1. 500.0 mL of 2.0 \times 10^{-4} \text{ M} \ Pb(NO_3)_2 \text{ solution is mixed with 800.0 mL of 3.0 \times 10^{-3} \text{ M} NaI solution. Do the necessary calculations to see if a precipitate will form or not.}

The possible precipitate would be PbI_2. (The NaNO_3 is soluble)

The equilibrium dissociation equation for the possible precipitate:

\[ \text{PbI}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2\text{I}^-(aq) \]

The [Pb^{2+}] right after mixing:

\[ [\text{Pb}^{2+}] = 2.0 \times 10^{-4} \text{ M} \times \frac{500.0 \text{ mL}}{1300.0 \text{ mL}} = 7.692 \times 10^{-5} \text{ M} \]

The [I^-] right after mixing:

\[ [\text{I}^-] = 3.0 \times 10^{-3} \text{ M} \times \frac{800.0 \text{ mL}}{1300.0 \text{ mL}} = 1.846 \times 10^{-3} \text{ M} \]

Trial \( K_{sp} = [\text{Pb}^{2+}][\text{I}^-]^2 \)

\[ = (7.692 \times 10^{-5})(1.846 \times 10^{-3})^2 = 2.6 \times 10^{-10} \]

\( K_{sp} = 8.5 \times 10^{-9} \) so Trial \( K_{sp} < K_{sp} \)

So there is NO PRECIPITATE.

2. If 5.5 grams of AgNO_3 solid is added to 50.0 mL of 2.0 \times 10^{-3} \text{ M} KIO_3 solution, will a precipitate of AgIO_3 form?

Possible precipitate would be AgIO_3

Equilibrium equation: \[ \text{AgIO}_3(s) \rightleftharpoons \text{Ag}^+(aq) + \text{IO}_3^-(aq) \]

Plan: g \rightarrow \text{mol} \rightarrow \text{M} [\text{AgNO}_3]

\[ 5.5 \text{ g AgNO}_3 \times \frac{1 \text{ mol}}{169.9 \text{ g}} = 0.03237 \text{ mol AgNO}_3 \]

\[ [\text{Ag}^+] = [\text{AgNO}_3] = \frac{0.03237 \text{ mol}}{0.0500 \text{ L}} = 0.6474 \text{ M} \]

\[ [\text{IO}_3^-] = [\text{KIO}_3] = 2.0 \times 10^{-3} \text{ M} \]

Trial \( K_{sp} = [\text{Ag}^+][\text{IO}_3^-] = (0.6474)(2.0 \times 10^{-3}) = 1.3 \times 10^{-3} \)

The actual \( K_{sp} \) of AgIO_3 is 3.2 \times 10^{-8} so Trial \( K_{sp} > K_{sp} \)

And there IS a Precipitate
3. Find the maximum possible [IO$_3^-$] in a solution in which [Pb$^{2+}$] = 3.0 x 10$^{-4}$ M.

**Equilibrium Equation for Precipitate:** Pb(IO$_3$)$_2$(s) ⇄ Pb$^{2+}$(aq) + 2 IO$_3^-$ (aq)

$$K_{sp} = [Pb^{2+}][IO_3^-]^2$$

$$[IO_3^-]^2 = \frac{K_{sp}}{[Pb^{2+}]}$$

$$[IO_3^-] = \sqrt{\frac{K_{sp}}{[Pb^{2+}]}} = \sqrt{\frac{3.7 \times 10^{-13}}{3.0 \times 10^{-4}}} = 3.5 \times 10^{-5} \text{ M}$$

4. If 0.20 M Na$_2$CO$_3$ solution is added slowly to a mixture of 0.010 M Ba(NO$_3$)$_2$ and 0.010 M AgNO$_3$, which precipitate would form first. Show all calculations in a logical way.

To find the [CO$_3^{2-}$] needed to start precipitation of BaCO$_3$:

$$\text{BaCO}_3(s) \rightleftharpoons \text{Ba}^{2+}(aq) + \text{CO}_3^{2-}(aq)$$

$$K_{sp} = [\text{Ba}^{2+}][\text{CO}_3^{2-}]$$

$$2.6 \times 10^{-9} = (0.010) \times [\text{CO}_3^{2-}]$$

$$[\text{CO}_3^{2-}] = \frac{2.6 \times 10^{-9}}{0.010} = 2.6 \times 10^{-7} \text{ M}$$

So the [CO$_3^{2-}$] needed to start precipitation of BaCO$_3$ is **2.6 x 10$^{-7}$ M**

To find the [CO$_3^{2-}$] needed to start precipitation of Ag$_2$CO$_3$:

$$\text{Ag}_2\text{CO}_3(s) \rightleftharpoons 2\text{Ag}^+(aq) + \text{CO}_3^{2-}(aq)$$

$$K_{sp} = [\text{Ag}^+]^2[\text{CO}_3^{2-}]$$

$$[\text{CO}_3^{2-}] = \frac{K_{sp}}{[\text{Ag}^+]^2}$$

$$[\text{CO}_3^{2-}] = \frac{8.5 \times 10^{-12}}{(0.010)^2} = 8.5 \times 10^{-8} \text{ M}$$

So you would need a lower [CO$_3^{2-}$] to start precipitation of Ag$_2$CO$_3$, so

**The Ag$_2$CO$_3$ will precipitate first!**