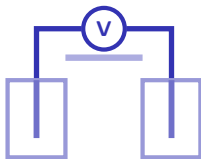


Electrochemistry

A-Level Chemistry ■ Topic 24

A-Level Chemistry



What we will cover

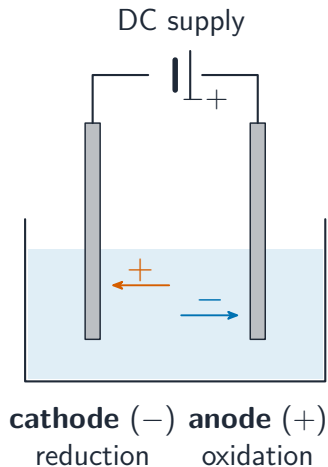
- 1 Electrolysis
- 2 Faraday calculations
- 3 Standard electrode potentials
- 4 Cell potentials & feasibility
- 5 The Nernst equation



electrolysis of water



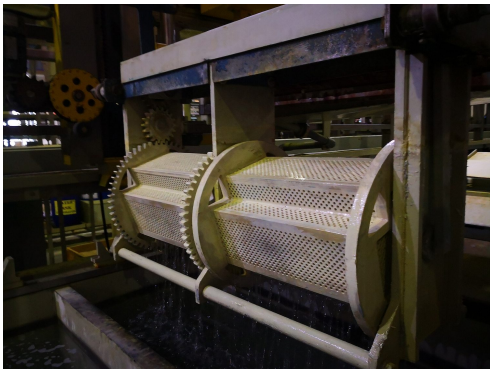
Electrolysis



Electrolysis breaks down an **electrolyte** :

- **cathode** (-): positive ions gain e^- (reduction)
- **anode** (+): negative ions lose e^- (oxidation)

Faraday calculations



Charge $Q = It$; moles of electrons = Q/F
(**Faraday constant**).

2.0 A for 30 min through CuSO_4

$$Q = 3600 \text{ C}; n(\text{e}^-) = 0.0373; n(\text{Cu}) = 0.0187 \Rightarrow m \approx 1.2 \text{ g}$$

Use the half-equation to find moles of product.

From charge to mass, and back to Avogadro

the charge

$Q = It$ – current in amperes, time in **seconds**

to moles of electrons

divide by the Faraday constant, $96\,500\text{ C mol}^{-1}$

to moles of product

divide by the number of electrons in the **half-equation** – 2
for $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

to mass

multiply by the molar mass

a worked figure

2.0 A for 30 minutes is 3600 C, so 0.0373 mol of electrons
and **0.0187 mol of copper** – about 1.2 g

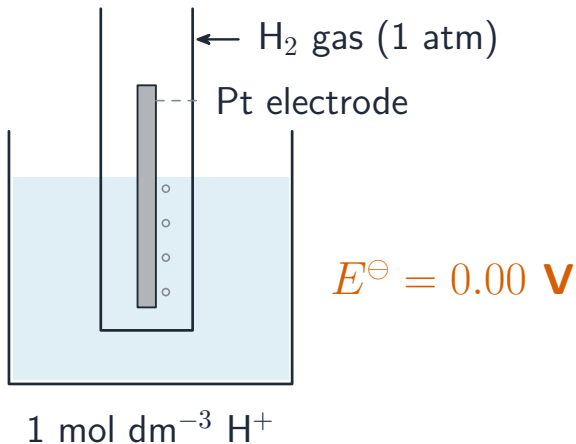
the Avogadro constant

measuring the mass deposited for a known charge lets you
work back to a value for it

The half-equation's electron count is the step candidates omit. Write it down before dividing, every time.



Standard electrode potentials



The **standard electrode potential** $E^{\ominus}$ shows how easily a half-cell is reduced.

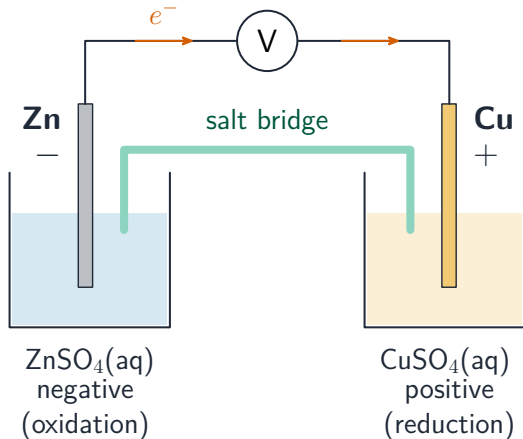
It is measured against the **standard hydrogen electrode**, defined as exactly 0.00 V.

