Electrochemistry Class 12 Formula Sheet

CBSE Class 12 Chemistry · NCERT · for boards, NEET and JEE revision

ChapterNotes

Conductance & Conductivity

$$G = 1/R \cdot \kappa = G \times (l/A) \cdot \text{cell constant } G^* = l/A$$

G in siemens (S); κ in S cm^{-1}; cell constant l/A in cm^{-1}.

- Conductance G is just 1/resistance; ions (not electrons) carry the current.
- Conductivity κ = conductance of a 1 cm^3 cube of solution.
- Cell constant l/A standardises the measurement to the liquid itself.

Molar Conductivity & Dilution

$$\Lambda_m = (\kappa \times 1000) / c \ (S\ cm^2\ mol^{-1})$$

κ in S cm^{-1}, c in mol L^{-1}; the ×1000 converts L → cm^3.

- On dilution κ DECREASES (fewer ions per cm^3) but Λ_m INCREASES.
- Strong electrolytes: $\Lambda_m = \Lambda^\circ_m - A\sqrt{c}$, extrapolate to c=0 for Λ°_m.
- Weak electrolytes shoot up steeply near infinite dilution → can't extrapolate.

Kohlrausch's Law

$$\Lambda^\circ_m = \nu_+ \lambda^\circ_+ + \nu_- \lambda^\circ_-$$

ν are stoichiometric ion counts - e.g. $CaCl_2$: $\Lambda^\circ_m = \lambda^\circ(Ca^{2+}) + 2\lambda^\circ(Cl^-)$.

- Finds Λ°_m of weak acids: $\Lambda^\circ(CH_3COOH) = \Lambda^\circ(CH_3COONa) + \Lambda^\circ(HCl) - \Lambda^\circ(NaCl)$.
- Never drop the stoichiometric multiplier (the 2 for two Cl^-).
- H^+ and OH^- have abnormally high λ° via Grotthuss proton-hopping.

Degree of Dissociation & Ka

$$\alpha = \Lambda_m / \Lambda^\circ_m \cdot K_a = c\,\alpha^2 / (1 - \alpha)$$

Measured Λ_m goes on top; for tiny α, $K_a \approx c\alpha^2$.

- α is the fraction of the weak electrolyte that has dissociated.
- Example: Λ_m = 39.1, Λ°_m = 391 → α = 0.1 (10% dissociated).
- Feed α into $K_a = c\alpha^2/(1-\alpha)$ for the dissociation constant.

Galvanic Cells & EMF

$$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode} \text{ (both as reduction potentials)}$$

Anode = oxidation (–); cathode = reduction (+). SHE is defined as 0 V.

- Notation: Zn | Zn^{2+} || Cu^{2+} | Cu; || is the salt bridge keeping neutrality.
- Daniell cell: 0.34 – (–0.76) = 1.10 V, positive → spontaneous.
- E° is intensive - multiplying the equation does NOT change E°.

Electrochemical Series

$$F_2\ (+2.87\ V) \text{ strongest oxidiser} \cdot Li^+/Li\ (-3.05\ V) \text{ strongest reducer}$$

Higher E° = stronger oxidising agent → that species is reduced (cathode).

- More positive E° → species grabs electrons → becomes the cathode.
- Most negative E° → gives electrons most easily → strongest reducing agent.
- Series predicts feasibility, displacement and which ion discharges first.

Nernst Equation

$$E_{cell} = E^\circ_{cell} - (0.0591/n) \log Q$$

n = electrons from the FULL balanced reaction; Q = [products]/[reactants].

- EMF rises with more reactant ion, falls as products build up.
- 0.0591 = 2.303RT/F at 298 K only - invalid at other temperatures.
- Single electrode: $E = E^\circ - (0.0591/n) \log(1/[M^{n+}])$.

ΔG°, K and EMF

$$\Delta G^\circ = -nFE^\circ_{cell} \cdot E^\circ_{cell} = (0.0591/n) \log K_c$$

F = 96500 C mol^{-1}; at equilibrium $E_{cell} = 0$ and Q = K_c.

- Positive E°_{cell} ⇒ negative ΔG° ⇒ spontaneous ⇒ K > 1.
- At equilibrium the cell is dead: $E_{cell} = 0$, Q = K_c.
- Lets you compute K, ΔG° or Ksp straight from a voltage.

Electrolysis & Faraday's Laws

$$w = M\,I\,t / (n\,F) \cdot 1\ F = 96500\ C = 1\ mol\ e^-$$

In electrolysis cathode is NEGATIVE, anode positive - but anode is still oxidation. t in seconds.

- First law: w ∝ charge, Q = I t; moles of e^- = It/F.
- Second law: same charge → masses ∝ equivalent masses (M/n).
- Selective discharge: higher reduction potential discharges first; overpotential can flip it (brine → Cl_2).

Batteries, Fuel Cells & Corrosion

$$H_2\text{–}O_2 \text{ fuel cell: } 2H_2 + O_2 \rightarrow 2H_2O \cdot \eta = \Delta G/\Delta H$$

Primary = single-use; secondary = rechargeable (lead storage, Ni-Cd).

- Fuel cell is fed reactant continuously; η = ΔG/ΔH can approach 100%.
- Rusting is electrochemical: Fe → Fe^{2+} at anode, O_2 reduced at cathode → $Fe_2O_3 \cdot xH_2O$.
- Sacrificial protection: a more negative-E° metal (Zn –0.76 V, Mg) corrodes instead of Fe (–0.44 V).

Quick revision

- Anode = Oxidation, Cathode = Reduction (in ALL cells)
- $E^0_{cell} = E^0_{cathode} - E^0_{anode}$; positive → spontaneous
- Nernst equation at 298K: $E = E^0 - (0.0591/n) \log Q$
- At equilibrium: E = 0, $E^0 = (0.0591/n) \log K$
- $\Delta G^0 = -nFE^0$ (link between thermodynamics and electrochemistry)
- κ decreases with dilution; Λ_m increases with dilution
- Kohlrausch: $\Lambda^0_m = \nu_+\lambda^0_+ + \nu_-\lambda^0_-$
- Faraday: m = MIt/nF
- Lead-acid: rechargeable, 6 cells × 2V = 12V car battery
- Fuel cells: high efficiency (~70%), clean energy, produce only water
- Corrosion = electrochemical; prevent by galvanisation, cathodic protection

Full chapter notes with solved examples and practice questions: https://chapternotes.in/cbse/class-12-chemistry/electrochemistry-class-12-notes/