CLASS 12 CHEMISTRY — FORMULA SHEET

Chapter 2 : Electrochemistry (NCERT)

2.1 Electrochemical Cells & Electrode Potential

Concept	Detail
Galvanic (voltaic) cell	Chemical energy → electrical energy (spontaneous)
Electrolytic cell	Electrical energy → chemical reaction (non-spontaneous)
Anode	Oxidation occurs ; negative in galvanic, positive in electrolytic
Cathode	Reduction occurs ; positive in galvanic, negative in electrolytic
Cell notation	Anode | anode soln || cathode soln | cathode
Standard hydrogen electrode (SHE)	Reference electrode ; $E^\circ = 0$ V
EMF of cell	$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ (both as reduction potentials)

2.2 Nernst Equation

Concept	Formula
Nernst equation (electrode)	$E = E^\circ - (RT/nF) \ln Q$
At 298 K (simplified)	$E = E^\circ - (0.0591/n) \log Q$
For a cell	$E_{cell} = E^\circ_{cell} - (0.0591/n) \log Q$
At equilibrium	$E_{cell} = 0$; $E^\circ_{cell} = (0.0591/n) \log K_c$
Relation E° and K	$\log K_c = (n E^\circ_{cell}) \div 0.0591$
Gibbs energy & EMF	$\Delta G^\circ = -n F E^\circ_{cell}$
F (Faraday constant)	F = 96500 C/mol

2.3 Conductance & Conductivity

Quantity	Formula
Resistance	$R = \rho (l/A)$; unit ohm (Ω)
Conductance	$G = 1/R$; unit siemens (S)
Conductivity (κ)	$\kappa = 1/\rho = G \times (l/A)$; unit S cm^{-1}
Cell constant	$G^* = l/A$; $\kappa = G \times$ cell constant
Molar conductivity	$\Lambda_m = (\kappa \times 1000) \div C$; unit S cm^2 mol^{-1}
Effect of dilution	κ decreases ; Λ_m increases on dilution

2.4 Kohlrausch's Law

Concept	Formula
Limiting molar conductivity	Λ°_m (at infinite dilution)
Kohlrausch's law	$\Lambda^\circ_m = \nu_+ \lambda^\circ_+ + \nu_- \lambda^\circ_-$

Concept	Formula
Degree of dissociation	$\alpha = \Lambda_m \div \Lambda^\circ_m$
Dissociation constant (weak)	$K_a = (C\,\alpha^2) \div (1 - \alpha)$
Application	Find Λ°_m for weak electrolytes (e.g. CH_3COOH)
Strong vs weak	Strong: Λ_m rises slightly on dilution ; Weak: rises sharply

2.5 Faraday's Laws of Electrolysis

Concept	Formula
First law	Mass deposited $m = Z \times Q = Z \times I \times t$
Electrochemical equivalent (Z)	$Z = M \div (n\,F)$
Charge	$Q = I \times t$ (current × time)
Moles of electrons	$= Q \div F = (I\,t) \div 96500$
Second law	Masses deposited ∝ equivalent masses (same charge)
Mass deposited	$m = (M \times I \times t) \div (n \times 96500)$

2.6 Batteries & Corrosion

Concept	Detail
Primary cell	Non-rechargeable (e.g. dry cell, mercury cell)
Secondary cell	Rechargeable (e.g. lead storage battery, Ni-Cd)
Lead storage battery	Anode: Pb ; Cathode: PbO_2 ; electrolyte: H_2SO_4
Fuel cell	Converts fuel energy directly (e.g. H_2–O_2 fuel cell)
Corrosion	Oxidation of metal (e.g. rusting of iron)
Rusting product	Hydrated iron(III) oxide, $Fe_2O_3 \cdot xH_2O$
Prevention	Galvanisation, painting, cathodic protection, alloying

★ QUICK MEMORY AIDS & TRICKS

Anode vs cathode	Anode = oxidation, Cathode = reduction (always — 'An Ox, Red Cat').
Cell EMF	E°cell = E°cathode − E°anode (both as reduction potentials).
Nernst (298 K)	E = E° − (0.0591/n) log Q. Concentration affects potential.
ΔG & EMF	ΔG° = −nFE°. Positive E° → negative ΔG° → spontaneous.
Molar conductivity	Λm = κ×1000/C. Increases on dilution (more ionisation).
Kohlrausch	Λ°m = sum of ionic contributions ; α = Λm/Λ°m.
Faraday	m = MIt/(nF) ; F = 96500 C/mol = charge of 1 mole electrons.
Lead battery	Pb anode, PbO_2 cathode, H_2SO_4 electrolyte — rechargeable.