

CLASS 12

CHEMISTRY

NCERT SOLUTIONS

NCERT Solutions for Class 12 Chemistry Chapter 2: Electrochemistry – Free PDF Download

Chapter 2 covers electrochemistry — galvanic cells, standard electrode potentials, the Nernst equation, Gibbs energy and equilibrium constants for redox reactions, conductivity and molar conductivity, Kohlrausch's law, and Faraday's laws of electrolysis. All 18 exercises (2.1–2.18) remain current under the 2026-...

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How to Approach This Chapter

Fix the anode/cathode and positive/negative electrode conventions for galvanic versus electrolytic cells clearly in your notes before attempting numericals — most conceptual errors in this chapter trace back to mixing these up. For Nernst equation and Faraday's law problems, write down every given quantity with correct units first, since most numerical errors come from unit mismatches rather than the formula itself.

Exam Weightage: How Important Is This Chapter?

Electrochemistry carries 9 marks out of the 70-mark theory paper (about 13%) in the CBSE Class 12 Chemistry board exam — one of the highest-weightage individual chapters in the syllabus.

Source basis: CBSE Class 12 Chemistry chapter-wise marks distribution, current marking scheme.

NCERT Exercise Solutions

2.1 Arrange the following metals in the order in which they displace each other from the solution of their salts: Al, Cu, Fe, Mg and Zn.

Ans: Based on standard electrode (reduction) potentials, the more reactive (more negative potential) metal displaces a less reactive one from its salt solution. The order, from most to least reactive: **Mg > Al > Zn > Fe > Cu**. So Mg displaces Al, Zn, Fe and Cu from their salt solutions; Al displaces Zn, Fe and Cu (but not Mg); and so on, with Cu displacing none of the others.

2.2 Given the standard electrode potentials, $K^+/K = -2.93V$, $Ag^+/Ag = 0.80V$, $Hg^{2+}/Hg = 0.79V$, $Mg^{2+}/Mg = -2.37V$, $Cr^{3+}/Cr = -0.74V$. Arrange these metals in their increasing order of reducing power.

Ans: A more negative standard electrode potential means the metal is a stronger reducing agent (more easily oxidised). Ranking the given potentials from most positive to most negative: $Ag(0.80) > Hg(0.79) > Cr(-0.74) > Mg(-2.37) > K(-2.93)$. Reducing power increases in the opposite direction, so the increasing order of reducing power is: **Ag < Hg < Cr < Mg < K**.

2.3 Depict the galvanic cell in which the reaction $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$ takes place. Further show: (i) Which electrode is negatively charged? (ii) The carriers of current in the cell. (iii) Individual reaction at each electrode.

Ans: The cell is represented as $Zn(s)|Zn^{2+}(aq)||Ag^+(aq)|Ag(s)$. (i) The **zinc electrode (anode) is negatively charged**, since it is the site of oxidation and supplies electrons to the external circuit. (ii) Current is carried by **electrons through the external wire** (from the zinc anode to the silver cathode) and by **ions through the salt bridge/electrolyte** (completing the internal circuit). (iii) At the anode (Zn):

$\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$. At the cathode (Ag): $2\text{Ag}^+(\text{aq}) + 2\text{e}^- \rightarrow 2\text{Ag(s)}$.

2.4 Calculate the standard cell potentials of galvanic cells in which the following reactions take place, and calculate the $\Delta_r G^\circ$ and equilibrium constant of the reactions: (i)

$2\text{Cr(s)} + 3\text{Cd}^{2+}(\text{aq}) \rightarrow 2\text{Cr}^{3+}(\text{aq}) + 3\text{Cd(s)}$ (ii) $\text{Fe}^{2+}(\text{aq}) + \text{Ag}^+(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{Ag(s)}$.

Ans: (i) Using $E^\circ(\text{Cd}^{2+}/\text{Cd}) = -0.40\text{V}$ and $E^\circ(\text{Cr}^{3+}/\text{Cr}) = -0.74\text{V}$: $E^\circ_{\text{cell}} = -0.40 - (-0.74) = +0.34\text{V}$. With $n=6$: $\Delta_r G^\circ = -nFE^\circ = -6 \times 96500 \times 0.34 \approx -1.97 \times 10^5\text{J}$, and from $\Delta_r G^\circ = -RT \ln K$, $K \approx 3.2 \times 10^{34}$ — an extremely large K , showing the reaction goes essentially to completion. (ii) Using $E^\circ(\text{Ag}^+/\text{Ag}) = 0.80\text{V}$ and $E^\circ(\text{Fe}^{3+}/\text{Fe}^{2+}) = 0.77\text{V}$: $E^\circ_{\text{cell}} = 0.80 - 0.77 = +0.03\text{V}$. With $n=1$: $\Delta_r G^\circ \approx -2895\text{J}$, giving $K \approx 3.2$ — a much smaller equilibrium constant, consistent with the small cell potential.

