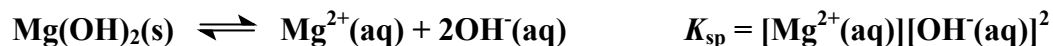


In the presence of excess hydroxide ion, Mg^{2+} can be precipitated as $\text{Mg}(\text{OH})_2(\text{s})$. What amount (in mol) of solid sodium hydroxide must be added to a 0.10 M solution of $\text{Mg}(\text{NO}_3)_2$ to just cause precipitation of $\text{Mg}(\text{OH})_2(\text{s})$. The solubility product constant of $\text{Mg}(\text{OH})_2$ is $7.1 \times 10^{-12} \text{ M}^3$.

The solubility equilibrium and product for $\text{Mg}(\text{OH})_2(\text{s})$ are:



With $[\text{Mg}^{2+}(\text{aq})] = 0.10 \text{ M}$, precipitation will occur when:

$$[\text{OH}^{-}(\text{aq})]^2 \geq \frac{K_{\text{sp}}}{[\text{Mg}^{2+}(\text{aq})]} = \frac{(7.1 \times 10^{-12})}{(0.10)} \text{ so } [\text{OH}^{-}(\text{aq})] \geq 8.4 \times 10^{-6} \text{ M}$$

(As the volume of the solution is not specified, the number of moles of $\text{NaOH}(\text{s})$ cannot be given. A 1 L solution would require $8.4 \times 10^{-6} \text{ mol}$.)

ANSWER: $8.4 \times 10^{-6} \text{ M}$

In a separate experiment, the $\text{Mg}(\text{OH})_2$ is precipitated by adding 0.10 mol of $\text{Mg}(\text{NO}_3)_2$ to 1.0 L of a 0.10 M NH_3 solution. What amount (in mol) of NH_4Cl must be added to this solution to just dissolve the precipitate? The $\text{p}K_{\text{a}}$ of NH_4Cl is 9.24.

With $[\text{Mg}^{2+}(\text{aq})] = 0.10 \text{ M}$, dissolution will start to occur when:

$$[\text{OH}^{-}(\text{aq})]^2 \leq \frac{K_{\text{sp}}}{[\text{Mg}^{2+}(\text{aq})]} = \frac{(7.1 \times 10^{-12})}{(0.10)} \text{ so } [\text{OH}^{-}(\text{aq})] \leq 8.4 \times 10^{-6} \text{ M}$$

This $[\text{OH}^{-}(\text{aq})]$ corresponds to $\text{pOH} = -\log_{10}([\text{OH}^{-}(\text{aq})]) = -\log_{10}(8.4 \times 10^{-6}) = 5.1$. Using $\text{pH} = 14.0 - \text{pOH}$, $\text{pH} = (14.0 - 5.1) = 8.9$.

When NH_4Cl is added, the solution contains an acid (NH_4Cl) and its conjugate base (NH_3). The solution contains an acid (NH_4Cl) and its conjugate base (NH_3). The Henderson-Hasselbalch equation can be used to work out the required $[\text{NH}_4\text{Cl}]$ with $[\text{NH}_3] = 0.10 \text{ M}$ and $\text{pH} = 8.9$:

$$\text{pH} = \text{p}K_{\text{a}} + \log_{10} \left(\frac{[\text{base}]}{[\text{acid}]} \right) = 9.24 + \log_{10} \left(\frac{0.10}{[\text{NH}_4\text{Cl}]} \right) = 8.9$$

$$\left(\frac{0.10}{[\text{NH}_4\text{Cl}]} \right) = 10^{(8.9 - 9.24)} \text{ so } [\text{NH}_4\text{Cl}] = 0.2 \text{ M}$$

This molarity is for a 1.0 L solution so that 0.2 mol of NH_4Cl are required.

ANSWER: 0.2 mol