Marks

- A melt containing Cr^{3+} is electrolysed for exactly 1 hour with a current of 0.54 A. Calculate the quantity of chromium that is deposited in this time at the electrode.

2

The number of moles of electrons delivered by a current *I* in time *t* is equal to $\frac{It}{F}$:

$$\text{number of moles of electrons} = \frac{It}{F} = \frac{(0.54\text{ A}) \times (60 \times 60\text{ s})}{96485\text{ C mol}^{-1}} = 0.0201\text{ mol}$$

The reduction reaction is $Cr^{3+} + 3e^- \rightarrow Cr$ so 3 mol of electrons are required for each mol of Cr. The number of moles of Cr produced is therefore:

number of moles of Cr = ⅓ × 0.0201 mol = 0.00672 mol

This corresponds to:

mass of Cr = number of moles × molar mass =

(0.00672 mol) × (52.00 g mol^{-1}) = 0.35 g

Answer: **0.35 g**

- An Ag electrode immersed in an aqueous solution containing $AgNO_3$ (0.010 M) and NaCN (1.00 M) has a potential of –0.66 V. Calculate the stability constant of the complex ion, $[Ag(CN)_2]^-$.

4

The standard reduction potential, *E*°, for the reaction $Ag^+(aq) + e^- \rightarrow Ag(s)$ is +0.80 V. The Nernst equation can be used to calculate the $[Ag^+(aq)]$ after CN^- is added. At this point, *E* = -0.66 V.

Using the Nernst equation for this 1 electron process:

$$E = E^\circ - \frac{RT}{nF}\ln Q$$

$$= (+0.80\text{ V}) - \frac{(8.314\text{ J K}^{-1}\text{ mol}^{-1}) \times (298\text{ K})}{1 \times 96485\text{ C mol}^{-1}} \ln\frac{1}{[Ag^+(aq)]} = -0.66\text{ V}$$

$$[Ag^+(aq)] = 2.0 \times 10^{-25}\text{ M}$$

A reaction table can be used to calculate work out the stability constant:

	$Ag^+(aq)$	$2CN^-(aq)$	$\rightleftharpoons$	$Ag(CN)_2^-(aq)$
Initial	**0.010**	**1.00**		**0**
Change	**-*x***	**-2*x***		**+*x***
Equilibrium	**2.0×10^{-25}**	**0.98**		**0.010**

Hence,

$$K_{stab} = \frac{[Ag(CN)_2^-(aq)]}{[Ag^+(aq)][CN^-(aq)]^2} = \frac{(0.010)}{(2.0 \times 10^{-25})(0.98)^2} = 5.1 \times 10^{-22}$$

Answer: **5.1×10^{-22}**