Chemistry 12

Solubility Equilibrium II

III . 5 - 6
pg. 88 - 95

Name:
Date:
Block:

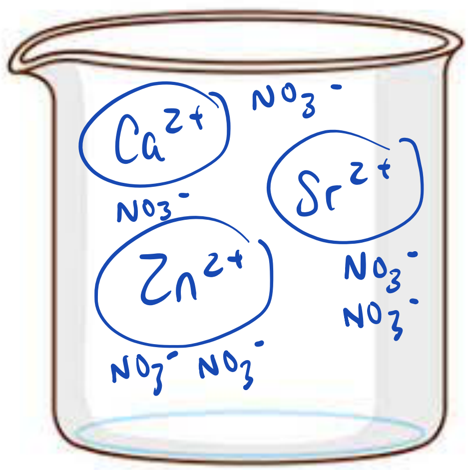
1. **Forming a Precipitate**
2. **Solubility Product Constant *(One Source of Ions)***

Forming a Precipitate

Example:

A solution may contain the ions **Ca^{2+}, Sr^{2+} and Zn^{2+}**. How would you precipitate the ions out of solution individually? Describe your answer using a flow chart.

- All are cations - therefore an addition of an anion will precipitate out these cations. (low solubility)
- There are also spectator ions (soluble) in the solution to help balance out the charge.



- What can precipitate out Ca^{2+}?

  SO_4^{2-}, OH^- (2), PO_4^{3-}, CO_3^{2-}, SO_3^{2-}

- What can precipitate out Sr^{2+}?

  (3) SO_4^{2-}, PO_4^{3-}, CO_3^{2-}, SO_3^{2-}

- What can precipitate out Zn^{2+}?

  (1) S^{2-}, OH^-, PO_4^{3-}, CO_3^{2-}, SO_3^{2-}

- **What needs to be added first?**

  S^{2-} (along with a soluble spectator ion)
  $Na_2S_{(aq)}$

As a flow chart:

Ca^{2+}, Sr^{2+}, Zn^{2+}

① Add $Na_2S_{(aq)}$ (spectator)

Ca^{2+}, Sr^{2+} | $ZnS_{(s)}$ ppt

② Add $NaOH_{(aq)}$

Sr^{2+} | $Ca(OH)_{2\,(s)}$ ppt

③ Add $Na_2SO_{4\,(aq)}$

$SrSO_{4\,(s)}$ ppt

Practice:

1. For each of the following solutions, describe a process to individually remove each ion. Be sure to list the compounds that you add in order, and the method of removing the precipitate.

   a. Ag^+ Ba^{2+} and Be^{2+}

   ① Add $NaCl_{(aq)}$

   Ba^{2+}, Be^{2+} | $AgCl_{(s)}$ ppt

   ② Add $Na_2SO_{4(aq)}$

   Be^{2+} | $BaSO_{4(s)}$ ppt

   ③ Add $NaOH_{(aq)}$

   $Be(OH)_{2(s)}$ ppt

   b. Br^-, SO_4^{2-} and S^{2-}

   Br^- : Ag^+, Pb^{2+}, Cu^+ ②

   SO_4^{2-} : Ag^+, Sr^{2+}, Ca^{2+}, Ba^{2+}, Pb^{2+} ③

   S^{2-} : (All others) Fe^{2+} ①

   ① Add $Fe(NO_3)_{2(aq)}$ ↑ soluble spectator

   Br^-, SO_4^{2-} | $FeS_{(s)}$ ppt

   ② Add $CuNO_{3(aq)}$

   SO_4^{2-} | $CuBr_{(s)}$ ppt

   ③ Add $AgNO_{3(aq)}$

   $Ag_2SO_{4(s)}$ ppt

   c. OH^-, PO_4^{3-} and S^{2-}

   ① Add $Sr(NO_3)_{2(aq)}$

   OH^-, S^{2-} | $Sr(PO_4)_{2(s)}$ ppt

   ② Add $Ba(NO_3)_{2(aq)}$

   S^{2-} | $Ba(OH)_{2(s)}$ ppt

   ③ Add $Al(NO_3)_{3(aq)}$

   $Al_2S_{3(s)}$ ppt

   d. Cr^{3+}, Ca^{2+} and Mg^{2+}

   ① Add $Na_2SO_{4(aq)}$

   Cr^{3+}, Mg^{2+} | $CaSO_{4(s)}$ ppt

   ② Add $Na_2S_{(aq)}$

   Mg^{2+} | $Cr_2S_{3(s)}$ ppt

   ③ Add $NaOH_{(aq)}$

   $Mg(OH)_{2(s)}$ ppt

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Solubility Product Constant K_{sp} *(One Source of Ions)*

