- Consider the following *unbalanced* reaction at 25 °C:

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

Calculate the standard cell potential. **3**

The two half cells and the standard reduction potentials are:

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $\quad E_{red}° = +0.77$ V
$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$ $\quad E_{red}° = +0.15$ V

As the Sn^{4+} / Sn^{2+} cell has the least positive reduction potential, it is reversed and becomes the oxidation half cell:

$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e$ $\quad E_{ox}° = -0.15$ V

Hence, the standard cell potential is:

$E° = (+0.77\text{ V}) + (-0.15\text{ V}) = +0.62$ V

Answer: **+0.62 V**

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

Using the half cell reactions above, the balanced equation is:

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

It involves the transfer of 2 electrons. The equilibrium constant for this 2 electron reaction is therefore given by:

$E° = (RT/nF)\ln K$

$\ln K = E° \times (nF / RT) = (0.62) \times (2 \times 96485) / (8.314 \times 298) = 48.3$

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

Answer: **9.4×10^{20}**