

Electrolysis

Answer key with worked solutions

A. Understanding

1. Concepts.

(a) Using electricity to drive a **non-spontaneous** redox reaction. (b) An electrolytic cell **uses** electrical energy, a galvanic cell **produces** it; in electrolysis the anode is **positive**, in a galvanic cell it is negative. (c) The **ions must be free to move** to the electrodes; in a solid ionic compound they are fixed in the lattice.

B. Predicting products

2. Molten compounds.

(a) Cathode **potassium**, anode **bromine** (b) cathode **magnesium**, anode **chlorine** (c) cathode **aluminium**, anode **oxygen**

3. Aqueous solutions.

(a) Cathode **copper**, anode **chlorine** (b) cathode **hydrogen**, anode **oxygen** (c) cathode **silver**, anode **oxygen** (d) cathode **hydrogen**, anode **iodine**

4. Half-equations.

(b) Cathode: $2\text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{H}_2 + 2\text{OH}^-$; anode: $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$. (c) Cathode: $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$; anode: $2\text{H}_2\text{O} \rightarrow \text{O}_2 + 4\text{H}^+ + 4\text{e}^-$.

5. Copper refining.

The **impure copper anode** is oxidized and dissolves: $\text{Cu} \rightarrow \text{Cu}^{2+} + 2\text{e}^-$. At the **pure copper cathode**, $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$ deposits pure copper. Less reactive impurities (silver, gold) do not dissolve and fall to the bottom as **anode sludge**; more reactive ones dissolve but stay in solution.

C. Faraday's law

6. Silver plating.

$$Q = (0.500 \text{ C/s})(30.0 \text{ min} \times 60 \text{ s/min}) = 900. \text{ C}; n(e^-) = \frac{900. \text{ C}}{96\,485 \text{ C/mol}} = 9.328 \times 10^{-3} \text{ mol}$$
$$1 \text{ mol } e^- \text{ per mol Ag: } m = 9.328 \times 10^{-3} \text{ mol} \times 107.87 \text{ g/mol} = \mathbf{1.01 \text{ g}}$$

7. Copper deposition time.

$$n(\text{Cu}) = \frac{2.00 \text{ g}}{63.55 \text{ g/mol}} = 0.03147 \text{ mol}; n(e^-) = 2 \times 0.03147 \text{ mol} = 0.06294 \text{ mol}$$
$$Q = 0.06294 \text{ mol} \times 96\,485 \text{ C/mol} = 6073 \text{ C}; t = \frac{6073 \text{ C}}{1.50 \text{ C/s}} = 4049 \text{ s} = \mathbf{67.5 \text{ min}}$$

8. Electrolysis of water.

$$Q = (0.500 \text{ C/s})(7200 \text{ s}) = 3600 \text{ C}; n(e^-) = \frac{3600 \text{ C}}{96\,485 \text{ C/mol}} = 0.03731 \text{ mol}$$
$$\text{H}_2 (2 e^- \text{ each}): n = 0.01866 \text{ mol}; V = \frac{(0.01866 \text{ mol})(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(298.15 \text{ K})}{1.00 \text{ atm}} = \mathbf{0.456 \text{ L}}$$
$$\text{O}_2 (4 e^- \text{ each}): n = 9.328 \times 10^{-3} \text{ mol}; V = \mathbf{0.228 \text{ L}} \text{ (half the volume of hydrogen, as expected from } 2 \text{H}_2\text{O} \longrightarrow 2 \text{H}_2 + \text{O}_2)$$

D. Challenge

9. Energy for aluminium.

$$(a) n(\text{Al}) = \frac{1.00 \times 10^3 \text{ g}}{26.98 \text{ g/mol}} = 37.06 \text{ mol}; n(e^-) = 3 \times 37.06 \text{ mol} = 111.2 \text{ mol}$$
$$Q = 111.2 \text{ mol} \times 96\,485 \text{ C/mol} = \mathbf{1.07 \times 10^7 \text{ C}}$$
$$(b) E = QV = (1.073 \times 10^7 \text{ C})(4.5 \text{ V}) = \mathbf{4.8 \times 10^7 \text{ J}}; \frac{4.83 \times 10^7 \text{ J}}{3.6 \times 10^6 \text{ J/kWh}} = \mathbf{13 \text{ kWh}} \text{ per kilogram}$$

(c) Aluminium production needs **huge amounts of electricity**, which is the main cost, so smelters are built where electricity is cheap and plentiful (and, with hydroelectric power, low in carbon emissions).