Collegedunia NCERT Formula Sheet

The Ultimate Formula Reference for Class 12 (12th) Chemistry — NCERT 2026-27

Chapter 2: Electrochemistry

Galvanic & Electrolytic Cells · Nernst Equation · Conductance · Kohlrausch's Law · Electrolysis · Batteries · Corrosion · JEE & NEET

Quick Reference — Constants & Symbols Used in This Chapter

Quantity / Symbol	**Value (SI / common unit)**	**Meaning**
Faraday constant, F	$96{,}487$ C mol^{-1} (≈ 96500)	Charge carried by 1 mole of electrons
Gas constant, R	8.314 J K^{-1} mol^{-1}	Used in Nernst, ΔG, equilibrium relations
Avogadro number, N_A	6.022×10^{23} mol^{-1}	Particles per mole; $F = N_A \cdot e$
Electronic charge, e	1.602×10^{-19} C	Charge on one electron (magnitude)
SHE reference	$E^\circ_{H^+/H_2} = 0.00$ V	Standard hydrogen electrode; all E° are vs SHE
$2.303\,RT/F$ at 298 K	0.0591 V	Nernst constant for one-electron change at 25 °C
E°_{cell}	V (volts)	Standard cell EMF: $E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$
κ (kappa)	S m^{-1} (or S cm^{-1})	Conductivity (specific conductance)
Λ_m	S m^2 mol^{-1} (or S cm^2 mol^{-1})	Molar conductivity
Λ°_m	S m^2 mol^{-1}	Limiting molar conductivity (infinite dilution)

1 Electrochemical Cells: Galvanic vs Electrolytic

NCERT Section 2.1 introduces two cell types. A **galvanic** (voltaic) cell converts the Gibbs energy of a spontaneous redox reaction into electrical energy. An **electrolytic** cell does the opposite: an external voltage drives a non-spontaneous reaction. The Daniell cell (Zn | Zn^{2+} || Cu^{2+} | Cu) gives 1.1 V at unit concentrations.

Daniell cell reaction & EMF

Overall: $\mathbf{Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)}$

Oxidation (anode): $\mathbf{Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-}$

Reduction (cathode): $\mathbf{Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)}$
At unit concentrations: $E^\circ_{\text{cell}} = 1.10$ V
The **anode is negative** and **cathode is positive** in a galvanic cell (opposite in electrolytic). The salt bridge maintains electrical neutrality and prevents polarisation.

Standard cell potential (EMF)

$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$
(both values are taken as **standard reduction potentials** vs SHE; do *not* reverse the sign of the anode value before subtracting).
If $E^\circ_{\text{cell}} > 0$, the cell reaction (as written) is **spontaneous**. A galvanic cell is built by pairing a half-cell with high E° (cathode) with one with low E° (anode).

Cell notation (IUPAC convention)

Anode is written on the **left**, cathode on the **right**. Single bar | marks a phase boundary (electrode | electrolyte). Double bar || marks the salt bridge.
Example: Zn(s) | Zn^{2+}(aq, 1 M) || Cu^{2+}(aq, 1 M) | Cu(s)

Standard electrode potential, E°

The potential of a half-cell measured **vs the SHE** when all species are at unit activity (1 M for ions, 1 bar for gases, pure solids/liquids). By IUPAC convention, all E° values are reported as **reduction potentials**. A more positive E° means a stronger oxidising agent.

2 Nernst Equation

NCERT Section 2.4 generalises the cell EMF to non-standard concentrations. The Nernst equation gives the potential of a half-cell, or the full cell, at any concentration and temperature.

Nernst equation — single electrode

For the reduction $\mathbf{M^{n+}(aq) + ne^- \longrightarrow M(s)}$:

$$E_{M^{n+}/M} = E^\circ_{M^{n+}/M} - \frac{RT}{nF}\ \ln \frac{[M]}{[M^{n+}]}$$

Since $[M] = 1$ for a pure solid: $E = E^\circ + \dfrac{RT}{nF}\ \ln [M^{n+}]$

n = number of electrons transferred per ion; $R = 8.314$ J K^{-1} mol^{-1}; T in kelvin; $F = 96500$ C mol^{-1}.

