Unit 3: Solubility

Jul 23-2:07 PM

I) Ionic & Covalent Solutions

What is a **solution**?

What are some examples of different types of solutions?
solid-solid:
liquid-solid:
liquid-liquid:
liquid-gas:
gas-gas:

Jul 23-2:27 PM

I) Ionic & Covalent Solutions

What is a **solution**?

a homogeneous mixture
homogeneous: one phase (looks like one thing)
mixture: two or more different types of particles

What are some examples of different types of solutions?
solid-solid: metal alloys such as brass, steel
liquid-solid: salt water
liquid-liquid: alcohol & water
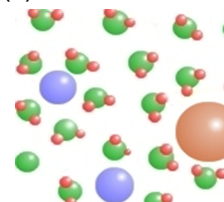liquid-gas: soda
gas-gas: air

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Solutions are made up of solute(s) and a solvent. Define each:
solute:

solvent:

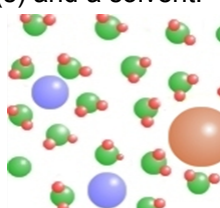What is the differencet between an ionic solution and a covalent solution?

Jul 23-2:30 PM

Solutions are made up of solute(s) and a solvent. Define each:
solute: the component(s) of a solution in lesser quantity

solvent: the component of a solution in greatest quantity

What is the difference between an ionic solution and a covalent solution?

ionic solution: an ionic compound (salt) dissolved (dissociated) in water to produce ions (charges)

covalent solution: a covalent (made up of non-metals) substance dissolved in water (sugar in water)

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How does a solid salt such as $NaCl_{(s)}$ dissolve in water? http://www.mhhe.com/physsci/chemistry/animations/chang_2e/molecular_view.swf
http://preparatorychemistry.com/Bishop_Solubility_frames.htm
http://preparatorychemistry.com/Bishop_Solubility_frames.htm

When a salt dissolves in water, it can be described as dissolving, but a more accurate term is **dissociation**. Why?

Jul 23-2:36 PM

solubilitynotes

How does a solid salt such as $NaCl_{(s)}$ dissolve in water? http://www.mhhe.com/physsci/chemistry/animations/chang_2e/molecular_view.swf

🔵 http://preparatorychemistry.com/Bishop_Solubility_frames.htm

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🔵

water molecules collide with the crystal lattice and knock Na^+ cations and Cl anions into solution

water molecules surround the cations and anions as a 'hydration shell'

When a salt dissolves in water, it can be described as dissolving, but a more accurate term is **dissociation**. Why?

NaCl does not dissolve in water as an NaCl neutral molecule. Instead, it dissociates (breaks apart) into Na^+ cations and Cl anions.

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$NaCl_{(aq)}$ is commonly used to depict table salt in solution (after dissociation has occurred). How come this is inaccurate?
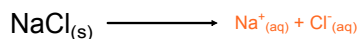
Example dissociation equation for a salt:

$NaCl_{(s)}$ ⟶
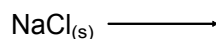
*In the solubility unit, always include states in equations

Jul 23-2:44 PM

$NaCl_{(aq)}$ is commonly used to depict table salt in solution (after dissociation has occurred). How come this is inaccurate?

When NaCl dissociates in water, it breaks apart into Na^+ and Cl^- ions. Thus, they are no longer together, contrary to what $NaCl_{(aq)}$ suggests. A more accurate description is $Na^+_{(aq)} + Cl^-_{(aq)}$.

Example dissociation equation for a salt:

$NaCl_{(s)}$ ⟶ $Na^+_{(aq)} + Cl^-_{(aq)}$

*In the solubility unit, always include states in equations

Jul 23-2:44 PM

Compounds that contain polyatomic **ions** are obviously ionic in nature (they are salts) and would dissociate in water to form ions. Write the dissociation equation for $K_2CO_{3(s)}$ disssolving in water:

Ions in solution are called _____.
They are what allow an ionic solution to _____.

Jul 23-2:50 PM

Compounds that contain polyatomic **ions** are obviously ionic in nature (they are salts) and would dissociate in water to form ions. Write the dissociation equation for $K_2CO_{3(s)}$ disssolving in water:

$K_2CO_{3(s)}$ ⟶ $2K^+_{(aq)} + CO_3^{2-}_{(aq)}$
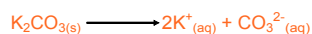
Ions in solution are called __electrolytes__.
They are what allow an ionic solution to __conduct electricity__.

Jul 23-2:50 PM

What is a covalent compound (also called 'molecular compound') composed of?

What is different about covalent compounds when they dissolve compared to ionic compounds?

Jul 23-3:30 PM

2

What is a covalent compound (also called 'molecular compound') composed of?

non-metals that share electrons

What is different about covalent compounds when they dissolve compared to ionic compounds?

Covalent compounds may dissolve into water (if they're polar) but do so as entire molecules. They do not dissociate into ions when dissolving like ionic compounds do.
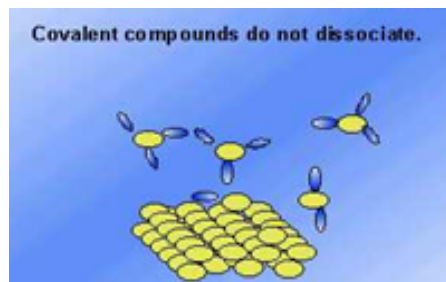
Jul 23-3:30 PM

Diagram of a covalent compound (sugar in this case) dissolving in water:

Covalent compounds do not dissociate.

Write the dissolving equation for the covalent sugar compound:
$$C_6H_{12}O_{6(s)} \longrightarrow$$
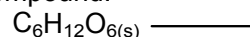
Jul 23-3:34 PM

Diagram of a covalent compound (sugar in this case) dissolving in water:

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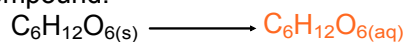
Jul 23-3:34 PM

Write the dissolving equation for CH_3OH (methanol) in water:

Do molecular (covalent) solutions conduct? Why or why not?

What do we call dissolved molecular compounds?

http://phet.colorado.edu/en/simulation/sugar-and-salt-solutions
micro

Jul 23-3:35 PM

Write the dissolving equation for CH_3OH (methanol) in water:
$$CH_3OH_{(s)} \longrightarrow CH_3OH_{(aq)}$$
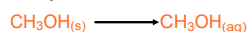
Do molecular (covalent) solutions conduct? Why or why not?

No, because ions are not created in the dissolving process. Only entire, neutral molecules dissolve.

What do we call dissolved molecular compounds?

non-electrolytes

http://phet.colorado.edu/en/simulation/sugar-and-salt-solutions
micro

Jul 23-3:35 PM

How can you tell the difference between ionic and covalent compounds?

Jul 23-4:09 PM

3

How can you tell the difference between ionic and covalent compounds?

ionic compounds start with a metal cation or NH_4^+ (ammonium)

covalent compounds start with a non-metal

acids often start with an H

Jul 23-4:09 PM

II) Solubility

What is a saturated solution?

How can you recognize a saturated solution visibly?

What kind of system is a saturated solution?

Is a saturated solution static or dynamic?

http://www.dlt.ncssm.edu/core/Chapter14-Gas_Phase-Solubility-Complex_Ion_Equilibria/Chapter14-Animations/Solubility_of_AgCl.html

Jul 23-3:40 PM

II) Solubility

What is a saturated solution?
a solution that has the maximum amount of solute dissolved

How can you recognize a saturated solution visibly?
it will probably have excess, undissolved solute at the bottom

What kind of system is a saturated solution?
an equilibrium system: the rate of dissolving of the solid solute equals the rate of crystallization of the dissolved solute
ex. $NaCl_{(s)} \rightleftharpoons Na^+_{(aq)} + Cl^-_{(aq)}$

Is a saturated solution static or dynamic?
dynamic due to the continuous forward and reverse reaction occurring in the system

http://www.dlt.ncssm.edu/core/Chapter14-Gas_Phase-Solubility-Complex_Ion_Equilibria/Chapter14-Animations/Solubility_of_AgCl.html

Jul 23-3:40 PM

Therefore, does the amount of undissolved salt at the bottom of the solution change?

Write a saturated solution equation for KBr:

http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_04&folder=saturated_solutions

Jul 23-3:47 PM

Therefore, does the amount of undissolved salt at the bottom of the solution change?
No, as it is being replenished as fast as it is being dissolved

Write a saturated solution equation for KBr:
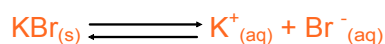
$$KBr_{(s)} \rightleftharpoons K^+_{(aq)} + Br^-_{(aq)}$$

http://www.wwnorton.com/college/chemistry/gilbert2/tutorials/interface.asp?chapter=chapter_04&folder=saturated_solutions

Jul 23-3:47 PM

Explain the difference between the rate of dissolving and the rate of precipitation for an **unsaturated** solution:

Write an unsaturated equation for KBr:

What's another name for the equation above?

Ho do the saturated and unsaturated equations differ?

Jul 23-3:49 PM

Explain the difference between the rate of dissolving and the rate of precipitation for an **unsaturated** solution:

rate of dissolving is greater than the rate of precipitation, thus eventually all the solute dissolves

Write an unsaturated equation for KBr:

$$KBr_{(s)} \longrightarrow K^+_{(aq)} + Br^-_{(aq)}$$
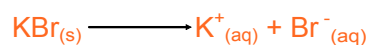
What's another name for the equation above?

dissociation equation

Ho do the saturated and unsaturated equations differ?

look at the arrow(s): single is unsaturated
double is saturated (equilibrium)

Jul 23-3:49 PM

Solubility: the solubility of a substance is a quantitative value. What is it?

Therefore, solubility can only be measured for what types of solutions?

What is 'concentration' in chemical terms?

