ANSWER KEY & MARKING SCHEME · CBSE CLASS 12

Electrochemistry

Chemistry · Chapter 2 · Use this with the Board Paper · Companion to Quick Drill

HOW TO USE

Attempt the Board Paper first (closed-book, full time). Then come here. For 2-mark+ questions, compare your answer to the model. For 3-4 mark questions, also consult the **Topper Templates** below — these show the exact step-by-step structure that scores full marks per CBSE marking-scheme conventions.

MODEL ANSWERS · BOARD PAPER

Section A — MCQ (1 mark each, 6 Qs)

Q1. Standard hydrogen electrode E° = ____ V (by convention). ***[1 mark]***

Ans: 0.00 V (international convention; reference for all E° values).

Q2. Faraday constant F = ____ C/mol. ***[1 mark]***

Ans: 96500 (precise: 96485). Charge per mole of electrons.

Q3. Nernst equation at 298K: E = E° – ____ × log Q. ***[1 mark]***

Ans: (0.0591/n).

Q4. Anode in a galvanic cell is the site of ____ . ***[1 mark]***

Ans: Oxidation (loss of electrons). Negative terminal in galvanic.

Q5. Molar conductivity Λ_m with concentration: ____ (increases / decreases). ***[1 mark]***

Ans: Decreases (counter-intuitive but true; max at infinite dilution).

Q6. Kohlrausch's law: Λ°_m = ____ + ____ (in symbols). ***[1 mark]***

Ans: $\nu_+\lambda°_+ + \nu_-\lambda°_-$.

Section B — Very Short Answer (2 marks each, 4 Qs)

Q7. Distinguish between galvanic and electrolytic cells. ***[2 marks]***

Ans: GALVANIC: spontaneous redox reaction (ΔG < 0, E° > 0); PRODUCES electrical energy. Example: battery, Daniell cell. Anode = NEGATIVE. ELECTROLYTIC: external source drives non-spontaneous reaction (ΔG > 0); CONSUMES electrical energy. Example: electrolysis, electroplating, chlor-alkali. Anode = POSITIVE (connected to + terminal of external source).

Q8. Write Nernst equation for the cell Zn(s) | $Zn^{2+}(c_1)$ || $Cu^{2+}(c_2)$ | Cu(s) at 298K. ***[2 marks]***

Ans: Cell reaction: $Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$ (n = 2). Q = $[Zn^{2+}]/[Cu^{2+}] = c_1/c_2$. E_cell = E°_cell – (0.0591/2) $\log(c_1/c_2)$. E°_cell = +1.10V.

Q9. How many coulombs are needed to deposit 1 mole of aluminium ($Al^{3+} + 3e^- \rightarrow Al$)? ***[2 marks]***

Ans: $Al^{3+} + 3e^- \rightarrow Al$. Need 3 moles of electrons. Q = 3 × 96500 = 289,500 C.

Q10. What is the cell notation for a galvanic cell with Zn as anode and Cu as cathode in 1M solutions? ***[2 marks]***

Ans: Zn(s) | Zn^{2+}(1M) || Cu^{2+}(1M) | Cu(s). Single | separates phases; double || represents salt bridge. Anode on left, cathode on right (convention).

Section C — Short Answer (3 marks each, 3 Qs)

Q11. Derive the Nernst equation at 298K for a half-cell reaction. ***[3 marks]***

Ans: Step 1 — Thermodynamic: ΔG = ΔG° + RT ln Q. For electrochemical reactions: ΔG = −nFE, ΔG° = −nFE°. Step 2 — Substitute: −nFE = −nFE° + RT ln Q. Step 3 — Divide by −nF: E = E° – (RT/nF) ln Q. Step 4 — At T = 298 K, (RT·ln10/F) = 0.0591V. So at 298K: E = E° – (0.0591/n) log Q. This is the Nernst equation in its working form.

Q12. Given E°_Ni^{2+}/Ni = −0.25V and E°_Cu^{2+}/Cu = +0.34V, identify cathode + anode + compute E°_cell + write cell notation. ***[3 marks]***

Ans: Higher E° = cathode. Cu (E° = +0.34V) is CATHODE; Ni (E° = −0.25V) is ANODE. E°_cell = E°_cathode – E°_anode = 0.34 – (−0.25) = +0.59V. Positive E° confirms spontaneous galvanic direction. Cell notation: Ni(s) | Ni^{2+}(1M) || Cu^{2+}(1M) | Cu(s).

