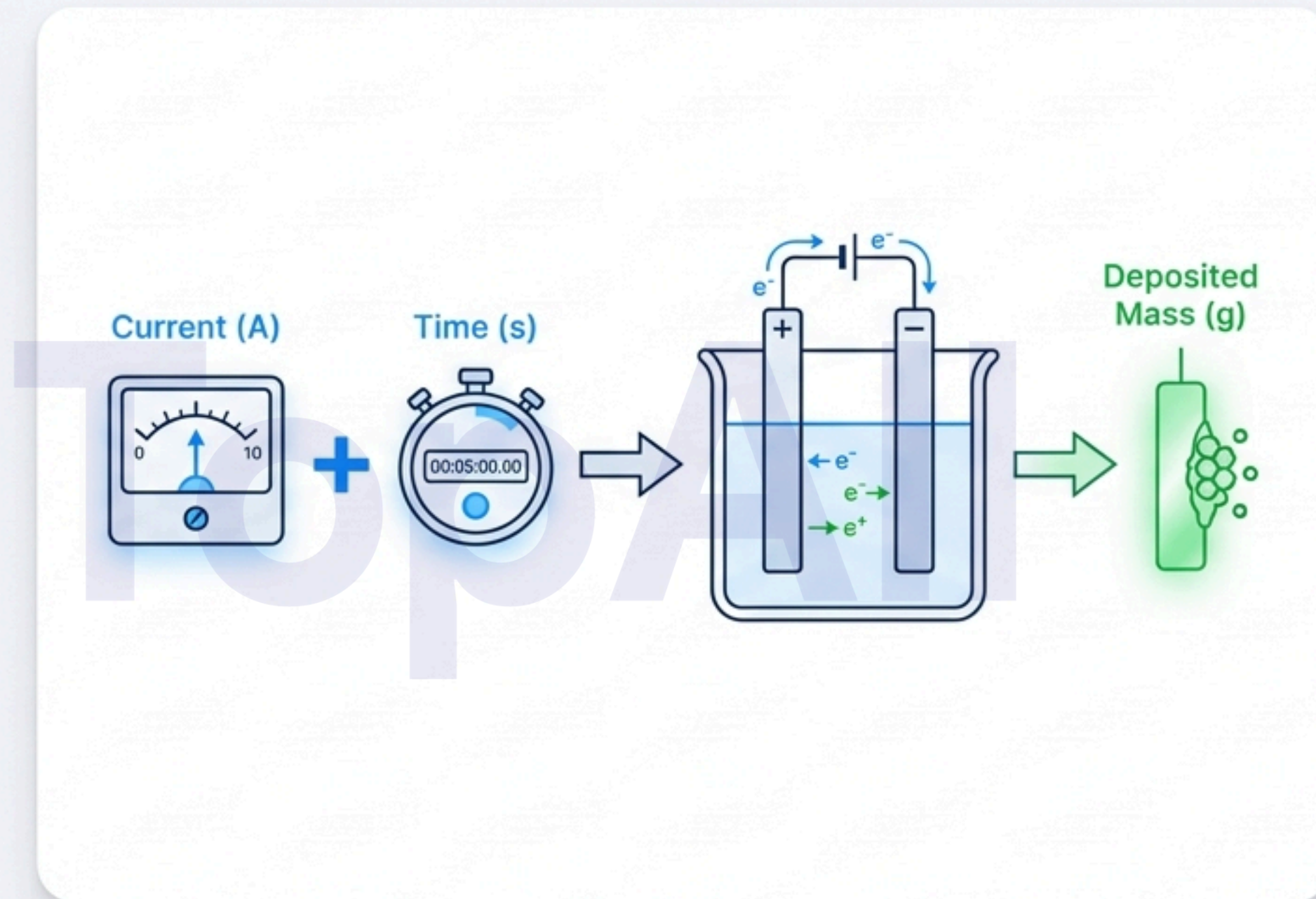


Faraday's Laws Current to Mass in 4 Steps

Master the quantitative aspects of electrolysis. Learn how to map electrical charge directly to the mass of deposited metals without memorizing complex derived formulas.



