Exercise walk-through

Cell Potential and Free Energy ▪ 9.9

eXercise

You must be able to

- calculate $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$
- use $\Delta G^\circ = -nFE^\circ_{\text{cell}}$
- link a positive E°_{cell} to a favorable cell

Method — Standard cell potential

$$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}},$$

using standard reduction potentials. A **positive** E°_{cell} means a favorable (galvanic) cell.

Method — A worked cell potential

Cu^{2+} ($E^\circ = +0.34$) as cathode, Zn^{2+} ($E^\circ = -0.76$) as anode:

$$E^\circ_{\text{cell}} = 0.34 - (-0.76) = +1.10\ \text{V}.$$

Method — Linking to free energy

$$\Delta G^\circ = -nFE^\circ_{\text{cell}},$$

where n is moles of electrons transferred and $F = 96\,500$ C/mol. A **positive** E° gives a **negative** ΔG° — favorable.

Method — A worked ΔG

For $n = 2$, $E^\circ = 1.10$ V: $\Delta G^\circ = -(2)(96\,500)(1.10) = -2.12 \times 10^5$ J $= -212$ kJ.

Practice 2.1 ▪ 1 mark

Write E°_{cell} in terms of the electrode potentials.

Solution 2.1

Method: Standard cell potential

Worked steps

$E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$ **B1**.

Practice 2.2 ▪ 2 marks

For cathode $+0.80$ V and anode -0.44 V, find E°_{cell}.

Solution 2.2

Worked steps

$0.80 - (-0.44) = +1.24$ V **M1** **A1**.

Practice 2.3 ▪ 1 mark

A cell has a positive E°_{cell}. Is it galvanic or electrolytic?

Solution 2.3

Method: Standard cell potential

Worked steps

Galvanic **B1**.

Exam-style 3.1 ▪ 1 mark

A galvanic (favorable) cell has an E°_{cell} that is

A positive

B negative

C zero

D undefined

Solution 3.1

Method: Standard cell potential

A positive ✓

B negative

C zero

D undefined

Worked steps

A — positive.

Exam-style 3.2 ▪ 2 marks

A cell has $E^\circ_{\text{cathode}} = +0.34$ V and $E^\circ_{\text{anode}} = -0.76$ V, with $n = 2$.
(a) Find E°_{cell}.
(b) Find ΔG° (use $F = 96\,500$ C/mol).

Solution 3.2

Worked steps

(a) $E^{\circ}_{\text{cell}} = 0.34 - (-0.76) = +1.10$ V M1 A1. (b) $\Delta G^{\circ} = -(2)(96\,500)(1.10) = -2.12 \times 10^{5}$ J $= -212$ kJ M1 M1 A1.

Exam-style 3.3 ▪ 2 marks

Explain why a positive E°_{cell} corresponds to a negative ΔG°.

Solution 3.3

Method: Linking to free energy

Worked steps

$\Delta G^\circ = -nFE^\circ$; with n and F positive, a positive E° makes ΔG° negative — a favorable reaction **M1** **A1**.