In a **saturated** solution, **equilibrium** is established between the dissolving and recrystallization of a salt.

ionic compound (low solubility)

recrystallization

$$AB_{(s)} \leftrightharpoons A^{+}_{(aq)} + B^{-}_{(aq)}$$

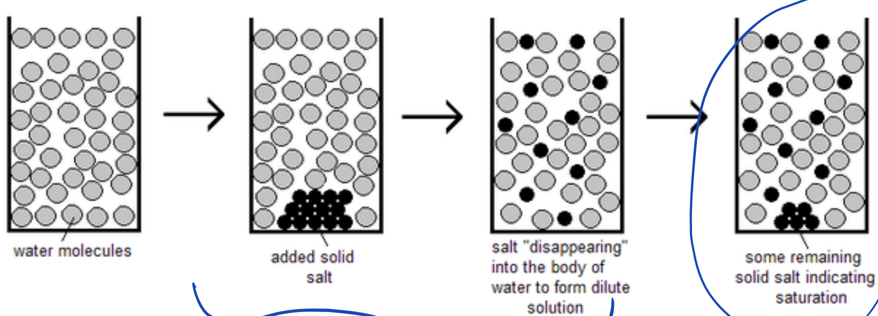
dissolving

Not included in Ksp expression

$$K_{eq} = K_{sp} = \frac{[P]}{[R]} = \frac{[A^+][B^-]}{1}$$

The solubility product constant, Ksp, is the ratio of the ion concentrations in a Saturated solution raised to the power of the coefficients in the equilibrium.

@ equilibrium



undersaturated

Point of Saturation = when a ppt just starts to form

Why aren't we using ICE TABLES?

Let's use an example with some mole ratios.

	$CD_{2\,(s)}$	$\leftrightharpoons$	$C^{2+}_{(aq)}$	+	$2D^{-}_{(aq)}$
Initial *(Where the stress is introduced)*					
Change *(How the system responds to the stress)*					
Equilibrium *(New equil'm concentrations)*					

Lab Simulation: Salts & Solubility

Day 1

HW

Ionization Equation Extra Practice:

- Represents the salt breaking apart into **ions**
  - $NaCl \rightarrow$ 1 Na^+ + 1 Cl^-
- If the salt were $CaCl_2$
  - $CaCl_2 \rightarrow$ 1 Ca^{2+} + 2 Cl^-

- Mole ratios represent the relative amounts of ions in solution

$NaCl_{(aq)}$	$\rightarrow$	$Na^+_{(aq)}$	+	$Cl^-_{(aq)}$
1.0M		1.0M		1.0M
$CaCl_{2\,(aq)}$	$\rightarrow$	$Ca^{2+}_{(aq)}$	+	2 $Cl^-_{(aq)}$
1.0M		1.0M		2.0M

x2

Practice:

