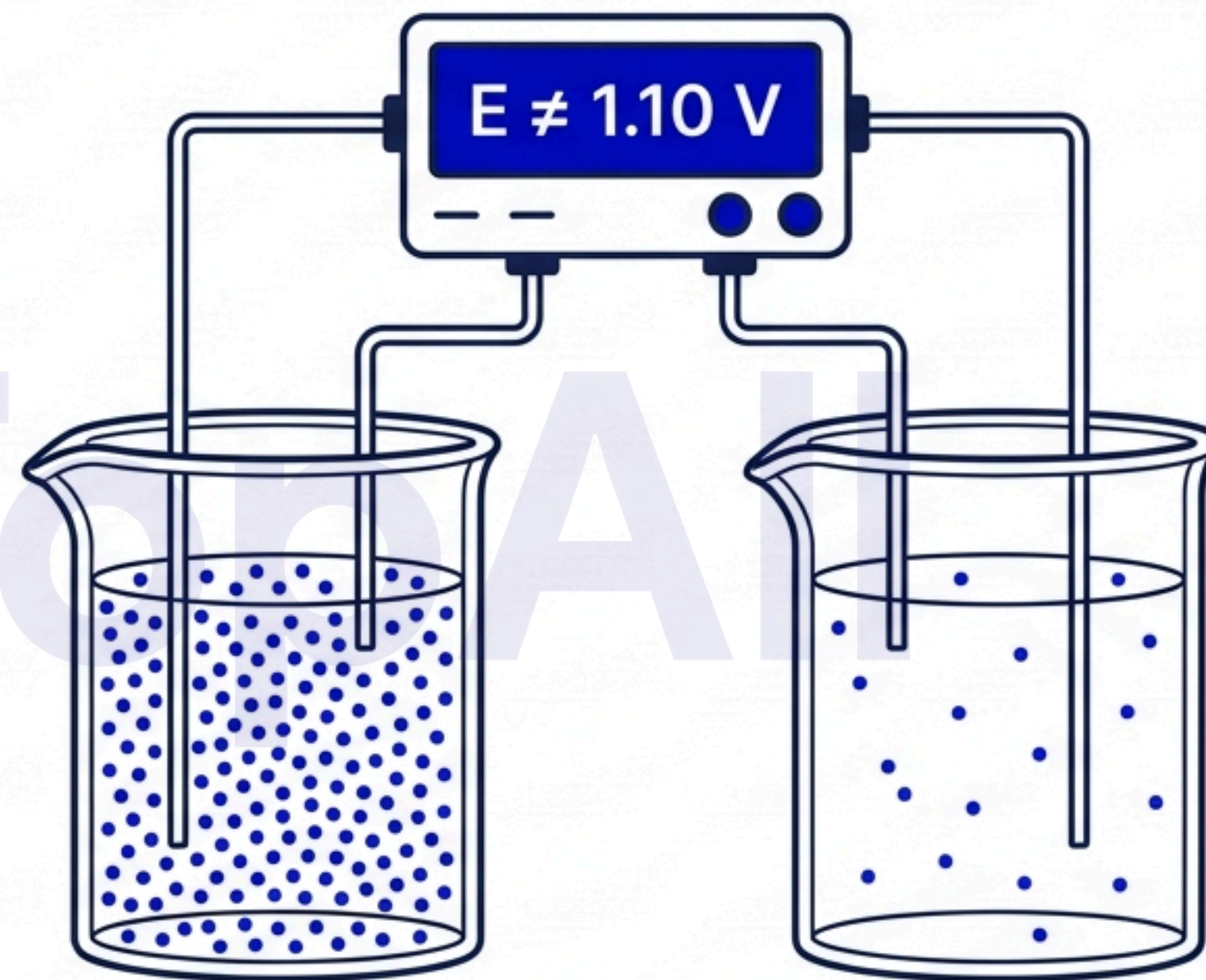


JEE MAIN + ADVANCED

Nernst Equation

Cell Potential Beyond
Standard Conditions



Why We Need Nernst

Standard Conditions



- Assumes strictly 1 M concentration for all aqueous species.
- Assumes 1 bar pressure for all gases.
- $E_{\text{cell}} = E^{\circ}_{\text{cell}}$ (Constant)

Real World Conditions

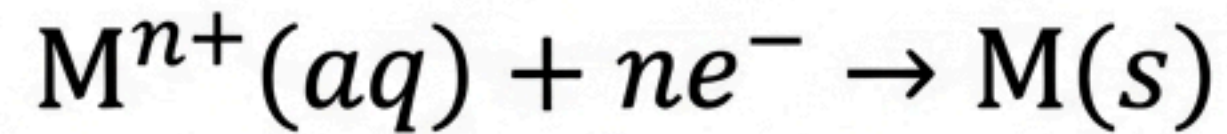


- Concentrations change as the cell operates.
- Reactants are consumed, products are formed.
- E_{cell} changes constantly until equilibrium ($E=0$).