Nernst equation — full cell

For cell reaction $aA + bB \longrightarrow cC + dD$ with n electrons transferred:

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF}\ \ln Q$$

where $Q = \dfrac{[C]^c[D]^d}{[A]^a[B]^b}$ (reaction quotient)

At $T = 298$ K, converting **ln** to **$\log_{10}$**:

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.0591}{n}\ \log Q \qquad \text{(V)}$$

As products build up (Q increases), E_{cell} falls. The cell stops working (dead battery) when $E_{\text{cell}} = 0$, at which point $Q = K_c$.

Daniell cell — Nernst form

$\mathrm{Zn} + \mathrm{Cu}^{2+} \longrightarrow \mathrm{Zn}^{2+} + \mathrm{Cu}$, $n = 2$, $E^\circ_{\text{cell}} = 1.10$ V

$$E_{\text{cell}} = 1.10 - \frac{0.0591}{2}\log\frac{[\mathrm{Zn}^{2+}]}{[\mathrm{Cu}^{2+}]} \quad \text{(at 298 K)}$$

Activities of pure solids Zn and Cu are taken as 1 and **do not appear in Q**. Raising $[\mathrm{Cu}^{2+}]$ or lowering $[\mathrm{Zn}^{2+}]$ **increases** the cell EMF.

Sign of the log term

Write Nernst as $E = E^\circ - \frac{0.0591}{n}\log Q$. A frequent slip is to put $Q = \frac{[\text{reactant}]}{[\text{product}]}$ — it must be $\frac{[\text{product}]}{[\text{reactant}]}$ (with stoichiometric powers). For a single reduction half-cell, the "product" is the reduced form (solid metal, activity 1), so $[M^{n+}]$ appears in the denominator of Q.

3 Gibbs Energy, Equilibrium Constant & Cell EMF

NCERT Section 2.5 links the electrical work done by a cell to the thermodynamic Gibbs energy of the reaction, and at equilibrium connects E° with K_c.

ΔG from cell EMF

$$\Delta_r G = -nFE_{\text{cell}}$$

$$\Delta_r G^\circ = -nFE^\circ_{\text{cell}}$$

where n = moles of electrons transferred per mole of reaction; $F = 96500$ C mol^{-1}. Units: J mol^{-1} (if E in V).

A galvanic cell delivers reversible electrical work $w_{\text{rev}} = \Delta G$. Spontaneity check: $E_{\text{cell}} > 0 \Leftrightarrow \Delta G < 0$.

Equilibrium constant from E°

At equilibrium $E_{\text{cell}} = 0$ and $Q = K_c$, so the Nernst equation gives:

$$E^\circ_{\text{cell}} = \frac{RT}{nF}\ln K_c = \frac{0.0591}{n}\log K_c \quad \text{(at 298 K)}$$

Rearranged: $\log K_c = \frac{n\,E^\circ_{\text{cell}}}{0.0591}$ (at 298 K)

Even a small E°_{cell} corresponds to a huge K_c when $n \geq 2$. Example: $E^\circ = 0.30$ V, $n = 2$ gives $\log K_c \approx 10.2$, i.e. $K_c \approx 10^{10}$.

ΔG° in terms of equilibrium constant

Combining $\Delta_r G^\circ = -nFE^\circ_{\text{cell}}$ with $E^\circ_{\text{cell}} = \frac{RT}{nF}\ln K_c$:

$$\Delta_r G^\circ = -RT\ln K_c = -2.303\,RT\log K_c$$

This is the bridge between electrochemistry and equilibrium thermodynamics. Large positive $E^\circ_{\text{cell}} \Rightarrow$ large negative $\Delta G^\circ \Rightarrow$ very large K_c (reaction goes essentially to completion).