FARADAY'S FIRST LAW

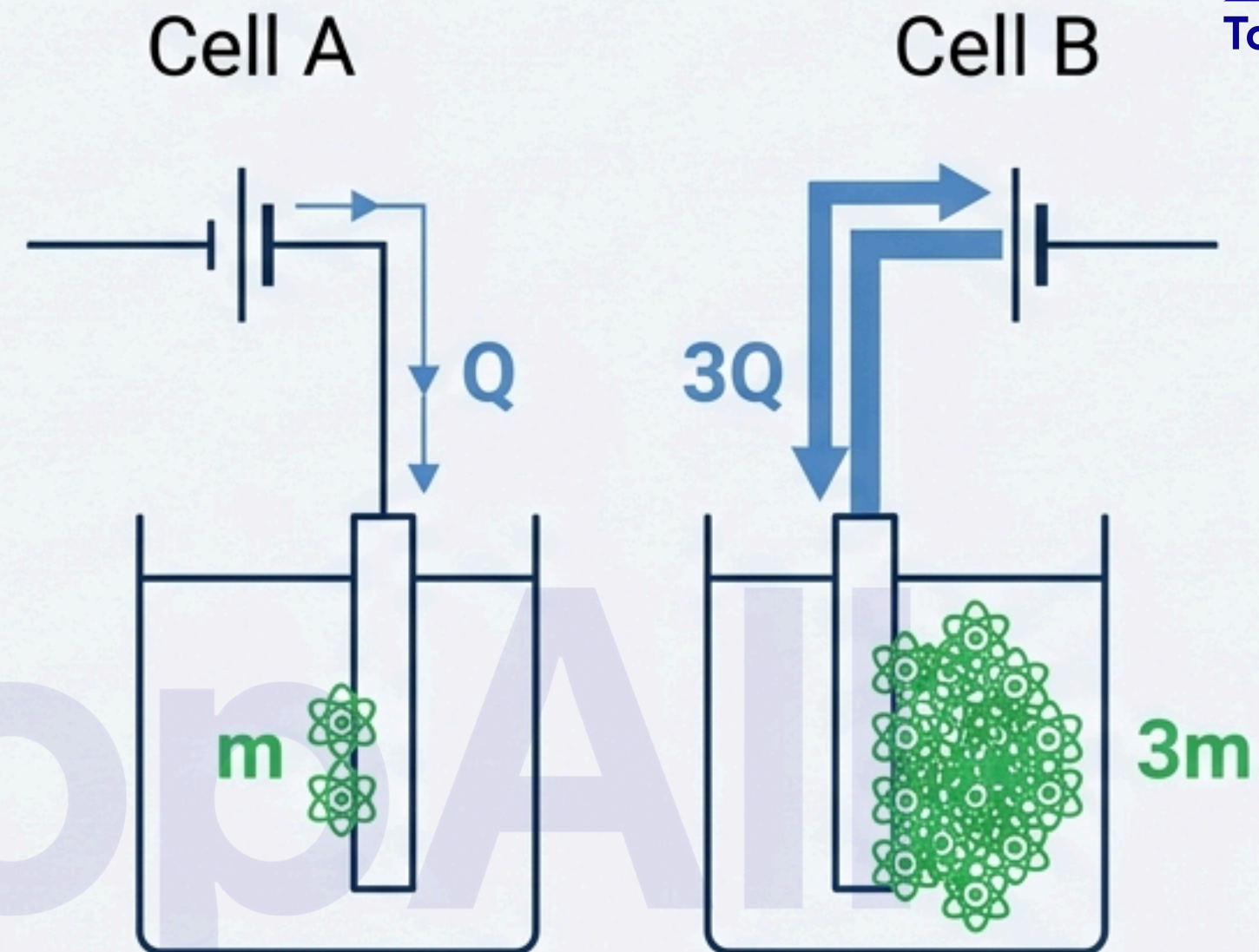
More Charge → More Reaction

The amount of chemical reaction at any electrode during electrolysis is **directly proportional** to the quantity of **electricity** (charge) passed through the electrolyte.

$$m \propto Q$$

$$m = ZQ$$

(Where Z is the electrochemical equivalent)

