

1. One Source vs. Two Source Solubility Problems
2. Challenging Solubility Problems
3. Prediction of Forming a Precipitate

One Source vs. Two Source Solubility Problems

One Source	Two Source
<ul style="list-style-type: none"> • Both ions come from the same salt (source) $\text{PbI}_{2(s)} \rightleftharpoons \text{Pb}^{2+}_{(aq)} + 2 \text{I}^{-}_{(aq)}$ <div style="text-align: center; margin: 10px 0;"> </div> $K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2$ $(s)(2s)^2 = 4s^3$	<ul style="list-style-type: none"> • Both ions come from a different salt (source) <div style="display: flex; justify-content: space-around; margin-bottom: 20px;"> <div style="text-align: center;"> $\text{Pb}(\text{NO}_3)_{2(s)} \rightleftharpoons \text{Pb}^{2+}_{(aq)} + 2 \text{NO}_3^{-}_{(aq)}$ <p style="margin: 0;">x</p> </div> <div style="text-align: center;"> $\text{KI}_{(s)} \rightleftharpoons \text{K}^{+}_{(aq)} + \text{I}^{-}_{(aq)}$ <p style="margin: 0;">y</p> </div> </div> $\text{PbI}_{2(s)} \rightleftharpoons \text{Pb}^{2+}_{(aq)} + 2 \text{I}^{-}_{(aq)}$ <div style="text-align: center; margin: 10px 0;"> </div> $K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2 = (x)(y)^2$
<ul style="list-style-type: none"> • Ion concentrations are related through mole ratio <ul style="list-style-type: none"> ❖ 1:1 ratio = $(s)(s) = s^2$ ❖ 1:2 ratio = $(s)(2s)^2 = 4s^3$ ❖ 1:3 ratio = $(s)(3s)^3 = 27s^4$ ❖ 2:3 ratio = $(2s)^2(3s)^3 = 108s^5$ 	<ul style="list-style-type: none"> • Related through K_{sp} ← from data booklet • Ex: Find the $[\text{I}^{-}]$ if $[\text{Pb}^{2+}] = 4.5 \times 10^{-3} \text{ M}$. ← x $K_{sp} = 8.5 \times 10^{-9} = (x)(y)^2$ $8.5 \times 10^{-9} = (4.5 \times 10^{-3})(y^2)$ $y = \sqrt{\frac{8.5 \times 10^{-9}}{4.5 \times 10^{-3}}}$ $= 1.37 \times 10^{-3} \text{ M} = [\text{I}^{-}]$

2 source:
can't
use
S²⁻!

Challenging Solubility Problems

1. A solution has a concentration of calcium ions equal to 2.5×10^{-2} M. What is the maximum concentration of sulphate ions allowed to be added without causing precipitation?

$$K_{sp} = 7.1 \times 10^{-5}$$



→ At saturation (equilibrium)



$$K_{sp} = [\text{Ca}^{2+}][\text{SO}_4^{2-}]$$

$$7.1 \times 10^{-5} = (2.5 \times 10^{-2})[\text{SO}_4^{2-}]$$

$$[\text{SO}_4^{2-}] = \boxed{2.8 \times 10^{-3} \text{ M}}$$

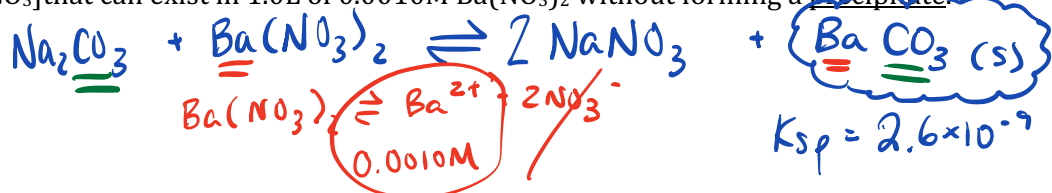
2. Determine the maximum $[\text{Na}_2\text{CO}_3]$ that can exist in 1.0L of 0.0010M $\text{Ba}(\text{NO}_3)_2$ without forming a precipitate.