Q13. Explain how Kohlrausch's law can be used to find Λ°_m of weak electrolyte CH_3COOH from data on strong electrolytes. ***[3 marks]***

> **Ans:** Direct measurement of Λ°_m for weak electrolyte is impossible (they don't fully dissociate even at high dilution). By Kohlrausch's law, ions contribute independently at infinite dilution. So Λ°_m(CH_3COOH) can be derived from strong-electrolyte data: Λ°_m(CH_3COOH) = Λ°_m(CH_3COONa) + Λ°_m(HCl) − Λ°_m(NaCl). The ions H^+ and CH_3COO^- are obtained by combining the three: CH_3COO^- (from CH_3COONa) + H^+ (from HCl) − Na^+/Cl^- pair (from NaCl).

Section D — Long Answer (4 marks each, 2 Qs)

Q14. (a) State Faraday's two laws of electrolysis. (b) Calculate the mass of copper deposited when 0.5 A current is passed through a $CuSO_4$ solution for 1 hour. (Atomic mass of Cu = 63.5 g/mol) ***[4 marks]***

> **Ans:** (a) FIRST LAW: Mass deposited at an electrode is proportional to the quantity of electricity passed. m = Z × Q = Z × I × t, where Z = equivalent weight / F. SECOND LAW: When the same quantity of electricity passes through different electrolytes, masses deposited are proportional to their equivalent weights: $m_1/m_2 = E_1/E_2$. (b) Q = I × t = 0.5 × 3600 = 1800 C. $Cu^{2+} + 2e^- \rightarrow Cu$, so equivalent weight = 63.5/2 = 31.75 g/eq. m = (Q × eq.wt)/F = (1800 × 31.75)/96500 = 57150/96500 = 0.592 g of Cu. Verify: moles of e^- = 1800/96500 = 0.0187; moles of Cu = 0.0187/2 = 0.00933; mass = 0.00933 × 63.5 = 0.592 g ✓

Q15. Case study: Corrosion of iron causes losses of nearly 2-3% of global GDP annually. (a) Describe the electrochemical mechanism of rusting. (b) State THREE methods for preventing corrosion, with the principle behind each. ***[3 marks]***

> **Ans:** (a) Rust mechanism: iron rusts via electrochemical process. Iron acts as ANODE: $Fe \rightarrow Fe^{2+} + 2e^-$. Oxygen at adjacent CATHODE sites is reduced: $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$. Fe^{2+} ions are further oxidised by atmospheric O_2 and combine with water to form hydrated iron(III) oxide: $Fe_2O_3 \cdot xH_2O$ = RUST. Three requirements: O_2, H_2O, electrolyte (CO_2 dissolved in water creates carbonic acid). (b) PREVENTION: (1) BARRIER coating — paint, oil, plastic, or tin-plating physically excludes O_2 and H_2O. Mechanism: physical isolation. (2) GALVANISATION — coat iron with zinc. Mechanism: Zn (E° = −0.76V) is more reactive than Fe (E° = −0.44V), so Zn oxidises preferentially, sacrificing itself. Protects even if scratched. (3) CATHODIC PROTECTION — connect a more reactive metal (Mg, Zn block) to the iron structure. Mechanism: the connected metal oxidises instead of the iron. Used in underground pipelines, ship hulls.

★ TOPPER ANSWER TEMPLATES

3 TEMPLATES · MEMORISE THE FORMAT

★ TOPPER TEMPLATE — 5-mark question: Derive Nernst equation for a half-cell + apply at non-standard concentrations

Annual

Step 1 ***[1 mark]***	**Start from thermodynamic relationship**	ΔG = ΔG° + RT ln Q (general result for any reaction at non-standard state). For electrochemical reaction, the work done is electrical: ΔG = −nFE and ΔG° = −nFE°. Substitute: −nFE = −nFE° + RT ln Q.
Step 2 ***[1 mark]***	**Solve for E**	Divide both sides by −nF: E = E° − (RT/nF) ln Q. This is the NERNST EQUATION in its general form. T = absolute temperature in Kelvin; F = Faraday constant = 96485 C/mol; R = 8.314 J/(mol·K); n = number of electrons transferred.
Step 3 ***[1 mark]***	**Substitute room-temperature constants**	At T = 298 K (25°C), the constants combine: RT/F × ln(10) = (8.314 × 298 / 96485) × 2.303 ≈ 0.0591 V. Substituting: E = E° − (0.0591/n) log Q. This is the room-temperature working form examiners expect in numerical questions.
Step 4 ***[1 mark]***	**Apply to a specific cell**	For $Cu^{2+} + 2e^- \rightarrow Cu$ at 0.1 M [Cu^{2+}]: Q = 1/[Cu^{2+}] = 1/0.1 = 10. E = 0.34 − (0.0591/2) × log(10) = 0.34 − 0.02955 ≈ 0.31 V. So reducing [Cu^{2+}] from 1M to 0.1M lowers the half-cell potential by ~30mV.
Step 5 ***[1 mark]***	**Application — cell EMF at non-standard state**	For a complete cell, E_cell = E°_cell − (0.0591/n) log Q where Q is the cell reaction quotient. EMF DROPS as the reaction proceeds (products build, reactants deplete, Q increases). When E_cell = 0, the cell has reached equilibrium — no more useful work. This is the basis of all battery design + Nernst-equation calculations CBSE tests.