Jul 23-3:54 PM

Solubility: the solubility of a substance is a quantitative value. What is it?

the molarity at saturation
(remember, saturation is when the maximum amount of solute has been dissolved)

Therefore, solubility can only be measured for what types of solutions?

saturated solutions (solutions at equilibrium)

What is 'concentration' in chemical terms?

how much solute is in a certain volume of solution

Jul 23-3:54 PM

Concentration is most commonly measured using what unit?

Concentration can also be measured using mass of solute per volume of solution, as you'll see in some examples.

Jul 23-3:59 PM

Concentration is most commonly measured using what unit?

Molarity (M)

$$Molarity = \frac{moles\ of\ solute}{litres\ of\ solution}$$
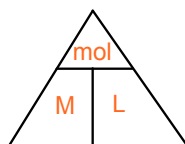
Concentration can also be measured using mass of solute per volume of solution, as you'll see in some examples.

Jul 23-3:59 PM

Every salt dissociates to some extent in water, but every salt has a different solubility in water. What does this mean?

Some salts can dissociate to a larger extent in water before becoming saturated. These salts are called _____ salts. Some salts dissociate only very little in water before saturating. These salts are called _____ salts.

Jul 23-4:02 PM

solubilitynotes

Every salt dissociates to some extent in water, but every salt has a different solubility in water. What does this mean?

there are no salts that will not dissolve in water but each salt has a unique molarity at saturation.

Some salts can dissociate to a larger extent in water before becoming saturated. These salts are called ____soluble____ salts. Some salts dissociate only very little in water before saturating. These salts are called _____low solubility_____ salts.

Solubility is also temperature dependent. A general rule for solid/liquid solutions is that solubility increases with increasing temperature.

Assignment 1:

Read Hebden pages 73-76 (check out the comic on p.76) and do Questions 1-5.

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Read Hebden pages 73-76 (check out the comic on p.76) and do Questions 1-5.

answers in the back of Hebden

III) Calculating Solubility and Ion Concentrations

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

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

dissoc eqn: $Na_2SO_{4(s)}$ $2Na^+_{(aq)} + SO_4^{2-}_{(aq)}$
0.20M   _____   _____

Calculate $[Fe^{3+}]$ and $[SO_4^{2-}]$ in a 0.50M solution of iron III sulphate:

III) Calculating Solubility and Ion Concentrations

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dissoc eqn: $Na_2SO_{4(s)}$ $2Na^+_{(aq)} + SO_4^{2-}_{(aq)}$
0.20M   0.40M   0.20M

Calculate $[Fe^{3+}]$ and $[SO_4^{2-}]$ in a 0.50M solution of iron III sulphate: $Fe_2(SO_4)_{3(s)} \longrightarrow 2Fe^{3+}_{(aq)} + 3SO_4^{2-}_{(aq)}$
0.50M   1.0M   1.5M

solubilitynotes

A saturated solution of $PbCl_2$ is found to contain 9.90g of $PbCl_2$ per litre of solution. Find $[Pb^{2+}]$.

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$$\frac{9.90g\ PbCl_2}{} \left| \frac{1mol\ PbCl_2}{278g\ PbCl_2} \right. = 0.035586mol\ PbCl_2$$

$$[PbCl_2] = \frac{0.035586mol}{1.00L} = 0.035586M\ PbCl_2$$

$$PbCl_{2(s)} \quad\quad Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$$
$$0.035586M \quad 0.0356M$$

$$[Pb^{2+}] = 0.0356M$$

Jul 23-4:22 PM

Write an equation showing the equilibrium in a saturated solution of silver carbonate. The solubility of silver carbonate is $1.29 \times 10^{-4}M$. Find $[Ag^+]$ in grams per litre.

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$$Ag_2CO_{3(s)} \rightleftharpoons 2Ag^+_{(aq)} + CO_3^{2-}_{(aq)}$$
$$1.29 \times 10^{-4}M \quad 2.58 \times 10^{-4}M$$
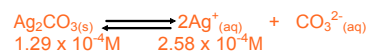
$$\frac{2.58 \times 10^{-4}mol\ Ag^+}{} \left| \frac{107.9g\ Ag^+}{1mol\ Ag^+} \right. = 0.0278\ g/L$$

Jul 23-4:27 PM

A 558mL solution of $Al_2(SO_4)_3$ has a concentration of 0.0342M. You then add 325mL of water to it. What is the new concentration?

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Dilution: $M_iV_i = M_fV_f$

$M_i = 0.0342M$
$V_i = 0.558L$
$V_f = 558 + 325 = 883mL = 0.883L$

$(0.0342)(0.558) = M_f(0.883)$

$M_f = 0.0216M$

$[Al_2(SO_4)_3] = 0.0216M$

Jul 23-4:32 PM

solubilitynotes

250mL of 0.30M K_2SO_4 and 250mL of 0.80M $MgCl_2$ are mixed, and no precipitate forms. Calculate the concentration of each ion in the final solution.

dilution:

$[K_2SO_4]_f = \dfrac{(0.30M)(0.25L)}{0.50L} = 0.15M$

$K_2SO_{4(s)} \longrightarrow 2K^+_{(aq)} + SO_4^{2-}_{(aq)}$
0.15M   0.30M   0.15M

$[MgCl_2]_f = \dfrac{(0.80M)(0.25L)}{0.50L} = 0.40M$

$MgCl_{2(aq)} \longrightarrow Mg^{2+}_{(aq)} + 2Cl^-_{(aq)}$
0.40M   0.40M   0.80M

*Hint: if equal volumes are mixed, the volume has been doubled. Therefore, molarities have been halved

$[K^+] = 0.30M$, $[Mg^{2+}] = 0.40M$, $[SO_4^{2-}] = 0.15M$, $[Cl^-] = 0.80M$

Jul 23-4:35 PM

Assignment 2
Hebden page 77, numbers 8 and 9
      page 78, number 13
        page 81, numbers 18a-d, 20abfg

Jul 23-4:41 PM

IV) Using the Solubility Table
The solubility table in the data booklet is used to predict whether a salt is **soluble** in water or **low solubility** in water. **Soluble** salts can dissociate and produce a solution of least 0.1M before becoming saturated. **Low solubility** salts dissociate very little as they become saturated before reaching 0.1M.
Identify three soluble salts, and three low solubility salts using the table. http://phet.colorado.edu/en/simulation/soluble-salts

Soluble   **Low Solubility**

Jul 24-1:31 PM

IV) Using the Solubility Table
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Identify three soluble salts, and three low solubility salts using the table. http://phet.colorado.edu/en/simulation/soluble-salts

Soluble   **Low Solubility**

sodium chloride   lead chloride
calcium sulphide   barium sulphate
potassium hydroxide   silver phosphate

Jul 24-1:31 PM

If two ions (a cation and an anion) that are low solubility with one another are put into the same solution, and the resulting concentrations of these ions exceed the solubility (molarity at saturation) for that salt, then some of those cations and anions will precipitate out of solution.

mix them together

the saturation molarity of Ag^+ and Cl^- ions remain in solution

K^+ NO_3^- K^+ Ag^+ K^+
Ag^+ NO_3^- NO_3^- Cl^- NO_3^-
K^+ NO_3^- Cl^- K^+ K^+
K^+ NO_3^- NO_3^-

Silver nitrate dissolves in water producing silver and nitrate ions.

Potassium chloride dissolves in water producing potassium and chloride ions.

http://www.absorblearning.com/media/attachment.action?quick=uu&att=2210
http://preparatorychemistry.com/Bishop_Solubility_frames.htm
http://www.wiley.com/college/chem/brady184764/resources/ch04/index_ch4_bysect.h

The extra Ag^+ and Cl^- ions precipitate out as $AgCl_{(s)}$, as it is low solubility

Jul 24-1:37 PM

Is it possible that a low solubility salt may not precipitate?

Is it possible that a soluble salt could precipitate?

Jul 24-1:48 PM

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solubilitynotes

Is it possible that a low solubility salt may not precipitate?

Yes. If the concentration of low solubility ions mixed with one another are below the solubility (molarity at saturation) for that salt, no precipitate will form.

Is it possible that a soluble salt could precipitate?

Yes. If the solubility (molarity at saturation) is exceeded when two ions are mixed, a precipitate will occur.

Jul 24-1:48 PM

There are two ways you can create a saturated solution for a salt. Explain two ways that a saturated solution of $PbCl_2$ could be created:
1) Direct Method:

2) Indirect Method:

Jul 24-1:46 PM

There are two ways you can create a saturated solution for a salt. Explain two ways that a saturated solution of $PbCl_2$ could be created:
1) Direct Method:

add $PbCl_2$ salt to water, and continue to stir, until excess $PbCl_2$ salt remains on the bottom

2) Indirect Method:

Mix two solutions together - one that contains Pb^{2+} and one that contains Cl^- - such as $Pb(NO_3)_{2(aq)}$ and $NaCl_{(aq)}$. If the molarities of Pb^{2+} and Cl^- exceed the solubility of $PbCl_2$, a precipitate will form and you'll have a saturated solution (which also contains the spectator ions NO_3^- and Na^+

Jul 24-1:46 PM

Use your table to predict whether each salt is soluble or low solubility:
1) sodium hydroxide
2) calcium sulphate
3) ammonium phosphate
4) rubidium sulphide
5) copper (I) iodide
6) copper (II) chloride
7) iron (II) sulphide
8) strontium hydroxide
9) magnesium sulphite
10) barium hydroxide

http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/crm3s2_3.swf

Jul 24-2:07 PM

Use your table to predict whether each salt is soluble or low solubility:
1) sodium hydroxide   S
2) calcium sulphate   LS
3) ammonium phosphate   S
4) rubidium sulphide   S
5) copper (I) iodide   LS
6) copper (II) chloride   S
7) iron (II) sulphide   LS
8) strontium hydroxide   S
9) magnesium sulphite   LS
10) barium hydroxide   LS

http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/crm3s2_3.swf

Jul 24-2:07 PM

Will a precipitate form if equal volumes of the following 0.2M solutions are mixed?

