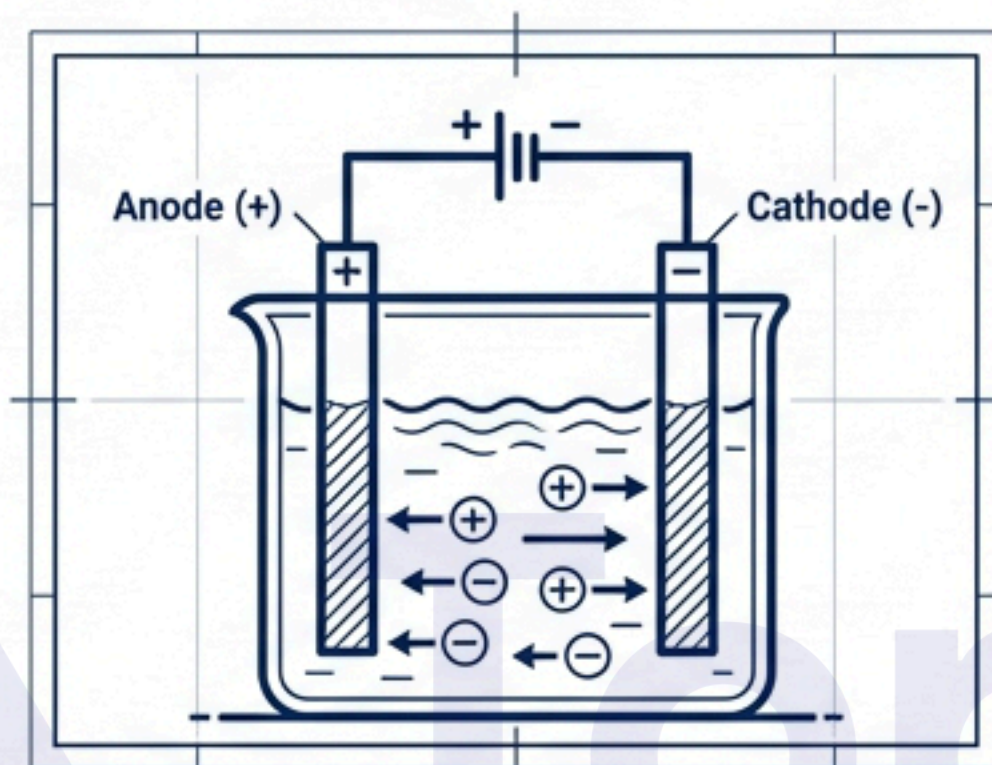


Faraday's Laws of Electrolysis



The Paradigm Shift

While galvanic cells spontaneously produce electricity, electrolytic cells consume electrical energy to force non-spontaneous chemical transformations.

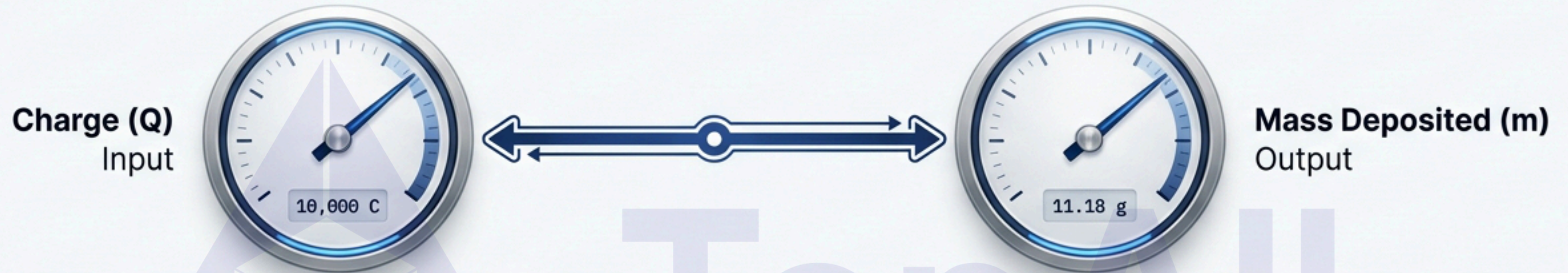
The Quantitative Focus

Michael Faraday's 1833 discoveries established the exact mathematical relationship between the electrical energy consumed and the chemical matter produced.

[Electrical Energy Input] → [Non-Spontaneous Chemical Transformation]

Electrolysis is the controlled use of electricity to drive chemical change; Faraday's laws allow us to quantify exactly how much change occurs.

The First Law of Electrolysis



The Law

The amount of chemical reaction (mass deposited or liberated) at any electrode is **directly proportional** to the **total quantity of electricity** passed through the electrolyte.