1. 0.25M KOH

   $KOH \rightarrow K^+ + OH^-$

   0.25M 0.25M

2. 0.75M $SrCl_2$

   $SrCl_2 \rightarrow Sr^{2+} + 2Cl^-$

   0.75M 1.5M

3. 0.20M NH_4NO_3

   $NH_4NO_3 \rightarrow NH_4^+ + NO_3^-$

   0.20M 0.20M

4. 0.45M Ag_2CrO_4

   $Ag_2CrO_4 \rightarrow 2Ag^+ + CrO_4^{2-}$

   0.90M 0.45M

5. 0.60M $Pb(IO_3)_2$

   $Pb(IO_3)_2 \rightarrow Pb^{2+} + 2IO_3^-$

   0.60M 1.2M

6. 0.011M $Mg(OH)_2$

   $Mg(OH)_2 \rightarrow Mg^{2+} + 2OH^-$

   0.011M 0.022M

7. 0.45M $BaCO_3$

   $BaCO_3 \rightarrow Ba^{2+} + CO_3^{2-}$

   0.45M 0.45M

8. 0.50M $(NH_4)_2SO_3$

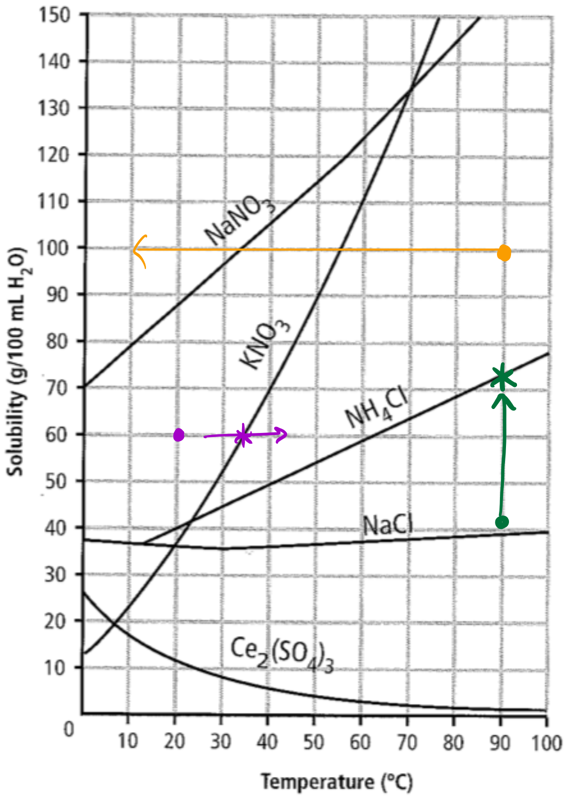
   $(NH_4)_2SO_3 \rightarrow 2NH_4^+ + SO_3^{2-}$

   1.0M 0.50M

Above line = solid (oversaturated)
At line = saturation point (equilibrium)
Below line = aqueous ions (undersaturated)

Solubility Curves

Consider the graph below:



a) At 10°C, which salt has the highest solubility?

Can dissolve the highest conc of ions

$NaNO_3$

b) At 10°C, which salt has the lowest solubility?

$Ce_2(SO_4)_3$

c) At 90°C, which salt has the highest solubility?

KNO_3

d) At 90°C, which salt has the lowest solubility?

$Ce_2(SO_4)_3$

e) If you put 40 g of NH_4Cl in 100 mL of water at 90°C, will you be able to form a saturated solution?

No (it's undersaturated, will stay as aqueous ions)

f) Approximately how many more grams of NH_4Cl could you add until it is saturated?

~35g (40g + 35g = 75g)

g) If you put 60 g of KNO_3 into 100 mL of water at 20°C and gradually heat the solution, what will you observe?

Solid KNO_3 will start to dissolve
(fully dissolved at 35°C)

h) If you dissolve 100 g of both $NaNO_3$ and KNO_3 in 100 mL of water at 90°C and then cool the mixture to 10°C, which salt will form crystals (ppt) first?

KNO_3

i) $Ce_2(SO_4)_3$ is an unusual substance as it does not follow the usual trend. What is unusual about $Ce_2(SO_4)_3$?

Solubility ↓ as temp ↑

The molar solubility of a substance is the molar concentration of a solute in a saturated solution.

MOLE RATIO WILL BE VERY IMPORTANT IN THIS UNIT!!

We need to write out the IONIZATION/DISSOCIATION equation to figure out the ratio.