$Pb(NO_3)_{2(aq)} + KI_{(aq)} \longrightarrow$
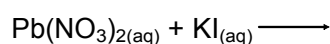
Which 1.0M soluble solutions could be mixed in order to produce a precipitate of $Mg(OH)_2$?

Jul 24-2:13 PM

9

solubilitynotes

Will a precipitate form if equal volumes of the following 0.2M solutions are mixed?

$Pb(NO_3)_{2(aq)} + KI_{(aq)} \longrightarrow$ $2KNO_{3(aq)} + PbI_{2(s)-}$
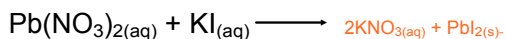
Yes, a ppt of PbI_2 will form. Mixing equal volumes of 0.2M solutions creates 0.1M solutions. So anything that is low solubility will form a precipitate.

Which 1.0M soluble solutions could be mixed in order to produce a precipitate of $Mg(OH)_2$?

any solution that has Mg^{2+} and is soluble - such as $Mg(NO_3)_2$ and any solution that has OH^- and is soluble - such as NaOH

$Mg(NO_3)_{2(aq)} + 2NaOH_{(aq)} \longrightarrow 2NaNO_{3(aq)} + Mg(OH)_{2(s)}$
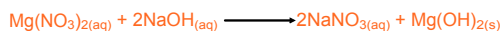
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Assignment 3
1) Describe the difference between a soluble salt and a low solubility salt.
2) If you mix two solutions together, and a low solubility ion combination results, explain when a precipitate would form and when it wouldn't.
3) Hebden page 83,
numbers 21 & 22
4) Hebden page 84,
number 24

Jul 24-2:20 PM

Assignment 3
1) Describe the difference between a soluble salt and a low solubility salt.
2) If you mix two solutions together, and a low solubility ion combination results, explain when a precipitate would form and when it wouldn't.
3) Hebden page 83, numbers 21 & 22
4) Hebden page 84, number 24

1) A soluble salt doesn't become saturated until it's made a solution of at least 0.1M. A low solubility salt becomes saturated before its molarity can get as high as 0.1M
2) If the low solubility ion concentrations exceed the solubility of the salt, it will precipitate. If the ion concentrations are below the solubility, a precipitate will not form.
3) answers in the back of Hebden
4) answers in the back of Hebden

Jul 24-2:20 PM

V) Formula, Complete Ionic, and Net Ionic Equations

These equations describe the process of mixing two soluble solutions together.

Suppose 0.2M $CaCl_{2(aq)}$ is mixed with an equal volume of 0.2M $AgNO_{3(aq)}$. Write the formula, complete ionic, and net ionic equations.

Formula Equation (double replacement equation):

Jul 24-2:28 PM

V) Formula, Complete Ionic, and Net Ionic Equations

These equations describe the process of mixing two soluble solutions together.

Suppose 0.2M $CaCl_{2(aq)}$ is mixed with an equal volume of 0.2M $AgNO_{3(aq)}$. Write the formula, complete ionic, and net ionic equations.

Formula Equation (double replacement equation):

$CaCl_{2(aq)} + 2AgNO_{3(aq)} \longrightarrow Ca(NO_3)_{2(aq)} + AgCl_{(s)}$
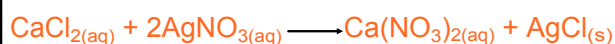
Jul 24-2:28 PM

Complete Ionic Equation
The formula equation is inaccurate for the most part, as the salts dissociate into ions in solution. This is reflected in the complete ionic equation:

Net Ionic Equation
Only ions that react are part of the Net.

Jul 24-2:32 PM

solubilitynotes

Complete Ionic Equation
The formula equation is inaccurate for the most part, as the salts dissociate into ions in solution. This is reflected in the complete ionic equation:

$Ca^{2+}_{(aq)} + 2Cl^-_{(aq)} + 2Ag^+_{(aq)} + 2NO_3^-_{(aq)} \longrightarrow Ca^{2+}_{(aq)} + 2NO_3^-_{(aq)} + 2AgCl_{(s}$

Net Ionic Equation
Only ions that react are part of the Net.

$Ag^+_{(aq)} + Cl^-_{(aq)} \longrightarrow AgCl_{(s)}$
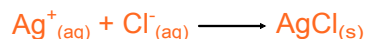
The example we did involved one low solubility ion combination. Sometimes, no low solubility ion combinations result. Sometimes, both ion combinations are low solubility.

Assignment 4

1) Aqueous magnesium choride is mixed with aqueous potassium hydroxide.

Assignment 4

1) Aqueous magnesium chloride is mixed with aqueous potassium hydroxide.
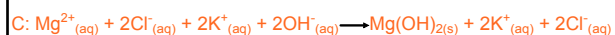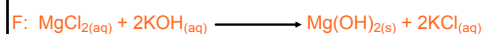
F: $MgCl_{2(aq)} + 2KOH_{(aq)} \longrightarrow Mg(OH)_{2(s)} + 2KCl_{(aq)}$

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

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

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2) Aqueous lithium carbonate is mixed with aqueous iron (II) sulphate.

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2) Aqueous lithium carbonate is mixed with aqueous iron (II) sulphate.
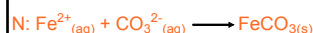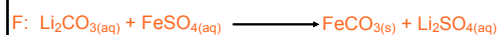
F: $Li_2CO_{3(aq)} + FeSO_{4(aq)} \longrightarrow FeCO_{3(s)} + Li_2SO_{4(aq)}$

C: $2Li^+_{(aq)} + CO_3^{2-}_{(aq)} + Fe^{2+}_{(aq)} + SO_4^{2-}_{(aq)} \longrightarrow FeCO_{3(s)} + 2Li^+_{(aq)} + SO_4^{2-}_{(aq)}$

N: $Fe^{2+}_{(aq)} + CO_3^{2-}_{(aq)} \longrightarrow FeCO_{3(s)}$

3) Aqueous copper (I) sulphate is mixed with aqueous strontium bromide.

Jul 24-2:40 PM

3) Aqueous copper (I) sulphate is mixed with aqueous strontium bromide.

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

C: $2Cu^+_{(aq)} + SO_4^{2-}{}_{(aq)} + Sr^{2+}_{(aq)} + 2Br^-_{(aq)} \longrightarrow 2CuBr_{(s)} + SrSO_{4(s)}$

N: $2Cu^+_{(aq)} + SO_4^{2-}{}_{(aq)} + Sr^{2+}_{(aq)} + 2Br^-_{(aq)} \longrightarrow 2CuBr_{(s)} + SrSO_{4(s)}$
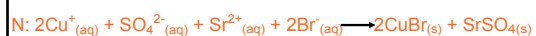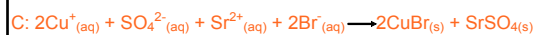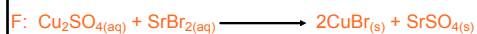
Jul 24-2:40 PM

4) Hebden page 87, numbers 25 e and i

Jul 24-2:40 PM

4) Hebden page 87, numbers 25 e and i

answers in the back of Hebden

Jul 24-2:40 PM

VI) Separating Mixtures of Ions by Precipitation

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

Jul 24-2:52 PM

VI) Separating Mixtures of Ions by Precipitation

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

add an anion that only precipitates one of the cations out

1) Add Cl^- as $NaCl_{(aq)}$ in order to precipitate Pb^{2+} as $PbCl_{2(s)}$

2) Add SO_4^{2-} as Na_2SO_4 in order to precipitate Ba^{2+} as $BaSO_{4(s)}$

Jul 24-2:52 PM

solubilitynotes

Assignment 5

1) Describe a method to separate Cl^- and OH^- that are in solution together.

2) A solution is known to contain Mg^{2+}, Ca^{2+}, and Pb^{2+}. Describe a method to separate the cations.

Jul 24-2:56 PM

Assignment 5

1) Describe a method to separate Cl^- and OH^- that are in solution together.

1) Add Mg^{2+} as $Mg(NO_3)_2$ to precipitate OH^- as $Mg(OH)_2$

2) Add Pb^{2+} as $Pb(NO_3)_2$ to precipitate Cl^- as $PbCl_2$

2) A solution is known to contain Mg^{2+}, Ca^{2+}, and Pb^{2+}. Describe a method to separate the cations.

1) Add Cl^- as $NaCl_{(aq)}$ to precipitate Pb^{2+} as $PbCl_2$

2) Add SO_4^{2-} as Na_2SO_4 to precipitate Ca^{2+} as $CaSO_4$

3) Add PO_4^{3-} as Na_3PO_4 to precipitate Mg^{2+} as $Mg_3(PO_4)_2$

Jul 24-2:56 PM

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3) You have a solution known to contain any or all of Cu^+, Ca^{2+}, Fe^{3+}, and Sr^{2+}. You have the following solutions available: 1M Na_2CO_3, 1M NaOH, 1M NaCl, and 1M Na_2S. What order would you add each to test for each cation?

4) Hebden page 90, numbers 28, 31, & 32

Jul 24-3:03 PM

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3) You have a solution known to contain any or all of Cu^+, Ca^{2+}, Fe^{3+}, and Sr^{2+}. You have the following solutions available: 1M Na_2CO_3, 1M NaOH, 1M NaCl, and 1M Na_2S. What order would you add each to test for each cation?