2.5 Write the Nernst equation and calculate the emf of the following cells at 298K: (i)

$\text{Mg(s)}|\text{Mg}^{2+}(0.001\text{M})||\text{Cu}^{2+}(0.0001\text{M})|\text{Cu(s)}$ (ii) $\text{Fe(s)}|\text{Fe}^{2+}(0.001\text{M})||\text{H}^+(1\text{M})|\text{H}_2(\text{g})(1\text{bar})|\text{Pt(s)}$ (iii)

$\text{Sn(s)}|\text{Sn}^{2+}(0.050\text{M})||\text{H}^+(0.020\text{M})|\text{H}_2(\text{g})(1\text{bar})|\text{Pt(s)}$ (iv)

$\text{Pt(s)}|\text{Br}^-(0.010\text{M})|\text{Br}_2(\text{l})||\text{H}^+(0.030\text{M})|\text{H}_2(\text{g})(1\text{bar})|\text{Pt(s)}$.

Ans: Using $E = E^\circ_{\text{cell}} - (0.0591/n) \log Q$ at 298K: (i) $E^\circ_{\text{cell}} = 0.34 - (-2.37) = 2.71\text{V}$, $Q = [\text{Mg}^{2+}]/[\text{Cu}^{2+}] = 10$, $E \approx 2.68\text{V}$. (ii) $E^\circ_{\text{cell}} = 0 - (-0.44) = 0.44\text{V}$, $Q = [\text{Fe}^{2+}]/[\text{H}^+]^2 = 0.001$, $E \approx 0.53\text{V}$. (iii) $E^\circ_{\text{cell}} = 0 - (-0.14) = 0.14\text{V}$, $Q = [\text{Sn}^{2+}]/[\text{H}^+]^2 = 125$, $E \approx 0.078\text{V}$. (iv) $E^\circ_{\text{cell}} = 0 - 1.09 = -1.09\text{V}$, $Q = 1/([\text{Br}^-]^2[\text{H}^+]^2) \approx 1.11 \times 10^5$, $E \approx -1.30\text{V}$.

2.6 In the button cells widely used in watches and other devices, the reaction

$\text{Zn(s)} + \text{Ag}_2\text{O(s)} + \text{H}_2\text{O(l)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag(s)} + 2\text{OH}^-(\text{aq})$ takes place. Determine $\Delta_r G^\circ$ and E° for the reaction.

Ans: Using $E^\circ(\text{Ag}_2\text{O}/\text{Ag}, \text{OH}^-) = 0.344\text{V}$ and $E^\circ(\text{Zn}^{2+}/\text{Zn}) = -0.76\text{V}$: $E^\circ_{\text{cell}} = 0.344 - (-0.76) = +1.104\text{V}$. With $n=2$: $\Delta_r G^\circ = -2 \times 96500 \times 1.104 \approx -213.1\text{kJ}$ — the large negative $\Delta_r G^\circ$ and positive E° confirm this reaction is strongly spontaneous, which is why it reliably powers button cells.

2.7 Define conductivity and molar conductivity for the solution of an electrolyte. Discuss their variation with concentration.

Ans: **Conductivity (κ)** is the conductance of a solution of unit length and unit cross-sectional area — the reciprocal of resistivity. **Molar conductivity (Λ_m)** is the conductivity of a solution divided by its molar concentration, $\Lambda_m = 1000\kappa/c$, representing the conducting power of all the ions produced by one mole of electrolyte. As concentration decreases, conductivity (κ) decreases (fewer ions per unit volume), but molar conductivity **increases** for both strong and weak electrolytes — sharply for weak electrolytes (as dissociation increases) and more gradually for strong electrolytes (as inter-ionic attractions weaken), approaching a limiting value Λ°_m at infinite dilution.

