Chemistry 12

Unit III - Solubility

KEY

I) Ionic vs. Molecular Solutions

Solution: a HOMOGENEOUS mixture (consisting of one phase) of two or more substances

e.g. salt and water, metal alloys, alcohol and water

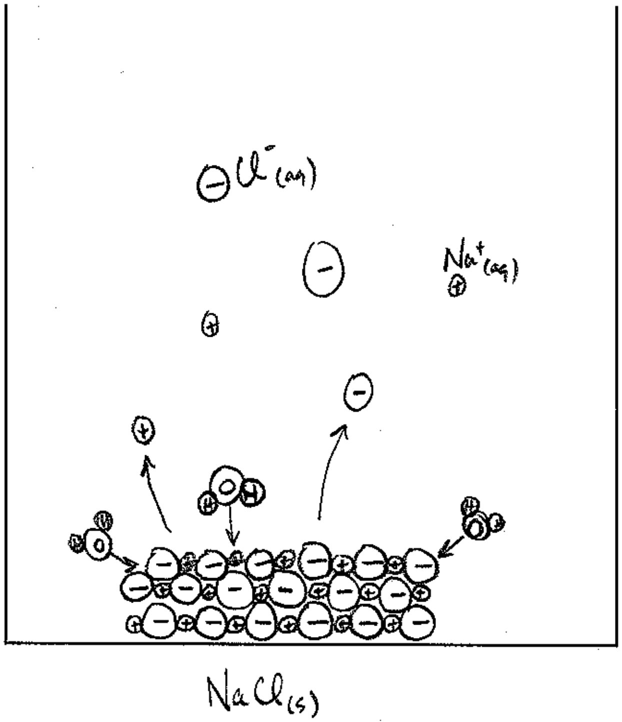
Solute: the component(s) of a solution in LESSER quantity

Solvent: the component of a solution in the GREATEST quantity

Salts (IONIC compounds) are made up of a METAL and a NONMETAL. Salts dissociate into IONS when dissolving into solution. Ions in solution are **electrolytes**. Because of the electrolytes, ionic solutions will conduct electricity better than pure water.

Dissociation Equation: $NaCl_{(s)} \Rightarrow Na^{+}_{(aq)} + Cl^{-}_{(aq)}$

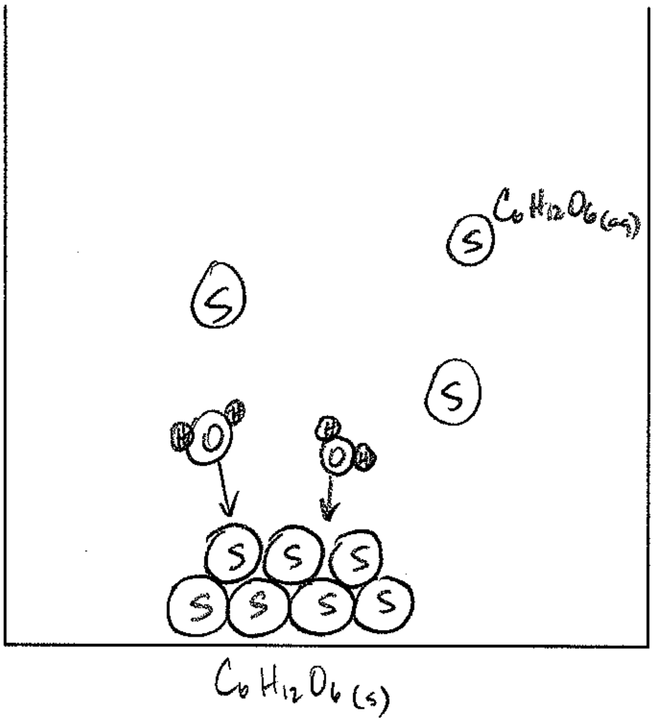
Diagram of $NaCl_{(s)}$ dissociating in water:



Water molecules collide w/ crystal, knocking ions into solution.

(Covalent)
Molecular Compounds are made up of NON METALS. They remain as a whole molecule when dissolving into solution. These neutral, dissolved molecules are 'non-electrolytes' because charged ions are not created. Certain molecular compounds **dissolve** in water, but no molecular compounds **dissociate** (split into ions) in water. Molecular solutions do not conduct electricity better than pure water.

Sugar (glucose) dissolving in water: $C_6H_{12}O_{6(s)} \Rightarrow C_6H_{12}O_{6(aq)}$
Diagram of sugar dissolving in water:



- Sugars possess many POLAR –OH groups which form H-bonds w/ water molecules to form a solution.

Water molecules collide with sugar (glucose) molecules, knocking them into solution.

II) Solubility

Saturated Solution: a solution in which excess solute remains undissolved (in solid state) on the bottom of the container, and the rate of dissolving (dissociation) EQUALS the rate of crystallization.

Saturated solutions are at **equilibrium** because the rate of dissolving equals the rate of re-crystallization (precipitation). In a saturated solution, the amount of undissolved solute on the bottom remains constant because an equal amount of salt is dissolving and re-crystallizing.

e.g. $KBr_{(s)} \Leftrightarrow K^+_{(aq)} + Br^-_{(aq)}$

double arrow

'Perfectly' Saturated Solution – no solid is observed on bottom, but addition of any more solid would **not** dissociate (dissolve).

– see Q6 in textbook, 'Gray area'

For a solution that is unsaturated, the rate of dissolving is greater than the rate of re-crystallization. Therefore, the solid crystals on the bottom eventually dissolve, and the solution is not at equilibrium.

e.g. $KBr_{(s)} \Rightarrow K^+_{(aq)} + Br^-_{(aq)}$

single arrow

Solubility: the solubility of a substance is the equilibrium concentration of that substance in solution at a given temperature. Concentration is measured using Molarity (M) = mol solute/L solution.

Every salt dissociates to some extent in water, but every salt has a different solubility in water. Some salts can dissociate a very large extent, while others barely dissociate at all. Solubility is also temperature dependent. A general rule for solid-liquid solutions is that **solubility increases with increasing temperature**.

High Solubility Salts can produce a solution greater than 0.1M before becoming saturated.

Low Solubility Salts create solutions that become saturated at ≤0.1M.

The solubility table in the data booklet is a tool utilized to predict if a salt is soluble or low solubility.

Assignment 1: Read pages 73-76 in Hebden and do Questions #1-7

III) Calculating Solubility and Ion Concentrations

Use stoichiometry to calculate the concentration of ions given the concentration of solute.

Eg. 1 Calculate $[Na^+]$ and $[SO_4^{2-}]$ in a 0.20M solution of $Na_2SO_{4(aq)}$

Dissoc. eq'n: $Na_2SO_{4(s)} \Rightarrow 2Na^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

$$\begin{array}{ccc} 0.20M & \underline{0.40M} & \underline{0.20M} \end{array}$$

Eg. 2 Calculate the $[Fe^{3+}]$ in a 0.50M solution of iron III sulfate.

$$Fe_2(SO_4)_{3(s)} \longrightarrow 2Fe^{3+}_{(aq)} + 3SO_4^{2-}{}_{(aq)}$$

$$\begin{array}{ccc} 0.50\text{ M} & \boxed{1.0\text{ M}} & 1.5\text{ M} \end{array}$$

Eg. 3 A saturated solution of $PbCl_2$ is found to contain 9.90g of $PbCl_2$ per liter of solution. Find $[Pb^{2+}]$. *change grams to moles and then write dissoc. eq'n and use stoich.

