Chemistry 12

Solubility Equilibrium III

Name:
Date:
Block:

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

- Both ions come from the **same salt** (source)

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

s s 2s (1:1, 1:2)

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

$(s)(2s)^2 = 4s^3$

- Ion concentrations are related through mole ratio
  - 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 = $(2s)^2(3s)^3 = 108s^5$

Two Source

- Both ions come from a **different** salt (source)

$Pb(NO_3)_{2(s)} \leftrightharpoons \mathbf{Pb^{2+}_{(aq)}} + 2\ NO_3^-{}_{(aq)}$ (x)

$KI_{(s)} \leftrightharpoons K^+_{(aq)} + \mathbf{I^-_{(aq)}}$ (y)

$PbI_{2(s)} \leftrightharpoons Pb^{2+}_{(aq)} + 2\ I^-_{(aq)}$ (x, y)

$K_{sp} = [Pb^{2+}][I^-]^2 = (x)(y)^2$

- Related through K_{sp} ← from data booklet
- Ex: Find the $[I^-]$ if $[Pb^{2+}] = 4.5 \times 10^{-3}$ M. ← (x)

$K_{sp} = 8.5 \times 10^{-9} = (x)(y)^2$

$8.5 \times 10^{-9} = (4.5 \times 10^{-3})(y^2)$

$y = \sqrt{\dfrac{8.5 \times 10^{-9}}{4.5 \times 10^{-3}}}$

$= 1.37 \times 10^{-3} M = [I^-]$

Challenging Solubility Problems

2 source: can't use S^2!

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? → At saturation (equilibrium)

$K_{sp} = 7.1 \times 10^{-5}$ salt: $CaSO_{4(s)}$

$$CaSO_{4(s)} \rightleftharpoons Ca^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

$$K_{sp} = [Ca^{2+}][SO_4^{2-}]$$

$$7.1 \times 10^{-5} = (2.5 \times 10^{-2})[SO_4^{2-}]$$

$$[SO_4^{2-}] = \boxed{2.8 \times 10^{-3} M}$$

2. Determine the maximum $[Na_2CO_3]$that can exist in 1.0L of 0.0010M $Ba(NO_3)_2$ 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

$$Na_2CO_3 + Ba(NO_3)_2 \rightleftharpoons 2\,NaNO_3 + BaCO_{3(s)}$$

$Ba(NO_3)_2 \rightleftharpoons Ba^{2+} + 2NO_3^-$ (0.0010M)

$K_{sp} = 2.6 \times 10^{-9}$

$$K_{sp} = 2.6 \times 10^{-9} = [Ba^{2+}][CO_3^{2-}]$$

$$2.6 \times 10^{-9} = (0.0010)[CO_3^{2-}]$$

$$[CO_3^{2-}] = 2.6 \times 10^{-6} M$$

$Na_2CO_3 \rightleftharpoons 2Na^+ + CO_3^{2-}$

$2.6 \times 10^{-6} M$ ← 2.6×10^{-6} (1:1)

3. What is the maximum $[Sr^{2+}]$ 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

$$Sr^{2+}{}_{(aq)} + Na_2SO_{4(aq)} \rightleftharpoons 2\,Na^+{}_{(aq)} + SrSO_{4(s)}$$

? M

$Na_2SO_4 \rightleftharpoons 2Na^+ + SO_4^{2-}$ (0.10M, 0.20M, 0.10M)

$SrSO_4 \rightleftharpoons Sr^{2+} + SO_4^{2-}$

$$K_{sp} = 3.4 \times 10^{-7} = [Sr^{2+}][SO_4^{2-}]$$

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

4. When 100.0 mL of 4.0×10^{-2} M $CaCl_2$ is added to 150.0 mL of 2.9×10^{-2} M NaOH, a precipitate just starts to form. What is the K_{sp} of this precipitate?

- Write a (balanced) double replacement reaction.

$CaCl_{2(aq)} + NaOH_{(aq)} \rightleftharpoons Ca(OH)_{2(s)} + 2NaCl_{(aq)}$

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

$Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^-$

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

$C_1V_1 = C_2V_2$

- Calculate the diluted concentrations of each ion.

$[CaCl_2]$

$C_1V_1 = C_2V_2$

$(4.0\times10^{-2})(100.0) = (C_2)(250.0)$

$C_2 = 1.6\times10^{-2}$ M $= [CaCl_2]$

$CaCl_2 \rightleftharpoons Ca^{2+} + 2Cl^-$

1.6×10^{-2} ; Ca^{2+}: 1.6×10^{-2} M

$[NaOH]$

$C_1V_1 = C_2V_2$

$(2.9\times10^{-2})(150) = (C_2)(250)$

$C_2 = 1.7\times10^{-2}$ M $= [NaOH]$

$NaOH \rightleftharpoons Na^+ + OH^-$

1.7×10^{-2} ; OH^-: 1.7×10^{-2}

- Calculate the K_{sp} value.

