

Chemistry 12
Tutorial 11
Solutions

1. 250.0 mL of $3.0 \times 10^{-4} \text{ M Ba(NO}_3)_2$ is mixed with 350.0 mL of $0.0020 \text{ M Na}_2\text{SO}_4$ solution.

- a) Determine which product could possibly be a precipitate.

The possible precipitate would be BaSO_4 . (The NaNO_3 is soluble)

- b) Write the equilibrium dissociation equation for the possible precipitate in (a).



- c) Calculate the $[\text{Ba}^{2+}]$ right after mixing

$$[\text{Ba}^{2+}] = 3.0 \times 10^{-4} \text{ M} \times \frac{250.0 \text{ mL}}{600.0 \text{ mL}} = 1.25 \times 10^{-4} \text{ M}$$

- d) Calculate the $[\text{SO}_4^{2-}]$ right after mixing

$$[\text{SO}_4^{2-}] = 0.0020 \text{ M} \times \frac{350.0 \text{ mL}}{600.0 \text{ mL}} = 1.167 \times 10^{-3} \text{ M}$$

- e) Calculate the *trial* K_{sp} .

$$\text{Trial } K_{sp} = [\text{Ba}^{2+}] [\text{SO}_4^{2-}] = (1.25 \times 10^{-4}) (1.167 \times 10^{-3}) = 1.46 \times 10^{-7}$$

- d) Which is greater, the trial K_{sp} or the real K_{sp} ?

$$K_{sp} = 1.1 \times 10^{-10} \text{ so Trial } K_{sp} > K_{sp}$$

- e) Would a precipitate form in this case? **Yes.**

2. Will a precipitate form if 3.8 grams of $\text{Ca}(\text{NO}_3)_2$ is added to 250.0 mL of 0.0050 M Na_2SO_4 solution? Calculate the Trial K_{sp} first.

Possible precipitate would be CaSO_4



Plan : g \rightarrow mol \rightarrow M [$\text{Ca}(\text{NO}_3)_2$]

$$3.8 \text{ g Ca}(\text{NO}_3)_2 \times \frac{1 \text{ mol}}{164.1 \text{ g}} = 0.02316 \text{ mol Ca}(\text{NO}_3)_2$$

$$[\text{Ca}^{2+}] = [\text{Ca}(\text{NO}_3)_2] = \frac{0.02316 \text{ mol}}{0.250 \text{ L}} = 0.09263 \text{ M}$$

$$[\text{SO}_4^{2-}] = [\text{Na}_2\text{SO}_4] = 0.0050 \text{ M}$$

$$\text{Trial } K_{\text{sp}} = [\text{Ca}^{2+}][\text{SO}_4^{2-}] = (0.09263)(0.0050) = 4.6 \times 10^{-4}$$

The actual K_{sp} of CaSO_4 is 7.1×10^{-5} so Trial $K_{\text{sp}} > K_{\text{sp}}$

And there IS a Precipitate

Calculate the maximum possible [Cl⁻] in a solution in which [Pb²⁺] is 1.0×10^{-3} M.



K_{sp} Expression: $K_{\text{sp}} = [\text{Pb}^{2+}][\text{Cl}^{-}]^2$

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

$$[\text{Cl}^{-}] = \sqrt{\frac{K_{\text{sp}}}{[\text{Pb}^{2+}]}} = \sqrt{\frac{1.2 \times 10^{-5}}{1.0 \times 10^{-3}}} = 0.11 \text{ M}$$

4. Calculate the $[\text{Pb}^{2+}]$ necessary to just start the precipitation of PbBr_2 from a solution in which the $[\text{Br}^-]$ is 0.0030 M.



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

$$[\text{Pb}^{2+}] = \frac{K_{sp}}{[\text{Br}^-]^2} = \frac{6.6 \times 10^{-6}}{(0.0030)^2} = 0.73 \text{ M}$$

5. Calculate the mass of NaI which must be added to 500.0 mL of a 2.0×10^{-4} M solution of $\text{Pb}(\text{NO}_3)_2$ in order to form a saturated solution of PbI_2 .



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

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

$$[\text{I}^-] = \sqrt{\frac{K_{sp}}{[\text{Pb}^{2+}]}} = \sqrt{\frac{8.5 \times 10^{-9}}{2.0 \times 10^{-4}}} = 6.519 \times 10^{-3} \text{ M}$$

$$\text{moles I}^- = \text{moles of NaI} = 6.519 \times 10^{-3} \text{ M} \times 0.5000 \text{ L} = 3.2596 \times 10^{-3} \text{ moles}$$

$$\text{mass of NaI} = 3.2596 \times 10^{-3} \text{ moles} \times \frac{149.9 \text{ g}}{1 \text{ mol}} = 0.49 \text{ g NaI}$$

6. If 0.1 M KCl is added dropwise to a beaker containing 0.10 M Ag^+ and 0.10 M Pb^{2+} , which precipitate would form first? Show all equations and calculations.

To get the $[\text{Cl}^-]$ needed to start precipitation of AgCl:



$$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

$$1.8 \times 10^{-10} = (0.10) \times [\text{Cl}^-]$$

$$[\text{Cl}^-] = \frac{1.8 \times 10^{-10}}{0.10} = 1.8 \times 10^{-9} \text{ M}$$

So the $[\text{Cl}^-]$ needed to just start precipitation with Ag^+ is 1.8×10^{-9} M

To get the $[\text{Cl}^-]$ needed to start precipitation of PbCl_2 :

Equilibrium Equation for Precipitate: $\text{PbCl}_{2(s)} \rightleftharpoons \text{Pb}^{2+}_{(aq)} + 2\text{Cl}^{-}_{(aq)}$

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

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

$$[\text{Cl}^-] = \sqrt{\frac{K_{sp}}{[\text{Pb}^{2+}]}} = \sqrt{\frac{1.2 \times 10^{-5}}{0.10}} = 1.1 \times 10^{-2} \text{ M}$$

So the $[\text{Cl}^-]$ needed to just start precipitation with Pb^{2+} is $1.1 \times 10^{-2} \text{ M}$

Since the $[\text{Cl}^-]$ will reach $1.8 \times 10^{-9} \text{ M}$ before it reaches $1.1 \times 10^{-2} \text{ M}$, the precipitate AgCl will form first.

Self-Test on Tutorial 11

- 500.0 mL of $2.0 \times 10^{-4} \text{ M}$ $\text{Pb}(\text{NO}_3)_2$ 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:



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:

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

Trial K_{sp} :

$$\begin{aligned}\text{Trial } K_{sp} &= [\text{Pb}^{2+}][\text{I}^-]^2 \\ &= (7.692 \times 10^{-5}) (1.846 \times 10^{-3})^2 = 2.6 \times 10^{-10}\end{aligned}$$

$$K_{sp} = 8.5 \times 10^{-9} \quad \text{so Trial } K_{sp} < K_{sp}$$

So there is **NO PRECIPITATE**.

2. If 5.5 grams of AgNO_3 solid is added to 50.0 mL of 2.0×10^{-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]

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

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

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

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

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

And there **IS** a Precipitate

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



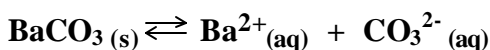
$$K_{\text{sp}} = [\text{Pb}^{2+}][\text{IO}_3^-]^2$$

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

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

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!