"saturated"

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

$$\begin{array}{ccc} 3.5586 \times 10^{-2}\text{ M} & 3.5586 \times 10^{-2}\text{ M} & 7.12 \times 10^{-2}\text{ M} \end{array}$$

$$\frac{9.90\text{ g } PbCl_2}{1.00\text{ L}} \times \frac{1\text{ mol } PbCl_2}{278.2\text{ g } PbCl_2} = 3.5586 \times 10^{-2}\ \frac{\text{mol}}{\text{L}} = 3.5586 \times 10^{-2}\text{ M}$$

$$\boxed{[Pb^{2+}] = 3.56 \times 10^{-2}\text{ M}}$$

Eg. 4 Write an equation showing the equilibrium in a saturated solution of silver carbonate. The solubility of Ag_2CO_3 is 1.29×10^{-4} M. What is the solubility in grams per 100.0 mL?

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

1.29×10^{-4} M

$$\frac{1.29 \times 10^{-4}\text{ mol } Ag_2CO_3}{1.00\text{ L}} \times \frac{275.8\text{ g } Ag_2CO_3}{1\text{ mol } Ag_2CO_3} \times \frac{1.00\text{ L}}{10\ (100\text{ mL})} = \boxed{3.56 \times 10^{-3}\ \frac{\text{g}}{100\text{ mL}}}$$

Eg. 5 250mL of 0.30M K_2SO_4 and 250mL of 0.80M $MgCl_2$ are mixed and no reaction occurs. Calculate the concentration of each ion in the final solution. *use dilution eq'n and then write dissoc. eq'ns and use stoich.

$$\underset{0.15\,M}{K_2SO_{4(s)}} \longrightarrow \underset{0.30\,M}{2K^+_{(aq)}} + \underset{0.15\,M}{SO_4{}^{2-}{}_{(aq)}} \qquad \underset{0.40\,M}{MgCl_{2(s)}} \longrightarrow \underset{0.40\,M}{Mg^{2+}_{(aq)}} + \underset{0.80\,M}{2Cl^-_{(aq)}}$$

$$[K_2SO_4]_f = \frac{M_iV_i}{V_f} = \frac{(0.30\,M)(0.25\,L)}{0.50\,L} = 0.15\,M \qquad [MgCl_2]_f = \frac{M_iV_i}{V_f} = \frac{(0.80\,M)(0.25\,L)}{0.50\,L} = 0.40\,M$$

$[K^+] = 0.30\,M$ $\quad [Mg^{2+}] = 0.40\,M$

$[SO_4{}^{2-}] = 0.15\,M$ $\quad [Cl^-] = 0.80\,M$

Eg. 6 Calculate the concentration of all ions when 200.0 mL of 0.600M $AlBr_3$ and 300.0 mL of 0.400M $BaBr_2$ are mixed together.

$$\underset{0.240\,M}{AlBr_{3(s)}} \longrightarrow \underset{0.240\,M}{Al^{3+}_{(aq)}} + \underset{0.720\,M}{3Br^-_{(aq)}} \qquad \underset{0.240\,M}{BaBr_{2(s)}} \longrightarrow \underset{0.240\,M}{Ba^{2+}_{(aq)}} + \underset{0.480\,M}{2Br^-_{(aq)}}$$

$$[AlBr_3]_f = \frac{M_iV_i}{V_f} = \frac{(0.600\,M)(0.2000\,L)}{0.5000\,L} = 0.240\,M \qquad [BaBr_2]_f = \frac{M_iV_i}{V_f} = \frac{(0.400\,M)(0.3000\,L)}{0.5000\,L} = 0.240\,M$$

$[Al^{3+}] = 0.240\,M$

$[Ba^{2+}] = 0.240\,M$

$[Br^-] = 0.720 + 0.480 = 1.200\,M$

Assignment 2: Hebden p. 77 # 8-11 / p. 78 #12-14 / p. 81 #18, 19, 20abfg (*try the rest of 20 for practice only)

IV) Using the Solubility Table

The solubility table in the data booklet is used to predict whether an ionic salt is **soluble** in water (>0.1M) or **low solubility** in water (≤0.1M). Any compound that is determined to be "low solubility," is assumed to form a precipitate.

Precipitate: a SOLID that forms in a solution due to the mixing of two ions that are low solubility when in combination.

Hints:

1. All nitrate (NO_3^-) compounds are soluble.
2. All alkali metal (first column) compounds are soluble.
3. All ammonium (NH_4^+) compounds are soluble.

Use your table to predict whether the following salts are soluble, or low solubility and form a precipitate (ppt.) in water.

1. Sodium hydroxide S
2. Ammonium acetate S
3. Calcium sulphate LS
4. Lead II chloride LS
5. Potassium chloride S
6. Calcium bromide S
7. Potassium carbonate S
8. Aluminum sulphate S
9. Lead (II) sulphide LS
10. Copper II sulphate S
11. Iron II sulphide LS
12. Barium hydroxide LS
13. Silver bromide LS
14. Magnesium carbonate LS
15. Copper II chloride S
16. Copper I chloride LS

Other types of questions you must be able to answer using the solubility table:

i) Be able to predict if a precipitate will form when two aqueous solutions are mixed:

$$Pb(NO_3)_{2(aq)} + 2KI_{(aq)} \Rightarrow 2KNO_{3(aq)} + PbI_{2(s)}$$

*Find possible products

* Is each product soluble (aq) or will it form a ppt (s)?

ii) Be able to determine which aqueous solutions would have to be mixed in order to produce a certain precipitate:
Which aqueous solutions could be mixed in order to produce a precipitate of $Mg(OH)_{2(s)}$? $Mg(NO_3)_2$ + $NaOH$ (among others)

Assignment 3: Hebden p. 83 #21-23 and p. 84 #24

V) Formula, Complete, and Net Ionic Equations

Eg. 1

Formula Equation:

$Na_2CO_{3(aq)} + CaCl_{2(aq)} \Rightarrow 2NaCl_{(aq)} + CaCO_{3(s)}$

Notice that no ions are written in a formula equation.

Complete Ionic Equation:

Spectator ions

$2Na^+_{(aq)} + CO_3^{2-}{}_{(aq)} + Ca^{2+}{}_{(aq)} + 2Cl^-_{(aq)} \Rightarrow 2Na^+_{(aq)} + 2Cl^-_{(aq)} + CaCO_{3(s)}$

Notice that only ions are written in a complete ionic equation, except for precipitates.
Build and balance the complete ionic equation using the formula equation.

Net Ionic Equation:

$Ca^{2+}{}_{(aq)} + CO_3^{2-}{}_{(aq)} \Rightarrow CaCO_{3\,(s)}$

Notice that only the ions that become a part of the precipitate are written as reactants (the 'crossed-out' ions are known as *spectator ions*).
Build and balance the net ionic equation using the complete ionic equation.

Eg. 2 Write the Formula, Complete Ionic, and Net Ionic Equations when the following aqueous compounds are mixed:
$LiBr_{(aq)}$ and $CaS_{(aq)}$

MOLECULAR: $2LiBr_{(aq)} + CaS_{(aq)} \longrightarrow Li_2S_{(aq)} + CaBr_{2(aq)}$

COMPLETE IONIC: $2Li^+_{(aq)} + 2Br^-_{(aq)} + Ca^{2+}_{(aq)} + S^{2-}_{(aq)} \longrightarrow 2Li^+_{(aq)} + S^{2-}_{(aq)} + Ca^{2+}_{(aq)} + 2Br^-_{(aq)}$

NET IONIC: NONE! * all ions cross-out.

Assignment 4: For each of the following, write formula, complete, and net ionic equations:

1. Aqueous magnesium chloride is mixed with aqueous potassium hydroxide.
2. Aqueous lithium carbonate is mixed with aqueous iron (II) sulphate.
3. Aqueous copper I sulphate is mixed with aqueous strontium bromide.

Assignment 5: Hebden p. 87 #25

-QUIZ 1-

VI) Separating Mixtures of Ions by Precipitation

Describe a method to separate Ba^{2+} and Pb^{2+} ions which are in solution together.

1. Find an anion that will form a precipitate (low solubility compound) with **only one** of the cations, and add it to the solution in the form of a sodium salt.

