

Cells, cell potential and electrolysis

AP Chemistry · Worksheet answers

Each answer shows the formula, the substitution and the final result with its unit in bold. Check the sign of E°_{cell} first, then n , because most lost marks come from those two steps.

Level: AP Chemistry · **Curriculum:** AP Chem Unit 9

Problem 1 Reading a lead-copper cell

- Lead has the lower E° (-0.13 V against $+0.34\text{ V}$), so lead is oxidised: electrode X is the anode, **$\text{Pb} \rightarrow \text{Pb}^{2+} + 2\text{e}^-$** .
- Electrons leave the anode and flow through the wire and voltmeter to the copper cathode: **from X to Y**.
- Pb^{2+} builds up in the anode beaker, so nitrate anions move **toward the lead half-cell** to keep it neutral.

Problem 2 Magnesium and silver cell

- Silver has the higher E° , so silver is the cathode.
- $E^\circ_{\text{cell}} = 0.80 - (-2.37) = \mathbf{+3.17\text{ V}}$.
- $\text{Mg} \rightarrow \text{Mg}^{2+} + 2\text{e}^-$ and $2\text{Ag}^+ + 2\text{e}^- \rightarrow 2\text{Ag}$ give **$\text{Mg} + 2\text{Ag}^+ \rightarrow \text{Mg}^{2+} + 2\text{Ag}$** . The silver potential is not doubled.

Problem 3 Will the iron nail react?

- Zn^{2+} is the species reduced: $E^\circ = -0.76 - (-0.44) = \mathbf{-0.32\text{ V}}$, not spontaneous.
- $E^\circ = 0.34 - (-0.44) = \mathbf{+0.78\text{ V}}$, spontaneous.
- Only the nail in copper(II) sulfate reacts: it becomes coated with reddish-brown copper and the blue solution fades toward pale green.

Problem 4 From current to electrons

- $t = 2.00 \times 3600 = 7200\text{ s}$.
- $Q = It = 0.250 \times 7200 = \mathbf{1800\text{ C}}$.
- $n(\text{e}^-) = 1800 / 96\,485 = \mathbf{0.0187\text{ mol}}$.
- $N = 0.018656 \times 6.022 \times 10^{23} = \mathbf{1.12 \times 10^{22}\text{ electrons}}$.

Problem 5 Balancing aluminium and copper

- Anode: $\text{Al} \rightarrow \text{Al}^{3+} + 3\text{e}^-$. Cathode: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$.

2. Multiply the anode by 2 and the cathode by 3 so each has 6 electrons: **$2\text{Al} + 3\text{Cu}^{2+} \rightarrow 2\text{Al}^{3+} + 3\text{Cu}$** .
3. $n = 6$. $E^\circ_{\text{cell}} = 0.34 - (-1.66) = \mathbf{+2.00\text{ V}}$.

Problem 6 The hydrogen electrode

1. **$1.0\text{ mol/dm}^3\text{ H}^+(\text{aq})$, H_2 gas at 1 atm, inert platinum electrode, 25°C** ; E° is defined as 0.00 V.
2. Copper has the higher E° , so the copper half-cell is the cathode and reads +0.34 V.
3. $\text{Zn} + 2\text{H}^+ \rightarrow \text{Zn}^{2+} + \text{H}_2$ has $E^\circ = 0.00 - (-0.76) = \mathbf{+0.76\text{ V}}$, so zinc does react and bubbles of hydrogen form.

Problem 7 Electrolysis of molten potassium bromide

1. A is joined to the positive terminal, so A is the **anode (+)**; B is the **cathode (-)**.
2. Anode: $2\text{Br}^- \rightarrow \text{Br}_2 + 2\text{e}^-$ (brown vapour). Cathode: $\text{K}^+ + \text{e}^- \rightarrow \text{K}$ (liquid potassium).
3. The reaction $2\text{K} + \text{Br}_2 \rightarrow 2\text{KBr}$ is spontaneous, so the reverse is not; **the battery supplies the energy that forces it**.

Problem 8 Free energy of a copper-silver cell

1. $E^\circ_{\text{cell}} = 0.80 - 0.34 = \mathbf{+0.46\text{ V}}$.
2. Two electrons move: $n = 2$.
3. $\Delta G^\circ = -nFE^\circ = -2 \times 96\,485 \times 0.46 = -88\,766\text{ J/mol} = \mathbf{-88.8\text{ kJ/mol}}$; negative, so spontaneous.

Problem 9 Equilibrium constant for iron and copper

1. $E^\circ_{\text{cell}} = 0.34 - (-0.44) = \mathbf{0.78\text{ V}}$; $n = 2$.
2. $\log K = 2 \times 0.78 / 0.0592 = 26.35$, so **$K \approx 2.2 \times 10^{26}$** .
3. K is enormous, so at equilibrium essentially all the Cu^{2+} has been converted.

