CHEMISTRY (CODE – 043)

Chapter-wise Test: Electrochemistry

(NCERT Class XII Chemistry, Unit 2: Electrochemistry)

SOLUTIONS AND MARKING SCHEME

CLASS XII (2026-27)

Time: 1½ hours **Max. Marks: 35**

Step-by-step solutions with the distribution of marks. Marks for each step are shown in square brackets [ ]. Any other correct method or answer should be given full credit.

Quick answer key – Section A

Q.	1	2	3	4	5	6	7	8
Ans.	B	C	D	A	B	A	C	B

SECTION A

Questions 1 to 8 are multiple choice questions. Only one of the choices is correct. Select and write the correct choice as well as the answer to these questions. Each question carries 1 mark.

1	**Answer: B. S m^2 mol^{-1}** $\Lambda_m = \kappa / c$. With κ in S m^{-1} and c in mol m^{-3}, Λ_m has the unit S m^{-1} / mol m^{-3} = S m^2 mol^{-1}. (S m^{-1} is the unit of conductivity, and Ω m is the unit of resistivity.)	1
2	**Answer: C. $Cr_2O_7^{2-}$** A species can oxidise Br^- only if its reduction potential is higher than $E°(Br_2/Br^-)$ = 1.09 V, and it must not oxidise Cl^-, so its $E°$ must be lower than 1.36 V. $Cr_2O_7^{2-}$ (1.33 V) satisfies 1.09 < 1.33 < 1.36. MnO_4^- (1.51 V) oxidises both; Fe^{3+} (0.77 V) and I_2 (0.54 V) oxidise neither.	1
3	**Answer: D** For a strong electrolyte Λ_m varies linearly with $c^{1/2}$ ($\Lambda_m = \Lambda°_m - Ac^{1/2}$), so the straight line A is a strong electrolyte and its intercept gives $\Lambda°_m$. For a weak electrolyte Λ_m rises very steeply at low concentration because the degree of dissociation increases sharply, so the curve B cannot be extrapolated reliably. ($\Lambda°_m$ of a weak electrolyte is obtained by Kohlrausch's law.)	1
4	**Answer: A. 6 F** $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$. Six moles of electrons are needed, i.e. 6 × 96500 C = 6 F.	1
5	**Answer: B. Lead storage battery** In a secondary cell the cell reaction can be reversed by passing current in the opposite direction. The lead storage battery and the Ni–Cd cell are the important examples. The dry cell and mercury cell are primary cells (reaction occurs only once), and a fuel cell works by continuous supply of reactants.	1
6	**Answer: A. −212.3 kJ mol^{-1}** $\Delta_r G° = -nFE°_{cell} = -2 \times 96500$ C $mol^{-1} \times 1.1$ V = −212300 J mol^{-1} = **−212.3 kJ mol^{-1}**. The value is negative, so the reaction is spontaneous.	1
7	**Answer: C. 108 g** $Ag^+ + e^- \rightarrow Ag$. One mole of electrons (1 F = 96500 C) deposits one mole of silver, i.e. 108 g (Faraday's first law).	1
8	**Answer: B. Both A and R are true, and R is not the correct explanation of A.** Conductivity (κ) decreases on dilution because the number of ions per unit volume decreases. Molar conductivity ($\Lambda_m = \kappa/c$) increases on dilution because the volume containing one mole of electrolyte increases (and, for weak electrolytes, the degree of dissociation increases). R is true, but it is not the reason for the fall in κ.	1

SECTION B

Question Nos. 9 to 11 are very short answer questions carrying 2 marks each.

9	**(A)** (i) The limiting molar conductivity of an electrolyte can be represented as the sum of the individual contributions of its cation and anion: $\Lambda^\circ_m = \nu_+\lambda^\circ_+ + \nu_-\lambda^\circ_-$. [1] (ii) $\Lambda^\circ_m(CH_3COOH) = \lambda^\circ(H^+) + \lambda^\circ(CH_3COO^-) = 349.6 + 40.9 =$ **390.5 S cm^2 mol^{-1}** [1] **(B)** (i) $\Lambda^\circ_m(CaCl_2) = \lambda^\circ(Ca^{2+}) + 2\lambda^\circ(Cl^-) = 119.0 + 2(76.3) =$ **271.6 S cm^2 mol^{-1}** [1] (ii) $\Lambda^\circ_m(MgSO_4) = 106.0 + 160.0 =$ **266.0 S cm^2 mol^{-1}** [1]	1 + 1
10	**(A)** Anode: $Pb(s) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2e^-$ [1] Cathode: $PbO_2(s) + SO_4^{2-}(aq) + 4H^+(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$ [1] **(B)** (i) Anode: $Zn(Hg) + 2OH^- \rightarrow ZnO(s) + H_2O + 2e^-$; Cathode: $HgO + H_2O + 2e^- \rightarrow Hg(l) + 2OH^-$ [1] (ii) The overall reaction, $Zn(Hg) + HgO(s) \rightarrow ZnO(s) + Hg(l)$, involves no ion in solution whose concentration changes during the life of the cell, so the cell potential (about 1.35 V) remains constant. [1]	1 + 1
11	**(A)** Cell constant $G^* = \kappa \times R = 1.29$ S m^{-1} $\times$ 100 Ω = **129 m^{-1}** (= 1.29 cm^{-1}) [1] $\kappa = G^*/R = 129$ m^{-1} / 200 Ω = **0.645 S m^{-1}** [1] **(B)** $\Lambda_m = \kappa \times 1000 / c$ (κ in S cm^{-1}, c in mol L^{-1}) [1] $\Lambda_m = (0.0015 \times 1000) / 0.01 =$ **150 S cm^2 mol^{-1}** [1]	1 + 1