A. 2.6×10^{-12} M

B. 2.6×10^{-9} M

C. 2.6×10^{-6} M

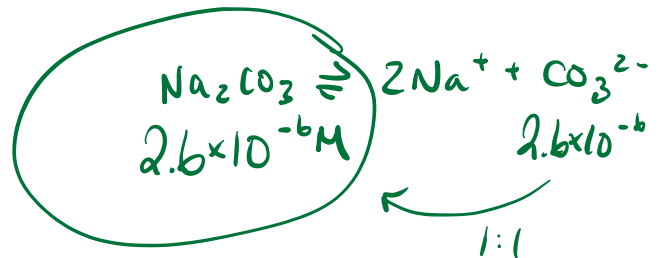
D. 5.1×10^{-5} M



$$K_{sp} = 2.6 \times 10^{-9} = [\text{Ba}^{2+}][\text{CO}_3^{2-}]$$

$$2.6 \times 10^{-9} = (0.0010)[\text{CO}_3^{2-}]$$

$$[\text{CO}_3^{2-}] = 2.6 \times 10^{-6} \text{ M}$$



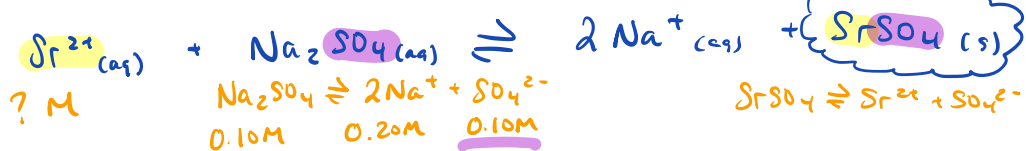
3. What is the maximum $[\text{Sr}^{2+}]$ that can exist in a solution of 0.10 M Na_2SO_4 ?

A. 3.4×10^{-7} M

B. 3.4×10^{-6} M

C. 1.7×10^{-6} M

D. 5.8×10^{-4} M



$$K_{sp} = 3.4 \times 10^{-7} = [\text{Sr}^{2+}][\text{SO}_4^{2-}]$$

$$[\text{Sr}^{2+}] = \frac{3.4 \times 10^{-7}}{0.10 \text{ M}} = \boxed{3.4 \times 10^{-6} \text{ M}}$$

4. When 100.0 mL of $4.0 \times 10^{-2} \text{ M CaCl}_2$ is added to 150.0 mL of $2.9 \times 10^{-2} \text{ M NaOH}$, a precipitate just starts to form. What is the K_{sp} of this precipitate?

- Write a (balanced) double replacement reaction.



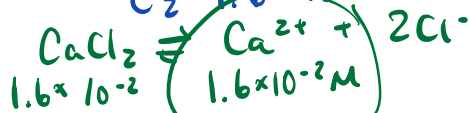
- What is the possible precipitate? Write the K_{sp} expression.



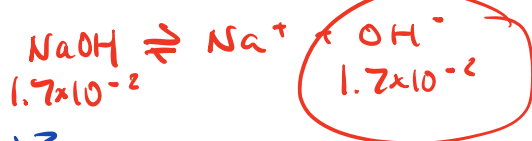
$$K_{sp} = [\text{Ca}^{2+}][\text{OH}^-]^2$$

$C_1V_1 = C_2V_2$ Calculate the diluted concentrations of each ion.

[CaCl₂]
 $C_1V_1 = C_2V_2$
 $(4.0 \times 10^{-2})(100.0) = (C_2)(250.0)$
 $C_2 = 1.6 \times 10^{-2} \text{ M} = [\text{CaCl}_2]$



[NaOH]
 $C_1V_1 = C_2V_2$
 $(2.9 \times 10^{-2})(150) = (C_2)(250)$
 $C_2 = 1.7 \times 10^{-2} \text{ M} = [\text{NaOH}]$



- Calculate the K_{sp} value.

$$K_{sp} = (1.6 \times 10^{-2})(1.7 \times 10^{-2})^2$$

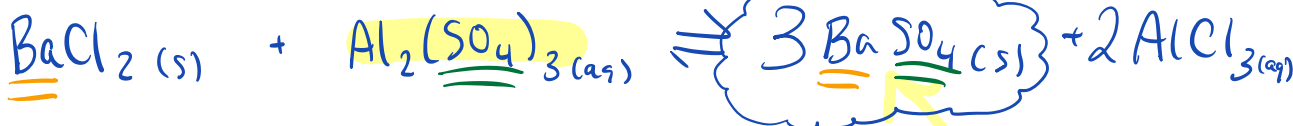
$$= 4.6 \times 10^{-6}$$

Solute (solid)

Solvent (liquid)

5. Up to 15.0g of BaCl_2 can be dissolved in 2.5L of $\text{Al}_2(\text{SO}_4)_3$ without a precipitate being formed. Find $[\text{Al}_2(\text{SO}_4)_3]$.

