Marks 7

- What is the solubility of $Cu(OH)_2$ in mol L^{-1}? K_{sp} ($Cu(OH)_2$) is 1.6×10^{-19} at 25 °C.

The dissolution reaction and associated solubility product are:

$$Cu(OH)_2(s) \rightleftharpoons Cu^{2+}(aq) + 2OH^-(aq) \qquad K_{sp} = [Cu^{2+}(aq)][OH^-(aq)]^2$$

If x mol dissolve in one litre, $[Cu^{2+}(aq)] = x$ M and $[OH^-(aq)] = 2x$. Hence:

$$K_{sp} = (x)(2x)^2 = 4x^3 = 1.6 \times 10^{-19}$$

$$x = 3.4 \times 10^{-7} \text{ M}$$

Answer: **3.4×10^{-7} M**

The overall formation constant for $[Cu(NH_3)_4]^{2+}$ is 1.0×10^{13}. Write the equation for the reaction of Cu^{2+} ions with excess ammonia solution.

$$Cu^{2+}(aq) + 4NH_3(aq) \rightleftharpoons [Cu(NH_3)_4]^{2+}(aq)$$

Calculate the value of the equilibrium constant for the following reaction.

$$Cu(OH)_2(s) + 4NH_3(aq) \rightleftharpoons [Cu(NH_3)_4]^{2+}(aq) + 2OH^-(aq)$$

This reaction can be considered to occur via (i) $Cu(OH)_2$ dissolving followed by (ii) the $Cu^{2+}(aq)$ ions that form being complexed by ammonia.

For the formation of $[Cu(NH_3)_4]^{2+}$, the equilibrium constant is:

$$K_{stab} = \frac{[Cu(NH_3)_4{}^{2+}]}{[Cu^{2+}][NH_3]^4} = 1.0 \times 10^{13}$$

For the reaction of $Cu(OH)_2(s)$ with $NH_3(aq)$, the equilibrium constant is:

$$K = \frac{[Cu(NH_3)_4{}^{2+}][OH^-(aq)]^2}{[NH_3]^4}$$

To obtain K, K_{sp} is multiplied by K_{stab}:

$$K = K_{sp} \times K_{stab}$$

$$= \cancel{[Cu^{2+}(aq)]}[OH^-(aq)]^2 \times \frac{[Cu(NH_3)_4{}^{2+}]}{\cancel{[Cu^{2+}]}[NH_3]^4} = \frac{[Cu(NH_3)_4{}^{2+}][OH^-(aq)]^2}{[NH_3]^4}$$

$$= (1.6 \times 10^{-19}) \times (1.0 \times 10^{13}) = 1.6 \times 10^{-6}$$

Answer: **1.6×10^{-6}**

Would you expect $Cu(OH)_2(s)$ to dissolve in 1 M NH_3 solution? Briefly explain your answer.

No. Equilibrium constant *K* is very small so the reaction lies heavily in favour of reactants.