ADD Cl^- as NaCl to ppt. $PbCl_{2(s)}$

2. Filter off the precipitate.
3. Find an anion that will precipitate out the Ba^{2+} ion.

add SO_4^{2-} as Na_2SO_4 to ppt. $BaSO_{4(s)}$

Assignment 4

1. F: $MgCl_{2(aq)} + 2KOH_{(aq)} \longrightarrow Mg(OH)_{2(s)} + 2KCl_{(aq)}$

   C: $Mg^{2+}_{(aq)} + 2Cl^{-}_{(aq)} + 2K^{+}_{(aq)} + 2OH^{-}_{(aq)} \longrightarrow 2K^{+}_{(aq)} + 2Cl^{-}_{(aq)} + Mg(OH)_{2\underline{\underline{(s)}}}$

   N: $Mg^{2+}_{(aq)} + 2OH^{-}_{(aq)} \longrightarrow Mg(OH)_{2(s)}$

2. F: $Li_2CO_{3(aq)} + FeSO_{4(aq)} \longrightarrow Li_2SO_{4(aq)} + FeCO_{3(s)}$

   C: $2Li^{+}_{(aq)} + CO_{3\,(aq)}^{2-} + Fe^{2+}_{(aq)} + SO_{4\,(aq)}^{2-} \longrightarrow 2Li^{+}_{(aq)} + SO_{4\,(aq)}^{2-} + FeCO_{3\underline{\underline{(s)}}}$

   N: $Fe^{2+}_{(aq)} + CO_{3\,(aq)}^{2-} \longrightarrow FeCO_{3(s)}$

3. F: $Cu_2SO_{4(aq)} + SrBr_{2(aq)} \longrightarrow SrSO_{4(s)} + 2CuBr_{(s)}$

   C: $2Cu^{+}_{(aq)} + SO_{4\,(aq)}^{2-} + Sr^{2+}_{(aq)} + 2Br^{-}_{(aq)} \longrightarrow SrSO_{4(s)} + 2CuBr_{(s)}$

   N: Same as Complete Ionic.

Assignment 6: Ion Separation Exercises

1. Describe a method to separate Cl^- and OH^- that are in solution together. Hint: Start by picking a cation that will only precipitate **one** of the anions out. Add the cation as a nitrate salt.

① Add Mg^{2+} (as $Mg(NO_3)_2$) to ppt OH^- as $Mg(OH)_{2(s)}$. FILTER.
② Add Pb^{2+} (as $Pb(NO_3)_2$) to ppt Cl^- as $PbCl_{2(s)}$.

2. A solution is known to contain one or more of the following ions:
Mg^{2+}, Ca^{2+}, Pb^{2+}
What ions could be added and in what order to determine which cations are present?

Add S^{2-} (as Na_2S) to ppt Pb^{2+} as $PbS_{(s)}$.
Add SO_4^{2-} (as Na_2SO_4) to ppt Ca^{2+} as $CaSO_{4(s)}$.
Add OH^- (as $NaOH$) to ppt Mg^{2+} as $Mg(OH)_{2(s)}$

3. You have a solution known to contain any or all of Cu^+, Ca^{2+}, Fe^{3+}, and Sr^{2+}. You have the following 'test' solutions available: 1M CO_3^{2-}, 1M SO_4^{2-}, 1M OH^-, 1M Cl^-, 1M S^{2-} In what order would you add each of the above test solutions to test for the presence of each cation?

Cation / Anion	Cu^+	Ca^{2+}	Fe^{3+}	Sr^{2+}
CO_3^{2-}	PPT	PPT	PPT	PPT.
SO_4^{2-}	SOLUBLE	PPT	SOLUBLE	PPT.
OH^-	PPT.	PPT.	PPT	SOLUBLE
Cl^-	PPT	SOLUBLE	SOLUBLE	SOLUBLE
S^{2-}	PPT	SOLUBLE	PPT	SOLUBLE

① Cl^-
② S^{2-}
③ OH^-
④ either CO_3^{2-}/SO_4^{2-}

Assignment 7: Hebden p. 90 #26-32, 36

VII) Quantitative Solubility

Remember that a saturated solution is in equilibrium because the rate of DISSOCIATION equals the rate of RE-CRYSTALLIZATION.

$$BaSO_{4(s)} \Leftrightarrow Ba^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$$

double arrow.

Because an equilibrium exists, there exists an equilibrium constant. We use K_{sp} which stands for the SOLUBILITY PRODUCT CONSTANT.

IMPORTANT NOTE: The solubility constant (K_{sp}) is NOT the same as solubility.

$$K_{sp} = \frac{[PRODUCTS]}{[REACTANTS]}$$

Remember: SOLIDS and LIQUIDS not included in K expressions.

Write the K_{sp} expressions for saturated solutions of $Ca(OH)_2$, $PbCl_2$, and Ag_3PO_4

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

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

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

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

$$Ag_3PO_{4(s)} \rightleftharpoons 3Ag^+{}_{(aq)} + PO_4{}^{3-}{}_{(aq)}$$

$$K_{sp} = [Ag^+]^3[PO_4{}^{3-}]$$

Salts which are **soluble** have large K_{sp} values as their ion concentrations are large. Low soublility compounds have small K_{sp} values (see p. 332 Hebden). The smaller the K_{sp} value, the LESS soluble the salt.

VIII) Type A Problems: Calculating K_{sp}

i. If the solubility of Ag_2S in water is 1.3×10^{-17}M (this means that 1.3×10^{-17} mol of Ag_2S dissolve for every litre of solution), calculate the K_{sp}.

$$\underset{1.3\times10^{-17}M}{Ag_2S_{(s)}} \Leftrightarrow \underset{2.6\times10^{-17}M}{2Ag^+{}_{(aq)}} + \underset{1.3\times10^{-17}M}{S^{2-}{}_{(aq)}}$$

$$K_{sp} = [Ag^+]^2[S^{2-}]$$

$$= (2.6\times10^{-17})^2(1.3\times10^{-17})$$

$$= 8.8\times10^{-51}$$

(very low K_{sp})

ii. Calculate the K_{sp} of MgS if 300.0g is the maximum that can be dissolved in 2.00L of water.

$$\frac{300.0\text{g MgS}}{2.00\text{ L}} \left| \frac{1\text{ mol MgS}}{56.4\text{g MgS}} \right. = 2.6596\text{ M MgS}$$

$$\underset{2.6596\text{ M}}{MgS_{(s)}} \rightleftharpoons \underset{2.6596\text{ M}}{Mg^{2+}_{(aq)}} + \underset{2.6596\text{ M}}{S^{2-}_{(aq)}}$$

$$K_{sp} = [Mg^{2+}][S^{2-}] = (2.6596\text{ M})^2 = 7.07 \text{ (Large } K_{sp})$$

MgS "soluble" on table. Why? 2.6596 M > 0.1M

iii. A 25.00mL sample of saturated ZnF_2 solution was filtered and then evaporated to dryness. The mass of the residue was 0.508g. Calculate the solubility product constant of ZnF_2.

$$\frac{0.508\text{g } ZnF_2}{0.02500\text{ L}} \left| \frac{1\text{ mol } ZnF_2}{103.4\text{g } ZnF_2} \right. = 0.1965\text{ M } ZnF_2$$

(would be soluble on table)

$$\underset{0.1965\text{ M}}{ZnF_{2(s)}} \rightleftharpoons \underset{0.1965\text{M}}{Zn^{2+}_{(aq)}} + \underset{0.3930\text{ M}}{2F^{-}_{(aq)}}$$

$$K_{sp} = [Zn^{2+}][F^-]^2 = (0.1965\text{ M})(0.3930\text{ M})^2 = \boxed{0.0304}$$

Assignment 8: Type A Exercises

1. At 25°C, only 0.00245g of $BaSO_4$ can be dissolved in 1.0L of H_2O. Calculate the K_{sp} for $BaSO_4$.
2. At 25°C, the solubility of Ag_3PO_4 is 1.8×10^{-5}M. Calculate the K_{sp} for Ag_3PO_4.
3. An experiment showed that a maximum of 1.49g of $AgBrO_3$ can dissolve in 1.00L of water at 25°C. What is the K_{sp} for $AgBrO_3$ at this temperature?
4. A saturated solution of CaF_2 contains 4.15×10^{-4}mol of CaF_2 in 2.0L of solution. What is the K_{sp} for CaF_2?