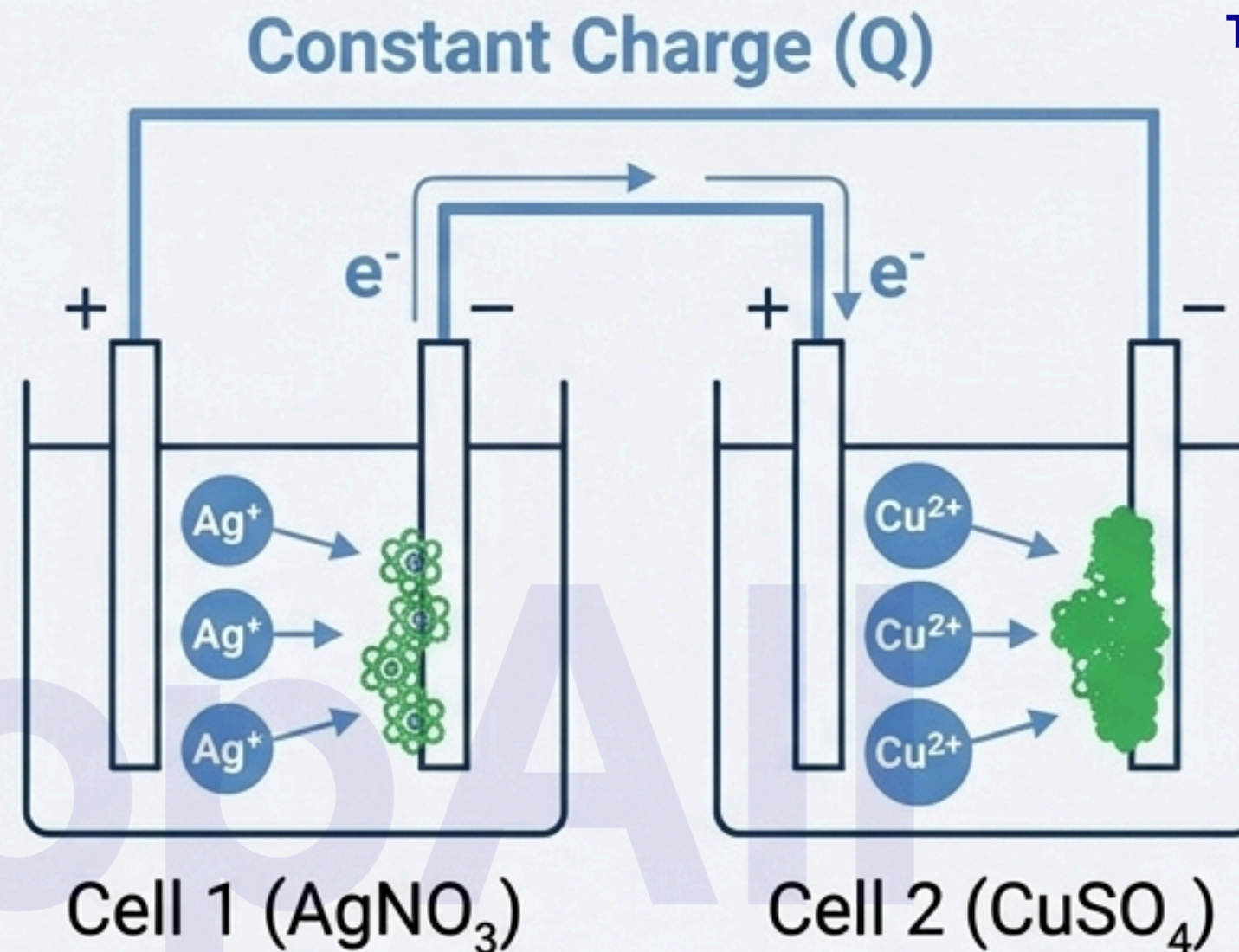


Key Takeaway: Charge is the ultimate currency of electrolysis. Double the coulombs, double the grams.

FARADAY'S SECOND LAW

Same Charge, Different Mass

When the same quantity of electricity passes through different electrolytic solutions connected in series, the masses of substances liberated are proportional to their chemical equivalent weights.



$$\frac{m_1}{m_2} = \frac{E_1}{E_2}$$

(Where E = Atomic Mass $\div$ Number of electrons required)



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Let's Test Them Now.**

$$E=mc^2$$



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Defining Electrical Charge

$$Q = It$$

Symbol: I
Metric: Current
Unit: Ampere (A)

Symbol: t
Metric: Time
Unit: Second (s)

Symbol: Q
Metric: Charge
Unit: Coulomb (C)

The Charge of a Mole

Electrons carry charge. To connect macroscopic physics to atomic chemistry, we calculate the exact charge held by one mole of electrons.

$$\text{Charge of } 1 \text{ e}^- = 1.602 \times 10^{-19} \text{ C}$$

$$1 \text{ Mole } (N_A) = 6.022 \times 10^{23}$$

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

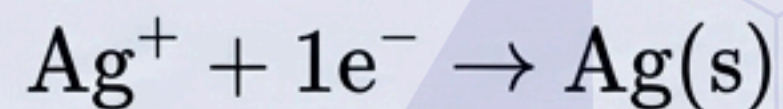
Standard JEE Approximation:

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



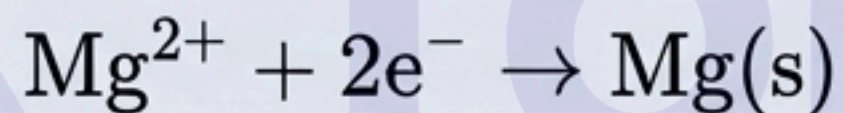
Counting Electrons (The Stoichiometry)

The amount of charge required for reduction depends on the oxidation state of the metal cation. Different metals demand different “prices” in Faraday units to yield one mole of solid metal.



