

Batteries, Fuel Cells and Corrosion

Answer key with worked solutions

A. Batteries and fuel cells

1. Batteries.

(a) A **primary** battery is used once; its reaction cannot practically be reversed. A **secondary** battery can be recharged. (b) An external current forces electrons in the opposite direction, driving the cell reaction **backwards** (electrolysis): Li^+ ions move back into the graphite electrode, restoring the reactants.

2. Fuel cell.

Anode: $2\text{H}_2 \longrightarrow 4\text{H}^+ + 4\text{e}^-$; cathode: $\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \longrightarrow 2\text{H}_2\text{O}$; overall:
 $2\text{H}_2 + \text{O}_2 \longrightarrow 2\text{H}_2\text{O}$

$$E^\circ_{\text{cell}} = 1.23\text{ V} - 0.00\text{ V} = \mathbf{1.23\text{ V}}$$

Advantage: the only exhaust is **water** (and it doesn't run down while fuelled). Challenge: hydrogen is hard to **store** and is often made from fossil fuels.

B. Corrosion

3. Three test tubes.

Tube 1: nail in tap water open to air: **rusts**. Tube 2: nail in boiled water (oxygen removed) with a layer of oil on top: **no rust** (no oxygen). Tube 3: nail in dry air with a drying agent (anhydrous calcium chloride), stoppered: **no rust** (no water). Only the tube with both water and oxygen shows rust.

4. Rusting half-equations.

Anode: $\text{Fe} \longrightarrow \text{Fe}^{2+} + 2\text{e}^-$; cathode (neutral water): $\text{O}_2 + 2\text{H}_2\text{O} + 4\text{e}^- \longrightarrow 4\text{OH}^-$

In acid: $E^\circ_{\text{cell}} = 1.23\text{ V} - (-0.44\text{ V}) = \mathbf{1.67\text{ V}}$

5. Coatings.

(a) Zinc is **more reactive** than iron ($E^\circ -0.76\text{ V}$ versus -0.44 V), so where both are exposed, zinc is oxidized instead of iron: **sacrificial protection**. (b) Tin is **less reactive** than iron, so at a scratch the iron becomes the **anode** and corrodes, faster than if it were bare.

C. Calculations

6. AA cell.

(a) $2.500\text{ A h} \times 3600\text{ s/h} = \mathbf{9000\text{ C}}$

(b) $\frac{9000\text{ C}}{96485\text{ C/mol}} = \mathbf{0.0933\text{ mol e}^-}$

(c) $2\text{ mol e}^- \text{ per mol Zn: } \frac{0.09328\text{ mol}}{2} \times 65.38\text{ g/mol} = \mathbf{3.05\text{ g Zn}}$

(d) $9000\text{ C} \times 1.5\text{ V} = 1.35 \times 10^4\text{ J}$, about **14 kJ**

7. Hydrogen used.

Charge: $1.00\text{ A} \times 3600\text{ s} = 3600\text{ C}$; electrons: $\frac{3600\text{ C}}{96485\text{ C/mol}} = 0.03731\text{ mol}$

H_2 (2 e^- each): $0.01866\text{ mol} \times 2.016\text{ g/mol} = \mathbf{0.0376\text{ g}}$

D. Challenge

8. Sacrificial anode.

Moles of Mg: $\frac{5000\text{ g}}{24.31\text{ g/mol}} = 205.7\text{ mol}$; electrons: $2 \times 205.7 = 411.4\text{ mol}$

Charge: $411.4\text{ mol} \times 96485\text{ C/mol} = 3.97 \times 10^7\text{ C}$

Time: $\frac{3.97 \times 10^7\text{ C}}{0.50\text{ A}} = 7.94 \times 10^7\text{ s}$; $\frac{7.94 \times 10^7\text{ s}}{3.156 \times 10^7\text{ s/year}} = \mathbf{\text{about } 2.5\text{ years}}$