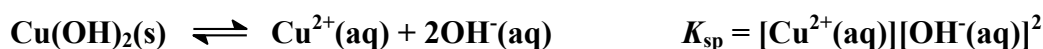


- What is the solubility of  $\text{Cu}(\text{OH})_2$  in  $\text{mol L}^{-1}$ ?  $K_{\text{sp}}(\text{Cu}(\text{OH})_2)$  is  $1.6 \times 10^{-19}$  at  $25^\circ\text{C}$ .

**Marks**  
7

The dissolution reaction and associated solubility product are:



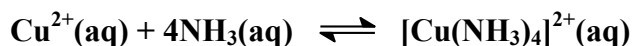
If  $x$  mol dissolve in one litre,  $[\text{Cu}^{2+}(\text{aq})] = x \text{ M}$  and  $[\text{OH}^{-}(\text{aq})] = 2x$ . Hence:

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

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

Answer:  $3.4 \times 10^{-7} \text{ M}$

The overall formation constant for  $[\text{Cu}(\text{NH}_3)_4]^{2+}$  is  $1.0 \times 10^{13}$ . Write the equation for the reaction of  $\text{Cu}^{2+}$  ions with excess ammonia solution.



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



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

For the formation of  $[\text{Cu}(\text{NH}_3)_4]^{2+}$ , the equilibrium constant is:

$$K_{\text{stab}} = \frac{[\text{Cu}(\text{NH}_3)_4]^{2+}}{[\text{Cu}^{2+}][\text{NH}_3]^4} = 1.0 \times 10^{13}$$

For the reaction of  $\text{Cu}(\text{OH})_2(\text{s})$  with  $\text{NH}_3(\text{aq})$ , the equilibrium constant is:

$$K = \frac{[\text{Cu}(\text{NH}_3)_4]^{2+}[\text{OH}^{-}(\text{aq})]^2}{[\text{NH}_3]^4}$$

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

$$\begin{aligned} K &= K_{\text{sp}} \times K_{\text{stab}} \\ &= [\text{Cu}^{2+}(\text{aq})][\text{OH}^{-}(\text{aq})]^2 \times \frac{[\text{Cu}(\text{NH}_3)_4]^{2+}}{[\text{Cu}^{2+}][\text{NH}_3]^4} = \frac{[\text{Cu}(\text{NH}_3)_4]^{2+}[\text{OH}^{-}(\text{aq})]^2}{[\text{NH}_3]^4} \\ &= (1.6 \times 10^{-19}) \times (1.0 \times 10^{13}) = 1.6 \times 10^{-6} \end{aligned}$$

Answer:  $1.6 \times 10^{-6}$

Would you expect  $\text{Cu}(\text{OH})_2(\text{s})$  to dissolve in 1 M  $\text{NH}_3$  solution? Briefly explain your answer.

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