5. A solution in equilibrium with solid Ag_2S on the bottom of the beaker was found to contain 1.6×10^{-16}M S^{2-} and 2.5×10^{-18}M Ag^+. Calculate the solubility product constant of Ag_2S.
Extra Practice: Hebden p. 95 #42, 46, 50-51.

IX) Type B Problems: Calculating Solubility

For these problems, you must use the K_{sp} chart on p. 333 in Hebden. Remember that **solubility** means the molarity of the salt at **equilibrium** (when it is **saturated**).

1. Calculate the solubility of $CaCO_3$ in water at 25°C.

$$\underset{}{CaCO_{3(s)}} \rightleftharpoons \underset{x\,M}{Ca^{2+}_{(aq)}} + \underset{x\,M}{CO^{2-}_{3(aq)}} \qquad \text{Let } x = [Ca^{2+}] = [CO_3^{2-}]$$

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

$$5.0 \times 10^{-9} = x^2$$

$$x = 7.071 \times 10^{-5}\ M = [Ca^{2+}]$$

$$\frac{7.071 \times 10^{-5}\ mol\ Ca^{2+}}{1.0\ L} \left| \frac{1\ mol\ CaCO_3}{1\ mol\ Ca^{2+}} \right. = \boxed{7.1 \times 10^{-5}\ M\ CaCO_3}$$

2. Calculate the solubility of PbI_2 in g/L at 25°C.

$$PbI_{2(s)} \rightleftharpoons \underset{x\,M}{Pb^{2+}_{(aq)}} + \underset{2x\,M}{2I^-_{(aq)}} \qquad \text{Let } x = [Pb^{2+}] \text{ then } 2x = [I^-]$$

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

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

$$8.5 \times 10^{-9} = 4x^3$$

$$2.125 \times 10^{-9} = x^3$$

$$x = 1.2856 \times 10^{-3}\ M = [Pb^{2+}] \overset{\text{using stoich.}}{=} [PbI_2]$$

$$\frac{1.2856 \times 10^{-3}\ mol\ PbI_2}{1.0\ L} \left| \frac{461.0\ g\ PbI_2}{1\ mol\ PbI_2} \right. = \boxed{0.59\ g\ PbI_2/L}$$

Assignment 8 - Type A Exercises

1. $$\frac{0.00245\text{ g BaSO}_4}{1.0\text{ L}}\left|\frac{1\text{ mol BaSO}_4}{233.4\text{ g BaSO}_4}\right. = 1.0497 \times 10^{-5}\text{ M BaSO}_4$$

$$\underset{1.0497\times10^{-5}\text{ M}}{\text{BaSO}_{4(s)}} \rightleftharpoons \underset{1.0497\times10^{-5}\text{ M}}{\text{Ba}^{2+}_{(aq)}} + \underset{1.0497\times10^{-5}\text{ M}}{\text{SO}_4{}^{2-}_{(aq)}}$$

$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4{}^{2-}] = (1.0497 \times 10^{-5})^2 = \boxed{1.1 \times 10^{-10}}$$

2. $$\underset{1.8\times10^{-5}\text{ M}}{\text{Ag}_3\text{PO}_{4(s)}} \rightleftharpoons \underset{5.4\times10^{-5}\text{ M}}{3\text{Ag}^{+}_{(aq)}} + \underset{1.8\times10^{-5}\text{ M}}{\text{PO}_4{}^{3-}_{(aq)}}$$

$$K_{sp} = [\text{Ag}^+]^3[\text{PO}_4{}^{3-}] = (5.4\times10^{-5})^3(1.8\times10^{-5}) = \boxed{2.8 \times 10^{-18}}$$

3. $$\frac{1.49\text{ g AgBrO}_3}{1.00\text{ L}}\left|\frac{1\text{ mol AgBrO}_3}{235.8\text{ g AgBrO}_3}\right. = 6.3189 \times 10^{-3}\text{ M}$$

$$\underset{6.3189\times10^{-3}\text{ M}}{\text{AgBrO}_{3(s)}} \rightleftharpoons \underset{6.3189\times10^{-3}\text{ M}}{\text{Ag}^{+}_{(aq)}} + \underset{6.3189\times10^{-3}\text{ M}}{\text{BrO}_3{}^{-}_{(aq)}}$$

$$K_{sp} = [\text{Ag}^+][\text{BrO}_3^-] = (6.3189 \times 10^{-3})^2 = \boxed{3.99 \times 10^{-5}}$$

4. $$[\text{CaF}_2] = \frac{\text{mol}}{V} = \frac{4.15\times10^{-4}\text{ mol}}{2.0\text{ L}} = 2.075 \times 10^{-4}\text{ M}$$

$$\underset{2.075\times10^{-4}\text{ M}}{\text{CaF}_{2(s)}} \rightleftharpoons \underset{2.075\times10^{-4}\text{ M}}{\text{Ca}^{2+}_{(aq)}} + \underset{4.15\times10^{-4}\text{ M}}{2\text{F}^{-}_{(aq)}}$$

$$K_{sp} = [\text{Ca}^{2+}][\text{F}^-]^2 = (2.075\times10^{-4})(4.15\times10^{-4})^2 = \boxed{3.6 \times 10^{-11}}$$

5. $$K_{sp} = [\text{Ag}^+]^2[\text{S}^{2-}] \qquad \text{Ag}_2\text{S}_{(s)} \rightleftharpoons 2\text{Ag}^{+}_{(aq)} + \text{S}^{2-}_{(aq)}$$

$$= (2.5\times10^{-18})^2(1.6\times10^{-16})$$

$$= \boxed{1.0 \times 10^{-51}}$$

3. How many grams of $PbBr_2$ can be dissolved in 250.0mL of water at 25°C?

$$PbBr_{2(s)} \rightleftharpoons \underset{x\text{ M}}{Pb^{2+}_{(aq)}} + \underset{2x\text{ M}}{2Br^-_{(aq)}}$$

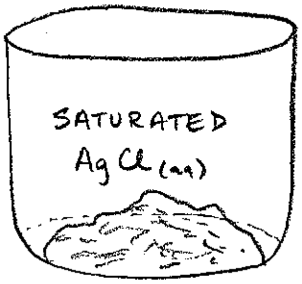
Let $x = [Pb^{2+}]$
then $2x = [Br^-]$

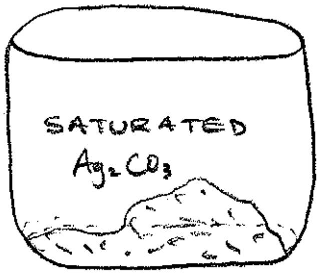
$$K_{sp} = [Pb^{2+}][Br^-]^2$$
$$6.6 \times 10^{-6} = x(2x)^2$$
$$6.6 \times 10^{-6} = 4x^3$$
$$1.65 \times 10^{-6} = x^3$$
$$x = 1.1817 \times 10^{-2}\text{ M} = [Pb^{2+}] = [PbBr_2]$$

$$\frac{1.1817 \times 10^{-2}\text{ mol }PbBr_2}{1.0\text{ L}} \left| \frac{1\text{ L}}{4\ (250\text{ mL})} \right| \frac{367.0\text{ g }PbBr_2}{1\text{ mol }PbBr_2}$$

$$= \boxed{1.1\text{ g }PbBr_2\text{ per 250 mL}}$$

4. Consider the following saturated solutions at 25°C:





Using calculations, identify the solution with the greater $[Ag^+]$.

$$AgCl_{(s)} \rightleftharpoons \underset{x\text{ M}}{Ag^+_{(aq)}} + \underset{x\text{ M}}{Cl^-_{(aq)}}$$

Let $x = [Ag^+] = [Cl^-]$

$$K_{sp} = [Ag^+][Cl^-]$$
$$1.8 \times 10^{-10} = x^2$$
$$x = 1.34 \times 10^{-5}\text{ M }Ag^+$$

$$Ag_2CO_{3(s)} \rightleftharpoons \underset{2x\text{ M}}{2Ag^+_{(aq)}} + \underset{x\text{ M}}{CO^{2-}_{3(aq)}}$$

Let $x = [CO_3^{2-}]$
then $2x = [Ag^+]$

$$K_{sp} = [Ag^+]^2[CO_3^{2-}]$$
$$8.5 \times 10^{-12} = (2x)^2(x)$$
$$x = 1.29 \times 10^{-4}\text{ M }CO_3^{2-}$$
$$2x = 2.57 \times 10^{-4}\text{ M }Ag^+$$

$\therefore Ag_2CO_3$ has greater $[Ag^+]$.

