CHEMISTRY · CLASS 12 · CHAPTER 2 · JEE AND NEET

Electrochemistry

Cells, potentials, conductance and electrolysis

A battery and an electroplating bath run the same chemistry in opposite directions. Learn which way the electrons want to go, and this chapter follows.

14 diagrams | 17 worked questions | JEE Advanced set | Fact sheet

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In short

- A galvanic cell turns a spontaneous redox reaction into electricity. An electrolytic cell uses electricity to drive a non-spontaneous one.
- $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$. A positive E°_{cell} means the reaction is spontaneous.
- The Nernst equation corrects the potential for concentration. It also links E° to the equilibrium constant and to ΔG°.
- On dilution, conductivity falls but molar conductivity rises. Kohlrausch's law gives Λ°_m for weak electrolytes.
- In electrolysis, the charge passed fixes the amount of product. One faraday is one mole of electrons.

0 How to Read This Set

- Read Sections 01 to 05 as one story: cell, potential, concentration, energy. Sections 06 to 08 are a second story: ions carrying current.
- Most questions here are numericals. Watch the units in every step.
- Check each section's badge. Green is for everyone. Magenta is mostly for JEE Advanced and boards.
- Orange boxes are traps. Five of them come straight from the mistakes our NEET analysis found.
- Try each worked question before reading its answer.

WHICH EXAM NEEDS WHAT

The rationalised NCERT book keeps this chapter in full. But the exam syllabi are narrower at the end.

- The **NEET 2026** and **JEE Main** syllabi end the unit with "Dry cell and lead accumulator; Fuel cells." They do not name the mercury cell, the nickel-cadmium cell or corrosion.
- The **JEE Advanced** syllabus lists "Batteries: Primary and Secondary, fuel cells; Corrosion."
- So Section 10 (other batteries) and Section 11 (corrosion) carry the magenta badge. NEET and JEE Main students can read them lightly. Board students need them.

PYQ ANALYSIS **See how NEET tested this chapter from 2015 to 2025** →

In our NEET analysis, about 68% of the questions from this chapter were numericals. Electrolysis and Faraday's laws alone gave a quarter of them. Frequency bands on this page describe NEET only; we have no JEE data for this chapter yet.

01 Electrochemical Cells

IN NEET AND JEE

Two kinds of cell

- **Electrochemistry** studies how chemical energy and electrical energy change into each other.
- A **galvanic (voltaic) cell** turns the chemical energy of a **spontaneous** redox reaction into electrical energy.
- An **electrolytic cell** uses electrical energy to make a **non-spontaneous** reaction happen.

The Daniell cell

- Zinc is dipped in zinc sulphate solution. Copper is dipped in copper sulphate solution.
- The reaction is:

$$\mathrm{Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)}$$

- With 1 M solutions, the cell gives **1.1 V**.
- Each metal in its solution is a **half-cell** (or redox couple).
- **Oxidation** happens at the **anode**. Here, zinc gives up electrons: $\mathrm{Zn \rightarrow Zn^{2+} + 2e^-}$.
- **Reduction** happens at the **cathode**. Here, copper ions take electrons: $\mathrm{Cu^{2+} + 2e^- \rightarrow Cu}$.
- In a galvanic cell, the anode is **negative** and the cathode is **positive**.
- Electrons flow through the outer wire from zinc to copper. The current flows the other way, from copper to zinc.

The salt bridge

- A **salt bridge** joins the two solutions. It is a U-tube filled with a solution of KCl or $\mathrm{NH_4NO_3}$, usually set in agar-agar gel.
- It completes the circuit.
- It keeps each solution electrically neutral, so the reaction can go on.

Writing a cell

- Write the anode on the left and the cathode on the right. A double line stands for the salt bridge:

$$\mathrm{Zn(s)\,|\,Zn^{2+}(aq)\,||\,Cu^{2+}(aq)\,|\,Cu(s)}$$

- The **cell potential** is the difference between the two electrode potentials:

$$E_{cell} = E_{right} - E_{left} = E_{cathode} - E_{anode}$$

- When no current is drawn, the cell potential is called the **emf** of the cell.

When an outside voltage pushes back

- Connect an external voltage source **against** the Daniell cell.
- If the external voltage is **less than 1.1 V**, the cell still works as a galvanic cell. Zinc dissolves and copper deposits.
- If it is **exactly 1.1 V**, no current flows. No reaction happens.
- If it is **more than 1.1 V**, the reaction reverses. Copper dissolves and zinc deposits. The cell is now an **electrolytic cell**.

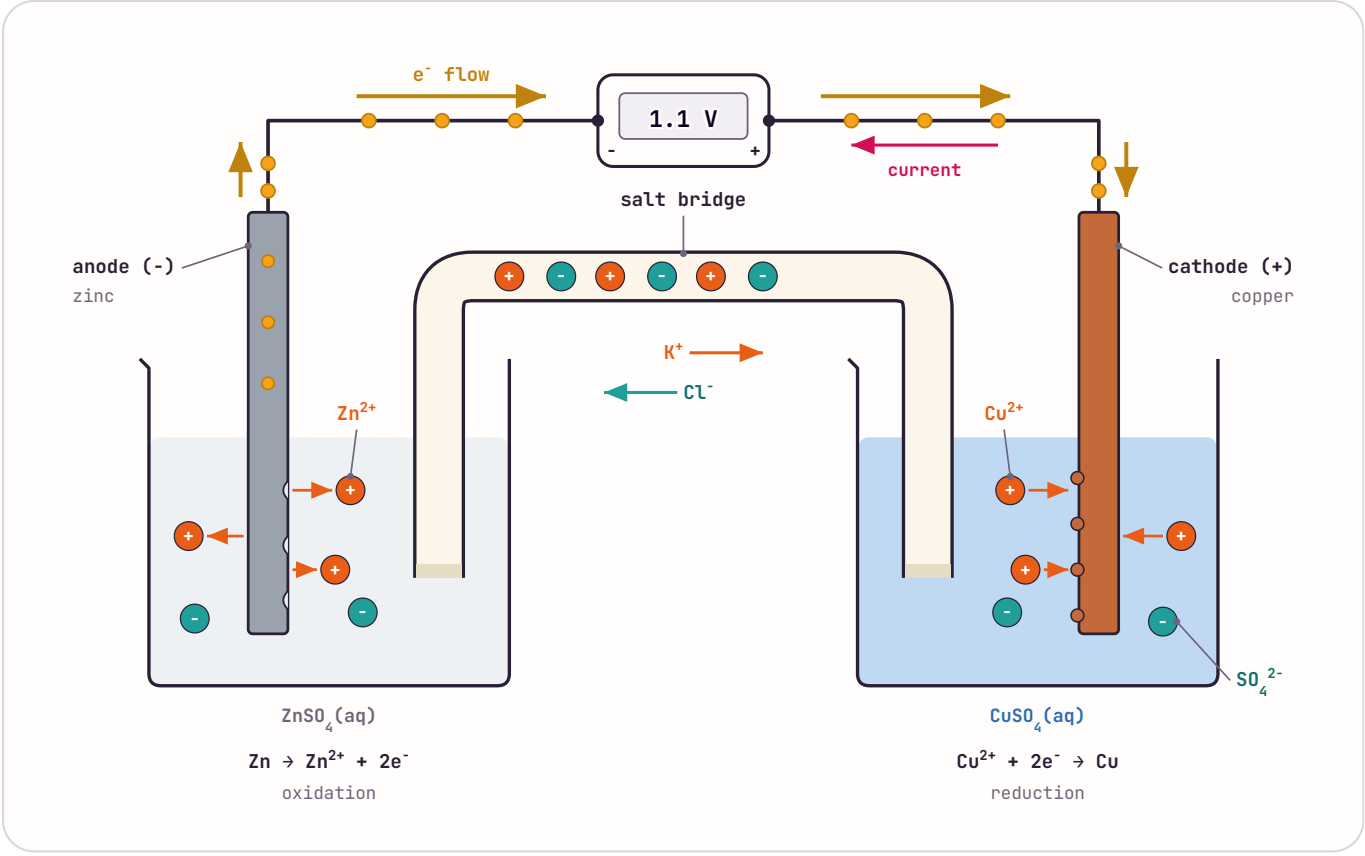


Fig 1 · Daniell cell. Zinc gives electrons; copper ions take them; the salt bridge keeps both sides neutral.

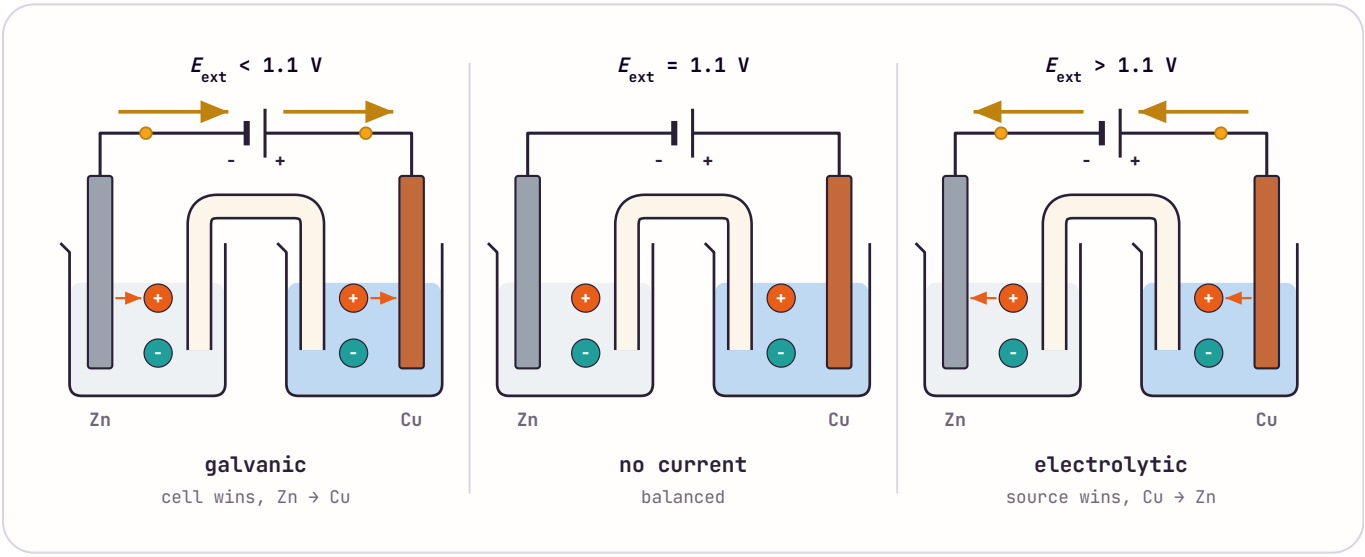


Fig 2 · Pushing back on the cell. Beat the cell's voltage, and the reaction runs backwards.



HAVING DIFFICULTY FOLLOWING WHICH WAY THE ELECTRONS GO?

Play it.

This chapter is a set of short games in the Logic Bloom app.
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SCAN



TRAP ALERT

TRAP ALERT

- In a galvanic cell the anode is negative. In an electrolytic cell the anode is positive. But in **both**, oxidation happens at the anode. Learn the reaction, not the sign.