$K_{sp} = (1.6\times10^{-2})(1.7\times10^{-2})^2$

$= 4.6\times10^{-6}$

5. Up to 15.0g of $BaCl_2$ can be dissolved in 2.5L of $Al_2(SO_4)_3$ without a precipitate being formed. Find $[Al_2(SO_4)_3]$.

solute (solid) — solvent (liquid)

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

$BaCl_{2(s)} + Al_2(SO_4)_{3(aq)} \rightleftharpoons 3BaSO_{4(s)} + 2AlCl_{3(aq)}$

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

$K_{sp} = [Ba^{2+}][SO_4^{2-}] = 1.1\times10^{-10}$

- Looking at the K_{sp} expression, is there an ion concentration value that could be determined? $([Al_2(SO_4)_3])$

$[BaCl_2] = \frac{15.0g}{2.5L} \times \frac{1 mol}{208.3g} = 0.029M$

$BaCl_2 \rightleftharpoons Ba^{2+} + 2Cl^-$

$0.029M$

$(0.029)[SO_4^{2-}] = 1.1\times10^{-10}$

$[SO_4^{2-}] = 3.8\times10^{-9}$ M

$Al_2(SO_4)_3 \rightleftharpoons 2Al^{3+} + 3SO_4^{2-}$

1.3×10^{-9} M ; 3.8×10^{-9}

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 Ksp 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.

$$X_2Y_{(s)} \leftrightharpoons 2\,X^{+} + Y^{2-}$$

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

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)

$$Na_2CO_{3\,(aq)} + MgCl_2 \rightleftharpoons 2NaCl_{(aq)} + MgCO_{3\,(s)}$$

- What are the concentrations of each of these ions?

$[MgCl_2]$

$(0.010)(12) = C_2(35)$

$C_2 = 0.0034\text{ M} = [MgCl_2]$

$MgCl_2 \rightleftharpoons Mg^{2+} + 2Cl^-$

Mg^{2+}: 0.0034 M

$[Na_2CO_3]$

$(0.020)(23) = C_2(35)$

$C_2 = 0.013\text{ M} = [Na_2CO_3]$

$Na_2CO_3 \rightleftharpoons 2Na^+ + CO_3^{2-}$

CO_3^{2-}: 0.013 M

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

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

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

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

Trial > Actual

∴ Yes, a ppt will form!

1000 mL

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

$$Pb(NO_3)_{2(aq)} + NaIO_{3(aq)} \rightleftharpoons Pb(IO_3)_{2(s)} + NaNO_{3(aq)}$$

$[Pb(NO_3)_2]$

$(6.3\times10^{-2})(8.5) = C_2(1008.5)$

$C_2 = 5.31\times10^{-4}M$

$Pb(NO_3)_2 \rightleftharpoons Pb^{2+} + 2NO_3^-$

$5.31\times10^{-4}M$

$[NaIO_3]$

$(1.2\times10^{-3})(1000) = C_2(1008.5)$

$C_2 = 1.18\times10^{-3}M$

$NaIO_3 \rightleftharpoons Na^+ + IO_3^-$

$1.8\times10^{-3}M$

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

$= (5.31\times10^{-4})(1.8\times10^{-3})$

$= 7.6\times10^{-10}$

Actual $K_{sp} = 3.7\times10^{-13}$

Trial > Actual

∴ ppt will form

(3) Will a precipitate form when 1.5 mL of 4.5 x 10^{-3} M ammonium bromate is added to 120.5 mL of 2.5x10^{-3} M silver nitrate?

$$NH_4BrO_{3(aq)} + AgNO_{3(aq)} \rightleftharpoons AgBrO_{3(s)} + NH_4NO_{3(aq)}$$

$[AgNO_3]$

$(2.5\times10^{-3})(120.5) = C_2(122.0)$

$C_2 = 2.47\times10^{-3}M$

$AgNO_3 \rightleftharpoons Ag^+ + NO_3^-$

$2.47\times10^{-3}M$

$[NH_4BrO_3]$

$(4.5\times10^{-3})(1.5) = C_2(122.0)$

$C_2 = 5.53\times10^{-5}M$

$NH_4BrO_3 \rightleftharpoons NH_4^+ + BrO_3^-$

$5.53\times10^{-5}M$

Trial $K_{sp} = [Ag^+][BrO_3^-]$

$= (2.47\times10^{-3})(5.53\times10^{-5})$

$= 1.37\times10^{-7}$

Actual $K_{sp} = 5.3\times10^{-5}$

Trial < Actual

∴ ppt will not form

Hebden Workbook Pg. 98 #56-69

Worksheet 3.3 #8-12