

Exercise walk-through

9.11

Electrolysis and Faraday's Law

AP Chemistry

Coupling, the Nernst shift, and electrolysis

coupled reactions

an unfavourable reaction ($\Delta G > 0$) can be driven by a favourable one ($\Delta G < 0$) **sharing a common intermediate**

the condition

the total ΔG of the pair must be negative – ATP hydrolysis is the biological case

the Nernst equation

away from standard conditions the cell potential shifts with concentration

qualitatively

as reactants are consumed and products build up, the potential **falls** toward zero – a dead battery is a cell at equilibrium

electrolysis

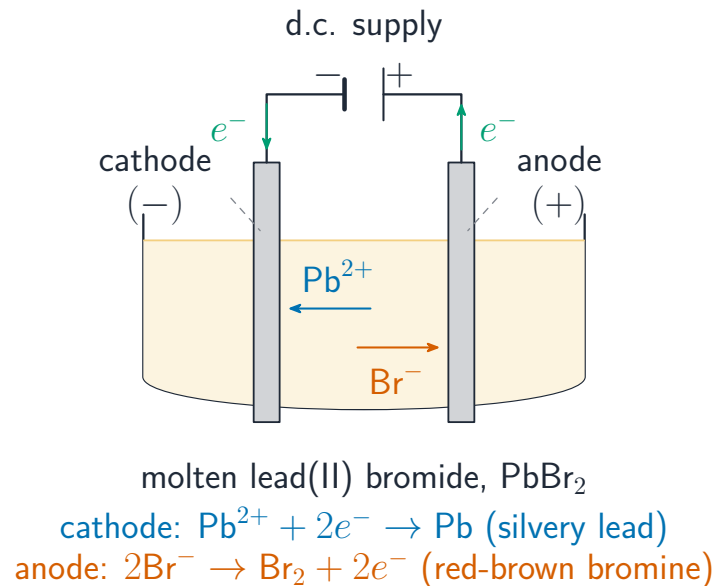
the **charge passed** determines how much substance is deposited – **Faraday's law**

the route

current $\times$ time $\rightarrow$ coulombs $\rightarrow$ moles of electrons $\rightarrow$ moles of substance $\rightarrow$ mass

A 2.0 A current for 30 minutes passes 3600 C, which is 0.0373 mol of electrons – and for Cu^{2+} that is **half** as many moles of copper.

Electrolysis



Electrolysis: ions move to the electrodes and are discharged

Exam tips

- A process is **thermodynamically favourable** when $\Delta G < 0$; combine enthalpy and entropy with $\Delta G = \Delta H - T\Delta S$ (match units – kJ vs J).
- Entropy **increases** solid→liquid→gas and when more gas moles are produced.
- Favourable does **not** mean fast – a high activation energy can make a $\Delta G < 0$ reaction extremely slow (kinetic control).
- In electrochemistry a **positive** $E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$ means a spontaneous (galvanic) cell; oxidation is at the **anode**, reduction at the **cathode**.
- In electrolysis the charge passed ($Q = It$) fixes the amount deposited (Faraday's law).

Exercises 2.1-2.3

2.1 Write the equation for charge in terms of current and time. [1]

2.2 Find the charge from 3.0 A for 200 s. [2]

2.3 How many moles of electrons are in 9650 C? [2]

Exercises 2.1·2.3

2.1 Write the equation for charge in terms of current and time. [1]

2.1

- $Q = It$

2.2 Find the charge from 3.0 A for 200 s. [2]

2.2

- $Q = (3.0)(200) = 600 \text{ C}$

2.3 How many moles of electrons are in 9650 C? [2]

2.3

- $\frac{9650}{96\,500} = 0.10 \text{ mol}$

Exercise 3.1

3.1 One mole of electrons carries a charge of about

[1]

☐

A 6.02×10^{23} C

B 96 500 C

C 1 C

D 8.314 C

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A 6.02×10^{23} C

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3.1 B

- 96 500 C (Faraday's constant)

Exercise 3.2 (a)

3.2 A current of 5.0 A flows for 1930 s through molten NaCl, depositing sodium ($\text{Na}^+ + e^- \rightarrow \text{Na}$).

(a) Find the total charge and the moles of electrons.

[3]

Exercise 3.2 (a)

3.2 A current of 5.0 A flows for 1930 s through molten NaCl, depositing sodium ($\text{Na}^+ + e^- \rightarrow \text{Na}$).

(a) Find the total charge and the moles of electrons. [3]

(a)

- $Q = (5.0)(1930) = 9650 \text{ C}$; $\text{mol } e^- = 9650/96\,500 = 0.10 \text{ mol}$

Exercise 3.2 (b)

3.2 A current of 5.0 A flows for 1930 s through molten NaCl, depositing sodium ($\text{Na}^+ + e^- \rightarrow \text{Na}$).

(b) Find the mass of sodium deposited ($M = 23 \text{ g/mol}$). [2]

Exercise 3.2 (b)

3.2 A current of 5.0 A flows for 1930 s through molten NaCl, depositing sodium ($\text{Na}^+ + e^- \rightarrow \text{Na}$).

(b) Find the mass of sodium deposited ($M = 23 \text{ g/mol}$). [2]

(b)

- $\text{Na}^+ + e^- \rightarrow \text{Na}$ (1:1), so 0.10 mol Na; mass = $0.10 \times 23 = 2.3 \text{ g}$

Exercise 3.3

3.3 A current deposits copper from Cu^{2+} . If 0.40 mol of electrons pass, find the moles of Cu deposited. [2]

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3.3

- $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}$, so $\text{mol Cu} = 0.40/2 = 0.20 \text{ mol}$