1M NaCl to test for Cu^+, then 1M Na_2S to test for Fe^{3+}, then 1M NaOH to test for Ca^{2+}, then 1M Na_2CO_3 to test for Sr^{2+}

4) Hebden page 90, numbers 28, 31, & 32

answers in the back of Hebden

Jul 24-3:03 PM

VII) Quantitative Solubility

A saturated solution is at equilibrium because the rate of _____ equals the rate of _____.
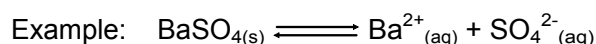
Example: $BaSO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$

Since equilibrium exists, we can work with an equilibrium constant (K_{eq}). The equilibrium constant for a saturated solution is called a K_{sp}. Write the K_{sp} equation for $BaSO_4$:

Jul 24-3:08 PM

VII) Quantitative Solubility

A saturated solution is at equilibrium because the rate of _____dissolving_____ equals the rate of _____precipitation_____.
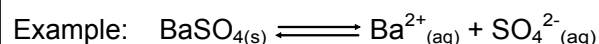
Example: $BaSO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$

Since equilibrium exists, we can work with an equilibrium constant (K_{eq}). The equilibrium constant for a saturated solution is called a K_{sp}. Write the K_{sp} equation for $BaSO_4$:

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

Jul 24-3:08 PM

solubilitynotes

The 'sp' from K_{sp} stands for _____.
The word 'product' is mathematical. A product is the result of a multiplication.
K_{sp} equations only contain multiplication (no division) because since the solid is always written on the left of the saturated solution equation, there will never be a denominator for the K_{sp} equation.

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

solid on left; not included in K_{sp} equation
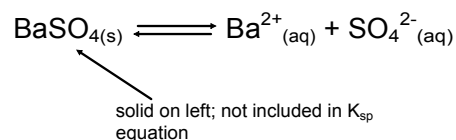
Jul 24-3:14 PM

Write the saturated solution equations (solid on the left) and the respective K_{sp} equations for $Ca(OH)_2$ and Ag_3PO_4.

Jul 24-3:18 PM

Write the saturated solution equations (solid on the left) and the respective K_{sp} equations for $Ca(OH)_2$ and Ag_3PO_4.

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

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

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

$$K_{sp} = [Ag^+]^3[PO_4^{3-}]$$
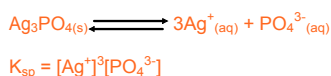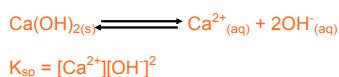
Jul 24-3:18 PM

Do **soluble** salts have large or small K_{sp} constants? Why? What about **low solubility** salts?

Notice that the K_{sp} table in the data booklet (p.5) only lists K_{sp} constants for low solubility salts. Notice that they are all small values, (much less than 1), as they should be.

Jul 24-3:23 PM

Do **soluble** salts have large or small K_{sp} constants? Why? What about **low solubility** salts?

Soluble salts have large K_{sp} values. K_{sp} is determined by multiplying ion concentrations. Soluble salts have larger ion concentrations, hence large K_{sp} constants.
Since low solubility salts have small ion concentrations, they will have small K_{sp} constants.

Notice that the K_{sp} table in the data booklet (p.5) only lists K_{sp} constants for low solubility salts. Notice that they are all small values, (much less than 1), as they should be.

Jul 24-3:23 PM

solubilitynotes

Jul 24-3:27 PM

VIII) Type A K_{sp} Problems: Calculating K_{sp}

K_{sp} constants can be calculated for soluble or low solubility salts, as long as there is information available about the respective **saturated** solution.

Example: If the solubility of Ag_2S in water is 1.3×10^{-17}M, calculate the K_{sp}.

$$Ag_2S_{(s)} \rightleftharpoons 2Ag^+_{(aq)} + S^{2-}_{(aq)}$$
$$1.3 \times 10^{-17}M \quad 2.6 \times 10^{-17}M \quad 1.3 \times 10^{-17}M$$

$K_{sp} = [Ag^+]^2[S^{2-}] = (2.6 \times 10^{-17})^2(1.3 \times 10^{-17}) = 8.8 \times 10^{-51}$

The K_{sp} is very small, supporting the fact that Ag_2S is a low solubility salt.

Jul 24-3:27 PM

Example: Calculate the K_{sp} of MgS if a maximum of 300.0g can be dissolved in 2.00L of water.

Jul 24-3:33 PM

Example: Calculate the K_{sp} of MgS if a maximum of 300.0g can be dissolved in 2.00L of water.

$$\frac{300.0g}{} \left| \frac{1mol\ MgS}{56.4g\ MgS} \right. = \frac{5.319mol\ MgS}{}$$

$[MgS] = \frac{5.319mol}{2.00L} = 2.66M$

$$MgS_{(s)} \rightleftharpoons Mg^{2+}_{(aq)} + S^{2-}_{(aq)}$$
$$2.66M \quad 2.66M \quad 2.66M$$
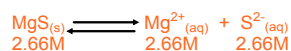
$K_{sp} = [Mg^{2+}][S^{2-}] = (2.66)(2.66) = 7.07$

This is a larger K_{sp} constant, as MgS is a soluble salt.

Jul 24-3:33 PM

Example: A 25.00mL sample of ZnF_2 saturated solution has the excess solid filtered off. Then, the solution is evaporated to dryness. The mass of the ZnF_2 salt that remained was 0.508g. Calculate the solubility product constant of ZnF_2.

Jul 24-3:38 PM

Example: A 25.00mL sample of ZnF_2 saturated solution has the excess solid filtered off. Then, the solution is evaporated to dryness. The mass of the ZnF_2 salt that remained was 0.508g. Calculate the solubility product constant of ZnF_2.

$$\frac{0.508g}{} \left| \frac{1mol}{103.4g} \right. = \frac{4.913 \times 10^{-3}mol}{}$$

$[ZnF_2] = \frac{4.913 \times 10^{-3}mol}{0.02500L} = 0.1965M$

$$ZnF_{2(s)} \rightleftharpoons Zn^{2+}_{(aq)} + 2F^-_{(aq)}$$
$$0.1965M \quad 0.1965M \quad 0.393M$$
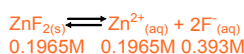
$K_{sp} = [Zn^{2+}][F^-]^2 = (0.1965)(0.393)^2 = 3.04 \times 10^{-2}$

Jul 24-3:38 PM

solubilitynotes

Assignment 6

1) At 25 degrees C, only 0.00245g of $BaSO_4$ can be dissolved in 1.0L of H_2O. Calculate the K_{sp} for $BaSO_4$.

$$\frac{0.00245g}{} \left| \frac{1mol}{233.4g} \right. = 1.0497 \times 10^{-5}mol$$

$$[BaSO_4] = \frac{1.0497 \times 10^{-5}mol}{1.0L} = 1.0497 \times 10^{-5}M$$

$$BaSO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$$
$$1.0497 \times 10^{-5}M \quad 1.0497 \times 10^{-5}M \quad 1.0497 \times 10^{-5}M$$
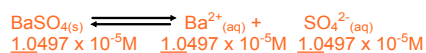
$$K_{sp} = [Ba^{2+}][SO_4^{2-}] = (1.0497 \times 10^{-5})^2 = 1.1 \times 10^{-10}$$

Jul 25-5:25 PM

Jul 25-5:25 PM

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2) At 25 degrees C, the solubility of Ag_3PO_4 is $1.8 \times 10^{-5}M$. Calculate the K_{sp} for Ag_3PO_4.

2) At 25 degrees C, the solubility of Ag_3PO_4 is $1.8 \times 10^{-5}M$. Calculate the K_{sp} for Ag_3PO_4.

$$Ag_3PO_{4(s)} \rightleftharpoons 3Ag^+_{(aq)} + PO_4^{3-}_{(aq)}$$
$$1.8 \times 10^{-5}M \quad 5.4 \times 10^{-5}M \quad 1.8 \times 10^{-5}M$$
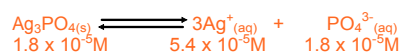
$$K_{sp} = [Ag^+]^3[PO_4^{3-}] = (5.4 \times 10^{-5})^3(1.8 \times 10^{-5}) = 2.8 \times 10^{-18}$$

Jul 25-5:27 PM

Jul 25-5:27 PM

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3) An experiment showed that a maximum of 1.49g of $AgBrO_3$ can dissolve in 1.00L of water at 25 degrees C. What is the K_{sp} for $AgBrO_3$ at this temperature?

3) An experiment showed that a maximum of 1.49g of $AgBrO_3$ can dissolve in 1.00L of water at 25 degrees C. What is the K_{sp} for $AgBrO_3$ at this temperature?

$$\frac{1.49g}{} \left| \frac{1mol}{235.8g} \right. = 6.319 \times 10^{-3}mol$$

$$[AgBrO_3] = \frac{6.319 \times 10^{-3}mol}{1.00L} = 6.319 \times 10^{-3}M$$

$$AgBrO_{3(s)} \rightleftharpoons Ag^+_{(aq)} + BrO_3^-_{(aq)}$$
$$6.319 \times 10^{-3}M \quad 6.319 \times 10^{-3}M \quad 6.319 \times 10^{-3}M$$
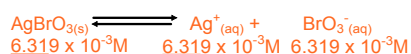
$$K_{sp} = [Ag^+][BrO_3^-] = (6.319 \times 10^{-3})^2 = 3.99 \times 10^{-5}$$

Jul 25-5:28 PM

Jul 25-5:28 PM

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?

Jul 25-5:29 PM

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?

$[CaF_2] = \dfrac{4.15 \times 10^{-4} mol}{2.0L} = 2.075 \times 10^{-4}M$

$CaF_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2F^-_{(aq)}$
$2.075 \times 10^{-4}M \quad 2.075 \times 10^{-4}M \quad 4.15 \times 10^{-4}M$
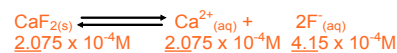
$K_{sp} = [Ca^{2+}][F^-]^2 = (2.075 \times 10^{-4})(4.15 \times 10^{-4})^2 = 3.6 \times 10^{-11}$

Jul 25-5:29 PM

5) A solution in equilibrium with solid Ag_2S on the bottom of the beaker was found to contain $1.6 \times 10^{-16}M$ S^{2-} and $2.5 \times 10^{-18}M$ Ag^+. Calculate the solubility product constant of Ag_2S.

Jul 25-5:30 PM

5) A solution in equilibrium with solid Ag_2S on the bottom of the beaker was found to contain $1.6 \times 10^{-16}M$ S^{2-} and $2.5 \times 10^{-18}M$ Ag^+. Calculate the solubility product constant of Ag_2S.