Combining half-cells

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

- find the **polarity**: the more negative electrode is the negative terminal, and electrons flow from it through the external circuit to the positive electrode.
- judge **reactivity**: a more positive $E^\ominus$ means a better **oxidising agent** (easily reduced).

Cell potentials and feasibility



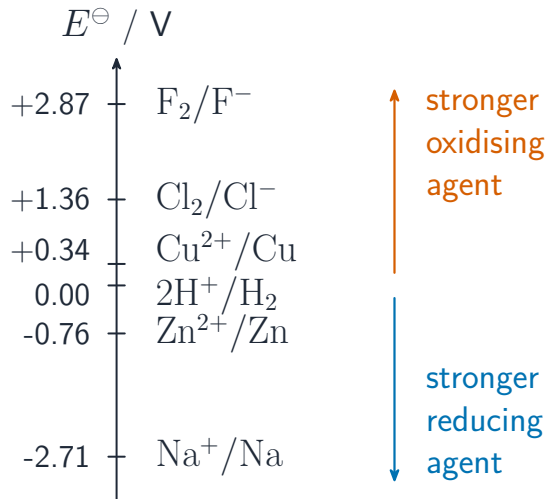
$$E_{\text{cell}}^{\ominus} = E^{\ominus}(\text{more+}) - E^{\ominus}(\text{less+}).$$

$$\text{Zn } (-0.76) \text{ \& Cu } (+0.34)$$

$$E_{\text{cell}}^{\ominus} = 0.34 - (-0.76) = 1.10 \text{ V}$$

A reaction is **feasible** when $E_{\text{cell}}^{\ominus} > 0$.
($\Delta G^{\ominus} = -nE^{\ominus}F$.)

The electrochemical series



- a more **positive** $E^{\ominus} \rightarrow$ a stronger **oxidising agent** (easily reduced)
- a more **negative** $E^{\ominus} \rightarrow$ a stronger **reducing agent**

The Nernst equation

$$E = E^{\ominus} + \frac{0.059}{z} \log \frac{[\text{oxidised}]}{[\text{reduced}]}$$

Away from standard conditions, raising the **oxidised** species' concentration makes E **more positive** (z = electrons in the half-equation).

Predicting the products

molten

metal at cathode
non-metal at anode

aqueous

also contains water,
so H_2 or O_2 may form

A molten electrolyte gives the metal and non-metal; aqueous also has water

Exam tips

- $E_{\text{cell}}^{\ominus} = E^{\ominus}(\text{more positive}) - E^{\ominus}(\text{less positive})$; a **positive** $E_{\text{cell}}^{\ominus}$ means the reaction is feasible.
- Standard conditions for $E^{\ominus}$: **298 K**, **1 mol dm⁻³**, **100 kPa**, platinum electrode – state them if asked.
- A **more negative** electrode potential means a **stronger reducing agent**; use the series to predict the direction.
- For electrolysis quantities use $Q = It$ and moles of electrons = Q/F .



Topic 24 – key takeaways

1

Electrolysis

cathode reduces; anode oxidises

2

Faraday

$$Q = It; \text{ moles of } e^- = Q/F$$

3

 $E^\ominus$

measured vs the hydrogen
electrode (0 V)

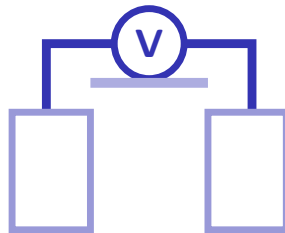
4

Cells

$$E_{\text{cell}}^\ominus > 0 \text{ means feasible}$$

Check yourself

- 1 What happens at the cathode?
- 2 What is the charge Q equal to?
- 3 What is the SHE defined as?
- 4 A cell is feasible when $E_{\text{cell}}^{\ominus}$ is what?



Answers 1. reduction (positive ions gain e^-) 2. It 3. 0.00 V 4. positive