Assignment 9: Type B Exercises

1. The K_{sp} at a certain temperature for $Ni(OH)_2$ is 1.6×10^{-16}. Calculate the solubility of $Ni(OH)_2$.
2. At a certain temperature, the K_{sp} for $CaSO_4$ is 2.4×10^{-5}. Calculate the solubility of $CaSO_4$ in g/L.
3. What is the equilibrium concentration of Cd^{2+} ions in a saturated solution made by shaking $CdS_{(s)}$ with water? $K_{sp} = 6.0 \times 10^{-27}$
4. Calculate the mass of $MgCO_3$ that could dissolve in 3.0L of water at 25°C
5. What is the $[Ag^+]$ in a saturated solution of Ag_2CO_3 at 25°C ?
6. What mass of $PbSO_4$ will dissolve in 5.0L of water at 25°C?

Extra Practice: Hebden p. 95 #43-45, 47-49, 52-55.

- QUIZ 2 -

X) Type C Problems: Trial K_{sp} ('Q' in Hebden)

When two solutions are mixed, a low solubility ion combination could be introduced and could form a precipitate if enough are present to saturate. A Trial K_{sp} can be calculated to deduce whether a precipitate will or will not form.

A precipitate will form only if the TRIAL K_{sp} is GREATER than the K_{sp}!

1. Will a precipitate form if 40.0mL of 8.0×10^{-3}M $Mg(NO_3)_2$ are mixed with 60.0mL of 1.0×10^{-2}M K_2CO_3 ?

$$Mg(NO_3)_{2(aq)} + K_2CO_{3(aq)} \Rightarrow \underbrace{MgCO_{3(s)}}_{\text{possible ppt.}} + 2KNO_{3(aq)}$$

$$[Mg(NO_3)_2]_f = \frac{M_iV_i}{V_f} = \frac{(8.0 \times 10^{-3}\,M)(0.0400\,L)}{0.100\,L} = 3.2 \times 10^{-3}\,M$$

$$[K_2CO_3]_f = \frac{M_iV_i}{V_f} = \frac{(1.0 \times 10^{-2}\,M)(0.0600\,L)}{0.100\,L} = 0.0060\,M$$

$$\underset{3.2 \times 10^{-3}\,M}{Mg(NO_3)_{2(s)}} \longrightarrow \underset{3.2 \times 10^{-3}\,M}{Mg^{2+}_{(aq)}} + \underset{6.4 \times 10^{-3}\,M}{2NO_3^-}$$

$$\underset{0.0060\,M}{K_2CO_{3(s)}} \longrightarrow \underset{0.0120\,M}{2K^+_{(aq)}} + \underset{0.0060\,M}{CO^{2-}_{3(aq)}}$$

$$MgCO_{3(s)} \overset{?}{\rightleftharpoons} Mg^{2+}_{(aq)} + CO^{2-}_{3(aq)}$$

$$\text{Trial } K_{sp} = (3.2 \times 10^{-3}\,M)(0.0060\,M) = 1.92 \times 10^{-5}$$

$$K_{sp}\,(\text{table}) = 6.8 \times 10^{-6}$$

Trial $K_{sp} > K_{sp}$ ∴ ppt WILL form.

* excess ions will continue to form ppt. until trial K_{sp} equals K_{sp}.

Assignment 9 : Type B Exercises

1. $$\underset{}{Ni(OH)_{2(s)}} \rightleftharpoons \underset{x\,M}{Ni^{2+}_{(aq)}} + \underset{2x\,M}{2OH^{-}_{(aq)}}$$

Let $x = [Ni^{2+}]$

then $2x = [OH^-]$

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

$1.6 \times 10^{-16} = x(2x)^2$

$1.6 \times 10^{-16} = 4x^3$

$x = 3.41995 \times 10^{-6}\ M = [Ni^{2+}] = [Ni(OH)_2]$

$\boxed{[Ni(OH_2)] = 3.4 \times 10^{-6}\ M}$

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

Let $x = [Ca^{2+}] = [SO_4{}^{2-}]$

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

$2.4 \times 10^{-5} = x^2$

$x = 4.899 \times 10^{-3}\ M = [Ca^{2+}] = [CaSO_4]$

$$\frac{4.899 \times 10^{-3}\ \text{mol } CaSO_4}{1.0\ L} \times \frac{136.2\ g\ CaSO_4}{1\ \text{mol } CaSO_4} = \boxed{0.67\ g\ CaSO_4 / L}$$

3. $$CdS_{(s)} \rightleftharpoons \underset{x\,M}{Cd^{2+}_{(aq)}} + \underset{x\,M}{S^{2-}_{(aq)}}$$

Let $x = [Cd^{2+}] = [S^{2-}]$

$K_{sp} = [Cd^{2+}][S^{2-}]$

$6.0 \times 10^{-27} = x^2$

$x = \boxed{7.7 \times 10^{-14}\ M = [Cd^{2+}]}$

4. $$MgCO_{3(s)} \rightleftharpoons \underset{x\,M}{Mg^{2+}_{(aq)}} + \underset{x\,M}{CO_3{}^{2-}_{(aq)}}$$

Let $x = [Mg^{2+}] = [CO_3{}^{2-}]$

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

$6.8 \times 10^{-6} = x^2$

$x = 0.002608\ M = [Mg^{2+}] = [MgCO_3]$

$\text{mol} = MV$

$= (0.002608\ M)(3.0\ L)$

$= 0.007823\ \text{mol } MgCO_3$

$$\frac{0.007823\ \text{mol}}{} \times \frac{84.3\ g}{1\ \text{mol}} = 0.66\ g\ MgCO_3$$

in 3.0 L

5. $$Ag_2CO_{3(s)} \rightleftharpoons \underset{2x\,M}{2Ag^{+}_{(aq)}} + \underset{x\,M}{CO_3{}^{2-}_{(aq)}}$$

Let $x = [CO_3{}^{2-}]$

then $2x = [Ag^+]$

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

$8.5 \times 10^{-12} = (2x)^2(x)$

$8.5 \times 10^{-12} = 4x^3$

$x = 1.2856 \times 10^{-4}\ M = [CO_3{}^{2-}]$

$\boxed{[Ag^+] = 2.6 \times 10^{-4}\ M}$

6. $$PbSO_{4(s)} \rightleftharpoons \underset{x\,M}{Pb^{2+}_{(aq)}} + \underset{x\,M}{SO_4{}^{2-}_{(aq)}}$$

Let $x = [Pb^{2+}]$

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

$1.8 \times 10^{-8} = x^2$

$x = [Pb^{2+}] = [PbSO_4] = 1.342 \times 10^{-4}\ M$

$\text{mol} = MV = (1.342 \times 10^{-4}\ \text{mol})(5.0\ L) =$

$6.708 \times 10^{-4}\ \text{mol } PbSO_4 \times \frac{303.3\ g}{1\ \text{mol}} = \boxed{0.20\ g}$