COMMON LOSS OF MARKS:

- Skipping the ΔG → −nFE substitution (−1 mark).
- Using log instead of ln without the 2.303 factor (−1 mark).
- Missing the (0.0591/n) shortcut at 298K (−0.5 mark).

★ TOPPER TEMPLATE — 4-mark question: Galvanic cell EMF computation from E° values + cell notation

90% of years

Step 1 *[1 mark]*	**Identify cathode (higher E°) and anode (lower E°)**	Given two half-cells, both written as REDUCTION potentials. The half-cell with HIGHER E° (more positive reduction potential) acts as CATHODE — gets reduced. The LOWER E° half-cell acts as ANODE — gets oxidised. Example: Zn^{2+}/Zn (E° = −0.76V) and Cu^{2+}/Cu (E° = +0.34V). Cu is cathode; Zn is anode.
Step 2 *[1 mark]*	**Compute cell EMF**	E°_cell = E°_cathode − E°_anode (both as REDUCTION potentials). Example: E°_cell = 0.34 − (−0.76) = +1.10 V. Positive E°_cell confirms the spontaneous direction. If E°_cell is negative, you've got the cathode/anode wrong — swap them.
Step 3 *[1 mark]*	**Write cell notation**	Convention: ANODE | anode-ion (concentration) || cathode-ion (concentration) | CATHODE. The | separates phases; the || represents the salt bridge. Example: Zn(s) | Zn^{2+}(1M) || Cu^{2+}(1M) | Cu(s). The double-bar emphasises that the salt bridge eliminates liquid junction potential.
Step 4 *[1 mark]*	**Identify electron flow and ion migration**	Electrons flow from anode (Zn) to cathode (Cu) in the external circuit. In the salt bridge, cations migrate toward the cathode half-cell; anions migrate toward the anode half-cell — completing the circuit and maintaining electrical neutrality. CBSE rewards explicit mention of both directions in 5-mark long-answers.

COMMON LOSS OF MARKS:

- Reversing cathode and anode (−2 marks).
- Adding E° values instead of subtracting cathode minus anode (−1 mark).
- Wrong cell notation order (−0.5 mark).

★ TOPPER TEMPLATE — 4-mark numerical: Faraday's laws — mass deposited / moles of gas evolved

80% of years

Step 1 *[1 mark]*	**Compute total charge Q passed**	Q = I × t (current in amperes × time in seconds). Example: I = 2 A, t = 30 min = 1800 s. Q = 2 × 1800 = 3600 C.
Step 2 *[1 mark]*	**Compute moles of electrons (or use Faraday's law)**	Moles of electrons = Q/F where F = 96500 C/mol. Example: 3600/96500 = 0.0373 mol of electrons. Alternative direct route via Faraday's first law: m = (Z × I × t) where Z = equivalent weight / 96500 (electrochemical equivalent).
Step 3 *[1 mark]*	**Apply to half-reaction to find product**	Identify half-reaction. Example: $Cu^{2+} + 2e^- \rightarrow Cu$ means 2 moles of electrons produce 1 mole of Cu. So 0.0373 mol electrons → 0.0373/2 = 0.0187 mol Cu = 0.0187 × 63.5 g = 1.18 g Cu deposited at the cathode.
Step 4 *[1 mark]*	**Quote final answer with units + verify**	1.18 g of Cu deposited. Verify: Cu equivalent weight = 63.5/2 = 31.75 g/eq. Mass = (3600/96500) × 31.75 = 1.18 g ✓. Always quote units (g for mass, L or mol for gas) — CBSE deducts 0.5 mark for missing units.

COMMON LOSS OF MARKS:

- Wrong n in moles-of-electrons step (−2 marks).
- Forgetting the half-reaction stoichiometry (−1 mark).
- Missing units (−0.5).

MARKING SCHEME — GENERAL NOTES

- Numerical questions: step marks for Q computation, moles of electrons, half-reaction stoichiometry, final mass with units.
- Derivation questions: ΔG → −nFE substitution + algebraic step + room-temperature shortcut.
- Cell EMF questions: must identify cathode/anode correctly; subtracting wrong direction loses 2 marks.
- Corrosion questions: must mention all THREE requirements (O_2 + H_2O + electrolyte) for full marks.