2.8 The conductivity of 0.20M solution of KCl at 298K is 0.0248Scm^{-1} . Calculate its molar conductivity.

Ans: $\Lambda_m = 1000 \times \kappa / c = 1000 \times 0.0248 / 0.20 = 124\text{Scm}^2\text{mol}^{-1}$.

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

Ans: Cell constant = $\kappa \times R = 0.146 \times 10^{-3} \times 1500 = 0.219\text{cm}^{-1}$.

2.10 The conductivity of sodium chloride at 298K has been determined at different concentrations, and the results are given below. Calculate Λ_m for all concentrations, and find the value of Λ°_m .

Ans: Using $\Lambda_m = \kappa / c$ (in SI units, κ in Sm^{-1} , c in molm^{-3}): at $c=0.001\text{M}$, $\Lambda_m \approx 123.7 \times 10^{-4}\text{Sm}^2\text{mol}^{-1}$; at 0.010M , $\approx 118.5 \times 10^{-4}$; at 0.020M , $\approx 115.8 \times 10^{-4}$; at 0.050M , $\approx 111.1 \times 10^{-4}$; at 0.100M , $\approx 106.7 \times 10^{-4}\text{Sm}^2\text{mol}^{-1}$ — molar conductivity rises steadily as concentration falls, as expected for a strong electrolyte. Extrapolating the plot of Λ_m against $\sqrt{c}$ to $c \rightarrow 0$ (Debye-Hückel-Onsager behaviour, since NaCl is fully dissociated) gives $\Lambda^\circ_m \approx 124 \times 10^{-4}\text{Sm}^2\text{mol}^{-1}$.

2.11 Conductivity of 0.00241M acetic acid is $7.896 \times 10^{-5} \text{Scm}^{-1}$. Calculate its molar conductivity. If Λ_m° for acetic acid is $390.5 \text{Scm}^2 \text{mol}^{-1}$, what is its dissociation constant?

Ans: $\Lambda_m = 1000 \times 7.896 \times 10^{-5} / 0.00241 \approx 32.76 \text{Scm}^2 \text{mol}^{-1}$. Degree of dissociation $\alpha = \Lambda_m / \Lambda_m^\circ = 32.76 / 390.5 \approx 0.0839$. Using $K_a = c\alpha^2 / (1 - \alpha)$: $K_a \approx 1.85 \times 10^{-5}$.

2.12 How much charge is required for the following reductions: (i) 1mol of Al^{3+} to Al? (ii) 1mol of Cu^{2+} to Cu? (iii) 1mol of MnO_4^- to Mn^{2+} ?

Ans: (i) $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$ needs 3 Faradays: $3 \times 96500 = 2.895 \times 10^5 \text{C}$. (ii) $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ needs 2 Faradays: $2 \times 96500 = 1.93 \times 10^5 \text{C}$. (iii) $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$ needs 5 Faradays: $5 \times 96500 = 4.825 \times 10^5 \text{C}$.

2.13 How much electricity in terms of Faraday is required to produce (i) 20.0g of Ca from molten CaCl_2 ? (ii) 40.0g of Al from molten Al_2O_3 ?

Ans: (i) Moles of Ca = $20.0 / 40.0 = 0.5 \text{mol}$; since $\text{Ca}^{2+} + 2\text{e}^- \rightarrow \text{Ca}$, required electricity = $2 \times 0.5 = 1$ Faraday. (ii) Moles of Al = $40.0 / 27.0 \approx 1.4815 \text{mol}$; since $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$, required electricity = $3 \times 1.4815 \approx 4.44$ Faradays.

2.14 How much electricity is required in coulombs for the oxidation of (i) 1mol of H_2O to O_2 ? (ii) 1mol of FeO to Fe_2O_3 ?

Ans: (i) From $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$, 1mol H_2O releases 2mol electrons: $2 \times 96500 = 1.93 \times 10^5 \text{C}$. (ii) Oxidising Fe^{2+} (in FeO) to Fe^{3+} (in Fe_2O_3) releases 1 electron per Fe atom: $1 \times 96500 = 9.65 \times 10^4 \text{C}$.

2.15 A solution of $\text{Ni}(\text{NO}_3)_2$ is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

Ans: $Q = It = 5 \times (20 \times 60) = 6000 \text{C}$. Moles of electrons = $6000 / 96500 \approx 0.0622 \text{mol}$. Since $\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$, moles of Ni = 0.0311mol . Mass of Ni = $0.0311 \times 58.7 \approx 1.82 \text{g}$.

