Class 12 Chemistry – Chapter 2: Electrochemistry

Revision notes as per the NCERT 2026-27 chapter plan (2.1 to 2.8) | For CBSE boards, JEE Main and NEET

2.1 – 2.2 Electrochemical cells and galvanic cells

Point	Galvanic (voltaic) cell	Electrolytic cell
Energy change	Chemical energy → electrical energy	Electrical energy → chemical energy
Reaction	Spontaneous redox reaction	Non-spontaneous; needs an external source
Anode (oxidation)	Negative electrode	Positive electrode
Cathode (reduction)	Positive electrode	Negative electrode
Electron flow	From anode to cathode through the external circuit (in both cells)	From anode to cathode through the external circuit (in both cells)

- **Daniell cell:** Zn(s) | Zn^{2+}(aq) || Cu^{2+}(aq) | Cu(s); $E°_{cell}$ = 1.10 V. Anode: Zn → Zn^{2+} + $2e^-$; cathode: Cu^{2+} + $2e^-$ → Cu.
- **Cell notation:** anode on the left, cathode on the right; a single line (|) marks a phase boundary and a double line (||) marks the salt bridge.
- **Salt bridge** (inert electrolyte such as KCl or KNO_3): completes the inner circuit and keeps both half-cells electrically neutral.
- **$E°_{cell} = E°_{cathode} - E°_{anode}$** (always use reduction potentials).
- **SHE:** Pt(s) | H_2(g, 1 bar) | H^+(aq, 1 M); E° = 0.00 V by convention.
- **Electrochemical series:** the more positive the E°, the stronger the oxidising agent (F_2/F^- = +2.87 V); the more negative the E°, the stronger the reducing agent (Li^+/Li = −3.05 V).

2.3 Nernst equation

- For aA + bB → cC + dD: **$E_{cell} = E°_{cell} - (RT/nF)$ ln Q**. At 298 K: **$E_{cell} = E°_{cell} - (0.059/n)$ log Q**, where Q = $[C]^c[D]^d / [A]^a[B]^b$.
- Half-cell: $E(M^{n+}/M) = E° + (0.059/n) \log [M^{n+}]$ at 298 K.
- At equilibrium E_{cell} = 0 and Q = K, so **$E°_{cell} = (0.059/n) \log K$** at 298 K.
- **$\Delta G° = -nFE°_{cell}$**, $\Delta G = -nFE_{cell}$ and $\Delta G° = -2.303$ RT log K. F = 96487 C mol^{-1} ≈ 96500 C mol^{-1}.
- In a concentration cell both electrodes are the same, so $E°_{cell}$ = 0 and the EMF comes only from the concentration difference.

2.4 Conductance of electrolytic solutions

Quantity	Formula / meaning	Unit
Conductance (G)	G = 1/R	S (siemens)
Conductivity (κ)	κ = 1/ρ = G × (l/A) = G × G*, where G* is the cell constant	S cm^{-1} or S m^{-1}
Molar conductivity (Λ_m)	Λ_m = κ/c; with κ in S cm^{-1} and M in mol L^{-1}, Λ_m = (κ × 1000)/M	S cm^2 mol^{-1}
Dilution effect	κ decreases and Λ_m increases on dilution	—

- **Strong electrolytes:** $\Lambda_m = \Lambda°_m - A\sqrt{c}$ (Debye–Hückel–Onsager); $\Lambda°_m$ is found by extrapolating to c = 0.
- **Weak electrolytes:** Λ_m rises steeply at low concentration because the degree of dissociation increases, so $\Lambda°_m$ cannot be found by extrapolation.
- **Kohlrausch law:** $\Lambda°_m = \nu_+\lambda°_+ + \nu_-\lambda°_-$. Example: $\Lambda°_m(CH_3COOH) = \Lambda°_m(CH_3COONa) + \Lambda°_m(HCl) - \Lambda°_m(NaCl)$.
- Degree of dissociation: **$\alpha = \Lambda_m/\Lambda°_m$**, and $K_a = c\alpha^2/(1 - \alpha)$.

2.5 Electrolytic cells and electrolysis

- **Faraday's first law:** mass deposited w ∝ charge Q, so w = Z I t (Z = electrochemical equivalent).

- **Faraday's second law:** for the same charge, masses deposited are proportional to equivalent weights (M/n).
- **w = (M × I × t) / (n × F)**; 1 faraday (96500 C) is the charge on 1 mol of electrons.
- **Worked example:** 2 A for 965 s through $CuSO_4$ solution (M = 63.5, n = 2): w = (63.5 × 2 × 965)/(2 × 96500) = 0.635 g of copper.