$Ag_2S_{(s)} \rightleftharpoons 2Ag^+_{(aq)} + S^{2-}_{(aq)}$
$2.5 \times 10^{-18}M \quad 1.6 \times 10^{-16}M$
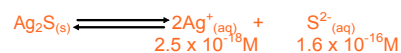
$K_{sp} = [Ag^+]^2[S^{2-}] = (2.5 \times 10^{-18})^2(1.6 \times 10^{-16}) = 1.0 \times 10^{-51}$

Jul 25-5:30 PM

IX) Type B K_{sp} Problems: Calculating Solubility

In Type B problems, you must calculate the solubility of a salt using the K_{sp} constant from p.5 of the data booklet.
Remember that **solubility** is the molarity at saturation.

Jul 25-6:02 PM

Example: Calculate the solubility of $CaCO_3$ in water at 25 degrees C.

Jul 25-6:04 PM

solubilitynotes

Example: Calculate the solubility of $CaCO_3$ in water at 25 degrees C.

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

$K_{sp} = [Ca^{2+}][CO_3^{2-}]$
K_{sp} from table $= 5.0 \times 10^{-9}$

Let $x = [Ca^{2+}]$
so x also equals $[CO_3^{2-}]$

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

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

$x = \underline{7.0}71 \times 10^{-5}$

so $[Ca^{2+}] = [CO_3^{2-}] = \underline{7.0}71 \times 10^{-5}$

$$CaCO_{3(s)} \rightleftharpoons Ca^{2+}_{(aq)} + CO_3^{2-}_{(aq)}$$
$\underline{7.0}71 \times 10^{-5}$ $\underline{7.0}71 \times 10^{-5}$M $\underline{7.0}71 \times 10^{-5}$M

Therefore, the solubility of $CaCO_3$, otherwise written as $[CaCO_3]$ equals 7.1×10^{-5}M

Jul 25-6:04 PM

Example: Calculate the solubility of PbI_2 in g/L.

Jul 25-6:05 PM

Example: Calculate the solubility of PbI_2 in g/L.

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

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

$K_{sp} = [Pb^{2+}][I^-]^2$
K_{sp} from table $= 8.5 \times 10^{-9}$

$8.5 \times 10^{-9} = (x)(2x)^2 = 4x^3$

$x = \sqrt[3]{\dfrac{8.5 \times 10^{-9}}{4}}$

$x = \underline{1.}28564 \times 10^{-3}$

$[Pb^{2+}] = \underline{1.}28564 \times 10^{-3}$M
$[I^-] = 2(\underline{1.}28564 \times 10^{-3}) = \underline{2.}57 \times 10^{-3}$M

$$PbI_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2I^-_{(aq)}$$
  x x $2x$
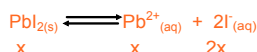
therefore, $[PbI_2] = \underline{1.}28564 \times 10^{-3}$M

$\dfrac{\underline{1.}28564 \times 10^{-3}\text{mol}}{} \Big| \dfrac{461.0\text{g}}{1\text{mol}}$

$= 0.59$ g/L

Jul 25-6:05 PM

Example: How many grams of $PbBr_2$ can be dissolved in 250.0mL of water at 25 degrees C?

Jul 26-2:25 PM

Example: How many grams of $PbBr_2$ can be dissolved in 250.0mL of water at 25 degrees C?

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

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

$K_{sp} = [Pb^{2+}][Br^-]^2$
K_{sp} from table $= 6.6 \times 10^{-6}$

$6.6 \times 10^{-6} = (x)(2x)^2 = 4x^3$

$x = \sqrt[3]{\dfrac{6.6 \times 10^{-6}}{4}}$

$x = \underline{1.}18 \times 10^{-2}$

$[Pb^{2+}] = \underline{1.}18 \times 10^{-2}$M
$[Br^-] = 2(\underline{1.}18 \times 10^{-2}) = \underline{2.}36 \times 10^{-2}$M

$$PbBr_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2Br^-_{(aq)}$$
  x x $2x$
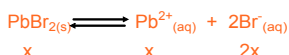
therefore, $[PbBr_2] = \underline{1.}18 \times 10^{-2}$M

$\dfrac{\underline{1.}18 \times 10^{-2}\text{mol}}{} \Big| \dfrac{367.0\text{g}}{1\text{mol}}$

$= \underline{4.}33$g/L

$\underline{4.}33$ divided by 4 = 1.1g $PbBr_2$

Jul 26-2:25 PM

Assignment 7

1) The K_{sp} at a certain temperature for $Ni(OH)_2$ is 1.6×10^{-16}. Calculate the solubility of $Ni(OH)_2$.

Jul 26-2:32 PM

solubilitynotes

Assignment 7

1) The K_{sp} at a certain temperature for $Ni(OH)_2$ is 1.6×10^{-16}. Calculate the solubility of $Ni(OH)_2$.

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

$[Ni^{2+}] = \underline{3.42} \times 10^{-6}M$

$[OH^-] = 2(3.42 \times 10^{-6}) = \underline{6.84} \times 10^{-6}M$

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

$$Ni(OH)_{2(s)} \rightleftharpoons Ni^{2+}_{(aq)} + 2OH^-_{(aq)}$$
$$\qquad\qquad x \qquad\qquad x \qquad\qquad 2x$$

$K_{sp} = [Ni^{2+}][OH^-]^2$
K_{sp} from table $= 1.6 \times 10^{-16}$

therefore, $[Ni(OH)_2] = 3.4 \times 10^{-6}M$

$1.6 \times 10^{-16} = (x)(2x)^2 = 4x^3$

$x = \sqrt[3]{\dfrac{1.6 \times 10^{-16}}{4}}$

$x = \underline{3.42} \times 10^{-6}$
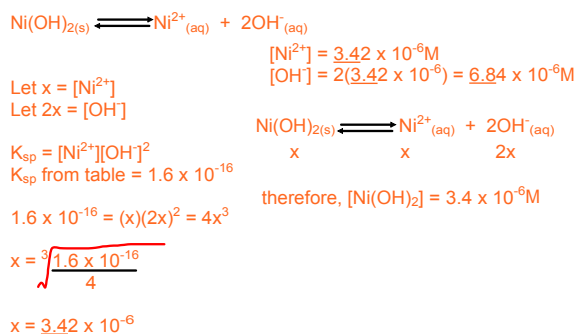
Jul 26-2:32 PM

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2) Find the solubility of $CaSO_4$ in g/L.

Jul 26-2:39 PM

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2) Find the solubility of $CaSO_4$ in g/L.

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

$K_{sp} = [Ca^{2+}][SO_4^{2-}]$
K_{sp} from table $= 7.1 \times 10^{-5}$

so $[Ca^{2+}] = [SO_4^{2-}] = \underline{8.426} \times 10^{-3}$

$$CaSO_{4(s)} \rightleftharpoons Ca^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$$
$$8.426 \times 10^{-3} \quad \underline{8.426} \times 10^{-3}M \quad \underline{8.426} \times 10^{-3}M$$

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

Therefore, $[CaSO_4] = \underline{8.426} \times 10^{-3}M$

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

$\dfrac{8.426 \times 10^{-3}mol}{} \;\Big|\; \dfrac{136.2g}{1\ mol}$

$x = \sqrt{7.1 \times 10^{-5}}$

$x = \underline{8.426} \times 10^{-3}$

$= 1.1\ g/L$
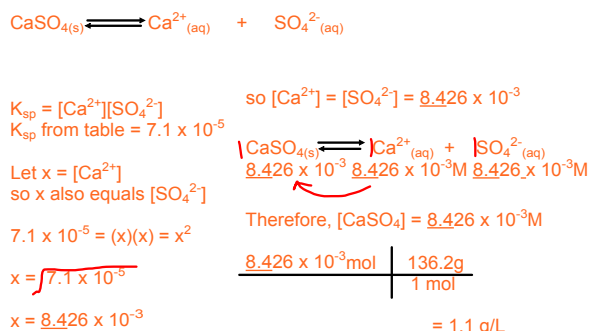
Jul 26-2:39 PM

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3) Which saturated solutions at 25 degrees C will have a greater $[Ag^+]$, AgCl or Ag_2CO_3?

Jul 26-2:45 PM

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3) Which saturated solutions at 25 degrees C will have a greater $[Ag^+]$, AgCl or Ag_2CO_3?

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

$K_{sp} = [Ag^+][Cl^-]$
K_{sp} from table $= 1.8 \times 10^{-10}$

Let $x = [Ag^+]$
so x also equals $[Cl^-]$
$1.8 \times 10^{-10} = (x)(x) = x^2$
$x = \sqrt{1.8 \times 10^{-10}}$
$x = \underline{1.34} \times 10^{-5}$
so $[Ag^+] = 1.3 \times 10^{-5}M$

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

$K_{sp} = [Ag^+]^2[CO_3^{2-}]$
K_{sp} from table $= 8.5 \times 10^{-12}$

Let $2x = [Ag^+]$
so $x = [CO_3^{2-}]$
$8.5 \times 10^{-12} = (2x)^2(x) = 4x^3$

$x = \underline{1.29} \times 10^{-4}$

so $[Ag^+] = 2(\underline{1.29} \times 10^{-4})$
$= 2.6 \times 10^{-4}M$

Ag_2CO_3 has a greater $[Ag^+]$
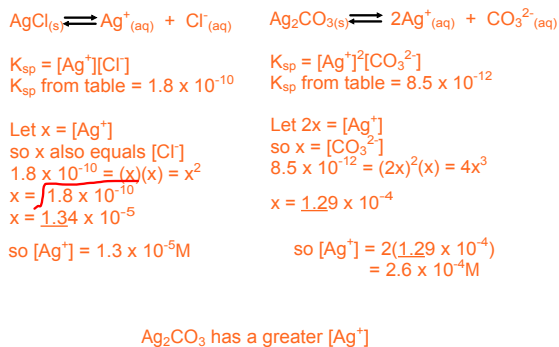
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4) Calculate the mass of $MgCO_3$ which could be dissolved in 3.0L of water at 25 degrees C.