JEE/NEET Extension — temperature dependence of EMF

The temperature coefficient of EMF: $\left(\frac{\partial E^\circ}{\partial T}\right)_P = \frac{\Delta_r S^\circ}{nF}$. Combined with $\Delta G^\circ = -nFE^\circ$ and $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$, this yields ΔH° and ΔS° from a single cell measurement at two temperatures — a standard JEE numerical.

4 Conductance & Conductivity of Electrolytes

NCERT Section 2.3 defines how electrolytic solutions carry current. Unlike metals, in solutions the charge carriers are **ions**, and conductivity rises with temperature (greater ion mobility) and depends on concentration.

Resistance, resistivity, cell constant

Resistance: $R = \rho \frac{l}{A}$ (ρ = resistivity, Ω m or Ω cm)

Conductance: $G = \frac{1}{R}$, unit **siemens** (S) $= \Omega^{-1}$

Conductivity (specific conductance): $\kappa = \frac{1}{\rho} = G \cdot \frac{l}{A}$

Cell constant: $G^* = \frac{l}{A}$ (unit m^{-1} or cm^{-1})

κ is the conductance of a solution between electrodes 1 m apart and of 1 m^2 area. The cell constant is measured by calibrating with a KCl solution of known κ.

Measurement of conductivity (Wheatstone bridge)

$$\kappa = G \times G^* = \frac{1}{R} \times \frac{l}{A}$$

For a calibrating KCl solution of known κ_{KCl} and measured resistance R_{KCl}:

$$\frac{l}{A} = \kappa_{\mathrm{KCl}} \cdot R_{\mathrm{KCl}}$$

An **AC source** is used (not DC) so the electrolyte is not electrolysed during measurement. Platinised-Pt electrodes minimise polarisation.

Electrolytic vs metallic (electronic) conductance

Metallic: electrons carry charge; κ *decreases* with T (lattice vibration); no chemical change.
Electrolytic: ions carry charge; κ *increases* with T (greater ion mobility, more dissociation); electrolysis occurs at electrodes.

5 Molar Conductivity & Kohlrausch's Law

NCERT Section 2.3.2 introduces molar conductivity Λ_m, which removes the dilution effect from κ and isolates the contribution of one mole of electrolyte. Kohlrausch's law (Section 2.3.3) then breaks Λ_m° into ion-specific contributions.

Molar conductivity

$$\Lambda_m = \frac{\kappa}{c}$$

SI: if κ is in S m^{-1} and c in mol m^{-3}, Λ_m is in S m^2 mol^{-1}.

Common form: $\Lambda_m\ (\mathrm{S\ cm^2\ mol^{-1}}) = \dfrac{1000 \cdot \kappa\ (\mathrm{S\ cm^{-1}})}{c\ (\mathrm{mol\ L^{-1}})}$

The factor of 1000 converts molarity (mol/L = mol/1000 cm^3) into cm-based units. Λ_m **always increases on dilution**; the reason differs for strong vs weak electrolytes.

Variation of Λ_m with concentration: Debye–Hückel–Onsager

For **strong electrolytes** (NCERT Eq. 2.27):

$$\Lambda_m = \Lambda_m^\circ - A\sqrt{c}$$

where Λ_m° = limiting (infinite-dilution) molar conductivity and A is a constant for a given solvent and temperature at fixed stoichiometry.

Plot of Λ_m vs $\sqrt{c}$ is a **straight line** with intercept Λ_m° — this is how Λ_m° is found experimentally for strong electrolytes. For weak electrolytes the plot is steeply curved near $c = 0$ and Λ_m° cannot be extrapolated this way.

Kohlrausch's law of independent migration of ions

$$\Lambda_m^\circ = \nu_+ \lambda_+^\circ + \nu_- \lambda_-^\circ$$

where $\lambda_+^\circ, \lambda_-^\circ$ = limiting molar ionic conductivities of cation/anion; ν_+, ν_- = number of cations/anions per formula unit.