02 Electrode Potential and the Electrochemical Series

IN NEET AND JEE FREQUENT *18.8% of NEET questions*

The standard hydrogen electrode (SHE)

- We cannot measure one electrode's potential alone. We can only measure a difference.
- So we pick one reference electrode and give it zero.
- The **standard hydrogen electrode** is a platinum electrode coated with **platinum black**.
- It dips in an acid solution, while pure hydrogen gas bubbles through it.
- Both the hydrogen gas (1 bar) and the H^+ ions (about 1 M) are at unit concentration.
- Its potential is taken as **zero at all temperatures**:

$$E^\circ(H^+/H_2) = 0.00\ V$$

- Pair any half-cell with the SHE, and the measured emf is that half-cell's potential.
  - With copper as the cathode, the emf is 0.34 V. So $E^\circ(Cu^{2+}/Cu) = +0.34\ V$.
  - With zinc, the emf shows $E^\circ(Zn^{2+}/Zn) = -0.76$ V.
- Platinum and gold are used as **inert electrodes** when the metal itself takes no part in the reaction.

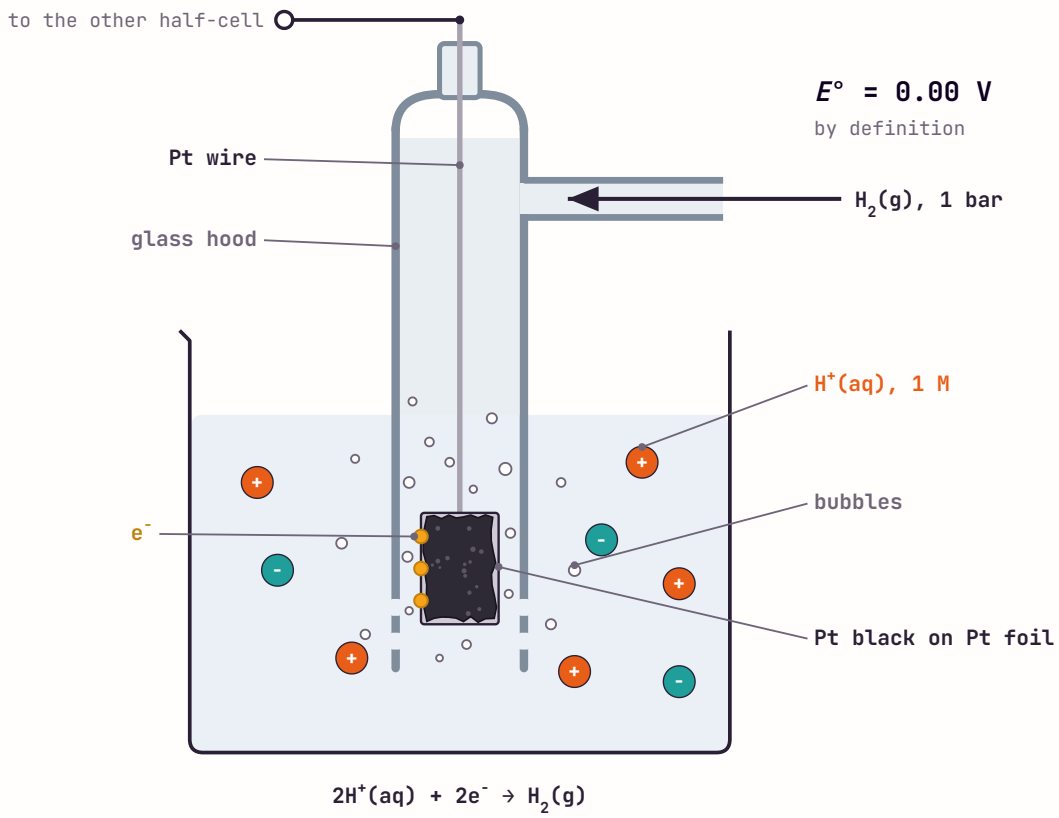


Fig 3 · Standard hydrogen electrode. Hydrogen bubbling over platinum black sets the zero of the scale.

Standard reduction potentials (NCERT values, 298 K)

HALF-REACTION	E° / V
$F_2 + 2e^- \rightarrow 2F^-$	+2.87
$Au^{3+} + 3e^- \rightarrow Au$	+1.40
$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.36
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	+1.23
$Br_2 + 2e^- \rightarrow 2Br^-$	+1.09
$Ag^+ + e^- \rightarrow Ag$	+0.80
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.77
$I_2 + 2e^- \rightarrow 2I^-$	+0.54
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34
$2H^+ + 2e^- \rightarrow H_2$	0.00
$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44
$Zn^{2+} + 2e^- \rightarrow Zn$	−0.76

HALF-REACTION	E° / V
$Mg^{2+} + 2e^- \rightarrow Mg$	−2.36
$Na^+ + e^- \rightarrow Na$	−2.71
$Li^+ + e^- \rightarrow Li$	−3.05

Reading the series

- A **higher (more positive)** E° means the oxidised form is a **stronger oxidising agent**. It accepts electrons easily.
- A **lower (more negative)** E° means the reduced form is a **stronger reducing agent**. It gives up electrons easily.
- So **fluorine** is the strongest oxidising agent here, and **lithium** is the strongest reducing agent.
- A species can oxidise anything **below** it in the table.
- Metals with negative E° are stronger reducing agents than hydrogen. They can displace H_2 from dilute acids.

Spontaneity

- For a cell, find $E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$.
- **Positive** E°_{cell}: the reaction is spontaneous as written.
- **Negative** E°_{cell}: the reverse reaction is spontaneous.

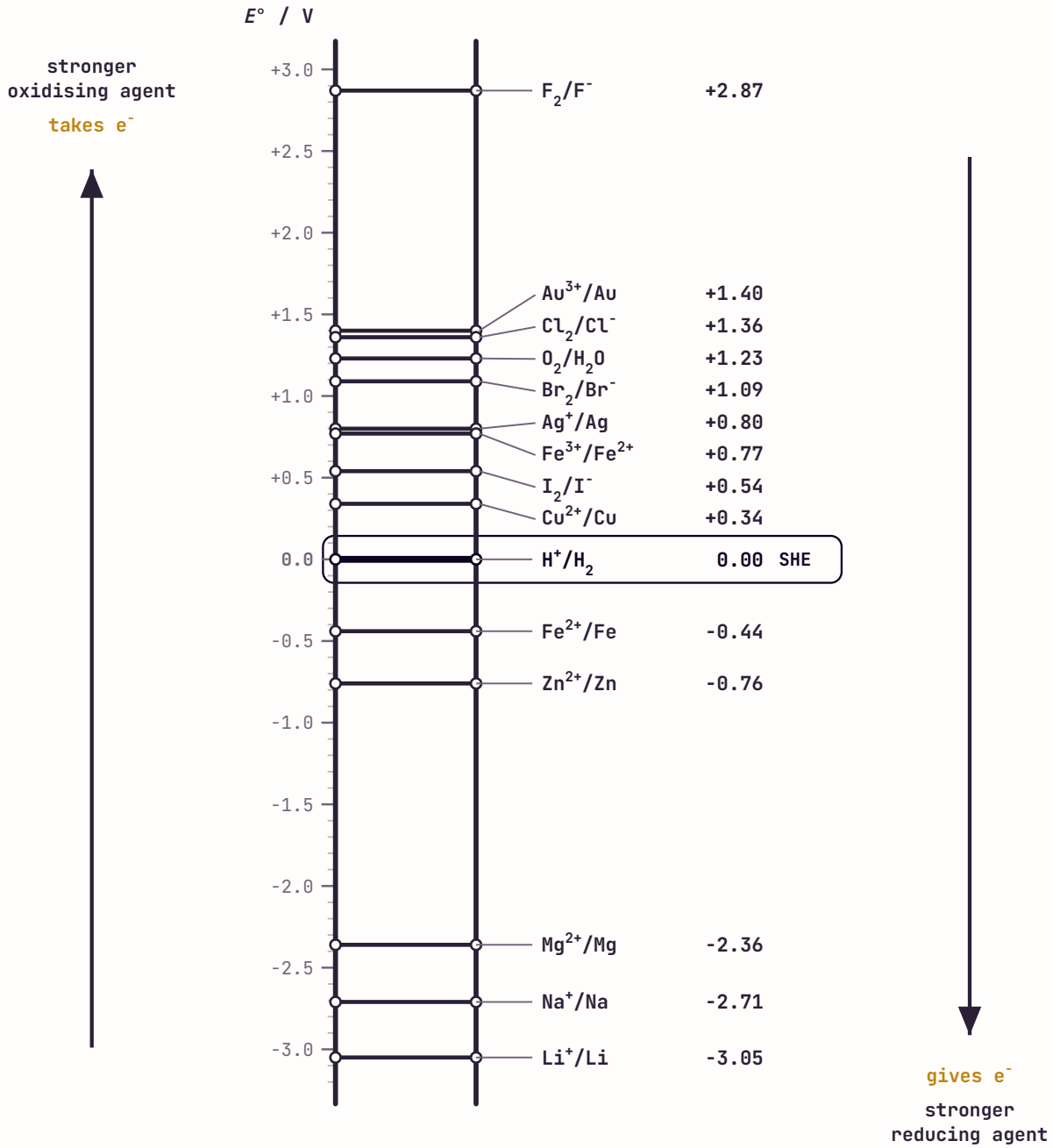


Fig 4 · The electrochemical series ladder. Higher rungs take electrons; lower rungs give them.



TRAP ALERT

TRAP ALERT

- Always subtract: cathode **minus** anode, using **reduction** potentials for both. Do not flip a sign and then subtract again. When both values are negative, write each sign out before subtracting. Our NEET analysis names this subtraction slip as a top mark-loser.



TRAP ALERT

TRAP ALERT

- Electrode potential does **not** change when you multiply a half-reaction by a number. $Ag^+ + e^- \rightarrow Ag$ is +0.80 V, and so is $2Ag^+ + 2e^- \rightarrow 2Ag$. Potential is an intensive property.

Q1 Worked question

NEET JEE MAIN LEVEL 2

Find E°_{cell} for $Zn(s)\,|\,Zn^{2+}\,\|\,Ag^+\,|\,Ag(s)$. Use $E^\circ(Zn^{2+}/Zn) = -0.76$ V and $E^\circ(Ag^+/Ag) = +0.80$ V.

1. 1.56 V
2. 0.04 V
3. 2.36 V
4. −1.56 V

Answer: 1

CORE PRINCIPLE Zinc is the anode, silver is the cathode. Subtract reduction potentials: cathode minus anode.

MATH $E^\circ_{cell} = 0.80 - (-0.76) = 1.56$ V.

TRAP Option 2 adds the two values as written. Option 3 doubles silver's potential because two Ag^+ ions react. Option 4 subtracts the wrong way round.

Q2 Worked question

NEET LEVEL 2

Use these values: $E^\circ(Fe^{3+}/Fe^{2+}) = +0.77$ V, $E^\circ(I_2/I^-) = +0.54$ V, $E^\circ(Br_2/Br^-) = +1.09$ V. Which statements are correct?