Jul 26-2:56 PM

19

solubilitynotes

4) Calculate the mass of $MgCO_3$ which could be dissolved in 3.0L of water at 25 degrees C.

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

$K_{sp} = [Mg^{2+}][CO_3^{2-}]$
K_{sp} from table = 6.8×10^{-6}

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

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

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

$x = \underline{2.61} \times 10^{-3}$

so $[Mg^{2+}] = [CO_3^{2-}] = \underline{2.61} \times 10^{-3}$

$MgCO_{3(s)} \rightleftharpoons Mg^{2+}_{(aq)} + CO_3^{2-}_{(aq)}$
$\underline{2.61} \times 10^{-3}$ $\underline{2.61} \times 10^{-3}M$ $\underline{2.61} \times 10^{-3}M$

Therefore, $[MgCO_3] = \underline{2.61} \times 10^{-3}M$

$$\frac{\underline{2.61} \times 10^{-3} mol}{} \left| \frac{84.3g}{1 \, mol} \right.$$

$= 0.\underline{21}98$ g/L x 3

$= 0.66g$
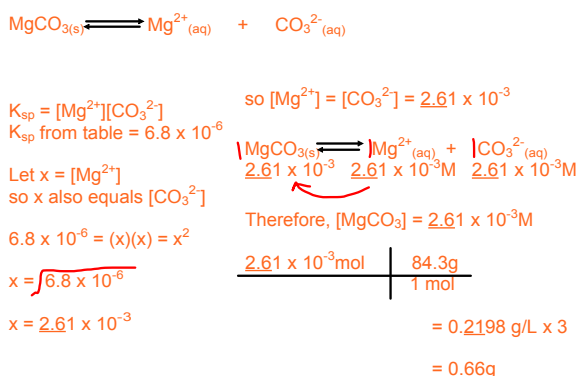
Jul 26-2:56 PM

X) Type C K_{sp} Problems: Trial K_{sp}

When two soluble solutions are mixed, a cation from one solution is introduced to an anion from another solution, and *visa versa*. One or both of the new ion combinations (cation and anion) could have low solubility. If this is the case, and there are too many of the low solubility ions present (more than what is necessary for a saturated solution), a precipitate will result. If there are not enough of the low solubility ions present to reach the saturation concentration, a precipitate will not form and the ions will stay dissolved in solution.

Jul 26-3:03 PM

A trial K_{sp} (same concept as trial K_{eq}) can be calculated for these mixtures to deduce whether a precipitate will form or not when the two solutions are mixed.
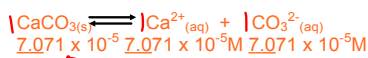If the trial K_{sp} is larger than the actual K_{sp}, the concentrations of low solubility ions is greater than saturation, so a precipitate will result. If the trial K_{sp} is smaller than the actual K_{sp}, the concentrations of low solubility ions has not yet reached saturation, therefore a precipitate will not form.

Jul 26-3:06 PM

Example: Will a precipitate form if 40.0mL of $8.0 \times 10^{-3}M$ $Mg(NO_3)_2$ is mixed with 60.0mL of $1.0 \times 10^{-2}M$ K_2CO_3?
$Mg(NO_3)_{2(aq)} + K_2CO_{3(aq)} \longrightarrow$
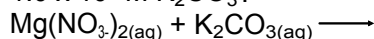
Jul 26-3:10 PM

Example: Will a precipitate form if 40.0mL of $8.0 \times 10^{-3}M$ $Mg(NO_3)_2$ is mixed with 60.0mL of $1.0 \times 10^{-2}M$ K_2CO_3?
$Mg(NO_3)_{2(aq)} + K_2CO_{3(aq)} \longrightarrow 2KNO_{3(aq)} + MgCO_{3(s)}$
possible ppt
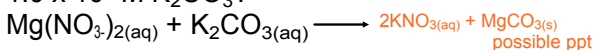
since two solutions are being mixed, dilution occurring:
$[Mg(NO_3)_2]_f = \frac{(8.0 \times 10^{-3})(0.0400)}{(0.1000)}$

$= 3.2 \times 10^{-3}M$

$[K_2CO_3]_f = \frac{(1.0 \times 10^{-2})(0.0600)}{(0.1000)}$

$= 6.0 \times 10^{-3}M$

$Mg(NO_3)_2 \rightleftharpoons Mg^{2+} + 2NO_3^-$
$3.2 \times 10^{-3}M$ $3.2 \times 10^{-3}M$

$K_2CO_3 \rightleftharpoons 2K^+ + CO_3^{2-}$
$6.0 \times 10^{-3}M$ $6.0 \times 10^{-3}M$

possible ppt is $MgCO_3$:
$MgCO_{3(s)} \rightleftharpoons Mg^{2+}_{(aq)} + CO_3^{2-}_{(aq)}$
$K_{sp} = [Mg^{2+}][CO_3^{2-}]$
K_{sp} from table = 6.8×10^{-6}
trial $K_{sp} = (3.2 \times 10^{-3})(6.0 \times 10^{-3})$
$= 1.9 \times 10^{-5}$

Trial $K_{sp} > K_{sp}$ therefore a ppt of $MgCO_3$ will form

Jul 26-3:10 PM

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

Jul 26-3:19 PM

solubilitynotes

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

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

since two solutions are being mixed, dilution occurring:

$[CaCl_2]_f = \dfrac{(0.00100)(0.0500)}{(0.1000)}$

$= 5.00 \times 10^{-4}M$

$[Na_2SO_4]_f = \dfrac{(1.00 \times 10^{-2})(0.0500)}{(0.1000)}$

$= 5.00 \times 10^{-3}M$

$CaCl_2 \rightleftharpoons Ca^{2+} + 2Cl^-$
$5.00 \times 10^{-4}M \quad 5.00 \times 10^{-4}M$

$Na_2SO_4 \rightleftharpoons 2Na^+ + SO_4^{2-}$
$5.00 \times 10^{-3}M \qquad 5.00 \times 10^{-3}M$

possible ppt is $CaSO_4$:
$CaSO_{4(s)} \rightleftharpoons Ca^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$
$K_{sp} = [Ca^{2+}][SO_4^{2-}]$
K_{sp} from table = 7.1×10^{-5}
trial $K_{sp} = (5.00 \times 10^{-4})(5.00 \times 10^{-3})$
$\qquad = 2.5 \times 10^{-6}$

Trial $K_{sp} < K_{sp}$ therefore a ppt of $CaSO_4$ will NOT form

Jul 26-3:19 PM

Assignment 8

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

Jul 26-3:26 PM

Assignment 8

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

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

since two solutions are being mixed, dilution occurring:

$[Zn(NO_3)_2]_f = \dfrac{(3.0 \times 10^{-10})(1.0)}{(2.0)}$

$= 1.5 \times 10^{-10}M$

$[Na_2SO_4]_f = \dfrac{(2.0 \times 10^{-11})(1.0)}{(2.0)}$

$= 1.0 \times 10^{-11}M$

$Zn(NO_3)_2 \rightleftharpoons Zn^{2+} + 2NO_3^-$
$5.00 \times 10^{-4}M \quad 5.00 \times 10^{-4}M$

$Na_2S \rightleftharpoons 2Na^+ + S^{2-}$
$5.00 \times 10^{-3}M \qquad 5.00 \times 10^{-3}M$

possible ppt is ZnS:
$ZnS_{(s)} \rightleftharpoons Zn^{2+}_{(aq)} + S^{2-}_{(aq)}$
$K_{sp} = [Zn^{2+}][S^{2-}]$
K_{sp} from table = 2.0×10^{-25}
trial $K_{sp} = (1.5 \times 10^{-10})(1.0 \times 10^{-11})$
$\qquad = 1.5 \times 10^{-21}$

Trial $K_{sp} > K_{sp}$ therefore a ppt of ZnS will form

Jul 26-3:26 PM

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

Jul 26-3:33 PM

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

$$SrS_{(aq)} + K_2SO_{4(aq)} \longrightarrow K_2S_{(aq)} + SrSO_{4(s)}$$
$$\text{possible ppt}$$

since two solutions are being mixed, dilution occurring:

$[SrS]_f = \dfrac{(6.7 \times 10^{-3})(2.0)}{(3.0)}$

$= 4.467 \times 10^{-3}M$

$[K_2SO_4]_f = \dfrac{(4.3 \times 10^{-4})(1.0)}{(3.0)}$

$= 1.433 \times 10^{-4}M$

$SrS \rightleftharpoons Sr^{2+} + S^{2-}$
$4.467 \times 10^{-3}M \quad 4.467 \times 10^{-3}M$

$K_2SO_4 \rightleftharpoons 2K^+ + SO_4^{2-}$
$1.433 \times 10^{-4}M \qquad 1.433 \times 10^{-4}M$

possible ppt is $SrSO_4$:
$SrSO_{4(s)} \rightleftharpoons Sr^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$
$K_{sp} = [Sr^{2+}][SO_4^{2-}]$
K_{sp} from table = 3.4×10^{-7}
trial $K_{sp} = (4.467 \times 10^{-3})(1.433 \times 10^{-4})$
$\qquad = 6.4 \times 10^{-7}$

Trial $K_{sp} > K_{sp}$ therefore a ppt of $SrSO_4$ will form

Jul 26-3:33 PM

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

Jul 26-3:42 PM

solubilitynotes

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?

