

Galvanic Cells and Electrode Potentials

Answer key with worked solutions

A. Understanding

1. Galvanic cell.

Oxidation happens at the **anode** (negative) and **reduction** at the **cathode** (positive). Electrons flow through the wire **from anode to cathode**. The **salt bridge** lets ions move between the half-cells so that the solutions stay electrically neutral and the current can continue.

B. Cell potentials

2. Zinc–silver cell.

(a) Ag^+/Ag has the more positive E° , so silver is the **cathode**; zinc is the **anode**. (b) $E^\circ_{\text{cell}} = +0.80 \text{ V} - (-0.76 \text{ V}) = +1.56 \text{ V}$ (c) $\text{Zn} + 2 \text{Ag}^+ \longrightarrow \text{Zn}^{2+} + 2 \text{Ag}$ (d) $\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq}) \parallel \text{Ag}^+(\text{aq}) \mid \text{Ag(s)}$

3. Nickel–silver cell.

$E^\circ_{\text{cell}} = +0.80 \text{ V} - (-0.25 \text{ V}) = +1.05 \text{ V}$; **nickel** is the anode.

4. Magnesium–copper cell.

$E^\circ_{\text{cell}} = +0.34 \text{ V} - (-2.37 \text{ V}) = +2.71 \text{ V}$ (magnesium is the anode).

5. Spontaneous?

(a) Zn^{2+} would be reduced (cathode) and Cu oxidized (anode): $E^\circ_{\text{cell}} = -0.76 \text{ V} - (+0.34 \text{ V}) = -1.10 \text{ V}$, so **not spontaneous**. (b) H^+ reduced, Zn oxidized: $E^\circ_{\text{cell}} = 0.00 \text{ V} - (-0.76 \text{ V}) = +0.76 \text{ V}$, so **spontaneous** (zinc dissolves in acid).

6. Strongest agents.

Strongest oxidizing agent: **F_2** (most positive E° , +2.87 V). Strongest reducing agent: **Mg** (most negative E° , -2.37 V, so Mg is most easily oxidized).

C. Free energy

7. ΔG° for the Daniell cell.

$$E^\circ_{\text{cell}} = +0.34 \text{ V} - (-0.76 \text{ V}) = +1.10 \text{ V}; n = 2 \text{ mol of electrons.}$$

$$\Delta G^\circ = -(2 \text{ mol}) \left(96\,485 \frac{\text{C}}{\text{mol}} \right) (1.10 \text{ V})$$

$$= -2.12 \times 10^5 \text{ J} = -212 \text{ kJ}$$

(mol cancels, and $\text{C} \times \text{V} = \text{J}$.)

D. Electrolysis

8. Silver plating.

$$t = 45.0 \text{ min} \times 60 \text{ s/min} = 2700 \text{ s}; Q = 1.50 \text{ C/s} \times 2700 \text{ s} = 4050 \text{ C}$$

$$n(e^-) = \frac{4050 \text{ C}}{96\,485 \text{ C/mol}} = 0.04198 \text{ mol}; 1 \text{ mol Ag per mol } e^-, \text{ so } n(\text{Ag}) = 0.04198 \text{ mol}$$

$$m = 0.04198 \text{ mol} \times 107.87 \text{ g/mol} = \mathbf{4.53 \text{ g}}$$

E. Challenge

9. Time to deposit 5.00 g Cu.

$$n(\text{Cu}) = \frac{5.00 \text{ g}}{63.55 \text{ g/mol}} = 0.07868 \text{ mol}; n(e^-) = 2 \times 0.07868 \text{ mol} = 0.1574 \text{ mol}$$

$$Q = 0.1574 \text{ mol} \times 96\,485 \text{ C/mol} = 1.518 \times 10^4 \text{ C}$$

$$t = \frac{Q}{I} = \frac{1.518 \times 10^4 \text{ C}}{3.00 \text{ C/s}} = 5.06 \times 10^3 \text{ s, and } 5.06 \times 10^3 \text{ s} \times \frac{1 \text{ min}}{60 \text{ s}} = \mathbf{84.3 \text{ min}}$$