(a) Fe^{3+} can oxidise I^- to I_2.

(b) Fe^{3+} can oxidise Br^- to Br_2.

(c) Br_2 can oxidise Fe^{2+} to Fe^{3+}.

(d) I_2 can oxidise Fe^{2+} to Fe^{3+}.

(e) Of the three oxidised forms, Br_2 is the strongest oxidising agent.

Choose the correct answer from the options given below.

1. (a), (b) and (e) only
2. (b), (c) and (d) only
3. (a), (c) and (e) only
4. (a), (d) and (e) only

Answer: 3

CORE PRINCIPLE An oxidising agent can oxidise species that sit **below** it in the series.

MATH Fe^{3+} (0.77) is above I_2 (0.54), so (a) works. Fe^{3+} is below Br_2 (1.09), so (b) fails. Br_2 is above Fe^{3+}, so (c) works. I_2 is below Fe^{3+}, so (d) fails.

TRAP Options 1 and 2 accept (b), which needs Fe^{3+} to beat a stronger oxidising agent. Option 4 accepts (d).

03 The Nernst Equation

IN NEET AND JEE REGULAR *9.3% of NEET questions*

For one electrode

- Standard potentials assume 1 M solutions. Real solutions are rarely 1 M.
- For $M^{n+}(aq) + ne^- \rightarrow M(s)$:

$$E = E^\circ - \frac{RT}{nF} \ln \frac{1}{[M^{n+}]}$$

- $R = 8.314\ J\ K^{-1}\ mol^{-1}$, T in kelvin, $F = 96487\ C\ mol^{-1}$, n = electrons in the half-reaction.
- The concentration of a pure solid is taken as 1.

For a whole cell, at 298 K

- At 298 K, $\frac{2.303RT}{F} = 0.059$ V. Changing ln to log:

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{n} \log Q$$

- For a general reaction $aA + bB \rightarrow cC + dD$:

$$Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

- For the Daniell cell:

$$E_{cell} = 1.10 - \frac{0.059}{2}\log\frac{[Zn^{2+}]}{[Cu^{2+}]}$$

- Raising $[Cu^{2+}]$ or lowering $[Zn^{2+}]$ raises the cell potential.

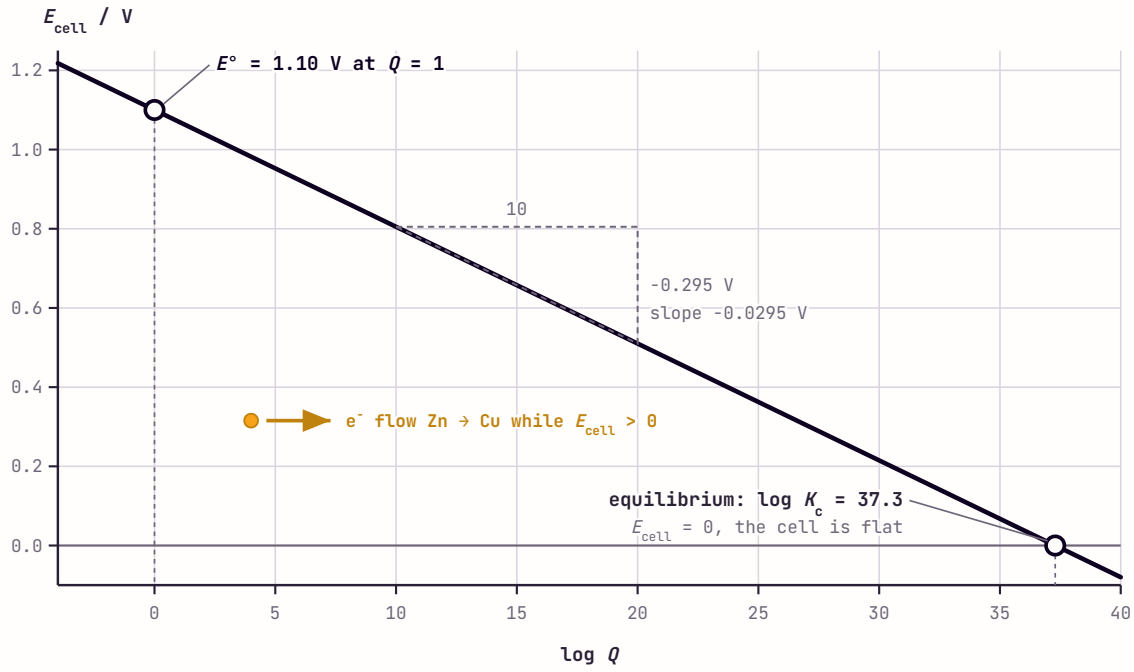


Fig 5 · Nernst line. The cell potential falls as Q grows, and reaches zero at equilibrium.



TRAP ALERT

TRAP ALERT

- The coefficients in the balanced equation become **powers** in Q. For $Mg + 2Ag^{+}$, $Q = \frac{[Mg^{2+}]}{[Ag^{+}]^2}$. Our NEET analysis found this missing-exponent slip in a 2022 question.

Q3 Worked question

NEET JEE MAIN LEVEL 2

For a Daniell cell, $E^{\circ}_{cell} = 1.10$ V. Find E_{cell} at 298 K when $[Zn^{2+}] = 0.1$ M and $[Cu^{2+}] = 0.001$ M.

1. 1.159 V
2. 1.041 V
3. 0.982 V
4. 1.100 V

Answer: 2

CORE PRINCIPLE Products go on top of Q. Here $n = 2$.

MATH $Q = \frac{0.1}{0.001} = 100$. $E = 1.10 - \frac{0.059}{2}\log 100 = 1.10 - 0.059 = 1.041$ V.

TRAP Option 1 inverts Q. Option 3 uses $n = 1$. Option 4 ignores concentration.

Q4 Worked question

JEE MAIN LEVEL 2

Find E_{cell} at 298 K for the cell reaction $Mg(s) + 2Ag^+(10^{-4}\ M) \rightarrow Mg^{2+}(10^{-3}\ M) + 2Ag(s)$. Take $E^\circ_{cell} = 3.17$ V.

1. 3.14 V
2. 2.88 V
3. 3.32 V
4. 3.02 V

Answer: 4

CORE PRINCIPLE Two electrons move, so $n = 2$. The coefficient 2 on Ag^+ becomes a square in Q.

MATH $Q = \frac{10^{-3}}{(10^{-4})^2} = 10^5$. $E = 3.17 - \frac{0.059}{2}(5) = 3.17 - 0.1475 \approx 3.02$ V.

TRAP Option 1 forgets to square $[Ag^+]$. Option 2 squares it but uses $n = 1$. Option 3 adds the correction instead of subtracting.

04 Equilibrium Constant from E°

IN NEET AND JEE

- As a cell runs, Q grows and E_{cell} falls.
- At equilibrium, $E_{cell} = 0$ and $Q = K_c$. Put these into the Nernst equation:

$$E^\circ_{cell} = \frac{2.303RT}{nF}\log K_c$$

- At 298 K:

$$\log K_c = \frac{nE^\circ_{cell}}{0.059}$$

- For the Daniell cell, $\log K_c = \frac{2 \times 1.1}{0.059} \approx 37.3$. So $K_c \approx 2 \times 10^{37}$. The reaction goes almost to completion.
- A small change in E° makes a huge change in K, because E° sits in the power.



THINK IT THROUGH

SANITY CHECK

- A positive E°_{cell} gives $K_c > 1$. A negative one gives $K_c < 1$. If your answer breaks this, recheck the sign.

Q5 Worked question

JEE MAIN LEVEL 2

A cell reaction involves 2 electrons and has $E^\circ_{cell} = 0.236$ V at 298 K. What is its equilibrium constant?

1. 10^8
2. 10^4
3. 10^{-8}
4. 10^{16}

Answer: 1

CORE PRINCIPLE $\log K = \frac{nE^\circ}{0.059}$.

MATH $\log K = \frac{2 \times 0.236}{0.059} = 8$. So $K = 10^8$.

TRAP Option 2 uses $n = 1$. Option 3 flips the sign. Option 4 squares the answer.

05 Cell Potential and Gibbs Energy

IN NEET AND JEE FREQUENT *15.6% of NEET questions, rising fast*

- The electrical work a cell does equals the fall in Gibbs energy:

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

- At standard conditions:

$$\Delta_r G^\circ = -nFE^\circ_{cell} = -RT\ln K$$

- A positive E_{cell} gives a negative $\Delta_r G$, so the reaction is spontaneous.

- For the Daniell cell: $\Delta_r G^\circ = -2 \times 96487 \times 1.1 = -212{,}271 \text{ J mol}^{-1} \approx -212.3 \text{ kJ mol}^{-1}$.
- E_{cell} is **intensive**: it does not depend on how much reacts. $\Delta_r G$ is **extensive**: double the equation, and n doubles, so $\Delta_r G$ doubles.

The three-way link

$$\Delta_r G^\circ = -nFE^\circ_{cell} = -2.303RT\log K$$

- Know any one, and you can find the other two.



TRAP ALERT

TRAP ALERT

- nFE gives joules. Divide by 1000 for kJ. Our NEET analysis lists J vs kJ as a repeated mark-loser.
- Also note: NCERT uses $F = 96487$ C mol^{-1}. Many questions round it to 96500. Use whatever value the question gives.

Q6 Worked question

NEET LEVEL 2

Find $\Delta_r G^\circ$ for the Daniell cell, with $E^\circ_{cell} = 1.10$ V. Take $F = 96500$ C mol^{-1}.

1. -106.2 kJ mol^{-1}
2. $+212.3$ kJ mol^{-1}
3. -212.3 kJ mol^{-1}
4. $-212{,}300$ kJ mol^{-1}

Answer: 3

CORE PRINCIPLE $\Delta_r G^\circ = -nFE^\circ$, with $n = 2$ for Zn and Cu^{2+}.

MATH $-2 \times 96500 \times 1.10 = -212{,}300 \text{ J mol}^{-1} = -212.3 \text{ kJ mol}^{-1}$.

TRAP Option 1 uses n = 1. Option 2 drops the minus sign. Option 4 labels joules as kilojoules.

READ NEXT **Stoichiometry and concentration practice: moles and molarity** →

06 Conductance of Electrolytic Solutions

IN NEET AND JEE | FREQUENT | *12.5% of NEET questions*

Two kinds of conduction

- **Metallic (electronic) conductance** is carried by electrons. It depends on:
  - the nature and structure of the metal
  - the number of valence electrons per atom
  - temperature: it **falls** as temperature rises.
- **Electrolytic (ionic) conductance** is carried by ions moving through a solution. It depends on:
  - the nature of the electrolyte
  - the size of the ions and how strongly they are solvated
  - the nature of the solvent and its viscosity
  - the concentration of the electrolyte
  - temperature: it **rises** as temperature rises.
- Pure water conducts very poorly. Adding an electrolyte raises the conductivity a lot.