Example: $\Lambda_m^\circ(\mathrm{CaCl_2}) = \lambda^\circ(\mathrm{Ca^{2+}}) + 2\,\lambda^\circ(\mathrm{Cl^-})$

At infinite dilution each ion moves **independently** of its counter-ion. This lets us compute Λ_m° for any electrolyte from a table of ionic conductivities.

Kohlrausch for a weak electrolyte (e.g. CH_3COOH)

$$\Lambda_m^\circ(\mathrm{CH_3COOH}) = \Lambda_m^\circ(\mathrm{CH_3COONa}) + \Lambda_m^\circ(\mathrm{HCl}) - \Lambda_m^\circ(\mathrm{NaCl})$$

Each term on the right is a **strong electrolyte** whose Λ_m° is obtained by the $\sqrt{c}$ extrapolation. This indirect route is the only way to get Λ_m° for weak electrolytes.

Degree of dissociation & dissociation constant (weak electrolyte)

$$\alpha = \frac{\Lambda_m}{\Lambda_m^\circ} \quad \text{(fraction dissociated at concentration } c\text{)}$$

For a weak monoprotic acid HA:

$$K_a = \frac{c\,\alpha^2}{1-\alpha} = \frac{c\,\Lambda_m^2}{\Lambda_m^\circ\,(\Lambda_m^\circ - \Lambda_m)}$$

Conductance measurements at varying c give α and hence K_a for weak acids and bases — a key NCERT application of Kohlrausch's law.

JEE/NEET Extension — transport number

Transport (transference) number of an ion: $t_+ = \dfrac{\lambda_+^\circ}{\Lambda_m^\circ}$ and $t_- = \dfrac{\lambda_-^\circ}{\Lambda_m^\circ}$, with $t_+ + t_- = 1$. It is the fraction of the total current carried by that ion.

6 Electrolysis & Faraday's Laws

NCERT Section 2.6 covers electrolysis: passing current through an electrolyte to drive a non-spontaneous redox reaction. Faraday quantified the relation between charge passed and the amount of substance deposited or liberated.

Faraday's first law of electrolysis

$$w = Z \cdot Q = Z \cdot I \cdot t$$

where w = mass deposited (g); $Q = It$ = charge passed (coulomb); Z = electrochemical equivalent (g/C).

$$Z = \frac{M}{nF} \quad \text{(mass per coulomb)}$$

The mass of substance liberated at an electrode is **directly proportional** to the quantity of charge passed.

Faraday's second law

When the *same* charge Q is passed through different electrolytes connected in series:

$\frac{w_1}{w_2} = \frac{E_1}{E_2}$ where $E = \frac{M}{n}$ is the equivalent weight (g/equivalent).

Masses deposited are in the ratio of their **equivalent weights**. One mole of electrons = 1 faraday = 96500 C and deposits 1 equivalent of substance.

Charge for n moles & moles deposited

Charge needed to deposit 1 mole of M^{n+} to M: $Q = nF$

Moles of metal deposited by charge Q: $\text{mol} = \frac{Q}{nF} = \frac{It}{nF}$

Mass deposited: $w = \frac{M \cdot I \cdot t}{nF}$

n = electrons per ion (Cu^{2+}: $n = 2$; Ag^{+}: $n = 1$; Al^{3+}: $n = 3$). For gases released (e.g. H_2, O_2), the volume at STP is $\frac{It}{nF} \times 22400$ cm^3.

Products of electrolysis depend on E° and overvoltage

Predicting the product at each electrode requires comparing **reduction potentials** of all species that could react, and accounting for **overvoltage** (extra voltage above E° needed in practice, especially for gases). In aqueous NaCl electrolysis the cathode reduces H^+/H_2O to H_2 (not Na^+), and the anode liberates Cl_2 rather than O_2 because of O_2 overvoltage on most anodes.

Charge counting: use n, not 1

A common slip is to write $w = M\,It/F$ and forget the n. Always divide by nF, where n is the number of electrons per ion. For depositing copper from Cu^{2+}, $n = 2$.

7 Batteries & Fuel Cells

NCERT Section 2.7 surveys commercial cells. **Primary** cells discharge irreversibly (cannot be recharged); **secondary** cells are rechargeable. **Fuel cells** are flow-through galvanic cells whose reactants are continuously fed.

