- In the electrolytic production of Al, what mass of Al can be deposited in 2.00 hours by a current of 1.8 A? **2**

The number of moles of electrons passed in 2.00 hours by a current of 1.8 A is:

number of moles of electrons = *It* / *F*
= (1.8 A)(2.00 × 60.0 × 60.0 s) / 96485 C mol^{-1}
= 0.13 mol

Aluminium is produced from Al_2O_3 which contains Al^{3+}. 3 electrons are needed to produce each Al so 3 mol of electrons are needed to produce 1 mol of Al. This quantity of electrons will therefore deposit:

number of moles of Al = 0.13 / 3 mol = 0.045 mol

As Al has a molar mass of 26.98 g mol^{-1}, this quantity corresponds to:

mass of Al = number of moles × molar mass
= 0.045 mol × 26.98 g mol^{-1} = 1.2 g

Answer: **1.2 g**

- What products would you expect at the anode and the cathode on electrolysis of a 1 M aqueous solution of NiI_2? Explain your answers. **2**

At the cathode, there are two possible reduction reactions:

$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ $\quad$ E^o = -0.24 V

$2H_2O(l) + 4H^+(aq) + 4e^- \rightarrow H_2(g) + OH^-(aq)$ $\quad$ E^o = -0.41 V

Reduction of $Ni^{2+}(aq)$ is easier, even without considering an overpotential for water.

At the anode, there are two possible oxidation reactions:

$2I^-(aq) \rightarrow I_2(g) + 2e^-$ $\quad$ E^o = -0.62 V

$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$ $\quad$ E^o = -0.82 V

Both reactions will have an overpotential but oxidation of iodine is easier and this will probably occur.

Overall, Ni(s) will be produced at the cathode and $I_2(g)$ will be produced at the anode.