The quantities

$$R = \rho\frac{l}{A}, \qquad \kappa = \frac{1}{\rho}, \qquad G = \frac{1}{R}$$

QUANTITY	SYMBOL	SI UNIT
Resistance	R	Ω (ohm)
Resistivity	ρ	Ω m
Conductance	G	S (siemens) = Ω^{-1}
Conductivity	κ (kappa)	S m^{-1}
Cell constant	$G^* = l/A$	m^{-1}

- Conductivity is the conductance of a solution between electrodes 1 m apart with 1 m^2 area.

Measuring conductivity

- A **conductivity cell** has two platinum electrodes coated with platinum black.

- Its resistance is measured with a **Wheatstone bridge**.
- An **AC** source is used, not DC. Direct current would electrolyse the solution and change its composition.
- The cell constant is found first, using a KCl solution of known conductivity:

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

Molar conductivity

- Conductivity depends on how many ions are present. To compare electrolytes fairly, we use **molar conductivity**:

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

- In SI units, Λ_m is in S m^2 mol^{-1}.
- In practice, κ is often in S cm^{-1} and c in mol L^{-1}. Then:

$$\Lambda_m \ (\mathrm{S\,cm^2\,mol^{-1}}) = \frac{\kappa \ (\mathrm{S\,cm^{-1}}) \times 1000}{c \ (\mathrm{mol\,L^{-1}})}$$

- The 1000 converts litres into cm^3.
- 1 S m^2 mol^{-1} = 10^4 S cm^2 mol^{-1}.

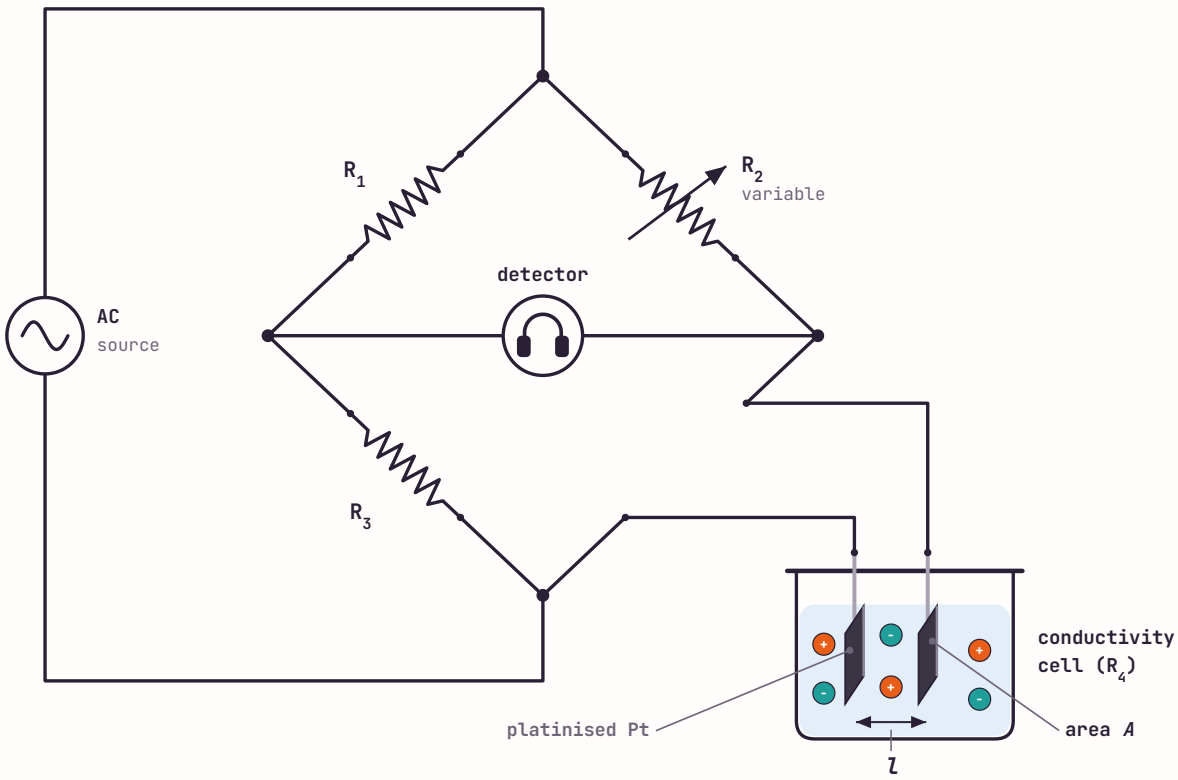


Fig 6 · Conductivity cell in a Wheatstone bridge. AC, not DC, so the solution is not electrolysed while we measure it.



TRAP ALERT

TRAP ALERT

- Two unit traps from our NEET analysis. First: use κ in S cm^{-1} before multiplying by 1000.
- Second: a cell constant given in m^{-1} must be converted. 100 m^{-1} = 1 cm^{-1}.

Q7 Worked question

NEET LEVEL 2

A conductivity cell has a cell constant of 1.25 cm^{-1}. Filled with a 0.05 M solution, its resistance is 250 Ω. Find the molar conductivity in S cm^2 mol^{-1}.

1. 0.10
2. 100
3. 10,000
4. 5.0

Answer: 2

CORE PRINCIPLE $\kappa = G^{*}/R$. Then $\Lambda_m = \kappa \times 1000/c$.

MATH $\kappa = \frac{1.25}{250} = 0.005$ S cm^{-1}. $\Lambda_m = \frac{0.005 \times 1000}{0.05} = 100$ S cm^2 mol^{-1}.

TRAP Option 1 forgets the 1000. Option 3 uses the cell constant as 125 m^{-1} without converting. Option 4 multiplies κ by 1000 but forgets to divide by c.

READ NEXT **Current Electricity notes: resistance, resistivity and the Wheatstone bridge** →

07 How Conductivity Changes with Concentration, and Kohlrausch's Law

IN NEET AND JEE FREQUENT *Kohlrausch's law, 12.5% of NEET questions*

Two opposite trends on dilution

- **Conductivity falls on dilution.** There are fewer ions per unit volume to carry current.
- **Molar conductivity rises on dilution.** It is the conductance of all the ions from one mole. On dilution, those ions move more freely, and weak electrolytes dissociate more.
- As concentration approaches zero, Λ_m approaches the **limiting molar conductivity**, Λ_m°.

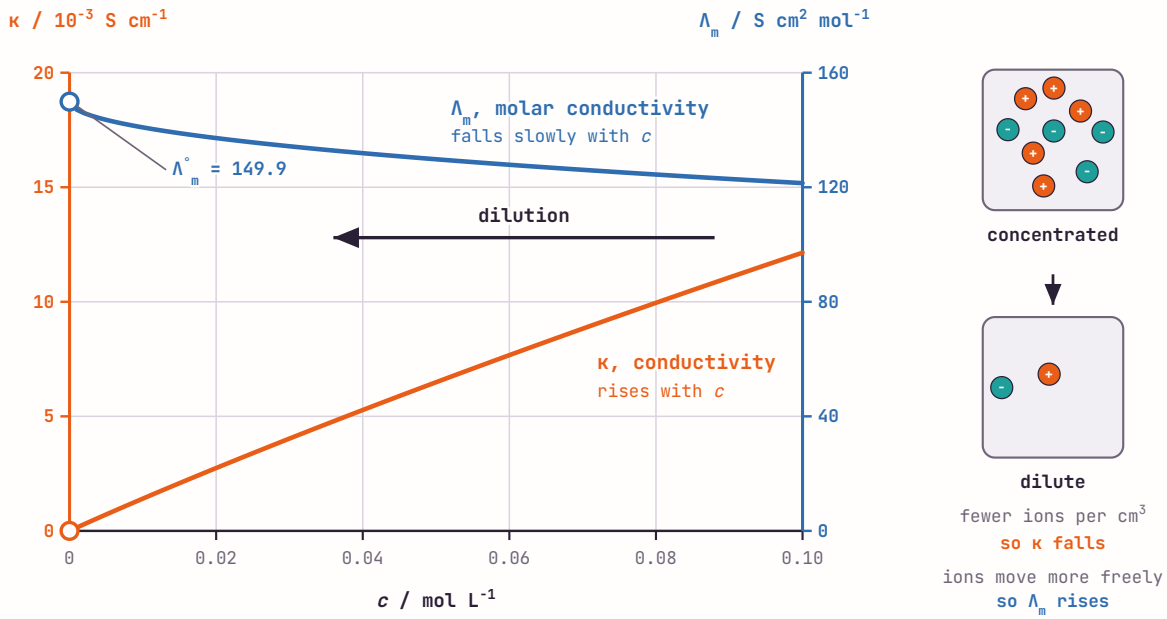


Fig 7 · Dilution trends. Dilute a solution: conductivity falls, molar conductivity rises.


TRAP ALERT

TRAP ALERT

- "On dilution, conductivity and molar conductivity both rise" is false. Our NEET analysis lists this dilution assumption among the top five traps.

Strong electrolytes

- Λ_m rises only slowly on dilution.
- It follows a straight line against $\sqrt{c}$:

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

- A depends on the type of electrolyte: the charges on its cation and anion (for example NaCl, $CaCl_2$, $MgSO_4$).
- Extrapolate the line to $c = 0$. The intercept is Λ_m°. The slope is $-A$.

Weak electrolytes

- Weak electrolytes (like acetic acid) have low Λ_m at high concentration.
- On dilution, Λ_m rises **steeply**, because the electrolyte dissociates more.
- The curve shoots up near zero. So Λ_m° **cannot** be found by extrapolation.
- The degree of dissociation is:

$$\alpha = \frac{\Lambda_m}{\Lambda_m^\circ}$$

- The dissociation constant then follows:

$$K_a = \frac{c\alpha^2}{1-\alpha}$$

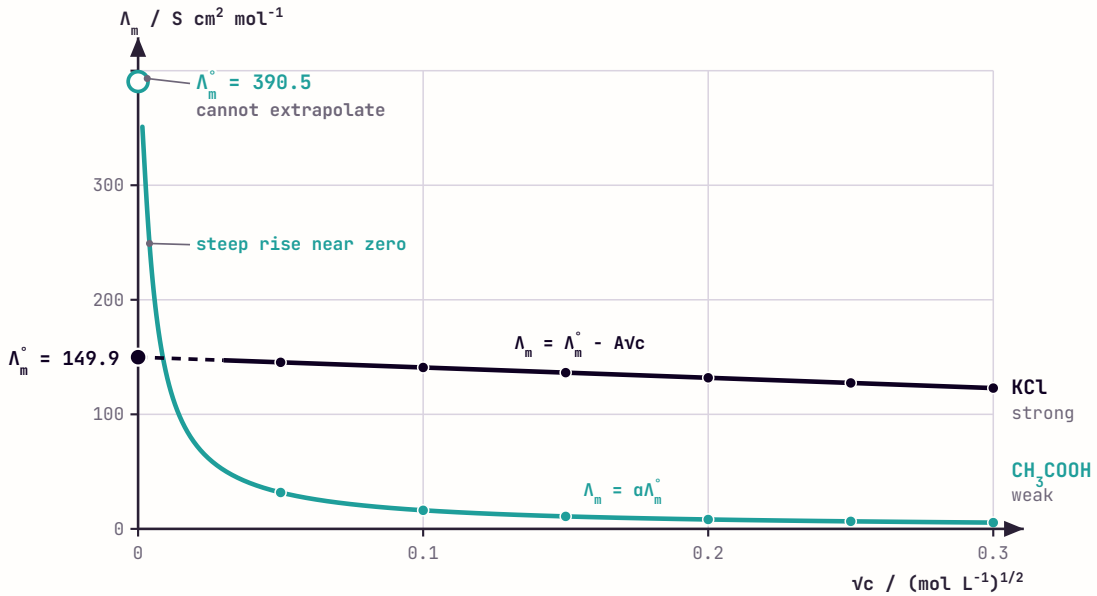


Fig 8 · Λ_m against $\sqrt{c}$. A strong electrolyte gives a line; a weak one gives a curve that shoots up.

