- An aqueous solution of $CuSO_4$ is electrolysed with a current of 1.00 A for 60 minutes. Calculate the mass of the products that are formed at the two electrodes. **4**

If a current of 1.00 A is passed for 60 minutes, the number of electrons delivered is,

$$\text{moles of } e^- = \frac{It}{F} = \frac{(1.00\,\text{A}) \times (60 \times 60\,\text{s})}{(96485\,\text{C mol}^{-1})} = 0.037 \text{ mol}$$

At the cathode, the two possible reduction reactions are:

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\quad E° = +0.34$ V
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$ $\quad E° = -0.83$ V

The reduction of $Cu^{2+}(aq)$ is much easier (E is less negative) so Cu(s) is formed on the cathode.

[This potential for the reduction of water corresponds to $[H^+(aq)] = 1$ M. Using the Nernst equation gives $E = -0.42$ V for a pH 7 solution. An overpotential of 0.4 to 0.6 V must be added so that the reduction potential is actually ~1 V. Thus it is even more difficult to reduce water than the $E°$ value suggests.]

Each mole of Cu(s) requires 2 moles of electrons so,

number of moles of Cu(s) = ½ × 0.037 mol = 0.019 mol

The atomic mass of Cu is 63.55 g mol^{-1} so the mass of copper deposited is,

mass of Cu(s) = number of moles × molar mass
= 0.019 mol × 63.55 g mol^{-1} = 1.2 g

At the anode, oxidation of H_2O occurs to produce $O_2(g)$.

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$ $\quad E° = -1.23$ V

As SO_4^{2-} contains S(VI), further oxidation is very difficult.

Production of each mole of $O_2(g)$ requires 4 moles of electrons so,

number of moles of $O_2(g)$ = ¼ × 0.037 mol = 0.0093 mol

The molar mass of O_2 is (2 × 16.00) = 32.00 g mol^{-1} so the mass of $O_2(g)$ formed is,

mass of $O_2(g)$ = number of moles × molar mass
= 0.0093 mol × 32.00 g mol^{-1} = 0.30 g