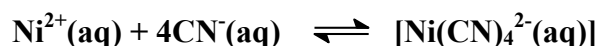


- Give the equilibrium concentration of $\text{Ni}^{2+}(\text{aq})$ ions in a solution formed by dissolving 0.15 mol of NiCl_2 in 0.500 L of 2.00 M KCN solution. The K_{stab} of $[\text{Ni}(\text{CN})_4]^{2-} = 1.7 \times 10^{30}$.

When the NiCl_2 is added to 0.500 L, the concentration of $\text{Ni}^{2+}(\text{aq})$:

$$\begin{aligned} \text{concentration} &= [\text{Ni}^{2+}(\text{aq})] = \text{number of moles} / \text{volume} \\ &= 0.15 \text{ mol} / 0.500 \text{ L} = 0.30 \text{ M} \end{aligned}$$

As K_{stab} is so large, essentially *all* of this will be complexed by the excess $\text{CN}^-(\text{aq})$ ions:



As essentially all of the $\text{Ni}^{2+}(\text{aq})$ becomes $[\text{Ni}(\text{CN})_4]^{2-}(\text{aq})$:

$$\begin{aligned} [\text{Ni}(\text{CN})_4]^{2-}(\text{aq}) &= 0.30 \text{ M} \\ [\text{CN}^-(\text{aq})] &= (2.00 - 4 \times 0.30) \text{ M} = 0.80 \text{ M} \end{aligned}$$

K_{stab} is the equilibrium constant for the reaction so:

$$K_{\text{stab}} = \frac{[\text{Ni}(\text{CN})_4]^{2-}(\text{aq})}{[\text{Ni}^{2+}(\text{aq})][\text{CN}^-(\text{aq})]^4}$$

If the *tiny* amount of uncomplexed $\text{Ni}^{2+}(\text{aq})$ has a concentration of x M then:

$$K_{\text{stab}} = \frac{(0.30)}{(x)(0.80)^4} = 1.7 \times 10^{30} \quad \text{so } x = 4.3 \times 10^{-31} \text{ M}$$

Answer: $4.3 \times 10^{-31} \text{ M}$