Kohlrausch's law of independent migration of ions

- At infinite dilution, every ion adds its own fixed share to the molar conductivity:

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

- ν_+ and ν_- are the numbers of cations and anions per formula unit.
- λ_+° and λ_-° are the limiting molar conductivities of the ions.

Limiting molar conductivities of ions at 298 K (NCERT values, S cm^2 mol^{-1})

ION	λ°	ION	λ°
H^+	349.6	OH^-	199.1
Na^+	50.1	Cl^-	76.3
K^+	73.5	Br^-	78.1
Ca^{2+}	119.0	CH_3COO^-	40.9
Mg^{2+}	106.0	SO_4^{2-}	160.0

- **Example:** $\Lambda_m^\circ(CH_3COOH) = 349.6 + 40.9 = 390.5$ S cm^2 mol^{-1}.

- **Why it matters for weak electrolytes:** combine strong electrolytes so that the unwanted ions cancel:

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

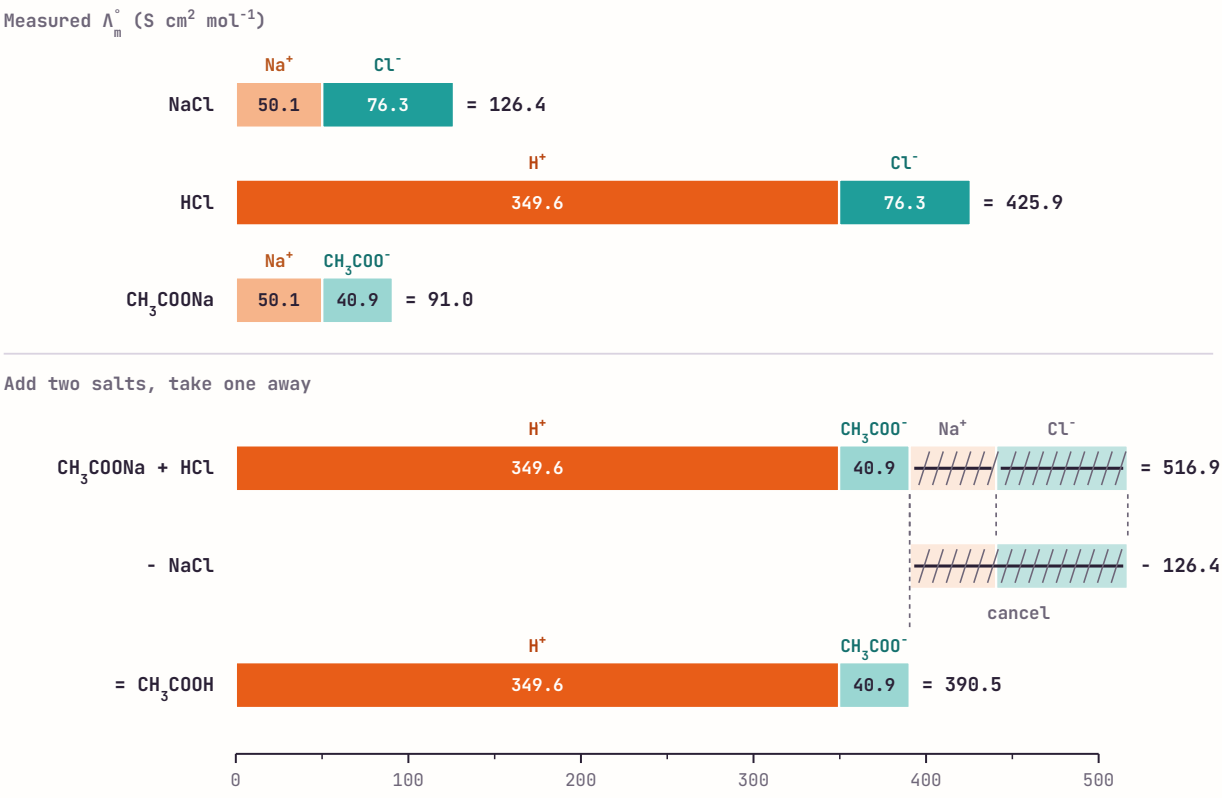
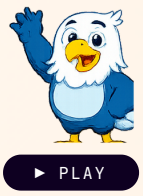


Fig 9 · Kohlrausch's law as bars. Add and subtract salts so that only the ions you want are left.



HAVING DIFFICULTY SEEING WHY THE Na^+ AND Cl^- CANCEL?

Play it.

This chapter is a set of short games in the Logic Bloom app.
logicbloom.co.in/download · Free to start.

SCAN

Q8 Worked question

NEET JEE MAIN LEVEL 2

Which statements about conductance are correct?

(a) The conductivity of a solution falls on dilution.

(b) The molar conductivity of a solution rises on dilution.

(c) For KCl, a plot of Λ_m against $\sqrt{c}$ is nearly a straight line.

(d) Λ_m° of acetic acid can be found by extrapolating its Λ_m against $\sqrt{c}$ graph.

(e) On dilution, the number of ions per unit volume increases.

Choose the correct answer from the options given below.

1. (a), (d) and (e) only
2. (b), (c) and (d) only
3. (a), (b), (c) and (d) only
4. (a), (b) and (c) only

Answer: 4

CORE PRINCIPLE Dilution lowers ions per volume, so κ falls. It frees the ions from one mole, so Λ_m rises. Weak electrolytes curve sharply near zero.

MATH

TRAP (d) fails because the weak-electrolyte curve cannot be extended to zero. (e) has the dilution effect backwards.

Q9 Worked question

JEE MAIN LEVEL 2

Given Λ_m° values (S cm^2 mol^{-1}): NH_4Cl = 129.8, NaOH = 248.1, NaCl = 126.4. Find Λ_m° for NH_4OH.

1. 504.3
2. 244.7
3. 251.5
4. 377.9

Answer: 3

CORE PRINCIPLE Choose the combination that leaves $NH_4^+ + OH^-$ and cancels $Na^+ + Cl^-$.

MATH $129.8 + 248.1 - 126.4 = 251.5$ S cm^2 mol^{-1}.

TRAP Option 1 adds all three. Option 2 subtracts the wrong salt. Option 4 forgets to remove NaCl.

08 Electrolysis and Faraday's Laws

IN NEET AND JEE | TOP TOPIC | *25% of NEET questions, the largest block in this chapter*

What happens

- In an **electrolytic cell**, an outside source of electricity drives a reaction.
- Example: electroplating copper. Copper from the anode dissolves as Cu^{2+}. Cu^{2+} at the cathode is deposited as copper metal.

Faraday's laws (1833)

- **First law:** the amount of chemical reaction at an electrode is proportional to the quantity of electricity passed.
- **Second law:** pass the same quantity of electricity through different electrolytes. The amounts of substances liberated are in proportion to their **chemical equivalent weights**.
  - Equivalent weight = atomic mass of the metal ÷ number of electrons needed to reduce its cation.
- Charge and current: $Q = It$, with I in amperes and t in **seconds**.
- **One faraday** is the charge on one mole of electrons:

$$1\,F = 96487\ \text{C mol}^{-1} \approx 96500\ \text{C mol}^{-1}$$

The working method

1. Find the charge: $Q = It$.
2. Find the moles of electrons: Q/F.
3. Use the half-reaction to change moles of electrons into moles of product.
4. Change moles into mass (or into gas volume).

- One mole of electrons deposits 1 mol of Ag ($Ag^+ + e^-$).
- The same charge deposits only $\frac{1}{2}$ mol of Cu ($Cu^{2+} + 2e^-$), or $\frac{1}{3}$ mol of Al ($Al^{3+} + 3e^-$).

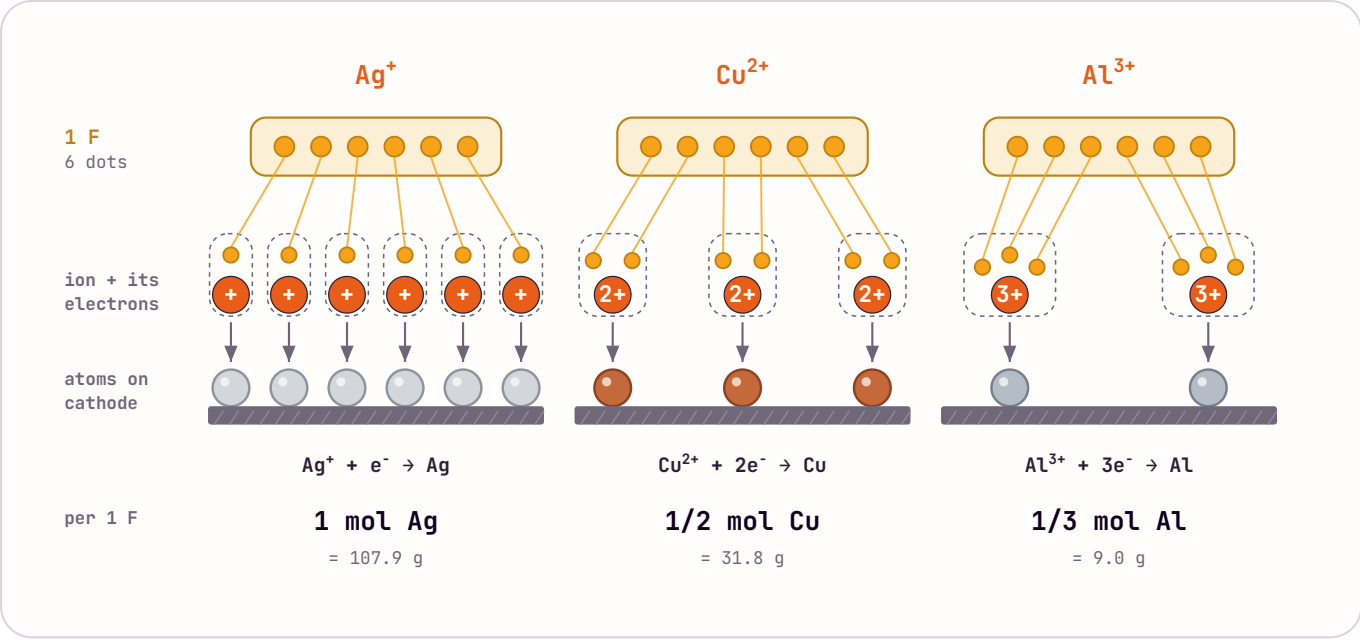


Fig 10 · One faraday, three metals. The same charge deposits fewer atoms when each ion needs more electrons.



