

Self-Test on Tutorial 11-KEY

1. 500.0 mL of 2.0×10^{-4} M $\text{Pb}(\text{NO}_3)_2$ solution is mixed with 800.0 mL of 3.0×10^{-3} 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:



The $[\text{Pb}^{2+}]$ 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 $[\text{I}^{-}]$ right after mixing:

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

$$\text{Trial Ksp} = [\text{Pb}^{2+}][\text{I}^{-}]^2$$

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

$$\text{Ksp} = 8.5 \times 10^{-9} \text{ so Trial Ksp} < \text{Ksp}$$

So there is NO PRECIPITATE.

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

Possible precipitate would be AgIO_3



Plan : g \rightarrow mol \rightarrow M [AgNO_3]

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

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

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

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

The actual Ksp of AgIO_3 is 3.2×10^{-8} so Trial Ksp $>$ Ksp

And there IS a Precipitate

3. Find the maximum possible $[\text{IO}_3^-]$ in a solution in which $[\text{Pb}^{2+}] = 3.0 \times 10^{-4} \text{ M}$.



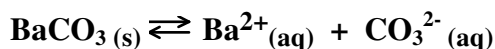
$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{IO}_3^-]^2$$

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

$$[\text{IO}_3^-] = \sqrt{\frac{K_{\text{sp}}}{[\text{Pb}^{2+}]}} = \sqrt{\frac{3.7 \times 10^{-13}}{3.0 \times 10^{-4}}} = 3.5 \times 10^{-5} \text{ M}$$

4. If 0.20 M Na_2CO_3 solution is added slowly to a mixture of 0.010 M $\text{Ba}(\text{NO}_3)_2$ and 0.010 M AgNO_3 , which precipitate would form first. Show all calculations in a logical way.

To find the $[\text{CO}_3^{2-}]$ needed to start precipitation of BaCO_3 :



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

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

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

So the $[\text{CO}_3^{2-}]$ needed to start precipitation of BaCO_3 is $2.6 \times 10^{-7} \text{ M}$

To find the $[\text{CO}_3^{2-}]$ needed to start precipitation of Ag_2CO_3 :



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

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

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

So you would need a lower $[\text{CO}_3^{2-}]$ to start precipitation of Ag_2CO_3 , so

The Ag_2CO_3 will precipitate first!