Marks

- Consider the following balanced redox reaction. 2

$$2In(s) + 3MnO_2(s) + 12H^+(aq) \rightarrow 2In^{3+}(aq) + 3Mn^{2+}(aq) + 6H_2O$$

If E^o = 1.568 V, what would be the measured potential of this cell at 298 K at the following concentrations?

$[H^+(aq)]$ = 0.25 M; $[In^{3+}(aq)]$ = 0.20 M; $[Mn^{2+}(aq)]$ = 0.42 M

The reaction quotient *Q* is given by:

$$Q = \frac{[In^{3+}(aq)]^2[Mn^{2+}]^3}{[H^+(aq)]^{12}}$$

The cell reaction involves a 6e⁻ process (2In → 2In(III) and 3Mn(IV) → 3Mn(II)). The cell potential can be calculated using the Nernst equation,

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

$$= (1.568\ \text{V}) - \frac{(8.314\ \text{J K}^{-1}\ \text{mol}^{-1}) \times (298\ \text{K})}{(6) \times (96485\ \text{C mol}^{-1})} \ln\left[\frac{(0.20)^2(0.42)^3}{(0.25)^{12}}\right] = 1.52\ \text{V}$$

Answer: **1.52 V**

- What is the value of the equilibrium constant for the following reaction at 298 K? 2

$$Cu^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Cu(s)$$

Relevant electrode potentials can be found on the data page.

The standard reduction potentials for $Cu^{2+}(aq)$ / Cu(s) and $Zn^{2+}(aq)$ / Zn(s) are +0.34 V and -0.76 V respectively. The $Zn^{2+}(aq)$ / Zn potential is the less positive and is reversed. The cell potential is therefore:

$$E^o_{cell} = ((0.34) - (-0.76))\ \text{V} = 1.10\ \text{V}$$

The equilibrium constant for this 2e⁻ process can be calculated using $E^o_{cell} = \frac{RT}{nF}\ln K$:

$$\ln K = \frac{(2) \times (96485\ \text{C mol}^{-1})}{(8.314\ \text{JK}^{-1}\ \text{mol}^{-1}) \times (298\ \text{K})} \times (1.10) = 85.7$$

$$K = e^{85.7} = 1.62 \times 10^{37}$$

Answer: **1.62×10^{37}**