- Write the double replacement reaction. (What is the solute? What is the solvent?)

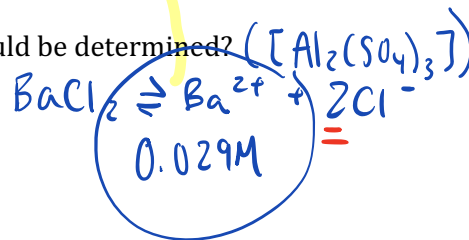


- What is the possible precipitate that could be formed? Write the K_{sp} expression and determine its value from the data booklet.

$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.1 \times 10^{-10}$$

- Looking at the K_{sp} expression, is there an ion concentration value that could be determined? ($[\text{Al}_2(\text{SO}_4)_3]$)

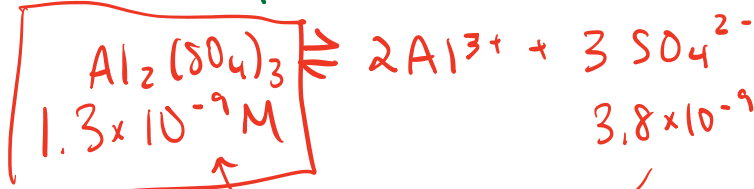
$$[\text{BaCl}_2] = \frac{15.0 \text{ g}}{2.5 \text{ L}} \times \frac{1 \text{ mol}}{208.3 \text{ g}} = 0.029 \text{ M}$$



$$(0.029)[\text{SO}_4^{2-}] = 1.1 \times 10^{-10}$$

$$[\text{SO}_4^{2-}] = 3.8 \times 10^{-9} \text{ M}$$

Worksheet 3.3 #1-7



÷ 3

Prediction of Forming a Precipitate

When **two different solutions** are mixed, we can predict whether a precipitate will form. The K_{sp} value represents the maximum product of the ion concentrations in a saturated solution.

Ⓢ Saturation / equilibrium

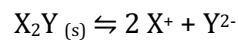
If an equilibrium is not present in solution, then we calculate a **trial ion product (TIP)** - (also called a **trial K_{sp} value** or **reaction quotient, Q**)

★

If Trial $K_{sp} >$ Actual K_{sp} - a precipitate forms. (oversaturated)

If Trial $K_{sp} <$ Actual K_{sp} - no precipitate forms. (undersaturated)

If Trial $K_{sp} =$ Actual K_{sp} - the solution is saturated.

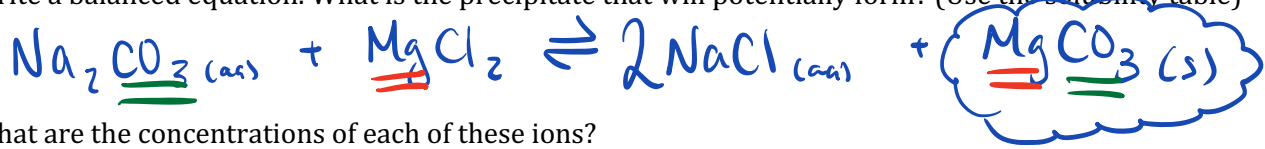


$$K_{sp} = [X^+]^2 [Y^{2-}]$$

Example.

(1) Will a precipitate form when 23 mL of 0.020 M Na_2CO_3 is added to 12 mL of 0.010 M $MgCl_2$?

- Write a balanced equation. What is the precipitate that will potentially form? (Use the solubility table)



- What are the concentrations of each of these ions?