2. Will a precipitate form if 50.0mL of 0.00100M $CaCl_2$ is added to 50.0mL of 0.0100M Na_2SO_4?

$$CaCl_{2(aq)} + Na_2SO_{4(aq)} \longrightarrow 2\,NaCl_{(aq)} + \underbrace{CaSO_{4(s)}}_{\text{possible ppt.}}$$

$$[CaCl_2]_f = \frac{M_iV_i}{V_f} = \frac{(0.00100\,M)(0.0500\,L)}{0.100\,L} = 5.0\times10^{-4}\,M$$

$$[Na_2SO_4]_f = \frac{M_iV_i}{V_f} = \frac{(0.0100\,M)(0.0500\,L)}{0.100\,L} = 5.0\times10^{-3}\,M$$

$$\underset{5.0\times10^{-4}\,M}{CaCl_{2(s)}} \longrightarrow \underset{5.0\times10^{-4}\,M}{Ca^{2+}_{(aq)}} + 2Cl^-_{(aq)}$$

$$\underset{5.0\times10^{-3}\,M}{Na_2SO_{4(s)}} \longrightarrow 2\,Na^+_{(aq)} + \underset{5.0\times10^{-3}\,M}{SO_4{}^{2-}{}_{(aq)}}$$

$$CaSO_{4(s)} \overset{?}{\rightleftharpoons} \underset{5.0\times10^{-4}\,M}{Ca^{2+}_{(aq)}} + \underset{5.0\times10^{-3}\,M}{SO_4{}^{2-}{}_{(aq)}}$$

$$\text{Trial } K_{sp} = [Ca^{2+}][SO_4{}^{2-}]$$
$$= (5.0\times10^{-4})(5.0\times10^{-3})$$
$$= 2.5\times10^{-6}$$

K_{sp} (TABLE) = 7.1×10^{-5}

Trial K_{sp} < K_{sp} ∴ ppt. will NOT form.

* not enough ions to form ppt. They stay 'happily' dissolved.

3. Does a ppt. form when 25.0 mL of 1.0×10^{-4}M $Zn(NO_3)_2$ is added to 45.0 mL of 2.4×10^{-5}M $Sr(OH)_2$? *K_{sp} ~~$(Zn(NO_3)_2)$~~ $(Zn(OH)_2)$ = 4.1×10^{-17}

$$Zn(NO_3)_{2(aq)} + Sr(OH)_{2(aq)} \longrightarrow Sr(NO_3)_{2(aq)} + \underbrace{Zn(OH)_{2(s)}}_{\text{possible ppt.}}$$

$$[Zn(NO_3)_2]_f = \frac{M_iV_i}{V_f} = \frac{(1.0\times10^{-4}M)(0.0250L)}{0.0700\,L} = 3.571\times10^{-5}\,M$$

$$[Sr(OH)_2]_f = \frac{M_iV_i}{V_f} = \frac{(2.4\times10^{-5}M)(0.0450L)}{0.0700\,L} = 1.543\times10^{-5}\,M$$

$$\underset{3.571\times10^{-5}\,M}{Zn(NO_3)_2} \longrightarrow \underset{3.571\times10^{-5}\,M}{Zn^{2+}_{(aq)}} + 2\,NO_3{}^-{}_{(aq)}$$

$$\underset{1.543\times10^{-5}\,M}{Sr(OH)_2} \longrightarrow Sr^{2+}_{(aq)} + \underset{3.086\times10^{-5}\,M}{2OH^-}$$

$$Zn(OH)_{2(s)} \overset{?}{\rightleftharpoons} \underset{3.571\times10^{-5}\,M}{Zn^{2+}_{(aq)}} + \underset{3.086\times10^{-5}\,M}{2\,OH^-_{(aq)}}$$

$$\text{Trial } K_{sp} = [Zn^{2+}][OH^-]^2$$
$$= (3.571\times10^{-5})(3.086\times10^{-5})^2$$
$$= 3.4\times10^{-14}$$

K_{sp} (TABLE) = 4.1×10^{-17}

Trial K_{sp} > K_{sp}

∴ ppt. will form.

Assignment 10: Type C Exercises

1. Will a precipitate form when 1.0L of 3.0×10^{-10}M $Zn(NO_3)_2$ is added to 1.0L of 2.0×10^{-11}M Na_2S?

2. Will a precipitate form when 2.0L of 6.7×10^{-3}M SrS is added to 1.0L of 4.3×10^{-4}M K_2SO_4 ?

3. Will a precipitate form when 1.0L of 5.0×10^{-4}M $MgCl_2$ is added to 1.0L of 2.0×10^{-2}M Na_2CO_3 ?

4. 1.0 L of 2.5×10^{-7}M of $Fe(NO_3)_2$ is mixed with 1.0 L of 1.0×10^{-5}M NaOH. Will a precipitate be observed?

Extra Practice: Hebden p. 98-99 #56, 59, 62-64, 66.

XI) Type D Problems: Determining the Conc. of a Specific Ion

Precipitate just starts to form when Trial K_{sp} becomes minutely greater than K_{sp}. Theoretically speaking, no ppt. forms when Trial $K_{sp} = K_{sp}$ (perfectly saturated solution).

1. What is the maximum $[Sr^{2+}]$ that can be dissolved in a 0.020M solution of K_2SO_4 without a precipitate of $SrSO_4$ forming?

$$\underset{0.020\,M}{K_2SO_4} \longrightarrow 2K^+ + \underset{0.020\,M}{SO_4^{2-}}$$

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

$$K_{sp} = [Sr^{2+}][SO_4^{2-}] \qquad K_{sp}\,(\text{table}) = 3.4 \times 10^{-7}$$

$$3.4 \times 10^{-7} = [Sr^{2+}](0.020\,M)$$

$$[Sr^{2+}] = \boxed{1.7 \times 10^{-5}\,M}$$

any more and a ppt. will begin to form.

Assignment 10 - Type C Exercises

1. $$Zn(NO_3)_{2\,(aq)} + Na_2S_{(aq)} \longrightarrow 2\,NaNO_{3\,(aq)} + \underset{\text{possible ppt}}{\underline{ZnS_{(s)}}}$$

$$[Zn(NO_3)_2]_f = \frac{M_iV_i}{V_f} = \frac{(3.0\times10^{-10}M)(1.0L)}{2.0\,L}$$
$$= 1.5\times10^{-10}M$$

$$[Na_2S]_f = \frac{M_iV_i}{V_f} = \frac{(2.0\times10^{-11}M)(1.0L)}{2.0\,L}$$
$$= 1.0\times10^{-11}M$$

$$\underset{1.5\times10^{-10}M}{Zn(NO_3)_2} \longrightarrow \underset{1.5\times10^{-10}M}{Zn^{2+}} + 2NO_3^-$$

$$\underset{1.0\times10^{-11}M}{Na_2S} \longrightarrow 2Na^+ + \underset{1.0\times10^{-11}M}{S^{2-}}$$

$$ZnS_{(s)} \rightleftharpoons Zn^{2+}_{(aq)} + S^{2-}_{(aq)}$$

$$\text{Trial } K_{sp} = [Zn^{2+}][S^{2-}] \qquad K_{sp}\,(\text{table}) = 2.0\times10^{-25}$$
$$= (1.5\times10^{-10})(1.0\times10^{-11})$$
$$= 1.5\times10^{-21}$$

Trial $K_{sp} > K_{sp}$ ∴ ppt. <u>WILL</u> form.