TRAP ALERT

TRAP ALERT

- One mole of a metal is not one faraday. Ca^{2+} needs **2 F** per mole of calcium. Our NEET analysis calls this n-factor slip a top mark-loser, and a 2020 question tested it directly.



TRAP ALERT

TRAP ALERT

- Time must be in seconds in $Q = It$. Convert minutes and hours first.

Q10 Worked question

NEET LEVEL 2

A current of 2.0 A flows through a $CuSO_4$ solution for 16 minutes 5 seconds. What mass of copper deposits at the cathode? (Cu = 63.5 g mol^{-1}, F = 96500 C mol^{-1})

1. 1.27 g
2. 0.635 g
3. 0.0106 g
4. 63.5 g

Answer: 2

CORE PRINCIPLE Charge → moles of electrons → moles of Cu (2 electrons each) → mass.

MATH $t = 965$ s. $Q = 2.0 \times 965 = 1930$ C $= 0.02$ F. Moles of Cu $= 0.01$. Mass $= 0.01 \times 63.5 = 0.635$ g.

TRAP Option 1 forgets Cu^{2+} needs 2 electrons. Option 3 uses 16.05 as seconds instead of converting. Option 4 assumes one mole deposits.

Q11 Worked question

NEET LEVEL 2

What charge is needed to deposit 20 g of calcium from molten $CaCl_2$? (Ca = 40 g mol^{-1}, F = 96500 C mol^{-1})

1. 96,500 C
2. 48,250 C
3. 193,000 C
4. 9,650 C

Answer: 1

CORE PRINCIPLE $Ca^{2+} + 2e^- \rightarrow Ca$. So each mole of calcium needs 2 F.

MATH 20 g = 0.5 mol Ca. Charge = $0.5 \times 2 \times 96500 = 96{,}500$ C.

TRAP Option 2 uses 1 F per mole. Option 3 forgets that 20 g is half a mole. Option 4 slips a decimal.

READ NEXT **mole calculations for electrolysis** →

09 Products of Electrolysis

IN NEET AND JEE | TOP TOPIC | *preferential discharge is in our NEET 2026 predictions*

- The products depend on the substance electrolysed, and on the electrodes.
- **Inert electrodes** (platinum, gold) take no part in the reaction. They only carry electrons.
- **Reactive electrodes** (like copper) can take part.

Molten NaCl

- Only Na^+ and Cl^- are present.
- Cathode: $Na^+ + e^- \rightarrow Na$. Anode: $Cl^- \rightarrow \frac{1}{2}Cl_2 + e^-$.
- So molten NaCl gives **sodium metal** and **chlorine**.

Aqueous NaCl

- Now water competes at both electrodes.
- **At the cathode**, two reductions compete:
  - $Na^+ + e^- \rightarrow Na$, $E^\circ = -2.71$ V
  - $H^+ + e^- \rightarrow \frac{1}{2}H_2$, $E^\circ = 0.00$ V (the H^+ comes from water)
  - The one with the **higher** E° wins. So **H_2** forms, not sodium.
- **At the anode**, two oxidations compete:
  - $Cl^- \rightarrow \frac{1}{2}Cl_2 + e^-$, $E^\circ = 1.36$ V
  - $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$, $E^\circ = 1.23$ V
  - By E° alone, O_2 should form. But oxygen needs an extra push, called **overpotential**. So **Cl_2** forms instead.
- Net result:

$$NaCl(aq) + H_2O(l) \rightarrow Na^+(aq) + OH^-(aq) + \frac{1}{2}H_2(g) + \frac{1}{2}Cl_2(g)$$

- The solution left behind contains NaOH.

Sulphuric acid (inert electrodes)

- **Dilute** H_2SO_4: water is oxidised at the anode, giving **O_2**.

- **Concentrated** H_2SO_4: sulphate ions are oxidised instead, giving **peroxodisulphate**: $2SO_4^{2-} \rightarrow S_2O_8^{2-} + 2e^-$, $E^\circ = 1.96$ V.
- Hydrogen forms at the cathode in both cases.

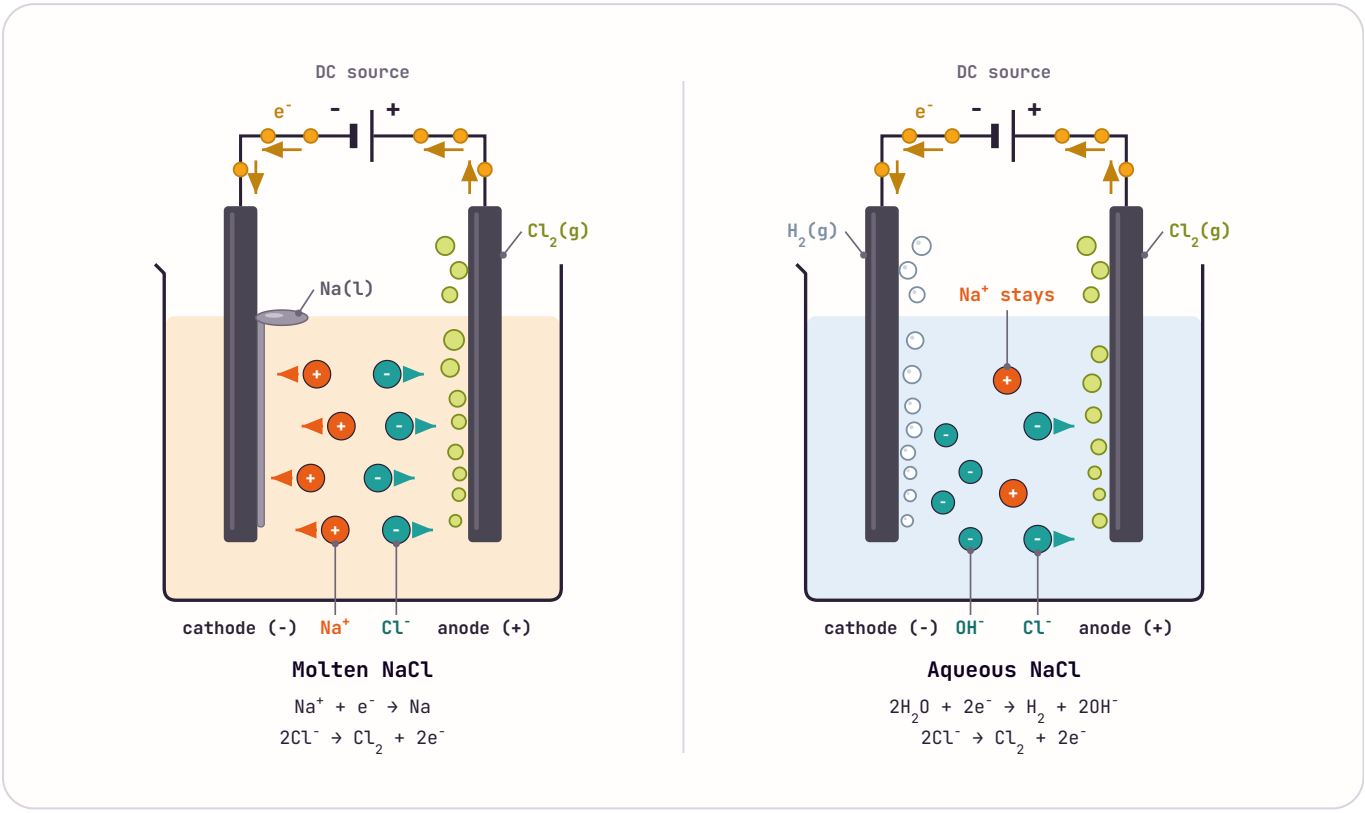


Fig 11 · Electrolysis of NaCl. Add water, and hydrogen replaces sodium at the cathode.



WAIT, REALLY?

MYTH

In electrolysis, the reaction with the more favourable E° always wins.

SCIENCE

Overpotential can change the winner. In brine, Cl_2 forms at the anode even though O_2 has the lower oxidation potential.

Q12 Worked question

NEET JEE MAIN LEVEL 2

Match each electrolysis (inert electrodes) with its products at the cathode and anode.

COLUMN I	COLUMN II (CATHODE, ANODE)
A. Molten NaCl	I. H_2, O_2
B. Aqueous NaCl	II. Na, Cl_2
C. Dilute H_2SO_4	III. H_2, $S_2O_8^{2-}$
D. Concentrated H_2SO_4	IV. H_2, Cl_2

Choose the correct answer from the options given below.

1. A-IV, B-II, C-I, D-III
2. A-II, B-IV, C-III, D-I
3. A-II, B-I, C-IV, D-III
4. A-II, B-IV, C-I, D-III

Answer: 4

CORE PRINCIPLE Water wins at the cathode whenever it is present. Overpotential gives Cl_2 in brine. Concentration decides the H_2SO_4 anode.

MATH

TRAP Option 1 swaps molten and aqueous NaCl. Option 2 swaps dilute and concentrated acid.

10 Batteries and Fuel Cells

IN NEET AND JEE RARE *6.3% of NEET questions, together with corrosion*

Green badge for the dry cell, lead storage battery and fuel cell. The mercury cell and nickel-cadmium cell are marked JEE Advanced mainly.

Primary batteries (cannot be recharged)

Dry cell (Leclanché cell)

- Anode: a **zinc** container.
- Cathode: a **carbon (graphite) rod** surrounded by powdered MnO_2 and carbon.
- The space between is filled with a moist paste of **NH_4Cl and $ZnCl_2$**.
- Anode: $Zn \rightarrow Zn^{2+} + 2e^-$
- Cathode: $MnO_2 + NH_4^+ + e^- \rightarrow MnO(OH) + NH_3$
- Potential: about **1.5 V**.

Mercury cell JEE ADVANCED MAINLY

- Used in small devices like hearing aids and watches.
- Anode: zinc-mercury amalgam. Cathode: a paste of HgO and carbon. Electrolyte: a paste of KOH and ZnO.
- Overall: $Zn(Hg) + HgO(s) \rightarrow ZnO(s) + Hg(l)$
- Potential: about **1.35 V**. It stays **constant** through the cell's life. The overall reaction has no ion whose concentration changes.