SECTION C

Question Nos. 12 to 15 are short answer questions carrying 3 marks each.

12	**(A)** (i) Zinc has the lower (more negative) electrode potential, so it is oxidised: **Zn is the anode** ($Zn \rightarrow Zn^{2+} + 2e^-$) and **Cu is the cathode** ($Cu^{2+} + 2e^- \rightarrow Cu$). Electrons flow in the external circuit from the zinc electrode (anode) to the copper electrode (cathode); the conventional current flows from Cu to Zn. [1] (ii) Zn(s) | Zn^{2+}(1 M) || Cu^{2+}(1 M) | Cu(s) [½]. $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} = 0.34 - (-0.76) =$ **1.10 V** [½] (iii) The salt bridge (a) completes the electrical circuit by allowing ions to flow between the two half-cells, and (b) maintains electrical neutrality of the two solutions. Without it the cell would stop working. [1] **(B)** (i) $E_{cell} = E^\circ_{cell} - (0.059/2) \log([Zn^{2+}]/[Cu^{2+}])$ [1] (ii) $E_{cell} = 1.10 - 0.0295 \times \log(0.1/0.01) = 1.10 - 0.0295 \times \log 10$ [1] $= 1.10 - 0.0295 =$ **1.0705 V ≈ 1.07 V** [1]	1 + 1 + 1

13	**(A)** (i) **Magnesium**: it has the most negative $E°$ (−2.36 V), so it loses electrons most easily. [1] (ii) $E°_{cell} = E°(\text{cathode, } Ag^+/Ag) - E°(\text{anode, } Fe^{2+}/Fe) = 0.80 - (-0.44) =$ **1.24 V** [1] (iii) $n = 2$. $\Delta_r G° = -nFE° = -2 \times 96500 \times 1.24 = -239320$ J mol^{-1} ≈ **−239.3 kJ mol^{-1}** [1] **(B)** (i) Metals with negative $E°$ (lower than $H^+/H_2 = 0.00$ V): **Fe, Zn and Mg**. [1] (ii) $E°_{cell} = 0.34 - (-2.36) =$ **2.70 V** [1] (iii) $n = 2$. $\Delta_r G° = -2 \times 96500 \times 2.70 = -521100$ J mol^{-1} = **−521.1 kJ mol^{-1}** [1]	1 + 1 + 1
14	**(A)** (i) $t = 80 \times 60 + 25 = 4825$ s. $Q = I \times t = 2$ A × 4825 s = **9650 C** (= 0.1 F) [1] (ii) $Cu^{2+} + 2e^- \rightarrow Cu$: 2 F deposit 63.5 g. Mass = (63.5 × 9650)/(2 × 96500) = **3.175 g** [1] (iii) $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$: 4 F liberate 1 mol O_2. Moles of O_2 = 0.1/4 = 0.025 mol; volume = 0.025 × 22.4 = **0.56 L (560 mL)** [1] **(B)** (i) At the cathode H^+/H_2O ($E° = 0.00$ V) is reduced in preference to Na^+ (−2.71 V): $H_2O(l) + e^- \rightarrow ½H_2(g) + OH^-(aq)$. [1] (ii) Oxidation of water to O_2 is a very slow process because of the high **overpotential** of oxygen, so Cl^- is oxidised instead: $Cl^- \rightarrow ½Cl_2 + e^-$. [1] (iii) $NaCl(aq) + H_2O(l) \rightarrow Na^+(aq) + OH^-(aq) + ½H_2(g) + ½Cl_2(g)$ [1]	1 + 1 + 1
15	**(A)** (i) Conductivity is the conducting power of the ions present in a unit volume of the solution. On dilution the number of ions per unit volume decreases, so κ decreases. [1] (ii) *Strong electrolyte:* it is completely ionised, so on dilution only the interionic attractions decrease and the ions move faster; Λ_m rises slowly ($\Lambda_m = \Lambda°_m - Ac^{1/2}$). *Weak electrolyte:* the degree of dissociation increases greatly on dilution, so the number of ions in the total volume increases steeply and Λ_m rises very rapidly at low concentration. [1 for the strong electrolyte, 1 for the weak electrolyte] **(B)** (i) $E(H^+/H_2) = E° - (0.059/2) \log (p(H_2) / [H^+]^2)$. With $E° = 0$ and $p = 1$ bar this becomes $E = -0.059 \log (1/[H^+]) = -0.059 \times$ pH [1] (ii) $E = -0.059 \times 10$ [1] = **−0.59 V** [1]	1 + 2