2.16 Three electrolytic cells A, B, C containing solutions of ZnSO_4 , AgNO_3 , and CuSO_4 respectively are connected in series. A steady current of 1.5 amperes was passed through them until 1.45g of silver deposited at the cathode of cell B. How long did the current flow? What mass of copper and zinc were deposited?

Ans: Moles of Ag deposited = $1.45 / 108 \approx 0.01343 \text{mol}$ = moles of electrons (since $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$). Charge passed = $0.01343 \times 96500 \approx 1295.6 \text{C}$. Time = $1295.6 / 1.5 \approx 864 \text{s}$ (≈ 14.4 minutes). Since the same charge passes through all three series-connected cells: moles of Cu deposited = $0.01343 / 2 \approx 0.00671 \text{mol}$, mass $\approx 0.00671 \times 63.5 \approx 0.426 \text{g}$; moles of Zn deposited = $0.01343 / 2 \approx 0.00671 \text{mol}$, mass $\approx 0.00671 \times 65.4 \approx 0.439 \text{g}$.

2.17 Using the standard electrode potentials given in the standard table, predict if the reaction between the following is feasible: (i) $\text{Fe}^{3+}(\text{aq})$ and $\text{I}^-(\text{aq})$ (ii) $\text{Ag}^+(\text{aq})$ and $\text{Cu}(\text{s})$ (iii) $\text{Fe}^{3+}(\text{aq})$ and $\text{Br}^-(\text{aq})$ (iv) $\text{Ag}(\text{s})$ and $\text{Fe}^{3+}(\text{aq})$ (v) $\text{Br}_2(\text{aq})$ and $\text{Fe}^{2+}(\text{aq})$.

Ans: Using $E^\circ(\text{Fe}^{3+}/\text{Fe}^{2+}) = 0.77 \text{V}$, $E^\circ(\text{I}_2/\text{I}^-) = 0.54 \text{V}$, $E^\circ(\text{Ag}^+/\text{Ag}) = 0.80 \text{V}$, $E^\circ(\text{Cu}^{2+}/\text{Cu}) = 0.34 \text{V}$, $E^\circ(\text{Br}_2/\text{Br}^-) = 1.09 \text{V}$ — a reaction is feasible only if $E^\circ_{\text{cell}} (= E^\circ_{\text{reduced species}} - E^\circ_{\text{oxidised species}}) > 0$. (i) $E^\circ_{\text{cell}} = 0.77 - 0.54 = +0.23 \text{V}$: **feasible** (Fe^{3+} oxidises I^- to I_2). (ii) $E^\circ_{\text{cell}} = 0.80 - 0.34 = +0.46 \text{V}$: **feasible** (Ag^+ oxidises Cu). (iii) $E^\circ_{\text{cell}} = 0.77 - 1.09 = -0.32 \text{V}$: **not feasible**. (iv) $E^\circ_{\text{cell}} = 0.77 - 0.80 = -0.03 \text{V}$: **not feasible**. (v) $E^\circ_{\text{cell}} = 1.09 - 0.77 = +0.32 \text{V}$: **feasible** (Br_2 oxidises Fe^{2+} to Fe^{3+}).

2.18 Predict the products of electrolysis in each of the following: (i) An aqueous solution of AgNO_3 with silver electrodes. (ii) An aqueous solution of AgNO_3 with platinum electrodes. (iii) A dilute solution of H_2SO_4 with platinum electrodes. (iv) An aqueous solution of CuCl_2 with platinum electrodes.

Ans: (i) With **silver electrodes**, the anode dissolves ($\text{Ag} \rightarrow \text{Ag}^+ + \text{e}^-$) while pure silver deposits at the cathode ($\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$) — the solution's concentration stays essentially unchanged; this is the basis of electrorefining. (ii) With inert **platinum electrodes**, Ag^+ is still reduced at the cathode (Ag deposits), but since Pt doesn't dissolve, water is oxidised at the anode, releasing **O_2 gas**. (iii) For **dilute H_2SO_4** , this

amounts to electrolysis of water: **H₂ gas at the cathode** and **O₂ gas at the anode**. (iv) For **CuCl₂** with Pt electrodes: **Cu deposits at the cathode** ($\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$) and **Cl₂ gas is released at the anode** ($2\text{Cl}^- \rightarrow \text{Cl}_2 + 2\text{e}^-$).

Class 12 Chemistry Chapter 2 – Notes and Extra Questions

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