Marks 3

- Zinc sulfate (0.50 g) is dissolved in 1.0 L of a 1.0 M solution of KCN. Write the chemical equation for the formation of the complex ion $[Zn(CN)_4]^{2-}$.

Calculate the concentration of Zn^{2+}(aq) in solution at equilibrium. Ignore any change in volume upon addition of the salt. K_{stab} of $[Zn(CN)_4]^{2-} = 4.2 \times 10^{19}$ M^{-4}.

The formula mass of $ZnSO_4$ is 65.39 (Zn) + 32.07 (S) + 4 × 16.00 (O) = 161.46.

0.50 g therefore corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{formula mass}} = \frac{0.50\,\text{g}}{161.46\,\text{g mol}^{-1}} = 0.0031\ \text{mol}$$

As $K_{stab} = 4.2 \times 10^{19}$ and is *very* large, the reaction essentially goes to completion. The reaction requires a 4:1 ratio CN^- : Zn^{2+}(aq) ions and as 0.0031 mol of Zn^{2+} and 1.0 mol of CN^- are present, CN^- is in excess.

Let the tiny amount of uncomplexed Zn^{2+}(aq) and its concentration in 1.0 L be:

$$\text{amount of } Zn^{2+}(aq) = x \text{ mol and } [Zn^{2+}(aq)] = \frac{\text{number of moles}}{\text{volume}} = \frac{x}{1.0}\ \text{M}$$

The amount of $[Zn(CN)_4]^{2+}$(aq) formed is therefore:

amount of $[Zn(CN)_4]^{2-}$(aq) = (0.0031 – x) ~ 0.0031 mol as x is so small.

Hence,

$$[[Zn(CN)_4]^{2-+}(aq)] \sim \frac{0.0031}{1.0} = 0.0031\ \text{M}$$

Formation of 0.0031 mol of $[Zn(CN)_4]^{2-}$(aq) requires (4 × 0.0031) = 0.012 mol of cyanide, leaving:

$$\text{amount of } CN^- = (1.0 - 0.012) = 0.99 \text{ mol and } [CN^-(aq)] = \frac{0.99}{1.0}\ \text{M}$$

Hence,

$$K_{stab} = \frac{[[Zn(CN)_4]^{2-}(aq)]}{[Zn^{2+}(aq)][CN^-(aq)]^4} = \frac{(0.0031)}{(x)(0.99)^4} = 4.2 \times 10^{19}$$

$$x = 7.7 \times 10^{-23} \text{ mol and so } [Zn^{2+}(aq)] = 7.7 \times 10^{-23}\ \text{M}$$

Answer: **7.8×10^{-23}**