Calculating the Mass of Electrolysis Product

Problem

Calculate the mass of aluminium produced by the electrolysis of molten aluminium chloride, if a current of 500 mA passes between the half-cells for 1.50 h.

What Is Required?

You need to calculate the mass of aluminium produced.

What Is Given?

You know the identity of the electrolyte, the current, and the time.
electrolyte: $AlCl_3(\ell)$
current: 500 mA
time: 1.50 h
You know the charge on one mole of electrons: 9.65×10^4 C/mol.

Plan Your Strategy

Use the current and the time to find the quantity of electric charge that passed from the anode to the cathode. From the charge, find the amount of electrons that passed through the circuit. Use the stoichiometry of the relevant half-reaction to relate the amount of electrons to the amount of aluminium produced. Use the molar mass of aluminium to convert the amount of aluminium to a mass of aluminium.

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Calculating the Mass of Electrolysis Product (continued)

Act on Your Strategy

To calculate the quantity of electrical charge in coulombs, convert the data to SI units:
1000 mA = 1 A

$$500\text{ mA} = \left(500\ \cancel{\text{mA}}\right)\left(\frac{1\text{ A}}{1000\ \cancel{\text{mA}}}\right)$$
$$= 0.500\text{ A}$$

$$1.50\text{ h} = \left(1.50\ \cancel{\text{h}}\right)\left(\frac{60\ \cancel{\text{min}}}{1\ \cancel{\text{h}}}\right)\left(\frac{60\text{ s}}{1\ \cancel{\text{min}}}\right)$$
$$= 5.4\times10^{3}\text{s}$$

$q = I\Delta t$
$q = (0.500\text{ A})(5400\text{ s})$
$= 2.700 \times 10^{3}\text{ C}$

Find the amount of electrons. One mole of electrons has a charge of 9.65×10^4 C:

$$\text{Amount of electrons } = 2700\ \cancel{\text{C}} \times \frac{1\text{ mol e}^-}{9.65\times10^4\ \cancel{\text{C}}}$$
$$= 0.0280\text{ mol e}^-$$

The half-reaction for the reduction of aluminium ions to aluminium atoms is:

$$Al^{3+} + 3e^- \rightarrow Al$$

Amount of aluminium formed:

$$= 0.0280\ \cancel{\text{mol e}^-} \times \frac{1\text{ mol Al}}{3\ \cancel{\text{mol e}^-}}$$
$$= 0.00933\text{ mol Al}$$

Convert the amount of aluminium to a mass:
$m_{\text{Al}} = nM_{\text{Al}}$
$= 0.252\text{ g}$

Check Your Solution

The answer is expressed in units of mass. To check your answer, use estimation. If the current were 1 A, then 1 mol of electrons would pass in 9.65×10^4 s. In this example, the current is less than 1 A, and the time is much less than 9.65×10^4 s. Therefore, much less than 1 mol of electrons would be used, and much less than 1 mol (27 g) of aluminium would be formed.

Using Dimensional Analysis:

$$m_{\text{Al}} = \left(\frac{0.500\ \cancel{\text{C}}}{1\ \cancel{\text{s}}}\times 5400\ \cancel{\text{s}}\right)\left(\frac{1\ \cancel{\text{mol e}^-}}{9.65\times10^4\ \cancel{\text{C}}}\right)\left(\frac{1\ \cancel{\text{mol}}\text{ Al}}{3\ \cancel{\text{mol e}^-}}\right)\left(\frac{27.0\text{ g}}{1\ \cancel{\text{mol}}}\right)$$
$$= 0.252\text{ g}$$

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