Electrolyte and electrodes	At cathode	At anode
Molten NaCl (inert electrodes)	Na	Cl_2
Aqueous NaCl (inert electrodes)	H_2 (water is reduced more easily than Na^+)	Cl_2 (O_2 formation needs an overpotential)
Dilute H_2SO_4 (Pt electrodes)	H_2	O_2
Aqueous $CuSO_4$ (Pt electrodes)	Cu deposits	O_2
Aqueous $CuSO_4$ (Cu electrodes)	Cu deposits	Cu dissolves
Aqueous $AgNO_3$ (Ag electrodes)	Ag deposits	Ag dissolves

2.6 Batteries

Battery	Anode / cathode / electrolyte	Key reactions and facts
Dry cell (primary)	Zn container / graphite rod surrounded by MnO_2 and carbon / paste of NH_4Cl and $ZnCl_2$	Anode: $Zn \rightarrow Zn^{2+} + 2e^-$. About 1.5 V. Cannot be recharged.
Mercury cell (primary)	Zn–Hg amalgam / HgO and carbon / moist paste of KOH and ZnO	Overall: Zn(Hg) + HgO → ZnO + Hg. Constant potential of about 1.35 V; used in watches and hearing aids.
Lead storage battery (secondary)	Pb / PbO_2 on a lead grid / about 38% H_2SO_4	Overall: $Pb + PbO_2 + 2H_2SO_4 \rightarrow 2PbSO_4 + 2H_2O$. About 2 V per cell. H_2SO_4 is consumed on discharge and regenerated on charging.
Nickel–cadmium cell (secondary)	Cd / $Ni(OH)_3$ / KOH	Overall: $Cd + 2Ni(OH)_3 \rightarrow CdO + 2Ni(OH)_2 + H_2O$. Longer life than the lead storage battery, but costlier.

2.7 Fuel cells

- Fuel cells convert the energy of combustion of fuels such as H_2 directly into electrical energy; reactants are supplied continuously from outside.
- **H_2–O_2 fuel cell** (KOH electrolyte, porous carbon electrodes with a catalyst): anode: $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$; cathode: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$; overall: $2H_2 + O_2 \rightarrow 2H_2O$.
- Advantages: no pollution from the cell reaction and higher efficiency than thermal power generation.

2.8 Corrosion

- Corrosion is an electrochemical process. At one spot (anode): $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$, $E° = -0.44$ V. At another spot (cathode): $O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$, $E° = +1.23$ V.
- Fe^{2+} is then oxidised further: $2Fe^{2+}(aq) + 2H_2O(l) + \frac{1}{2}O_2(g) \rightarrow Fe_2O_3(s) + 4H^+(aq)$. Rust is hydrated $Fe_2O_3 \cdot xH_2O$.
- **Prevention:** paint or oil coating, galvanisation (zinc coating), electroplating, and cathodic protection with a sacrificial anode such as Mg or Zn.

Quick formula sheet

Topic	Formula
EMF of a cell	$E°_{cell} = E°_{cathode} - E°_{anode}$
Nernst equation (298 K)	$E = E° - (0.059/n) \log Q$
Free energy and equilibrium	$\Delta G° = -nFE° = -2.303\ RT \log K$; $E° = (0.059/n) \log K$
Molar conductivity	$\Lambda_m = (\kappa \times 1000)/M$ (S cm^2 mol^{-1})

Topic	Formula
Kohlrausch law and degree of dissociation	$\Lambda^\circ_m = \nu_+\lambda^\circ_+ + \nu_-\lambda^\circ_-$; $\alpha = \Lambda_m/\Lambda^\circ_m$
Faraday’s law	$w = (M\ I\ t)/(n\ F)$; $F = 96500\ C\ mol^{-1}$

Common mistakes to avoid

- Using oxidation potentials in $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$; always use reduction potentials.
- Taking the wrong value of n; it must match the number of electrons in the balanced equation.
- Mixing units of κ and Λ_m (S cm^{-1} versus S m^{-1}) in conductance numericals.
- Extrapolating the Λ_m versus $\sqrt{c}$ graph to find Λ°_m of a weak electrolyte; use Kohlrausch law for weak electrolytes instead.
- Ignoring water at the electrodes when solving electrolysis of aqueous solutions.

Always check these notes against your latest NCERT textbook and the official CBSE curriculum for the current session, as some portions of the syllabus are revised from time to time.