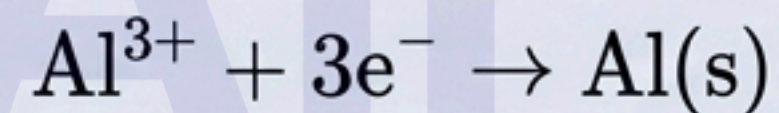
Requirement: 1 Mole of e^-

Charge: 1 F (96500 C)



Requirement: 2 Moles of e^-

Charge: 2 F (193000 C)



Requirement: 3 Moles of e^-

Charge: 3 F (289500 C)

NCERT APPLICATION

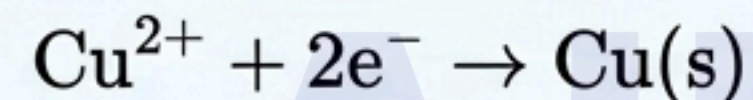
Calculating Deposited Mass

Problem: A solution of CuSO_4 is electrolysed for 10 minutes with a current of **1.5 amperes**. What is the mass of copper deposited at the cathode?
(Atomic mass Cu = 63 g/mol)

1. Find Charge (C):

$$t = 10 \text{ min} \times 60 \text{ s} = 600 \text{ s}$$

$$Q = 1.5 \text{ A} \times 600 \text{ s} = 900 \text{ C}$$

**2. Determine Stoichiometry:**

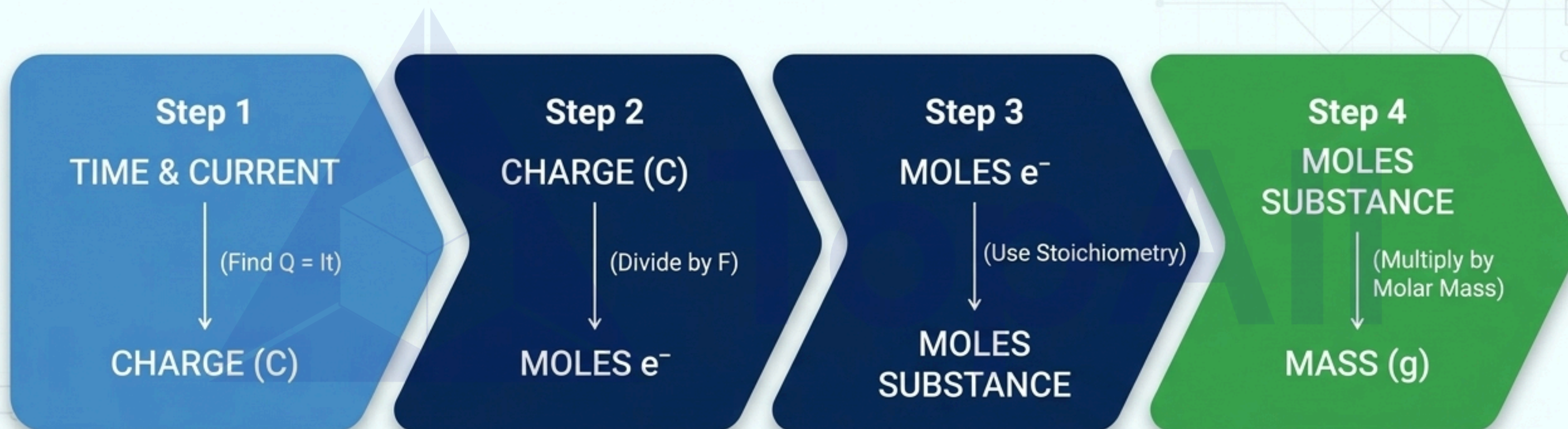
(Requires $2F = 2 \times 96487 \text{ C}$ to deposit 63g of Cu)

**3. Calculate Mass:**

$$\left[\frac{63 \text{ g}}{2 \times 96487 \text{ C}} \right] \times 900 \text{ C}$$

Final Result: 0.2938 g Cu deposited.

The 4-Step Electrolysis Engine



⚠ JEE Trap:

Always convert time from minutes or hours into seconds before using $Q=It$. Failing to convert is the #1 reason for incorrect answers in Faraday calculations.



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