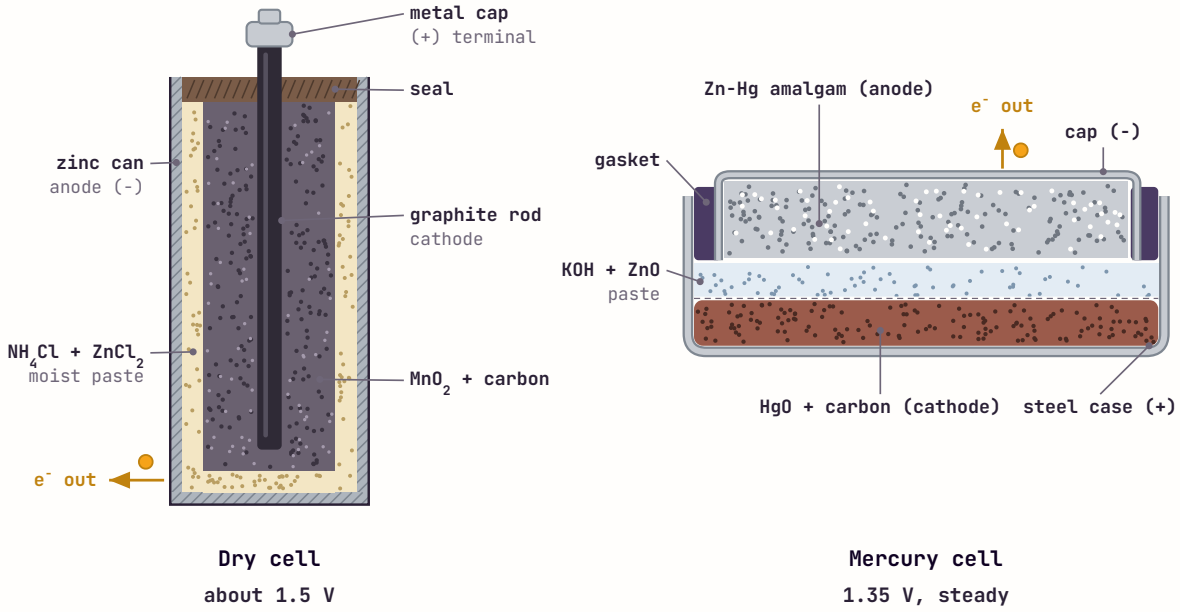


Fig 12 · Dry cell and mercury cell. Two primary cells; the mercury cell holds a steady 1.35 V.

Secondary batteries (can be recharged)

Lead storage battery (lead accumulator)

- Used in cars and inverters.
- Anode: **lead**. Cathode: a grid of lead packed with **PbO_2**. Electrolyte: **38% sulphuric acid**.
- On discharge:
  - Anode: $Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$
  - Cathode: $PbO_2 + SO_4^{2-} + 4H^+ + 2e^- \rightarrow PbSO_4 + 2H_2O$
  - Overall: $Pb + PbO_2 + 2H_2SO_4 \rightarrow 2PbSO_4 + 2H_2O$
- On charging, the reaction runs in reverse. $PbSO_4$ at each electrode turns back into Pb and PbO_2.



THINK IT THROUGH

SANITY CHECK

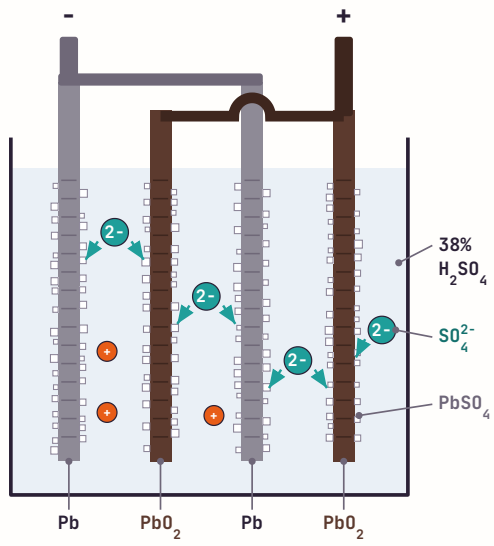
- As a lead battery discharges, sulphuric acid is used up and water forms. So the acid's density falls. That is how a battery's charge is checked.

Nickel-cadmium cell JEE ADVANCED MAINLY

- Longer life than the lead storage cell, but more expensive.
- Overall on discharge: $Cd + 2Ni(OH)_3 \rightarrow CdO + 2Ni(OH)_2 + H_2O$

Fuel cells

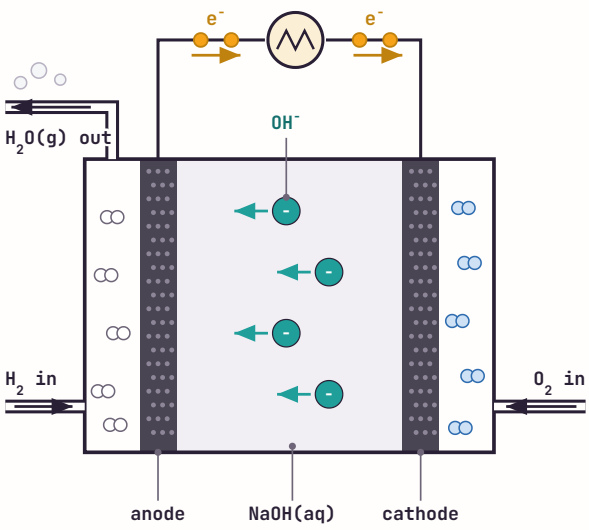
- A **fuel cell** turns the energy of burning a fuel, like hydrogen, directly into electricity.
- The H_2-O_2 fuel cell was used in the **Apollo space programme**. The water it made was used for drinking by the astronauts.
- Electrodes: porous carbon with a catalyst such as finely divided platinum or palladium. Electrolyte: concentrated aqueous NaOH.
- Cathode: $O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$
- Anode: $2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$
- Overall: $2H_2 + O_2 \rightarrow 2H_2O$
- Efficiency: about **70%**, compared with about **40%** for thermal power plants.
- Fuel cells run as long as fuel is supplied. They are pollution-free.



Lead storage battery

$Pb + SO_4^{2-} \rightarrow PbSO_4 + 2e^-$

$PbO_2 + SO_4^{2-} + 4H^+ + 2e^- \rightarrow PbSO_4 + 2H_2O$



H_2-O_2 fuel cell

$2H_2 + 4OH^- \rightarrow 4H_2O + 4e^-$

$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$

Fig 13 · Lead battery and fuel cell. One stores energy and is recharged; the other runs as long as fuel flows.

Q13 Worked question

NEET LEVEL 2

COLUMN I	COLUMN II
A. Dry cell	I. Moist NH_4Cl and $ZnCl_2$ paste around MnO_2 and carbon
B. Mercury cell	II. Used in Apollo; its product was drinking water
C. Lead storage battery	III. 38% H_2SO_4; can be recharged
D. H_2-O_2 fuel cell	IV. About 1.35 V, constant through its life

Choose the correct answer from the options given below.

1. A-III, B-IV, C-I, D-II
2. A-I, B-IV, C-III, D-II
3. A-I, B-II, C-III, D-IV
4. A-IV, B-I, C-III, D-II

Answer: 2

CORE PRINCIPLE Match each device to its electrolyte or its signature fact.

MATH

TRAP Option 1 swaps the dry cell and the lead battery. Option 3 swaps the mercury cell and the fuel cell.

11 Corrosion

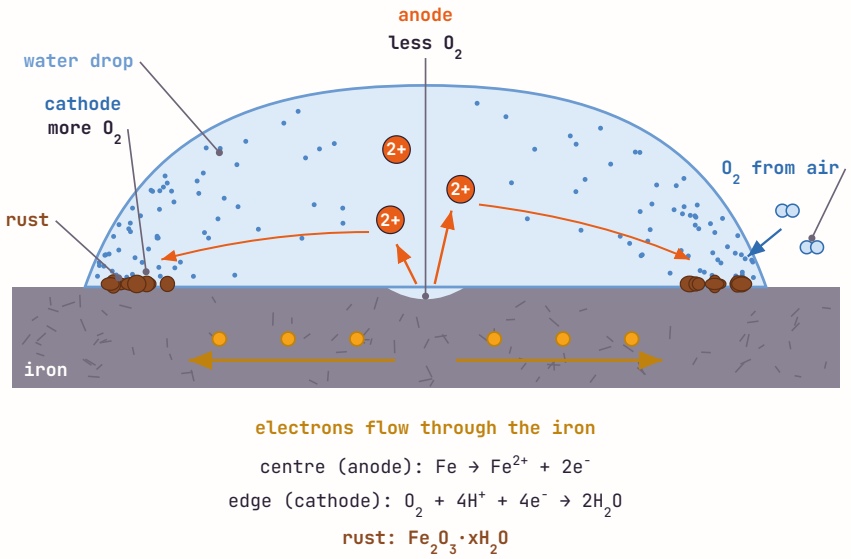
JEE ADVANCED MAINLY

In NCERT and the board syllabus; not named in the NEET 2026 or JEE Main syllabus.

- **Corrosion** slowly coats a metal surface with oxides or other salts of the metal.
- Examples: rusting of iron, tarnishing of silver, a green coating on copper and bronze.
- **Rusting is an electrochemical process.** A water droplet on iron acts as a tiny cell.
  - **Anode** (a spot where iron oxidises): $2Fe \rightarrow 2Fe^{2+} + 4e^-$, $E^\circ(Fe^{2+}/Fe) = -0.44$ V
  - **Cathode** (another spot on the surface): $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$, $E^\circ = 1.23$ V
  - The H^+ comes from H_2CO_3, formed when CO_2 from the air dissolves in water.
  - Overall: $2Fe + O_2 + 4H^+ \rightarrow 2Fe^{2+} + 2H_2O$, $E^\circ_{cell} = 1.67$ V
- Fe^{2+} is oxidised further by air to rust, $Fe_2O_3 \cdot xH_2O$.

Preventing corrosion

- **Barrier protection:** paint, or chemicals like bisphenol, keep air and water away.
- **Coating with other metals:** tin or zinc covers the iron.
- **Sacrificial electrode:** a more reactive metal (Mg, Zn) corrodes first and saves the iron. **Galvanising** coats iron with zinc for this reason.



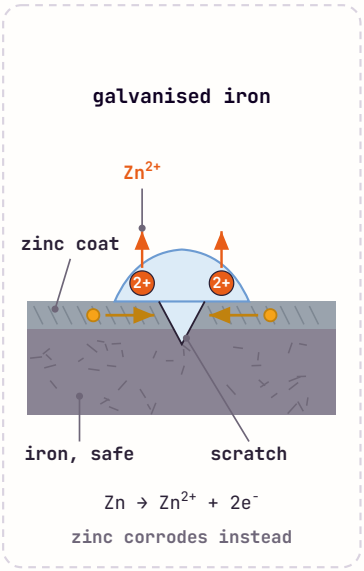


Fig 14 · A water drop on iron. Rusting is a tiny cell; zinc gives itself up to save the iron.



THINK IT THROUGH

SANITY CHECK

- Zinc protects iron even if the zinc coat is scratched. Zinc has the lower E°, so it oxidises first.

12 JEE Advanced Challenge Set

JEE ADVANCED MAINLY

- Four problems. Each needs a constructed path, a second chapter, calculus, or a result that defies the first guess. Take 6 to 8 minutes each.

Q14 Worked question

JEE ADVANCED · NUMERICAL · LEVEL 3

At 298 K, $E^\circ(\mathrm{Ag^+/Ag}) = +0.80$ V and $E^\circ(\mathrm{AgCl/Ag, Cl^-}) = +0.22$ V. Take $\frac{2.303RT}{F} = 0.059$ V. Find $\mathrm{p}K_{sp} = -\log K_{sp}$ for AgCl, correct to two decimal places.

