Self-Test on Tutorial 11-KEY

1. 500.0 mL of 2.0×10^{-4} M $Pb(NO_3)_2$ 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:

   $$\mathbf{PbI_{2\,(s)} \rightleftarrows Pb^{2+}_{(aq)} + 2I^{-}_{(aq)}}$$

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

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

   The $[I^-]$ right after mixing:

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

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

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

   Ksp = 8.5×10^{-9} so Trial Ksp < 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 $AgIO_3$

   Equilibrium equation: $\mathbf{AgIO_{3\,(s)} \rightleftarrows Ag^{+}_{(aq)} + IO_3^{-}{}_{(aq)}}$

   Plan : g → mol → M $[AgNO_3]$

   $$5.5\text{ g } AgNO_3 \times \frac{1\text{ mol}}{169.9\text{ g}} = 0.03237\text{ mol } AgNO_3$$

   $$[Ag^+] = [AgNO_3] = \frac{0.03237\text{ mol}}{0.0500\text{ L}} = 0.6474\text{ M}$$

   $$[IO_3^-] = [KIO_3] = 2.0 \times 10^{-3}\text{ M}$$

   Trial Ksp $= [Ag^+][IO_3^-] = (0.6474)(2.0 \times 10^{-3}) = 1.3 \times 10^{-3}$

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

   And there <u>IS</u> a Precipitate

3. Find the maximum possible $[IO_3^-]$ in a solution in which $[Pb^{2+}] = 3.0 \times 10^{-4}$ M.

Equilibrium Equation for Precipitate: $Pb(IO_3)_{2(s)} \rightleftarrows Pb^{2+}_{(aq)} + 2\, IO_3^-{}_{(aq)}$

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

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

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

4. If 0.20 M Na_2CO_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 $BaCO_3$:

$$BaCO_{3\,(s)} \rightleftarrows Ba^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$$

$$K_{sp} = [Ba^{2+}][CO_3^{2-}]$$

$$2.6 \times 10^{-9} = (0.010) \times [CO_3^{2-}]$$

$$[CO_3^{2-}] = \frac{2.6 \times 10^{-9}}{0.010} = 2.6 \times 10^{-7}\ M$$

So the $[CO_3^{2-}]$ needed to start precipitation of $BaCO_3$ is **2.6×10^{-7} M**

To find the $[CO_3^{2-}]$ needed to start precipitation of Ag_2CO_3:

$$Ag_2CO_{3\,(s)} \rightleftarrows 2Ag^+_{(aq)} + CO_3^{2-}{}_{(aq)}$$

$$Ksp = [Ag^+]^2[CO_3^{2-}]$$

$$[CO_3^{2-}] = \frac{Ksp}{[Ag^+]^2}$$

$$[CO_3^{2-}] = \frac{8.5 \times 10^{-12}}{(0.010)^2} = 8.5 \times 10^{-8}\ M$$

So you would need a lower $[CO_3^{2-}]$ to start precipitation of Ag_2CO_3, so

The Ag_2CO_3 will precipitate first!