2. $$SrS_{(aq)} + K_2SO_{4\,(s)} \longrightarrow K_2S_{(aq)} + \underset{\text{possible ppt.}}{\underline{SrSO_{4\,(s)}}}$$

$$[SrS]_f = \frac{M_iV_i}{V_f} = \frac{(6.7\times10^{-3}M)(2.0L)}{3.0\,L}$$
$$= 4.4667\times10^{-3}\,M$$

$$[K_2SO_4]_f = \frac{M_iV_i}{V_f} = \frac{(4.3\times10^{-4}M)(1.0L)}{3.0\,L}$$
$$= 1.4333\times10^{-4}\,M$$

$$\underset{4.4667\times10^{-3}M}{SrS} \longrightarrow \underset{4.4667\times10^{-3}M}{Sr^{2+}} + S^{2-}$$

$$\underset{1.4333\times10^{-4}M}{K_2SO_4} \longrightarrow 2K^+ + \underset{1.4333\times10^{-4}M}{SO_4^{2-}}$$

$$SrSO_{4\,(s)} \rightleftharpoons Sr^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

$$\text{Trial } K_{sp} = [Sr^{2+}][SO_4^{2-}] \qquad K_{sp}\,(\text{table}) = 3.4\times10^{-7}$$
$$= (4.4667\times10^{-3})(1.4333\times10^{-4})$$
$$= 6.40\times10^{-7}$$

Trial $K_{sp} > K_{sp}$ ∴ ppt. <u>WILL</u> form.

3. $MgCl_{2\,(aq)} + Na_2CO_{3\,(aq)} \longrightarrow 2NaCl_{(aq)} + \underset{\text{possible ppt}}{\underline{MgCO_{3\,(s)}}}$

$[MgCl_2]_f = \frac{M_iV_i}{V_f} = \frac{(5.0\times10^{-4}M)(1.0L)}{2.0L} = 2.5\times10^{-4}M$

$[Na_2CO_3]_f = \frac{M_iV_i}{V_f} = \frac{(2.0\times10^{-2}M)(1.0L)}{2.0L} = 1.0\times10^{-2}M$

$\underset{2.5\times10^{-4}M}{MgCl_2} \longrightarrow \underset{2.5\times10^{-4}M}{Mg^{2+}} + 2Cl^-$

$\underset{1.0\times10^{-2}M}{Na_2CO_3} \longrightarrow 2Na^+ + \underset{1.0\times10^{-2}M}{CO_3^{2-}}$

$$MgCO_{3\,(s)} \rightleftharpoons Mg^{2+}_{(aq)} + CO^{2-}_{3\,(aq)}$$

$\text{Trial } K_{sp} = [Mg^{2+}][CO_3^{2-}]$ $\quad K_{sp}\text{ (table)} = 6.8\times10^{-6}$

$= (2.5\times10^{-4})(1.0\times10^{-2})$

$= 2.5\times10^{-6}$

Trial $K_{sp} < K_{sp}$ ∴ ppt. will not form.

4. $Fe(NO_3)_{2\,(aq)} + 2NaOH_{(aq)} \longrightarrow 2NaNO_{3\,(aq)} + \underset{\text{possible ppt.}}{\underline{Fe(OH)_{2\,(s)}}}$

$[Fe(NO_3)_2]_f = \frac{M_iV_i}{V_f} = \frac{(2.5\times10^{-7}M)(1.0L)}{2.0L} = 1.25\times10^{-7}M$

$[NaOH]_f = \frac{M_iV_i}{V_f} = \frac{(1.0\times10^{-5})(1.0L)}{2.0L} = 5.0\times10^{-6}M$

$\underset{1.25\times10^{-7}M}{Fe(NO_3)_2} \longrightarrow \underset{1.25\times10^{-7}M}{Fe^{2+}} + 2NO_3^-$

$\underset{5.0\times10^{-6}M}{NaOH} \longrightarrow Na^+ + \underset{5.0\times10^{-6}M}{OH^-}$

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

$\text{Trial } K_{sp} = [Fe^{2+}][OH^-]^2$ $\quad K_{sp}\text{ (table)} = 4.9\times10^{-17}$

$= (1.25\times10^{-7})(5.0\times10^{-6})^2$

$= 3.125\times10^{-18}$

Trial $K_{sp} < K_{sp}$ ∴ ppt will not form.

2. If Cu^{2+} is slowly added to a solution of 0.020M KIO_3, at what $[Cu^{2+}]$ does a precipitate of $Cu(IO_3)_2$ form?

$$\underset{0.020\,M}{KIO_3} \longrightarrow K^+ + \underset{0.020\,M}{IO_3^-}$$

$$Cu(IO_3)_{2(s)} \rightleftharpoons Cu^{2+}_{(aq)} + 2\,IO^-_{3(aq)}$$

$K_{sp} = [Cu^{2+}][IO_3^-]^2$ $\quad$ K_{sp} (table) $= 6.9 \times 10^{-8}$

$6.9 \times 10^{-8} = [Cu^{2+}][0.020]^2$

$[Cu^{2+}] = \boxed{1.7 \times 10^{-4}\ M}$ or greater!

3. $Fe(NO_3)_2$ is slowly added to a 2.0L solution of 0.010M Na_2S. What mass of $Fe(NO_3)_2$ would be required to just start precipitation of FeS? $\quad Fe(NO_3)_2 + Na_2S \longrightarrow 2\,NaNO_3 + \underline{\underline{FeS}}_{(s)}$

$$\underset{0.010\,M}{Na_2S} \longrightarrow 2\,Na^+ + \underset{0.010\,M}{S^{2-}}$$

$$FeS_{(s)} \rightleftharpoons Fe^{2+}_{(aq)} + S^{2-}_{(aq)}$$

$K_{sp} = [Fe^{2+}][S^{2-}]$ $\quad$ K_{sp} (table) $= 6.0 \times 10^{-19}$

$6.0 \times 10^{-19} = [Fe^{2+}](0.010)$

$[Fe^{2+}] = 6.0 \times 10^{-17}\,M = [Fe(NO_3)_2]$ (STOICH)

$Fe(NO_3)_2 \longrightarrow Fe^{2+} + 2\,NO_3^-$

$\underset{Fe(NO_3)_2}{mol} = MV = (6.0 \times 10^{-17}\,M)(2.0\,L)$

$= 1.2 \times 10^{-16}$ Mol $Fe(NO_3)_2$

$$\frac{1.2 \times 10^{-16}\ mol\ Fe(NO_3)_2 \mid 179.8\,g}{\mid 1\,mol}$$

$\boxed{= 2.2 \times 10^{-14}\ g\ Fe(NO_3)_2}$

4. Removing pollution by precipitation:

Waste water from a mining operation was found to have a $[Cd^{2+}]$ exceeding environmental standards. Before discharging the water into an adjacent river, the $[Cd^{2+}]$ had to be lowered to 1.0×10^{-5}M. What $[OH^-]$ would be required to bring the $[Cd^{2+}]$ to an acceptable level? $K_{sp} = 2.2 \times 10^{-14}$ for $Cd(OH)_2$ $\quad$ * assume volume of pond known.

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

* Using OH^- ions to ppt. Cd^{2+} ions and lower $[Cd^{2+}]$

* DANGER: OH^- ↑ pH.

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

$2.2 \times 10^{-14} = (1.0 \times 10^{-5}\,M)[OH^-]^2$

$[OH^-] = \boxed{4.7 \times 10^{-5}\ M}$ → useful, of course to know volume of 'pond'!

* think: Trial Ksp vs. Ksp. is needed

There is $> 1.0 \times 10^{-5}\,M$ Cd^{2+} in the water. If you add $4.7 \times 10^{-5}\,M$ OH^- to it, ppt will form since Trial Ksp > Ksp.

Assignment 11: Type D Exercises

1. What $[S^{2-}]$ must be present to just start precipitation of CuS from a 0.20M solution of $CuCl_2$?

2. What $[Cu^{2+}]$ would be required to just start precipitation of $Cu(IO_3)_2$ from a 0.20M solution of KIO_3?

3. $AgNO_3$ is added to a 0.10M solution of NaCl. What $[Ag^+]$ must be present in order to observe a precipitate?

