → (2) D. → (1)

220. (*d*) A. → (4) B. → (3) C. → (1) D. → (2)

221. (*a*) A. → (4) B. → (3) C. → (5) D. → (2)

A. **Leclanche cell** The electrode reaction occurs at Leclanche cell are

At anode $Zn(s) \longrightarrow Zn^{2} + 2e^-$

At cathode

$$MnO_2 + NH_4^+ + e^- \longrightarrow MnO(OH) + NH_3$$

B. Ni-Cd cell is rechargeable. So, it has more life time.

C. Fuel cell produces energy due to combustion. So, fuel cell converts energy of combustion into electrical energy e.g.

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

D. Mercury cell does not involve any ion in solution and is used in hearing aids.

222. (*b*) Given, $Q = 2F$, Atomic mass of Cu = 63.5 u

Valency of the metal $Z = 2$

We have, $CuSO_4 \longrightarrow Cu^{2+} + SO_4^{2-}$

$$\underset{1\,mol}{Cu^{2+}} + \underset{\substack{2\,mol\\2F}}{2e^-} \longrightarrow \underset{1\,mol\,=\,63.5\,g}{Cu}$$

Alternatively, $w = zQ = \dfrac{E}{F} \cdot 2F = 2E = \dfrac{2 \times 63.5}{2} = 63.5\text{ g}$

223. (*c*) According to Debye Huckel Onsager equation,

$$\lambda_C = \lambda_\infty - B\sqrt{C}$$

where, λ_C = limiting equivalent conductivity at concentration C

λ_∞ = limiting equivalent conductivity at infinite dilution

C = concentration

224. (*a*) If E°_{cell} is positive, reaction is spontaneous, otherwise it is non-spontaneous.

$$Mn^{2+} + 2e^- \longrightarrow Mn;\ E^\circ = -1.18\text{V}$$
$$2Mn^{2+} \longrightarrow 2Mn^{3+} + 2e^-;\ E^\circ = -1.51\text{ V}$$
$$3Mn^{2+} \longrightarrow Mn + 2\,Mn^{3+};$$
$$E^\circ = E^\circ_{oxi} + E^\circ_{red} = -1.51 + (-1.18) = -2.69\text{V}$$

Negative E°_{cell} of the reaction indicates that reaction is non-spontaneous.

225. (*d*) Higher the SRP, better the oxidising agent. Among the given $E^\circ_{MnO_4^-/Mn^{2+}}$ is highest, hence MnO_4^- is the strongest oxidising agent.

226. (*d*) $\Delta G = -nEF = -\dfrac{2 \times 0.059 \times 96500}{1000}\text{ kJ} = -11.4\text{ kJ}$

227. (*d*) A cell reaction is spontaneous, if $\Delta G^\circ < 0$

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

Thus, $E^\circ_{cell} > 0$

$E^\circ_{cell} = E^\circ_{oxi} + E^\circ_{oxi}$

Reduction

$$X + Y^{2+} \longrightarrow X^{2+} + Y$$

Oxidation

(a) $X = Ni$, $Y = Fe$

$$Ni + Fe^{2+} \longrightarrow Ni^{2+} + Fe;$$

$E^\circ_{Ni/Ni^{2+}} = +\,0.23$ V

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

Thus, $E^\circ_{cell} = E^\circ_{Ni/Ni^{2+}} + E^\circ_{Fe^{2+}/Fe}$

$= -0.21$ V, $E^\circ_{cell} < 0$, reaction is non-spontaneous.

(b) $X = Ni$, $Y = Zn$,

$$Ni + Zn^{2+} \longrightarrow Ni^{2+} + Zn$$

$E^\circ_{Ni/Ni^{2+}} = 0.23$ V, $E^\circ_{Zn^{2+}/Zn} = -\,0.76$ V

$E^\circ_{cell} = -0.53$ V,

$E^\circ_{cell} < 0$, reaction is non-spontaneous.

(c) $X = Fe$, $Y = Zn$, $Fe + Zn^{2+} \longrightarrow Fe^{2+} + Zn$

$E^\circ_{cell} = E^\circ_{Fe/Fe^{2+}} + E^\circ_{Zn^{2+}/Zn} = 0.44 - 0.76 = -0.32$ V

$E^\circ_{cell} < 0$, reaction is non-spontaneous.

(d) $X = Zn$, $Y = Ni$,

$$Zn + Ni^{2+} \longrightarrow Zn^{2+} + Ni$$

$E^\circ_{cell} = E^\circ_{Zn/Zn^{2+}} + E^\circ_{Ni^{2+}/Ni} = 0.76 - 0.23 = 0.53$ V

The reaction is spontaneous.