SECTION D		
Question No. 16 is a case-based/data-based question carrying 4 marks.		
16	(i) Zn and Mg have more negative electrode potentials than iron (−0.76 V and −2.36 V against −0.44 V), so they are more easily oxidised than iron. They act as the anode and corrode in place of iron (**sacrificial protection**), while the iron acts as the cathode and does not rust. [1] (ii) Tin and copper have higher (less negative or positive) electrode potentials than iron, so iron is the more easily oxidised metal; in contact with them the iron acts as the anode and is oxidised ($Fe \rightarrow Fe^{2+} + 2e^-$), which leads to rust. [1] (iii) **(A)** Anode: $2Fe(s) \rightarrow 2Fe^{2+} + 4e^-$ ($E° = -0.44$ V) [1]. Cathode: $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$ ($E° = 1.23$ V) [1]. (Overall: $2Fe + O_2 + 4H^+ \rightarrow 2Fe^{2+} + 2H_2O$, $E°_{cell} = 1.67$ V.) **(B)** Sacrificial protection: a more reactive metal (such as Mg or Zn) is connected to the iron object; it corrodes itself and saves the iron. [1] Rust is hydrated ferric oxide, **$Fe_2O_3 \cdot xH_2O$**. [1]	1 + 1 + 2

SECTION E		
Question No. 17 is a long answer question carrying 5 marks.		
17	**(A)** (i) $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$ ($n = 2$) [1] (ii) $E°_{cell} = 0.80 - (-0.76) =$ **1.56 V** [1] (iii) $E_{cell} = E° - (0.059/2) \log([Zn^{2+}]/[Ag^+]^2) = 1.56 - 0.0295 \times \log(0.01/(0.01)^2) = 1.56 - 0.0295 \times \log 100 = 1.56 - 0.0295 \times 2 =$ **1.501 V ≈ 1.50 V** [1] (iv) $\Delta_r G° = -nFE° = -2 \times 96500 \times 1.56 = -301080$ J mol^{-1} ≈ **−301.1 kJ mol^{-1}** [1] (v) The EMF **increases**. Increasing $[Ag^+]$ (a reactant) decreases the reaction quotient $[Zn^{2+}]/[Ag^+]^2$, so the log term decreases and E increases (Le Chatelier's principle). [1] **(B)** (i) $\kappa = 0.016$ S m^{-1} = 0.016/100 = 1.6×10^{-4} S cm^{-1}. $\Lambda_m = \kappa \times 1000/c = (1.6 \times 10^{-4} \times 1000)/0.01 =$ **16 S cm^2 mol^{-1}** [1] (ii) $\Lambda°_m = \lambda°(H^+) + \lambda°(CH_3COO^-) = 349.6 + 40.9 =$ **390.5 S cm^2 mol^{-1}** [1] (iii) $\alpha = \Lambda_m/\Lambda°_m = 16/390.5 =$ **0.041** (4.1%) [1] (iv) $K_a = c\alpha^2/(1 - \alpha) \approx c\alpha^2 = 0.01 \times (0.041)^2 = 0.01 \times 1.68 \times 10^{-3} \approx$ **1.7×10^{-5} mol L^{-1}** [1] (v) For a weak electrolyte Λ_m rises very steeply as the concentration approaches zero (the degree of dissociation increases sharply), so the plot does not become a straight line and $\Lambda°_m$ cannot be found by extrapolation; it is obtained using Kohlrausch's law. [1]	5

Topic coverage of the paper

Topic (NCERT section)	Questions
Electrochemical cells, galvanic cells, electrode potential (2.1, 2.2)	Q2 (table), Q12 (A) (figure), Q13 (table)
Nernst equation, EMF, Gibbs energy (2.3)	Q6, Q12 (B), Q13 (table), Q15 (B), Q17 (A)
Conductance of electrolytic solutions; Kohlrausch's law (2.4)	Q1, Q3 (graph), Q8, Q9 (table), Q11, Q15 (A), Q17 (B) (table)
Electrolytic cells, Faraday's laws, products of electrolysis (2.5)	Q4, Q7, Q14
Batteries and fuel cells (2.6, 2.7)	Q5, Q10
Corrosion (2.8)	Q16 (table)

Figure/table-based questions (7): Q2, Q3, Q9, Q12, Q13, Q16 and Q17.