$MgCl_{2(aq)} + Na_2CO_{3(aq)} \longrightarrow 2NaCl_{(aq)} + MgCO_{3(s)}$
possible ppt

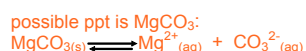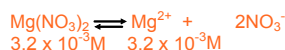since two solutions are being mixed, dilution occurring:

$[MgCl_2]_f = \dfrac{(5.0 \times 10^{-4})(1.0)}{(2.0)}$

$= 2.5 \times 10^{-4}$M

$[Na_2CO_3]_f = \dfrac{(2.0 \times 10^{-2})(1.0)}{(2.0)}$

$= 1.0 \times 10^{-2}$M

$MgCl_2 \rightleftharpoons Mg^{2+} + 2Cl^-$
2.5×10^{-4}M $\quad 2.5 \times 10^{-4}$M

$Na_2CO_3 \rightleftharpoons 2Na^+ + CO_3^{2-}$
1.0×10^{-2}M $\qquad\qquad 1.0 \times 10^{-2}$M

possible ppt is $MgCO_3$:
$MgCO_{3(s)} \rightleftharpoons Mg^{2+}_{(aq)} + CO_3^{2-}_{(aq)}$
$K_{sp} = [Mg^{2+}][CO_3^{2-}]$
K_{sp} from table $= 6.8 \times 10^{-6}$
trial $K_{sp} = (2.5 \times 10^{-4})(1.0 \times 10^{-2})$
$= 2.5 \times 10^{-6}$

Trial $K_{sp} < K_{sp}$ therefore a ppt of $CaSO_4$ will NOT form

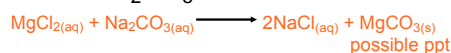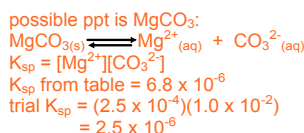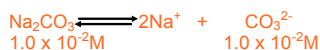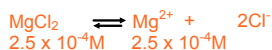Jul 26-3:42 PM

XI) Type D K_{sp} Problems: Determining the Concentration of a Specific Ion in a Saturated Solution

Recall that there are two ways to make a saturated solution: the direct method and the indirect method. Explain each using $Pb(OH)_2$ as an example.

Direct Method:

Indirect Method:

Jul 26-3:50 PM

XI) Type D K_{sp} Problems: Determining the Concentration of a Specific Ion in a Saturated Solution

Recall that there are two ways to make a saturated solution: the direct method and the indirect method. Explain each using $Pb(OH)_2$ as an example.

Direct Method: add $Pb(OH)_2$ salt to water and stir until there is excess $Pb(OH)_2$ solid on the bottom

Indirect Method:

mix two solutions - one that contains Pb^{2+} and one that contains OH^- with large enough concentrations to exceed the solubility of $Pb(OH)_2$. Then a ppt of $Pb(OH)_2$ forms and hence, a saturated solution

Jul 26-3:50 PM

It is important to recognize that ion stoichiometry for a saturated solution is only a reality when the saturated solution is made from dissolving the salt in water directly. What would be the proportion of $[OH^-]$ compared to $[Pb^{2+}]$ in this case?

If a saturated solution of a salt is made by combining two different solutions (indirectly), the $[Pb^{2+}]$ relative to the $[OH^-]$ could be anything, depending on the molarities of the solutions mixed. Therefore, you can't rely on 2 OH^- : 1 Pb^{2+} stoichiometry because the ions came from different solutions. However, the K_{sp} can still be relied upon, as you'll see in Type D problems.

Jul 26-3:54 PM

It is important to recognize that ion stoichiometry for a saturated solution is only a reality when the saturated solution is made from dissolving the salt in water directly. What would be the proportion of $[OH^-]$ compared to $[Pb^{2+}]$ in this case? 2 OH^- : 1 Pb^{2+}

If a saturated solution of a salt is made by combining two different solutions (indirectly), the $[Pb^{2+}]$ relative to the $[OH^-]$ could be anything, depending on the molarities of the solutions mixed. Therefore, you can't rely on 2 OH^- : 1 Pb^{2+} stoichiometry because the ions came from different solutions. However, the K_{sp} can still be relied upon, as you'll see in Type D problems.
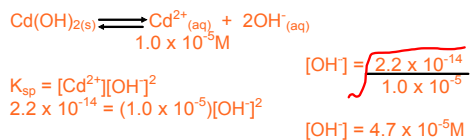
Jul 26-3:54 PM

Example: 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?

Jul 26-3:57 PM

solubilitynotes

Example: 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?

$K_2SO_{4(aq)} \longrightarrow 2K^+_{(aq)} + SO_4^{2-}_{(aq)}$
0.020M   0.020M

$SrSO_{4(s)} \rightleftharpoons Sr^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$
                    ?   0.020M

$K_{sp} = [Sr^{2+}][SO_4^{2-}]$
K_{sp} from table = 3.4×10^{-7}

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

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

Jul 26-3:57 PM

Example: If Cu^{2+} is slowly added to a solution of 0.020M KIO_3, at what $[Cu^{2+}]$ does a precipitate just start to form?

Jul 26-4:01 PM

Example: If Cu^{2+} is slowly added to a solution of 0.020M KIO_3, at what $[Cu^{2+}]$ does a precipitate just start to form?

$KIO_{3(aq)} \longrightarrow K^+_{(aq)} + IO_3^-_{(aq)}$
0.020M   0.020M

$Cu(IO_3)_{2(s)} \rightleftharpoons Cu^{2+}_{(aq)} + 2IO_3^-_{(aq)}$
                          ?   0.020M

$K_{sp} = [Cu^{2+}][IO_3^-]^2$
K_{sp} from table = 6.9×10^{-8}

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

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

Jul 26-4:01 PM

Example: $Fe(NO_3)_2$ 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?

Jul 26-4:05 PM

Example: $Fe(NO_3)_2$ 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?

$Fe(NO_3)_2 + Na_2S \longrightarrow 2NaNO_3 + FeS_{(s)}$
                                              possible ppt

$Na_2S \longrightarrow 2Na^+ + S^{2-}$
0.010M   0.010M $Fe(NO_3)_2 \longrightarrow Fe^{2+} + 2NO_3^-$
                                     $6.0 \times 10^{-17}M$ $6.0 \times 10^{-17}M$

$FeS_{(s)} \rightleftharpoons Fe^{2+}_{(aq)} + S^{2-}_{(aq)}$
                                     mol $Fe(NO_3)_2 = (6.0 \times 10^{-17}M)(2.0L)$
                                                $= 1.2 \times 10^{-16}mol$
$K_{sp} = [Fe^{2+}][S^{2-}]$
K_{sp} from table = 6.0×10^{-19}

$6.0 \times 10^{-19} = [Fe^{2+}][0.010]$ $\dfrac{1.2 \times 10^{-16}mol}{} \,\bigg|\, \dfrac{179.8g}{1mol}$

$[Fe^{2+}] = 6.0 \times 10^{-17}M$ $= 2.2 \times 10^{-14}g$

Jul 26-4:05 PM

Example: 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 $[Cd^{2+}]$ in the tank must be lowered to $1.0 \times 10^{-5}M$. What $[OH^-]$ would be required to bring the $[Cd^{2+}]$ to an acceptable level? $K_{sp} = 2.2 \times 10^{-14}$ for $Cd(OH)_2$

Jul 28-1:19 PM

solubilitynotes

Example: 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 $[Cd^{2+}]$ in the tank must 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 $Cd(OH)_2$

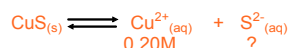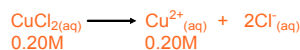$$Cd(OH)_{2(s)} \rightleftharpoons Cd^{2+}_{(aq)} + 2OH^-_{(aq)}$$
$$1.0 \times 10^{-5}M$$

$$K_{sp} = [Cd^{2+}][OH^-]^2$$
$$2.2 \times 10^{-14} = (1.0 \times 10^{-5})[OH^-]^2$$

$$[OH^-] = \sqrt{\frac{2.2 \times 10^{-14}}{1.0 \times 10^{-5}}}$$

$$[OH^-] = 4.7 \times 10^{-5}M$$

Jul 28-1:19 PM

Assignment 9

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

Jul 28-1:25 PM

Assignment 9

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

$$CuCl_{2(aq)} \longrightarrow Cu^{2+}_{(aq)} + 2Cl^-_{(aq)}$$
$$0.20M \qquad 0.20M$$

$$CuS_{(s)} \rightleftharpoons Cu^{2+}_{(aq)} + S^{2-}_{(aq)}$$
$$0.20M \qquad ?$$

$$K_{sp} = [Cu^{2+}][S^{2-}]$$
$$K_{sp} \text{ from table} = 6.0 \times 10^{-37}$$
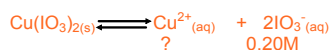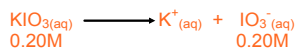
$$6.0 \times 10^{-37} = (0.20)[S^{2-}]$$

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

Jul 28-1:25 PM

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

Jul 28-1:32 PM

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

$$KIO_{3(aq)} \longrightarrow K^+_{(aq)} + IO_3^-_{(aq)}$$
$$0.20M \qquad 0.20M$$

$$Cu(IO_3)_{2(s)} \rightleftharpoons Cu^{2+}_{(aq)} + 2IO_3^-_{(aq)}$$
$$? \qquad 0.20M$$

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

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

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

Jul 28-1:32 PM

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

Jul 28-1:34 PM

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

$NaCl_{(aq)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)}$
0.10M   0.10M

$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$
                    ?   0.10M

$K_{sp} = [Ag^+][Cl^-]$
K_{sp} from table = 1.8×10^{-10}

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

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

Jul 28-1:34 PM

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

Jul 28-1:38 PM

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

$CaBr_{2(aq)} \longrightarrow Ca^{2+}_{(aq)} + 2Br^-_{(aq)}$
0.010M   0.010M

$CaSO_{4(s)} \rightleftharpoons Ca^{2+}_{(aq)} + SO_4^{2-}_{(aq)}$
                    0.010M

$K_{sp} = [Ca^{2+}][SO_4^{2-}]$
K_{sp} from table = 7.1×10^{-5}

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

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

Jul 28-1:38 PM

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

Jul 28-1:46 PM

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

$K_2CrO_{4(aq)} \longrightarrow 2K^+_{(aq)} + CrO_4^{2-}_{(aq)}$
5.0×10^{-4}M 5.0×10^{-4}M

$BaCrO_{4(s)} \rightleftharpoons Ba^{2+}_{(aq)} + CrO_4^{2-}_{(aq)}$
                    5.0×10^{-4}M

$K_{sp} = [Ba^{2+}][CrO_4^{2-}]$
K_{sp} from table = 1.2×10^{-10}

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

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

Jul 28-1:46 PM

XII) Hard Water

Hard water is caused by a high concentration of _____ and/or _____ ions in water. It's a problem in areas of the world rich in limestone ($CaCO_3$) and/or magnesite ($MgCO_3$). Water from rivers, lakes, and streams pass over limestone rock, and Ca^{2+} ends up in water by two processes.