Solubility = "s" = the [ ] of ions in a saturated solution

$BaCO_{3\,(s)} \leftrightharpoons \underset{s}{\underline{1}}\,Ba^{2+}_{(aq)} + \underset{s}{\underline{1}}\,CO_3^{2-}{}_{(aq)}$ *Ratio of ions:* 1 : 1 from table $K_{sp} = [Ba^{2+}][CO_3^{2-}]$ $= (s)(s) = s^2 = 2.6 \times 10^{-9}$ *Solubility* = $\sqrt{2.6 \times 10^{-9}}$ [ ] $= 5.1 \times 10^{-5}\,M = [Ba^{2+}]$ $= [CO_3^{2-}]$	$Fe(OH)_{2\,(s)} \leftrightharpoons \underset{s}{\underline{1}}\,Fe^{2+}_{(aq)} + \underset{2s}{\underline{2}}\,OH^-_{(aq)}$ *Ratio of ions:* 1 : 2 $K_{sp} = [Fe^{2+}][OH^-]^2$ $= (s)(2s)^2 = (s)(2s)(2s)$ $= 4s^3 = 4.9 \times 10^{-17}$ *Solubility* = $\sqrt[3]{\frac{4.9 \times 10^{-17}}{4}} = 2.3 \times 10^{-6}\,M = s$ $[Fe^{2+}] = 2.3 \times 10^{-6}\,M$ $[OH^-] = 2s = 4.6 \times 10^{-6}\,M$
$Fe(OH)_{3\,(s)} \leftrightharpoons \underset{s}{\underline{1}}\,Fe^{3+}_{(aq)} + \underset{3s}{\underline{3}}\,OH^-_{(aq)}$ *Ratio of ions:* 1 : 3 $K_{sp} = [Fe^{3+}][OH^-]^3$ $= (s)(3s)^3 = (s)(3s)(3s)(3s)$ $= 27s^4 = 2.6 \times 10^{-39}$ *Solubility* = $\sqrt[4]{\frac{2.6 \times 10^{-39}}{27}} = 9.9 \times 10^{-11}\,M$	$Sr_3(PO_4)_{2\,(s)} \leftrightharpoons \underset{3s}{\underline{3}}\,Sr^{2+}_{(aq)} + \underset{2s}{\underline{2}}\,PO_4^{3-}{}_{(aq)}$ *Ratio of ions:* 3 : 2 $K_{sp} = 1.0 \times 10^{-31} = [Sr^{2+}]^3[PO_4^{3-}]^2$ $= (3s)^3(2s)^2 = 108s^5$ *Solubility* = $\sqrt[5]{\frac{1.0 \times 10^{-31}}{108}} = 2.5 \times 10^{-7}\,M$

Solubility (M) → K_{sp}

(1) The molar solubility of $CaSO_4$ is 8.4×10^{-3} M at a particular temperature. Calculate its K_{sp}.

- What is the equation representing the equilibrium?

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

- Write the K_{sp} expression and substitute the concentration of ions into the K_{sp} expression:

$$K_{sp} = [Ca^{2+}][SO_4^{2-}] = (s)(s) = (8.4 \times 10^{-3})^2 = \boxed{7.1 \times 10^{-5}}$$

*Ksp has no units!

(2) The solubility of lead (II) chloride is 4.4 g/L. Calculate its K_{sp}.

$$[PbCl_2] = \frac{4.4\ \cancel{g}}{1\ L} \times \frac{1\ mol}{278.2\ \cancel{g}} = 0.016M = s$$

$$PbCl_{2(s)} \rightleftharpoons \underset{s}{Pb^{2+}_{(aq)}} + \underset{2s}{2Cl^-_{(aq)}}$$

$$K_{sp} = [Pb^{2+}][Cl^-]^2 = (s)(2s)^2 = 4s^3 = 4(0.016)^3 = \boxed{1.6 \times 10^{-5}}$$

(3) A student prepares a saturated solution by dissolving 5.5×10^{-5} mol of $Al(OH)_3$ in 500. mL of solution. Calculate the K_{sp} of $Al(OH)_3$.

$$[Al(OH)_3] = \frac{5.5 \times 10^{-5}\ mol}{0.5000\ L} = 1.1 \times 10^{-4} M = s$$

$$Al(OH)_{3(s)} \rightleftharpoons \underset{s}{Al^{3+}_{(aq)}} + \underset{3s}{3OH^-_{(aq)}}$$

$$K_{sp} = 27s^4 = 27(1.1 \times 10^{-4})^4 = \boxed{4.0 \times 10^{-15}}$$

(4) A student evaporated 150. mL of a saturated solution of magnesium phosphate. If 0.16g of solute remains, calculate the K_{sp}.