Answer: 9.83

CORE PRINCIPLE Build a cell whose overall reaction is the dissolving of AgCl. Then its E° gives K_{sp} through $\log K = \frac{nE^\circ}{0.059}$.

MATH Cathode: $\mathrm{AgCl} + e^- \rightarrow \mathrm{Ag} + \mathrm{Cl^-}$ (0.22 V). Anode: $\mathrm{Ag} \rightarrow \mathrm{Ag^+} + e^-$ (reduction potential 0.80 V). Sum: $\mathrm{AgCl} \rightarrow \mathrm{Ag^+} + \mathrm{Cl^-}$, which is the K_{sp} reaction. $E^\circ = 0.22 - 0.80 = -0.58$ V. $\log K_{sp} = \frac{1 \times (-0.58)}{0.059} = -9.83$. So $\mathrm{p}K_{sp} = 9.83$.

TRAP The overall reaction is not a redox reaction, so students assume electrochemistry cannot help. The negative E° is not an error: AgCl dissolving is not spontaneous at standard conditions.

Q15 Worked question

JEE ADVANCED · ONE OR MORE CORRECT · LEVEL 3

Aqueous NaCl is electrolysed between inert platinum electrodes. Which of the following is/are correct? (One or more of the following may be correct.)

(A) Hydrogen gas forms at the cathode.

(B) Chlorine forms at the anode, although water has the lower standard oxidation potential, because of oxygen's overpotential.

(C) The solution becomes more basic as electrolysis goes on.

(D) If molten NaCl were used instead, hydrogen would still form at the cathode.

Answer: (A), (B) and (C)

CORE PRINCIPLE Compare competing half-reactions at each electrode. Then allow for overpotential at the anode.

MATH Cathode: H^+ (0.00 V) beats Na^+ (−2.71 V), so H_2. Anode: O_2 (1.23 V) should win over Cl_2 (1.36 V), but overpotential makes Cl_2 form. Na^+ and OH^- stay behind, so the solution turns basic.

TRAP (D) forgets that molten NaCl has no water. Sodium metal forms there. Students who skip overpotential reject (B).

Q16 Worked question

JEE ADVANCED NUMERICAL LEVEL 3

For a cell reaction with $n = 2$, $E_{cell} = 1.10$ V at 298 K, and its temperature coefficient is $\left(\frac{\partial E}{\partial T}\right)_P = -4.0 \times 10^{-4}$ V K^{-1}. Take $F = 96500$ C mol^{-1}. Find ΔH for the reaction in kJ mol^{-1}, to the nearest integer.

Answer: −235

CORE PRINCIPLE This joins electrochemistry with thermodynamics. $\Delta G = -nFE$. Since $\Delta S = -\frac{\partial(\Delta G)}{\partial T}$, we get $\Delta S = nF\frac{\partial E}{\partial T}$. Then $\Delta H = \Delta G + T\Delta S$.

MATH $\Delta G = -2 \times 96500 \times 1.10 = -212.3$ kJ mol^{-1}. $\Delta S = 2 \times 96500 \times (-4.0 \times 10^{-4}) = -77.2$ J K^{-1} mol^{-1}. $\Delta H = -212.3 + 298 \times (-0.0772) = -212.3 - 23.0 \approx -235$ kJ mol^{-1}.

TRAP Using $\Delta H = \Delta G - T\Delta S$ gives −189. Leaving ΔS in J while ΔG is in kJ gives a wildly wrong number. The sign of $\frac{\partial E}{\partial T}$ sets the sign of ΔS.

Q17 Worked question

JEE ADVANCED LEVEL 3

A conductivity cell has a cell constant of 1.00×10^2 m^{-1}. Filled with 1.0×10^{-3} M acetic acid at 298 K, its resistance is 2.08×10^4 Ω. Use $\lambda^\circ(H^+) = 349.6$ and $\lambda^\circ(CH_3COO^-) = 40.9$ S cm^2 mol^{-1}. Use $K_a = \frac{c\alpha^2}{1-\alpha}$. Find K_a of acetic acid.

1. 1.52×10^{-5}
2. 1.52×10^{-11}
3. 1.73×10^{-5}
4. It cannot be found, because α comes out greater than 1

Answer: 3

CORE PRINCIPLE This joins conductance with ionic equilibrium. Find κ, then Λ_m, then α from Kohlrausch's law, then K_a.

MATH $G^* = 100$ m^{-1} $= 1.00$ cm^{-1}. $\kappa = \frac{1.00}{2.08 \times 10^4} = 4.81 \times 10^{-5}$ S cm^{-1}. $\Lambda_m = \frac{4.81 \times 10^{-5} \times 1000}{1.0 \times 10^{-3}} = 48.1$ S cm^2 mol^{-1}. $\Lambda_m^\circ = 390.5$. $\alpha = \frac{48.1}{390.5} = 0.123$. $K_a = \frac{10^{-3} \times (0.123)^2}{1 - 0.123} = 1.73 \times 10^{-5}$.

TRAP Option 4 appears if 100 m^{-1} is read as 100 cm^{-1}. That gives $\alpha \approx 12$, which is impossible. Option 2 forgets the 1000. Option 1 drops $(1 - \alpha)$, even though $\alpha = 0.12$ is not small enough to ignore.

13 Quick Review: What NEET Repeats

Frequency bands come from our NEET analysis. JEE data for this chapter is coming.

TOP TOPIC — FARADAY'S LAWS

Mass deposited by charge Q?

Moles of e^- = Q/F; divide by electrons per ion; times molar mass.

Why it recurs: A quarter of all questions.

TOP TOPIC — FARADAY'S LAWS

Faradays per mole of Ca from Ca^{2+}?

2 F.

Why it recurs: A named trap; tested in 2020.

TOP TOPIC — FARADAY'S LAWS

Unit of t in $Q = It$?

Seconds.

Why it recurs: A named trap.

TOP TOPIC — ELECTROLYSIS PRODUCTS

Aqueous NaCl: cathode and anode products?

H_2 and Cl_2.

Why it recurs: Preferential discharge is predicted.

TOP TOPIC — ELECTROLYSIS PRODUCTS

Why Cl_2 and not O_2 at the anode in brine?

Overpotential of oxygen.

Why it recurs: The classic "E° says otherwise" question.

FREQUENT — ELECTRODE POTENTIAL

E°_{cell} formula?

$E^\circ_{cathode} - E^\circ_{anode}$, both as reduction potentials.

Why it recurs: A named subtraction trap.

FREQUENT — ELECTRODE POTENTIAL

Strongest oxidising and reducing agents in NCERT's table?

F_2 (+2.87 V) and Li (−3.05 V).

Why it recurs: Series-reading questions.

FREQUENT — ELECTRODE POTENTIAL

Does E° double when the half-equation is doubled?

No. It is intensive.

Why it recurs: A common hidden trap.

FREQUENT — GIBBS ENERGY

$\Delta_r G^\circ$ from E°?

$-nFE^\circ$, in J; divide by 1000 for kJ.

Why it recurs: Asked almost every year since 2019.

FREQUENT — GIBBS ENERGY

What sign of E°_{cell} means spontaneous?

Positive, so ΔG° is negative.

Why it recurs: The spontaneity link.

FREQUENT — CONDUCTANCE

Cell constant?

$G^* = l/A = R\kappa$.

Why it recurs: Tested in 2023; due again.

FREQUENT — CONDUCTANCE

Λ_m from κ in S cm^{-1}?

$\kappa \times 1000/c$.

Why it recurs: A named unit trap.

FREQUENT — CONDUCTANCE

What happens to κ and Λ_m on dilution?

κ falls; Λ_m rises.

Why it recurs: A named dilution trap.

FREQUENT — KOHLRAUSCH'S LAW

Λ°_m of CH_3COOH from strong electrolytes?

CH_3COONa + HCl − NaCl.

Why it recurs: Predicted for weak electrolytes.

FREQUENT — KOHLRAUSCH'S LAW

Degree of dissociation from conductance?

$\alpha = \Lambda_m/\Lambda^\circ_m$.

Why it recurs: Links conductance to equilibrium.

REGULAR — NERNST EQUATION

What happens to coefficients in Q?

They become powers.

Why it recurs: A named exponent trap; 2022.

REGULAR — NERNST EQUATION

Value of $2.303RT/F$ at 298 K?

0.059 V.

Why it recurs: Every Nernst numerical uses it.

REGULAR — EQUILIBRIUM CONSTANT

$\log K$ from E°?

$nE^\circ/0.059$.

Why it recurs: Last tested 2019; due for a return.

RARE — BATTERIES

Electrolyte in a lead storage battery?

38% H_2SO_4.

Why it recurs: Device facts reappear in match questions.

RARE — FUEL CELLS

Product of the H_2-O_2 fuel cell, and its efficiency?

Water; about 70%.

Why it recurs: The Apollo fact.

★ Electrochemistry · Fact Sheet

One page for the night before the exam.

CELLS

Oxidation at anode, reduction at cathode, always

Galvanic: spontaneous, chemical → electrical; anode negative · Electrolytic: driven, anode positive

Daniell: Zn | Zn^{2+} || Cu^{2+} | Cu, 1.1 V · salt bridge: KCl or NH_4NO_3 in agar-agar

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$ · SHE = 0.00 V · Cu^{2+}/Cu +0.34 · Zn^{2+}/Zn −0.76

NERNST AND ENERGY (298 K)

$$E_{cell} = E^\circ_{cell} - \frac{0.059}{n}\log Q$$

Coefficients become powers

$\log K_c = \frac{nE^\circ}{0.059}$ · Daniell $K_c \approx 2 \times 10^{37}$

$\Delta_r G^\circ = -nFE^\circ = -2.303RT\log K$ · Daniell

$\Delta_r G^\circ \approx -212$ kJ mol^{-1}

F = 96487 C mol^{-1} (≈ 96500) · E intensive, ΔG extensive

CONDUCTANCE

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

$\kappa = G^* / R \cdot G^* = l/A$

1 S m^2 mol^{-1} = 10^4 S cm^2 mol^{-1} · 100 m^{-1} = 1 cm^{-1}

Dilution: κ down, Λ_m up · strong: $\Lambda_m = \Lambda^\circ_m - A\sqrt{c}$

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

Kohlrausch: $\Lambda^\circ_m = \nu_+\lambda^\circ_+ + \nu_-\lambda^\circ_-$ · H^+ 349.6, CH_3COO^- 40.9

ELECTROLYSIS

$$Q = It \ \textbf{(t in s)} \cdot \textbf{moles e}^- = Q/F$$

Ag^+: 1 F/mol · Cu^{2+}: 2 F/mol · Al^{3+}: 3 F/mol

Molten NaCl: Na, Cl_2 · Aqueous NaCl: H_2, Cl_2 (overpotential)

Dilute H_2SO_4: O_2 at anode · Concentrated: $S_2O_8^{2-}$

DEVICES

Dry cell ~1.5 V · mercury cell ~1.35 V, constant

Lead battery 38% H_2SO_4

H_2-O_2 fuel cell: NaOH, Pt/Pd catalyst, ~70% efficient, Apollo

Rust: Fe anode, O_2 cathode, E°_{cell} = 1.67 V · galvanising = zinc sacrificial coat

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