Hard water is not harmful to health, but can cause pipes to clog and hinders soap from creating suds

Jul 28-1:50 PM

25

solubilitynotes

XII) Hard Water

Hard water is caused by a high concentration of __Ca²⁺__ and/or __Mg²⁺__ ions in water. It's a problem in areas of the world rich in limestone ($CaCO_3$) and/or dolomite ($CaMg(CO_3)_2$). Water from rivers, lakes, and streams pass over limestone rock, and Ca^{2+} ends up in water by two processes.

Hard water is not harmful to health, but can cause pipes to clog and hinders soap from creating suds:

$$Ca^{2+} + NaC_{17}H_{35}COO \rightarrow Ca(C_{17}H_{35}COO)_2 + 2Na^+$$
$$\text{soap} \qquad \text{soap scum}$$
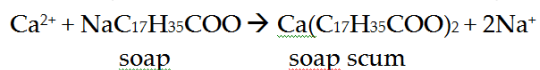
Jul 28-1:50 PM

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1) Permanent Hard Water is formed by acidic water (water containing H^+) due to acid rain. It is deemed 'permanent' because the reaction is not reversible:

$$CaCO_{3(s)} + 2H^+_{(aq)} \longrightarrow Ca^{2+}_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$
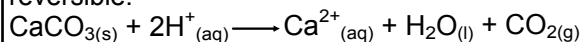
How do you 'soften' permanently hard water?

Jul 28-1:57 PM

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1) Permanent Hard Water is formed by acidic water (water containing H^+) due to acid rain. It is deemed 'permanent' because the reaction is not reversible:

$$CaCO_{3(s)} + 2H^+_{(aq)} \longrightarrow Ca^{2+}_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$
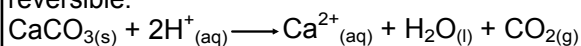
How do you 'soften' permanently hard water?

add an anion such as CO_3^{2-} that is low solubility with Ca^{2+}, hence precipitating it out of the water

Jul 28-1:57 PM

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2) Temporary Hard Water is due to the CO_2 dissolved in regular rainwater. It is deemed 'temporary' because it's a reversible reaction.
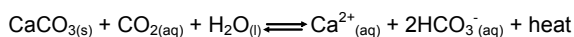
$$CaCO_{3(s)} + CO_{2(aq)} + H_2O_{(l)} \rightleftharpoons Ca^{2+}_{(aq)} + 2HCO_3^-_{(aq)} + heat$$

How do you soften temporary hard water?

http://www.absorblearning.com/media/attachment.action?quick=v1&att=2224
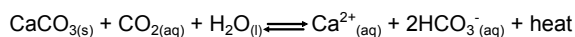
Jul 28-1:58 PM

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2) Temporary Hard Water is due to the CO_2 dissolved in regular rainwater. It is deemed 'temporary' because it's a reversible reaction.

$$CaCO_{3(s)} + CO_{2(aq)} + H_2O_{(l)} \rightleftharpoons Ca^{2+}_{(aq)} + 2HCO_3^-_{(aq)} + heat$$

How do you soften temporary hard water?

Increase the temperature of the water in order to cause a shift to the right (endothermic direction), thereby causing the $Ca^{2+}_{(aq)}$ to precipitate out of solution as $CaCO_{3(s)}$

http://www.absorblearning.com/media/attachment.action?quick=v1&att=2224

Jul 28-1:58 PM

Assignment 10

Read Hebden pages 103 & 104 and do questions 76-80 on page 104

Jul 28-2:09 PM

**solubilitynotes**

Assignment 10

Read Hebden pages 103 & 104 and do questions 76-80 on page 104

answers in the back of Hebden

Jul 28-2:09 PM

XIII) Altering the Solubility of a Salt

Saturated solution equations (solid on left, ions on the right) are simply equilibrium equations. In the equilibrium unit, we studied how concentrations changes can cause a shift

$$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$$
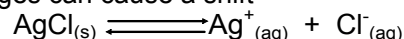
What shift would occur if $[Ag^+]$ is increased?

How would the amount of $AgCl_{(s)}$ change?

Since more $AgCl_{(s)}$ is being produced, we can say that the 'solubility of AgCl' has _____.

Jul 28-2:15 PM

XIII) Altering the Solubility of a Salt

Saturated solution equations (solid on left, ions on the right) are simply equilibrium equations. In the equilibrium unit, we studied how concentrations changes can cause a shift

$$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$$
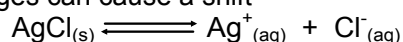
What shift would occur if $[Ag^+]$ is increased?
shift left
How would the amount of $AgCl_{(s)}$ change?
increase
Since more $AgCl_{(s)}$ is being produced, we can say that the 'solubility of AgCl' has __decreased__.

Jul 28-2:15 PM

By adding more Ag^+ to increase $[Ag^+]$ or more Cl^- to increase $[Cl^-]$, the equilibrium shifts _____, resulting in more $AgCl_{(s)}$.
This is called the _____,
because an ion that is already part of the equilibrium (common to the equilibrium) is being added. Keep in mind that Ag^+ would be added with a spectator anion such as NO_3^-, and Cl^- would be added with a spectator cation, such as Na^+.

Jul 28-2:22 PM

By adding more Ag^+ to increase $[Ag^+]$ or more Cl^- to increase $[Cl^-]$, the equilibrium shifts _left_, resulting in more $AgCl_{(s)}$.
This is called the _common ion effect_____,
because an ion that is already part of the equilibrium (common to the equilibrium) is being added. Keep in mind that Ag^+ would be added with a spectator anion such as NO_3^-, and Cl^- would be added with a spectator cation, such as Na^+.

Jul 28-2:22 PM

What kind of concentration change would have to occur to cause a shift right?

Thus, you must add an ion that will be low solubility, and hence precipitate out either Ag^+ or Cl^-.
What ion could you add (with applicable spectator) to precipitate out Ag^+?

What ion could you add (with applicable spectator) to precipitate out Cl^-?

Either of these methods would cause a shift right, causing more $AgCl_{(s)}$ to dissolve, thereby _____ the solubility of AgCl.

Jul 28-2:25 PM

solubilitynotes

What kind of concentration change would have to occur to cause a shift right?

decrease in $[Ag^+]$ or $[Cl^-]$

Thus, you must add an ion that will be low solubility, and hence precipitate out either Ag^+ or Cl^-.

What ion could you add (with applicable spectator) to precipitate out Ag^+?

OH as NaOH

What ion could you add (with applicable spectator) to precipitate out Cl^-?

Pb^{2+} as $Pb(NO_3)_2$

Either of these methods would cause a shift right, causing more $AgCl_{(s)}$ to dissolve, thereby _____increasing_____ the solubility of AgCl.

Jul 28-2:25 PM

Example: In which of the following 0.10M solutions would CaC_2O_4 be least soluble in and most soluble in: NaOH, KCl, $Ca(NO_3)_2$? Start by writing an equilibrium reaction for calcium oxalate.

Jul 28-2:29 PM

Example: In which of the following 0.10M solutions would CaC_2O_4 be least soluble in and most soluble in: NaOH, KCl, $Ca(NO_3)_2$? Start by writing an equilibrium reaction for calcium oxalate.

$$CaC_2O_{4(s)} \rightleftharpoons Ca^{2+}_{(aq)} + C_2O_{4(aq)}$$
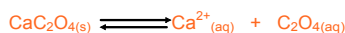
NaOH
OH^- is low solubility with Ca^{2+}, so it will ppt Ca^{2+} as $Ca(OH)_{2(s)}$ Thus, $[Ca^{2+}]$ decreases, causing a shift right so the solubility of CaC_2O_4 increases

MOST SOLUBLE

KCl
K^+ and Cl^- will not affect the equilibrium

$Ca(NO_3)_2$
Ca^{2+} is a common ion, so $[Ca^{2+}]$ increases, causing a shift left, creating more $CaC_2O_{4(s)}$ so the solubility of CaC_2O_4 decreases

LEAST SOLUBLE

Jul 28-2:29 PM

Example: In which of the following 0.10M solutions will $PbCl_2$ be most soluble?
HCl, $MgCl_2$, $AgNO_3$, NH_4NO_3

Jul 28-2:36 PM

Example: In which of the following 0.10M solutions will $PbCl_2$ be least soluble?
HCl, $MgCl_2$, $AgNO_3$, NH_4NO_3

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

HCl
0.10M Cl^- is a common ion so a shift right occurs therefore the solubility of $PbCl_2$ decreases

$MgCl_2$
0.20M Cl^- is a common ion so a shift right occurs, more so than HCl due to the larger M of Cl^-, therefore the solubility of $PbCl_2$ decreases more than HCl

$AgNO_3$
Ag^+ ppts with Cl^- causing $[Cl^-]$ to decrease, thus a shift right, so solubility of $PbCl_2$ increases

NH_4NO_3
neither NH_4^+ or NO_3^- affect the equilibrium

Jul 28-2:36 PM

Assignment 11

1) 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

Jul 28-2:43 PM

solubilitynotes

Assignment 11

1) 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

Jul 28-2:43 PM

2) 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

Jul 28-2:44 PM

3) Do Hebden page 108, numbers 81, 82, 84-86

Jul 28-2:45 PM