4. What is the maximum $[SO_4^{2-}]$ that can be present in 0.010M $CaBr_2$ solution without a precipitate forming?

5. What $[Ba^{2+}]$ must be present to just start a precipitation in a 5.0 x 10^{-4}M solution of K_2CrO_4?

Extra Practice: Hebden p. 98-99 #57-58, 60, 65, 67 (*Tough* 68-69)

- QUIZ 3 -

XII) Hard Water

More common

Hard water is caused by Ca^{2+} and/or Mg^{2+} ions dissolved in water. It frequently occurs in areas rich in limestone ($CaCO_3$). Ca^{2+} can be released from limestone in two ways:

1. Acid Rain

$$\underset{\text{acid}}{CaCO_3 + 2H^+} \Rightarrow Ca^{2+}{}_{(aq)} + H_2O + CO_2$$

- can form sinkholes if limestone is just below ground, or caverns if limestone is way below ground.

2. Due to CO_2 dissolved in rainwater

- water leaks through cracks.

$$CaCO_{3(s)} + CO_2 + H_2O \Leftrightarrow Ca^{2+}{}_{(aq)} + 2HCO_3^- + \text{heat}$$

* re-crystallization forms STALACTITES as $[CO_2]$ low in caves.

Two types of hard water:

a) PERMANENT hard water contains only Ca^{2+} and/or Mg^{2+}

b) TEMPORARY hard water contains Ca^{2+}/Mg^{2+} and HCO_3^-

DISADVANTAGES OF HARD WATER: I) Bitter taste. II) Cannot use hard water with SOAPS since the *stearate ion* ('soap' ion) precipitates with Ca^{2+}/Mg^{2+} so one would require WAY more soap than usual (expensive) AND one would have to deal with the curd-like precipitate that forms when either of these cations form ppt with the stearate ion.

ADVANTAGES??? I) Some people prefer the taste. II) Excellent source of useful electrolytes.

Assignment 11 - Type D Exercises

1. $\underset{0.20\,M}{CuCl_2} \longrightarrow \underset{0.20\,M}{Cu^{2+}} + 2Cl^-$

$CuS_{(s)} \rightleftharpoons Cu^{2+}_{(aq)} + S^{2-}_{(aq)}$

$K_{sp} = [Cu^{2+}][S^{2-}]$ $\quad K_{sp}\,(\text{table}) = 6.0 \times 10^{-37}$

$6.0 \times 10^{-37} = (0.20)[S^{2-}]$

$[S^{2-}] = \boxed{3.0 \times 10^{-36}\,M}$

2. $\underset{0.20\,M}{KIO_3} \longrightarrow K^+ + \underset{0.20\,M}{IO_3^-}$

$Cu(IO_3)_{2\,(s)} \rightleftharpoons Cu^{2+}_{(aq)} + 2IO_3^-{}_{(aq)}$

$K_{sp} = [Cu^{2+}][IO_3^-]^2$ $\quad K_{sp}\,(\text{table}) = 6.9 \times 10^{-8}$

$6.9 \times 10^{-8} = [Cu^{2+}][0.20]^2$

$[Cu^{2+}] = \boxed{1.7 \times 10^{-6}\,M}$

3. $\underset{0.10\,M}{NaCl} \longrightarrow \underset{0.10\,M}{Na^+} + \underset{0.10\,M}{Cl^-}$ $\qquad AgNO_3 \longrightarrow Ag^+ + NO_3^-$

$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$

$K_{sp} = [Ag^+][Cl^-]$ $\quad K_{sp}\,(\text{table}) = 1.8 \times 10^{-10}$

$1.8 \times 10^{-10} = [Ag^+][0.10]$

$[Ag] = \boxed{1.8 \times 10^{-9}\,M}$

4. $\underset{0.010\,M}{CaBr_2} \longrightarrow \underset{0.010\,M}{Ca^{2+}} + 2Br^-$

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

$K_{sp} = [Ca^{2+}][SO_4^{2-}]$ $\quad K_{sp}\,(\text{table}) = 7.1 \times 10^{-5}$

$7.1 \times 10^{-5} = (0.010\,M)[SO_4^{2-}]$

$[SO_4^{2-}] = \boxed{7.1 \times 10^{-3}\,M}$

5. $\underset{5.0 \times 10^{-4}\,M}{K_2Cr_2O_4} \longrightarrow 2K^+ + \underset{5.0 \times 10^{-4}\,M}{CrO_4^{2-}}$

$BaCrO_{4\,(s)} \rightleftharpoons Ba^{2+}_{(aq)} + CrO_4^{2-}{}_{(aq)}$

$K_{sp} = [Ba^{2+}][CrO_4^{2-}]$ $\quad K_{sp}\,(\text{table}) = 1.2 \times 10^{-10}$

$1.2 \times 10^{-10} = [Ba^{2+}](5.0 \times 10^{-4}\,M)$

$[Ba^{2+}] = \boxed{2.4 \times 10^{-7}\,M}$

How do you soften hard water?

Permanent hard water:

Use washing soda (Na_2CO_3)

$Ca^{2+}_{(aq)} + CO_3^{2-}_{(aq)} \Leftrightarrow CaCO_{3(s)}$ (gets rid of Ca^{2+})

'*Permanent*' because the only way to get rid of the offending Ca^{2+}/Mg^{2+} ions is to carry out a precipitation reaction. (filter the ppt. out of tap)

Temporary hard water:

Use **heat** to reverse the direction of the formation of temporary hard water.

$Ca^{2+}_{(aq)} + 2HCO_3^- + \text{heat} \Leftrightarrow CaCO_{3(s)} + CO_2 + H_2O$

The above reaction can clog pipes by production of solid calcium carbonate causing **boiler scale**. (trade-off)

Assignment 12: Hebden p. 104 #76-80

XIII) Altering the Solubility of a Salt

The amount of solid present in a saturated solution can be altered by adding ions to the solution. Depending on the ions added, more solid can dissolve, or more can be produced.

Decreasing Solubility (causing more solid to be PRODUCED):

$AgCl_{(s)} \Leftrightarrow Ag^+_{(aq)} + Cl^-_{(aq)}$

By adding Ag^+ as $AgNO_3$ or Cl^- as NaCl, the equilibrium is driven LEFT, resulting in more $AgCl_{(s)}$.

This is called the COMMON ION EFFECT.

Increasing Solubility (causing more solid to be DISSOLVED (DISSOCIATED)):

$$AgCl_{(s)} \Leftrightarrow Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$$

By adding an ion that will precipitate out one of the product ions, the equilibrium is driven RIGHT resulting in more $AgCl_{(s)}$ dissolving.

$$AgCl_{(s)} \Leftrightarrow Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$$

+

$SO_4^{2-}{}_{(aq)}$

↓

$Ag_2SO_{4(s)}$

1. In which of the following 0.10M solutions would CaC_2O_4 be least soluble in and most soluble in? $CaC_2O_{4(s)} \rightleftharpoons Ca^{2+} + C_2O_4^{2-}$

NaOH, MOST	KCl,	$Ca(NO_3)_2$ LEAST

2. In which of the following 0.10M solutions would $PbCl_2$ be most soluble? Least soluble? $PbCl_{2(s)} \rightleftharpoons Pb^{2+} + 2Cl^{-}$

HCl, ✓	$MgCl_2$, ✓ LEAST	$AgNO_3$, MOST	NH_4NO_3

3. In which of the following 0.10M solutions would $Sr(OH)_2$ be least soluble in?

[✓ A.] $Sr(NO_3)_2$
C. NaCl
B. MgS
D. KBr

4. In which of the following 0.10M solutions would NaCl be most soluble in?

A. H_2O
C. NH_4Cl
[✓ B.] $AgNO_3$
D. HNO_3

Assignment 13: Hebden p. 108 #81-86