Problem 10 Nickel plating a bracket

1. $t = 1500\text{ s}$; $Q = 1.50 \times 1500 = \mathbf{2250\text{ C}}$.
2. $n(\text{e}^-) = 2250 / 96\,485 = 0.02332\text{ mol}$; $n(\text{Ni}) = 0.01166\text{ mol} = \mathbf{1.17 \times 10^{-2}\text{ mol}}$.
3. $m = 0.01166 \times 58.69 = \mathbf{0.684\text{ g}}$.

Problem 11 How long to deposit silver?

1. $n(\text{Ag}) = 5.00 / 107.87 = \mathbf{0.04635\text{ mol}}$.
2. One electron per atom, so $Q = 0.04635 \times 96\,485 = \mathbf{4472\text{ C}}$.
3. $t = Q / I = 4472 / 2.00 = 2236\text{ s} = \mathbf{37.3\text{ min}}$.

Problem 12 Cell voltage at other concentrations

1. $Q = [\text{Zn}^{2+}] / [\text{Cu}^{2+}]$.
2. $Q = 0.050 / 0.80 = 0.0625$, $\log Q = -1.204$.
3. $E = 1.10 - (0.0592 / 2)(-1.204) = 1.10 + 0.0356 = \mathbf{1.14\text{ V}}$. Q is below 1, so E is above E° .

Problem 13 Reading a Nernst graph

1. E° is the value at $\log Q = 0$: **1.10 V**.
2. Slope = $(1.04 - 1.16) / (2 - (-2)) = -0.12 / 4 = \mathbf{-0.030\text{ V}}$ per unit of $\log Q$.
3. $n = 0.0592 / 0.030 = 1.97$, so **$n = 2$** .

Problem 14 Galvanic against electrolytic

Galvanic: chemical energy becomes electrical energy; anode negative; ΔG° negative (E°_{cell} positive); electrons come from the oxidation at the anode itself.

Electrolytic: electrical energy drives a chemical change; anode positive; ΔG° positive (E°_{cell} negative); electrons are pumped by the external power supply. **Oxidation is at the anode in both.**

Problem 15 Galvanised and tin-plated steel

1. Zinc has a lower E° than iron, so in the scratch zinc is oxidised in preference: it acts as a **sacrificial anode** and the iron stays the cathode.
2. Iron has a lower E° than tin, so iron becomes the anode and **corrodes faster** than it would bare.

Problem 16 Permanganate and iron(II) titration cell

1. Neither half-cell has a solid metal that takes part, so an **inert electrode** is needed to carry electrons.
2. $\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$, with $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^- \times 5$: **$\text{MnO}_4^- + 8\text{H}^+ + 5\text{Fe}^{2+} \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} + 5\text{Fe}^{3+}$** (charge +17 on each side).
3. $E^\circ_{\text{cell}} = 1.51 - 0.77 = \mathbf{+0.74\text{ V}}$; $n = 5$; $\Delta G^\circ = -5 \times 96\,485 \times 0.74 = -356\,995\text{ J/mol} = \mathbf{-357\text{ kJ/mol}}$.

Problem 17 Aluminium smelting

1. $Q = 5.00 \times 10^4 \times 3600 = 1.80 \times 10^8\text{ C}$; $n(\text{e}^-) = 1.80 \times 10^8 / 96\,485 = 1865.6\text{ mol}$; $n(\text{Al}) = 1865.6 / 3 = 621.9\text{ mol}$; $m = 621.9 \times 26.98 = 16\,778\text{ g} = \mathbf{16.8\text{ kg}}$.
2. $n(\text{O}_2) = 1865.6 / 4 = \mathbf{466\text{ mol}}$.
3. Water is reduced at -0.83 V , a less negative potential than -1.66 V , so **hydrogen forms**.

at the cathode before aluminium does.

Problem 18 A concentration cell

1. Both half-cells are the same couple, so $E^{\circ}_{\text{cell}} = 0$; the voltage comes from $Q \neq 1$.
2. Oxidation raises the concentration of the dilute side, so the 0.0010 mol/dm^3 half-cell is the anode.
3. $Q = 0.0010 / 1.0 = 10^{-3}$, so $E = 0 - (0.0592 / 2)(-3) = +0.089 \text{ V}$.
4. The two concentrations become equal, $Q = 1$, so $E \rightarrow 0$.

Problem 19 Spot the errors

1. The student **doubled the silver potential** and used $n = 1$ although two electrons are transferred.
2. $E^{\circ}_{\text{cell}} = 0.80 - (-0.76) = 1.56 \text{ V}$.
3. $\Delta G^{\circ} = -2 \times 96\,485 \times 1.56 = -301\,033 \text{ J/mol} = -301 \text{ kJ/mol}$.

Problem 20 Two cells in series

1. $n(\text{Cu}) = 0.640 / 63.55 = 0.010071 \text{ mol}$; two electrons each, so $n(e^{-}) = 0.0201 \text{ mol}$.
2. $\text{Ag}^{+} + e^{-} \rightarrow \text{Ag}$, so $n(\text{Ag}) = 0.020142 \text{ mol}$ and $m = 0.020142 \times 107.87 = 2.17 \text{ g}$.
3. $Q = 0.020142 \times 96\,485 = 1943 \text{ C}$; $I = 1943 / 1800 = 1.08 \text{ A}$.