Dry cell (Leclanché, primary)

Anode (Zn cup): $\mathrm{Zn} \longrightarrow \mathrm{Zn^{2+}} + 2e^-$

Cathode (graphite in MnO_2/NH_4Cl paste): $\mathrm{MnO_2} + \mathrm{NH_4^+} + e^- \longrightarrow \mathrm{MnO(OH)} + \mathrm{NH_3}$

EMF ≈ 1.5 V

Cannot be recharged because the cell reactions are not easily reversible. NH_3 produced complexes with Zn^{2+} as $[\mathrm{Zn(NH_3)_4}]^{2+}$.

Mercury cell (primary)

Anode: $\mathrm{Zn(Hg)} + 2\mathrm{OH^-} \longrightarrow \mathrm{ZnO(s)} + \mathrm{H_2O} + 2e^-$

Cathode: $\mathrm{HgO(s)} + \mathrm{H_2O} + 2e^- \longrightarrow \mathrm{Hg(l)} + 2\mathrm{OH^-}$

Overall: $\mathrm{Zn(Hg)} + \mathrm{HgO(s)} \longrightarrow \mathrm{ZnO(s)} + \mathrm{Hg(l)}$; EMF ≈ 1.35 V (constant)

Used in hearing aids and watches because the cell EMF stays constant during its life (no soluble ions in the net reaction).

Lead-acid storage battery (secondary)

Anode: $\mathrm{Pb(s)} + \mathrm{SO_4^{2-}} \longrightarrow \mathrm{PbSO_4(s)} + 2e^-$

Cathode: $\mathrm{PbO_2(s)} + \mathrm{SO_4^{2-}} + 4\mathrm{H^+} + 2e^- \longrightarrow \mathrm{PbSO_4(s)} + 2\mathrm{H_2O}$

Overall (discharge): $\mathbf{Pb + PbO_2 + 2H_2SO_4 \longrightarrow 2PbSO_4 + 2H_2O}$; $E^\circ_{\text{cell}} \approx 2.041$ V per cell
On **recharging**, the reaction is reversed ($PbSO_4$ converts back to Pb at anode and PbO_2 at cathode) by applying an external voltage greater than 2 V. Six cells in series give a 12 V car battery.

Nickel-cadmium cell (secondary)

Anode: $\mathbf{Cd(s) + 2OH^- \longrightarrow Cd(OH)_2(s) + 2e^-}$
Cathode: $\mathbf{NiO_2(s) + 2H_2O + 2e^- \longrightarrow Ni(OH)_2(s) + 2OH^-}$
Overall: $\mathbf{Cd + NiO_2 + 2H_2O \longrightarrow Cd(OH)_2 + Ni(OH)_2}$
Longer life than lead-acid and lighter, but Cd is toxic. Used in portable electronics (now largely replaced by Li-ion).

H_2–O_2 fuel cell (alkaline electrolyte)

Anode: $\mathbf{2H_2(g) + 4OH^-(aq) \longrightarrow 4H_2O(l) + 4e^-}$
Cathode: $\mathbf{O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)}$
Overall: $\mathbf{2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)}$; $E^\circ_{\text{cell}} \approx 1.23$ V; $\Delta_r G^\circ \approx -474$ kJ mol^{-1}
Reactants are fed **continuously**; the cell produces electricity as long as H_2 and O_2 are supplied. Used in the Apollo space programme. Efficiency ∼70% (much higher than thermal plants) and the only product is water — ecofriendly.

Why secondary cells can be recharged

A secondary cell uses electrode reactions whose products **stay attached to the electrodes** as insoluble solids (e.g. $PbSO_4$, $Cd(OH)_2$). An external voltage greater than the cell EMF reverses the spontaneous direction, regenerating the original reactants. Primary cells produce soluble or gaseous products that diffuse away, so the reverse reaction cannot be re-established.

