Marks

- A voltaic cell is constructed with a $Cr_2O_7^{2-}/Cr^{3+}$ (in acidic solution) half cell and a Sn/Sn^{2+} half cell. Measurement shows that the Sn electrode is negative. Write the balanced half equations and the overall spontaneous reaction. **3**

reduction half equation	**$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(l)$** **(reduction always occurs at the cathode which is positively charged)**
oxidation half equation	**$Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$** **(oxidation always occurs at the anode which is negatively charged)**
overall reaction	**$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 3Sn(s) \rightarrow 2Cr^{3+}(aq) + 7H_2O(l) + 3Sn^{2+}(aq)$**

- How many hours will it take to produce 1.00 kg of aluminium metal from a molten Al^{3+} salt, using a current of 100 A? **3**

Production to Al requires reduction of Al^{3+}: $Al^{3+} + 3e^- \rightarrow Al$.

Thus, production of 1 mol of Al requires 3 mol of electrons.

The atomic mass of Al is 26.98 so 1.00 kg corresponds to:

$$\text{number of moles of Al} = n = \frac{\text{mass}}{\text{atomic mass}} = \frac{(1.00 \times 10^3\ \text{g})}{(26.98\ \text{g mol}^{-1})} = 37.1\ \text{mol}$$

Hence, (3 × 37.1) mol = 111 mol of electrons are required. As number of moles of electrons delivered by a current is given by $\frac{I \times t}{F}$, the time taken to deliver this amount using a current of 100 A is:

$$t = \frac{nF}{I} = \frac{(111\ \text{mol} \times 96485\ \text{C mol}^{-1})}{(100\ \text{A})} = 107000\ \text{seconds} = \frac{10700}{(60 \times 60)} = 29.8\ \text{hours}$$

Answer: **29.8 hours**