

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

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

$$\begin{aligned} \text{number of moles of electrons} &= It / F \\ &= (1.8 \text{ A})(2.00 \times 60.0 \times 60.0 \text{ s}) / 96485 \text{ C mol}^{-1} \\ &= 0.13 \text{ mol} \end{aligned}$$

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:

$$\text{number of moles of Al} = 0.13 / 3 \text{ mol} = 0.045 \text{ mol}$$

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

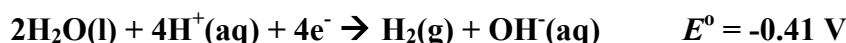
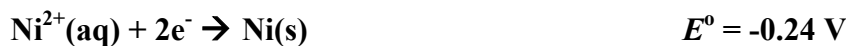
$$\begin{aligned} \text{mass of Al} &= \text{number of moles} \times \text{molar mass} \\ &= 0.045 \text{ mol} \times 26.98 \text{ g mol}^{-1} = 1.2 \text{ g} \end{aligned}$$

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.

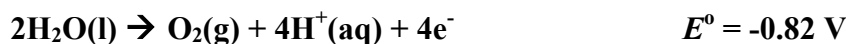
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At the cathode, there are two possible reduction reactions:



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

At the anode, there are two possible oxidation reactions:



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

Overall, $\text{Ni}(\text{s})$ will be produced at the cathode and $\text{I}_2(\text{g})$ will be produced at the anode.