The Nernst Equation bridges the gap, allowing us to calculate potential at ANY concentration.

General Electrode Equation

For a general reduction half-reaction:



STANDARD
ELECTRODE POTENTIAL

GAS CONSTANT
(8.314 J K⁻¹ mol⁻¹)

TEMPERATURE
IN KELVIN


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

MOLES OF ELECTRONS
TRANSFERRED

FARADAY CONSTANT
(~96487 C mol⁻¹)

MOLAR CONCENTRATION
OF THE ION



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$$E=mc^2$$



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The Universal Form

General Cell Equation

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

Q = Reaction Quotient

Standard Exam Form (at 298 K)

By substituting R, F, T=298K and converting $\ln$ to $\log_{10}$ (multiplier 2.303):

$$E_{\text{cell}} = E_{\text{cell}}^{\circ} - \frac{0.059}{n} \log Q$$

Use this form for 99% of JEE numericals at 298 K.

Building Q (Reaction Quotient)

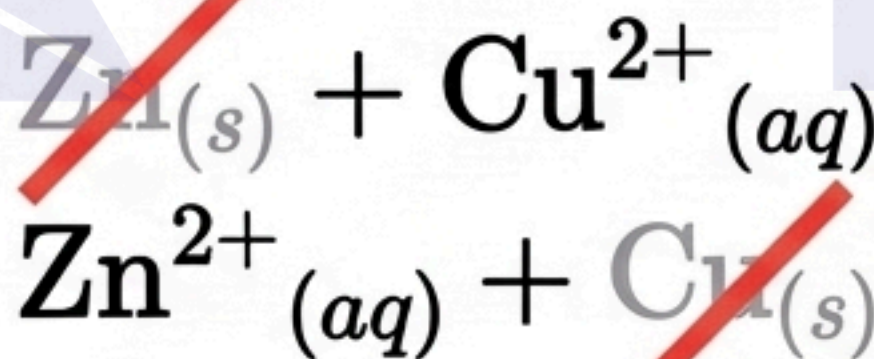
1. Write the Cell Reaction

Example:
The Daniell Cell



2. Eliminate Solids & Pure Liquids

Their concentration is taken as unity (1).



3. Formulate Q

Products over Reactants.

$$Q = \frac{[\text{Zn}^{2+}]}{[\text{Cu}^{2+}]}$$

The Concentration Effect

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



Increase Reactant

Increase $[\text{Cu}^{2+}]$

E_{cell} Rises

Q decreases, subtracted term becomes smaller (or negative).



Increase Product

Increase $[\text{Zn}^{2+}]$

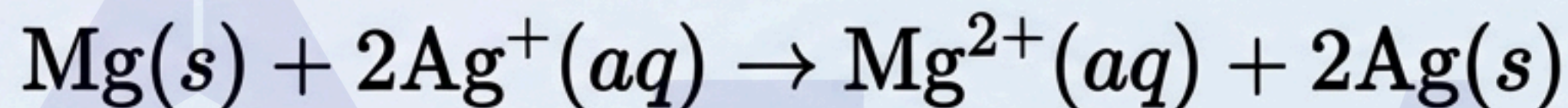
E_{cell} Falls

Q increases, subtracted term becomes larger.

JEE Example: Exponent Trap

Calculate E_{cell} for: $\text{Mg}(s) \mid \text{Mg}^{2+}(0.130M) \parallel \text{Ag}^{+}(0.0001M) \mid \text{Ag}(s)$

Given $E^{\circ}_{\text{cell}} = 3.17 \text{ V}$



Look at the stoichiometry!

$$Q = \frac{[\text{Mg}^{2+}]}{[\text{Ag}^{+}]^2}$$

Emphasize the exponent 2 on Ag^{+} .
This is the most common student error.

$$E = 3.17 - \frac{0.059}{2} \log \frac{0.130}{(0.0001)^2}$$

$$E_{\text{cell}} = 2.96 \text{ V}$$

Formula Card & Execution

The 5-Step Execution

1. Balance the overall cell reaction.
2. Find n (total electrons cancelled).
3. Write Q (Products / Reactants).
4. Remove pure solids and liquids ($= 1$).
5. Substitute into: $E = E^\circ - \frac{0.059}{n} \log Q$

JEE Trap

Stoichiometric coefficients MUST become powers in your Q expression.



Key Takeaways

- Standard state = 1M, Real state = Nernst.
- High reactant concentration = Higher voltage.
- Always balance the reaction first to find n and Q exponents.



You've Learned the Theory.

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