$[MgCl_2]$
 $(0.010)(12) = C_2(35)$
 $C_2 = 0.0034 M = [MgCl_2]$
 $MgCl_2 \rightleftharpoons Mg^{2+} + 2Cl^-$
 $0.0034 M$

$[Na_2CO_3]$
 $(0.020)(23) = C_2(35)$
 $C_2 = 0.013 M = [Na_2CO_3]$
 $Na_2CO_3 \rightleftharpoons 2Na^+ + CO_3^{2-}$
 $0.013 M$

- Calculate the value of TIP (Trial K_{sp})

$$\begin{aligned} \text{Trial } K_{sp} &= [Mg^{2+}][CO_3^{2-}] \\ &= (0.0034 M)(0.013 M) \\ &= 4.4 \times 10^{-5} \end{aligned}$$

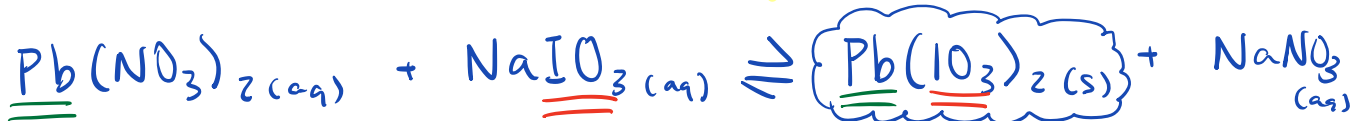
- Compare the TIP (Trial K_{sp}) with the real K_{sp} . Will a precipitate form?

$$\text{Actual } K_{sp} = 6.8 \times 10^{-6}$$

Trial > Actual
 \therefore Yes, a ppt will form!

(2) Will a precipitate form when 8.5 mL of 6.3×10^{-2} M lead (II) nitrate is added to 1.0 L of 1.2×10^{-3} M sodium iodate?

1000 mL



$[\text{Pb}(\text{NO}_3)_2]$
 $(6.3 \times 10^{-2})(8.5) = C_2(1008.5)$
 $C_2 = 5.31 \times 10^{-4} \text{ M}$
 $\text{Pb}(\text{NO}_3)_2 \rightleftharpoons \text{Pb}^{2+} + 2\text{NO}_3^-$
 $5.31 \times 10^{-4} \text{ M}$

$[\text{NaIO}_3]$
 $(1.2 \times 10^{-3})(1000) = C_2(1008.5)$
 $C_2 = 1.18 \times 10^{-3} \text{ M}$
 $\text{NaIO}_3 \rightleftharpoons \text{Na}^+ + \text{IO}_3^-$
 $1.8 \times 10^{-3} \text{ M}$

Trial $K_{sp} = [\text{Pb}^{2+}][\text{IO}_3^-]^2$
 $= (5.31 \times 10^{-4})(1.8 \times 10^{-3})^2$
 $= 7.6 \times 10^{-10}$

Actual $K_{sp} = 3.7 \times 10^{-13}$

Trial > Actual
 \therefore ppt will form

(3) Will a precipitate form when 1.5 mL of 4.5×10^{-3} M ammonium bromate is added to 120.5 mL of 2.5×10^{-3} M silver nitrate?



$[\text{AgNO}_3]$
 $(2.5 \times 10^{-3})(120.5) = C_2(122.0)$
 $C_2 = 2.47 \times 10^{-3} \text{ M}$
 $\text{AgNO}_3 \rightleftharpoons \text{Ag}^+ + \text{NO}_3^-$
 $2.47 \times 10^{-3} \text{ M}$

$[\text{NH}_4\text{BrO}_3]$
 $(4.5 \times 10^{-3})(1.5) = C_2(122.0)$
 $C_2 = 5.53 \times 10^{-5} \text{ M}$
 $\text{NH}_4\text{BrO}_3 \rightleftharpoons \text{NH}_4^+ + \text{BrO}_3^-$
 $5.53 \times 10^{-5} \text{ M}$

Trial $K_{sp} = [\text{Ag}^+][\text{BrO}_3^-]$
 $= (2.47 \times 10^{-3})(5.53 \times 10^{-5})$
 $= 1.37 \times 10^{-7}$

Actual $K_{sp} = 5.3 \times 10^{-5}$

Trial < Actual
 \therefore ppt will not form