8 Corrosion

NCERT Section 2.8 describes corrosion as a spontaneous electrochemical process. Atmospheric water containing dissolved O_2 and CO_2 sets up tiny galvanic cells on the iron surface, oxidising iron to rust.

Rusting of iron (electrochemical mechanism)

Anodic region (oxidation of Fe):

$\mathbf{Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-}$ $(E^\circ = -0.44$ V)

Cathodic region (reduction of O_2 in acidic H_2CO_3 from dissolved CO_2):

$\mathbf{O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)}$ $(E^\circ = +1.23$ V)

Overall: $\mathbf{2Fe + O_2 + 4H^+ \longrightarrow 2Fe^{2+} + 2H_2O}$; $E^\circ_{\text{cell}} = 1.67$ V

Subsequent atmospheric oxidation:

$\mathbf{4Fe^{2+} + O_2 + (4{+}2x)H_2O \longrightarrow 2Fe_2O_3 \cdot xH_2O + 8H^+}$

Rust is hydrated ferric oxide $Fe_2O_3 \cdot xH_2O$. The reaction needs **both O_2 and H_2O**; dry air or oxygen-free water alone does not rust iron.

Prevention of corrosion

Barrier methods: paint, grease, oil (block O_2/H_2O contact).
Galvanisation: coat iron with a layer of Zn (since $E^\circ_{\text{Zn}^{2+}/\text{Zn}} = -0.76$ V, more negative than Fe, Zn oxidises preferentially).
Cathodic protection (sacrificial anode): connect a block of Mg or Zn to the iron object (pipelines, ship hulls); the more active metal corrodes, leaving Fe intact.
Tinning (Sn coating): protects only while the coating is intact; once scratched, iron corrodes faster

(Sn is less active than Fe).

OIL RIG & cell-polarity memory aids

OIL RIG: **O**xidation **I**s **L**oss (of e$^-$), **R**eduction **I**s **G**ain.
AN OX, RED CAT: **AN**ode = **OX**idation; **RED**uction = **CAT**hode (in every cell, galvanic or electrolytic). Galvanic polarity: anode $-$ve, cathode $+$ve. Electrolytic polarity: anode $+$ve, cathode $-$ve (driven by external source).

9 Quick Reference — Summary Tables

Galvanic vs Electrolytic Cell

Feature	**Galvanic (voltaic)**	**Electrolytic**
Energy conversion	Chemical $\longrightarrow$ electrical	Electrical $\longrightarrow$ chemical
Spontaneity	Spontaneous ($\Delta G < 0$)	Non-spontaneous ($\Delta G > 0$)
Anode polarity	Negative ($-$)	Positive ($+$)
Cathode polarity	Positive ($+$)	Negative ($-$)
Anode reaction	Oxidation	Oxidation
Cathode reaction	Reduction	Reduction
Example	Daniell cell, dry cell	Electrolysis of brine, electroplating

Selected Standard Electrode Potentials (Reduction, 298 K, vs SHE)

Half-cell	E° **/ V**	**Half-cell**	E° **/ V**
F_2/F^-	$+2.87$	$2H^+/H_2$ (SHE)	0.00
MnO_4^-/Mn^{2+} (acid)	$+1.51$	Pb^{2+}/Pb	-0.13
Cl_2/Cl^-	$+1.36$	Sn^{2+}/Sn	-0.14
$Cr_2O_7^{2-}/Cr^{3+}$	$+1.33$	Ni^{2+}/Ni	-0.25
O_2/H_2O (acid)	$+1.23$	Fe^{2+}/Fe	-0.44
Ag^+/Ag	$+0.80$	Zn^{2+}/Zn	-0.76
Fe^{3+}/Fe^{2+}	$+0.77$	Al^{3+}/Al	-1.66
Cu^{2+}/Cu	$+0.34$	Mg^{2+}/Mg	-2.37

Higher (more positive) E° means a stronger oxidising agent (the oxidised form). Lower (more negative) E° means a stronger reducing agent (the reduced form). Use this table to predict the direction of any redox reaction.