$$Mg_3(PO_4)_{2(s)} \rightleftharpoons \underset{3s}{3Mg^{2+}_{(aq)}} + \underset{2s}{2PO_4^{3-}{}_{(aq)}}$$

$$[Mg_3(PO_4)_2] = \frac{0.16g}{0.150\ L} \times \frac{1\ mol}{262.9g} = 0.00406M = s$$

$$K_{sp} = [Mg^{2+}]^3[PO_4^{3-}]^2 = (3s)^3(2s)^2 = 108s^5 = 108(0.00406)^5 = \boxed{1.2 \times 10^{-10}}$$

(5) Calculate the K_{sp} of silver oxalate if the solubility is 0.033 g/L.

$$Ag_2C_2O_{4(s)} \rightleftharpoons \underset{2s}{2Ag^+_{(aq)}} + \underset{s}{C_2O_4^{2-}{}_{(aq)}}$$

$$[Ag_2C_2O_4] = \frac{0.033g}{1\ L} \times \frac{1\ mol}{303.8g} = 1.1 \times 10^{-4} M = s$$

$$K_{sp} = [Ag^+]^2[C_2O_4^{2-}] = (2s)^2(s) = 4s^3 = 4(1.1 \times 10^{-4})^3 = \boxed{5.3 \times 10^{-12}}$$

(6) A compound has a solubility of 7.1×10^{-5} M at 25°C. According to its K_{sp}, the compound is:

A. CuS B. AgBr C. $CaCO_3$ D. $CaSO_4$

All 1:1 ratio

$$K_{sp} = s^2$$

$$= (7.1 \times 10^{-5})^2 = 5.0 \times 10^{-9}$$

↳ matches Ksp of $CaCO_3$

K_{sp} → Solubility (M)

$K_{sp} = 4.9 \times 10^{-17}$

(1) Calculate the molar solubility of iron (II) hydroxide from its K_{sp}.

$Fe(OH)_{2\,(s)} \rightleftharpoons Fe^{2+}_{(aq)} + 2OH^-_{(aq)}$

s, s, 2s

$K_{sp} = [Fe^{2+}][OH^-]^2 = (s)(2s)^2 = 4s^3$

$4.9 \times 10^{-17} = 4s^3$

$s = \sqrt[3]{\frac{4.9 \times 10^{-17}}{4}}$

$= \boxed{2.3 \times 10^{-6} M = [Fe(OH)_2]}$

(2) Calculate the molar solubility of iron (III) hydroxide from its K_{sp}.

$K_{sp} = 2.6 \times 10^{-39}$

$Fe(OH)_{3\,(s)} \rightleftharpoons Fe^{3+}_{(aq)} + 3OH^-_{(aq)}$

s, s, 3s

$K_{sp} = 27s^4$

$2.6 \times 10^{-39} = 27s^4$

$s = \sqrt[4]{\frac{2.6 \times 10^{-39}}{27}}$

$= \boxed{9.9 \times 10^{-11} M = [Fe(OH)_3]}$

(3) Which of the following substances has the lowest solubility? — smallest s

A. ~~BaS~~   B. CuS   C. FeS   D. ZnS

S, LS, LS, LS

K_{sp} = 6.0×10^{-37}, 6.0×10^{-19}, 2.0×10^{-25}

All 1:1 ratio

$K_{sp} = s^2$

(4) How many moles of dissolved solute are present in 100.0mL of a saturated $SrCO_3$ solution?

A. 5.6×10^{-11} mol   B. 2.4×10^{-6} mol   C. 2.4×10^{-5} mol   D. 2.4×10^{-4} mol

$K_{sp} = 5.6 \times 10^{-10}$

$SrCO_{3\,(s)} \rightleftharpoons Sr^{2+}_{(aq)} + CO_3^{2-}_{(aq)}$

s, s

$\sqrt{K_{sp}} = \sqrt{s^2} = \sqrt{5.6 \times 10^{-10}}$

$s = 2.4 \times 10^{-5} M$

$0.1000 \cancel{L} \times \frac{2.4 \times 10^{-5} mol}{1 \cancel{L}}$

$= \boxed{2.4 \times 10^{-6} mol}$

__Worksheet 3.2__

(__Hebden Workbook Pg. 95 #42 – 55__) extra practice

$g \cdot L^{-1} = \frac{g}{L}$