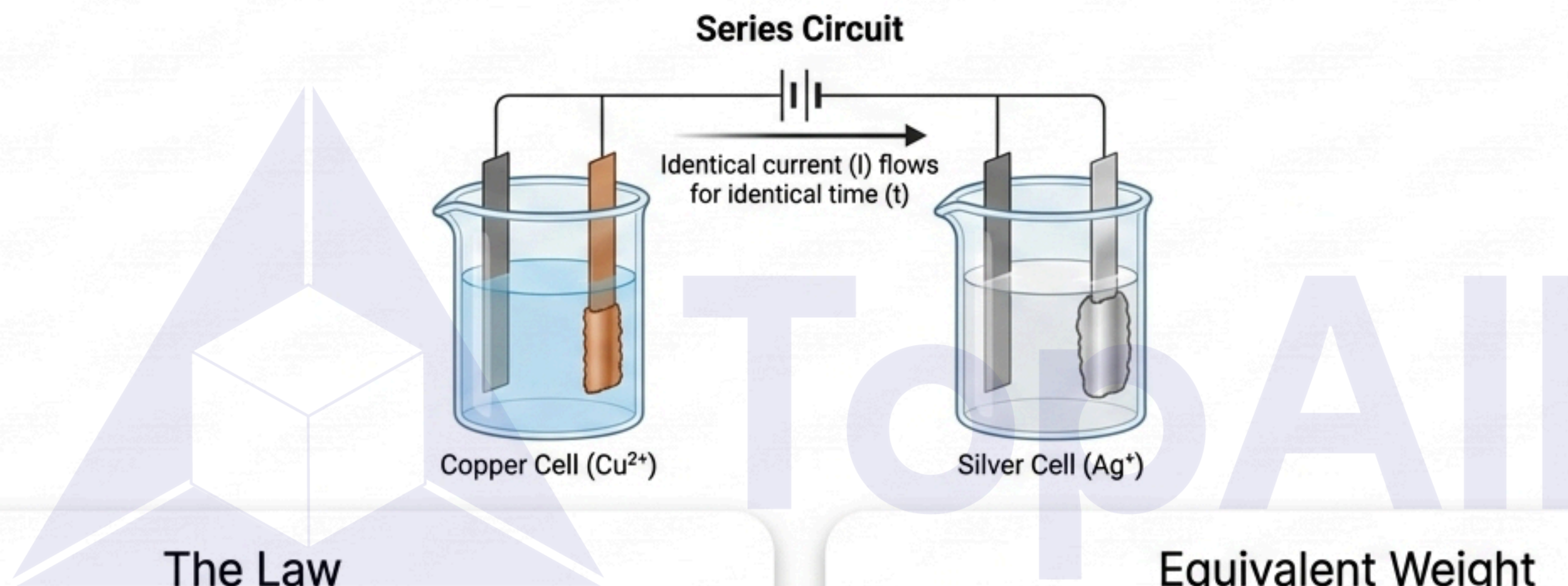
Industrial Application

This linear predictability is the foundation of industrial **electroplating**, allowing engineers to **dictate exact microscopic coating thicknesses** by controlling the charge.

$$m \propto Q \rightarrow m = ZQ \quad (\text{where } Z \text{ is the electrochemical equivalent})$$

More electrical charge passed through the cell equals a directly proportional increase in chemical reaction.

The Second Law of Electrolysis



When the same quantity of electricity passes through different electrolytes, the masses liberated are strictly proportional to their chemical equivalent weights.

Equivalent Weight

The atomic mass of the metal divided by the number of electrons required to reduce the cation.

$$m_1 / E_1 = m_2 / E_2 \text{ or } m_1 / m_2 = E_1 / E_2$$

Identical charge yields different chemical masses, governed entirely by the atomic stoichiometry of the specific metal.



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Quantity of Electricity

Equation Anatomy

$$Q = I \times t$$

Charge (Q)

The total aggregate of electricity passed through the system.
Measured in Coulombs [C].

Current (I)

The continuous flow rate of the electricity.
Measured in Amperes [A].

Time (t)

The duration of the electrical flow. Must always be measured in Seconds [s].

$$Q \text{ [Coulombs]} = I \text{ [Amperes]} \times t \text{ [Seconds]}$$

Total charge is the strict mathematical product of current intensity and time duration. Always convert minutes/hours to seconds before calculating.

Core Constant

One Faraday and Its Meaning

1 e⁻ charge (1.6021×10^{-19} C)

—**x**→

1 Mole
($N_A: 6.02 \times 10^{23}$)

—**=**→

1 Faraday
(96487 C)

The Definition

One Faraday (F) is the exact magnitude of electrical charge carried by one mole of electrons.

The Stoichiometric Link

$\text{Ag}^+(\text{aq}) + \text{e}^- \rightarrow \text{Ag}(\text{s})$ requires 1 mole of electrons = 1F.

$\text{Mg}^{2+}(\text{l}) + 2\text{e}^- \rightarrow \text{Mg}(\text{s})$ requires 2 moles of electrons = 2F.

$\text{Al}^{3+}(\text{l}) + 3\text{e}^- \rightarrow \text{Al}(\text{s})$ requires 3 moles of electrons = 3F.

$$1F = N_A \times \text{e}^- \approx 96487 \text{ C mol}^{-1} \text{ (Often approximated as } 96500 \text{ C mol}^{-1}\text{)}$$

Faraday's constant bridges invisible subatomic electron flow with macroscopic, measurable mole quantities.

Copper Deposition Example

Step 1: Input Data

A solution of CuSO_4 is electrolysed for 10 minutes with a current of 1.5 amperes.

$$I = 1.5 \text{ A}$$

$$t = 600 \text{ s}$$

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

Step 2: Cell Reaction



To deposit 1 mole of Copper (63 g), we require 2 moles of electrons (2F).

$$2 \times 96487 \text{ C yields } 63 \text{ g.}$$

Step 3: Calculation

$$\text{Mass} = \frac{63 \text{ g} \times 900 \text{ C}}{2 \times 96487 \text{ C}}$$

Result = 0.2938 g of Cu deposited at the cathode.

Calculate total charge → Determine required Faradays based on cation charge → Solve for actual mass deposited.

Charge Dictates Chemical Yield



Electrical Domain

Current (**Amperes**) and Time (**Seconds**) dictate the total **Coulombs** injected into the system.

Chemical Domain

The cation's oxidation state dictates the required **Faradays** (**1F** per mole of electrons needed for reduction).

Material Domain

The atomic mass and equivalent weight dictate the final physical mass deposited on the electrode.

$$\text{Mass} = (\text{Molar Mass} \times I \times t) / (n \times F) \text{ (where } n = \text{number of electrons exchanged)}$$

Electrolysis is a perfectly quantifiable conversion mechanism, turning variable electrical energy into precise, predictable chemical yields.



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