Marks 4

- Calculate the initial cell potential for the following *unbalanced* reaction at 25 °C from the standard electrode potentials. Assume the concentration of all species is initially 1 M.

$$Fe^{3+}(aq) + Sn^{2+}(aq) \rightarrow Fe^{2+}(aq) + Sn^{4+}(aq)$$

The balanced equation is:

$$\mathbf{2Fe^{3+}(aq) + Sn^{2+}(aq) \rightarrow 2Fe^{2+}(aq) + Sn^{4+}(aq)}$$

The relevant half cell reactions and potentials are:

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $E^o = +0.77$ V
$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$ $E^o = -0.15$ V (oxidation so sign changed)

Hence,

$E^o_{cell} = (+0.77 - 0.15)$ V = +0.62 V

Answer: **+0.62 V**

Calculate the equilibrium constant, K, for the reaction at 25 °C.

Using $E^o = \frac{RT}{nF}\ln K$ and $T = (25 + 273)$ K = 298 K,

$$0.62 \text{ V} = \frac{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K})}{2 \times (96485 \text{ C mol}^{-1})} \ln K$$

$K = 9.4 \times 10^{20}$

Note that the reaction involves 2 electrons ($n = 2$) as each Sn^{2+} is oxidised to Sn^{4+} and two Fe^{3+} are reduced to two Fe^{2+}.

Answer: **9.4×10^{20}**