Marks

- A concentration cell containing aqueous solutions of $Cu(NO_3)_2$ and solid copper metal is constructed so that the Cu^{2+} ion concentration in the cathode half-cell is 0.66 M. Calculate the concentration of the Cu^{2+} ion in the anode half-cell if the cell potential for the concentration cell at 25 °C is 0.03 V. **2**

The cathode and anode reactions are:

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ (cathode)
$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ (anode)

The standard electrode potential $E°$ = 0 V and the potential can be calculated using the Nernst equation for this 2 electron reaction, n = 2:

$$E = E° - \frac{RT}{nF}\ln Q = -\frac{RT}{nF}\ln\left(\frac{[Cu^{2+}(aq)]_{anode}}{[Cu^{2+}(aq)]_{cathode}}\right)$$

$$= -\frac{(8.314 \text{ J K}^{-1}\text{ mol}^{-1})(298 \text{ K})}{(2 \times 96485 \text{ C mol}^{-1})}\ln\left(\frac{[Cu^{2+}(aq)]_{anode}}{0.66}\right) = +0.03 \text{ V}$$

This gives $[Cu^{2+}(aq)]_{anode}$ = 0.06 M.

Answer: **0.06 M**

- In **acid solution**, dichromate ion oxidises iron(II) to iron(III) as illustrated in the partial equation: **3**

$$Fe^{2+} + Cr_2O_7^{2-} \rightarrow Fe^{3+} + Cr^{3+}$$

Write a balanced equation for this reaction.

The half reactions are:

$Fe^{2+} \rightarrow Fe^{3+} + e^-$
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

where H^+ has been added to the $Cr_2O_7^{2-}$ / Cr^{3+} couple to give H_2O.

To balance the electrons, the first reaction needs to be multiplied by 6. Hence:

$6Fe^{2+} + Cr_2O_7^{2-} + 14H^+ \rightarrow 6Fe^{3+} + 2Cr^{3+} + 7H_2O$

What would happen to the cell potential if the concentration of Cr^{3+} were increased?

It would decrease. If $[Cr^{3+}]$ is increased, Le Châtelier's principle predicts that the reaction will shift towards reactants, reducing the cell potential.