
# Unit 3: Solubility Equilibrium

Chem 11 Review
Preparation for Solubility Equilibrium

## Chem 11 Review
## Preparation for Solubility Equilibrium

**CHEM 11 Review**

It is expected that the student understands the concept of:
1. Strong electrolytes,
2. Weak electrolytes and
3. Nonelectrolytes.

CHEM 0011 – Unit 9 - Solubility of Ionic Compounds

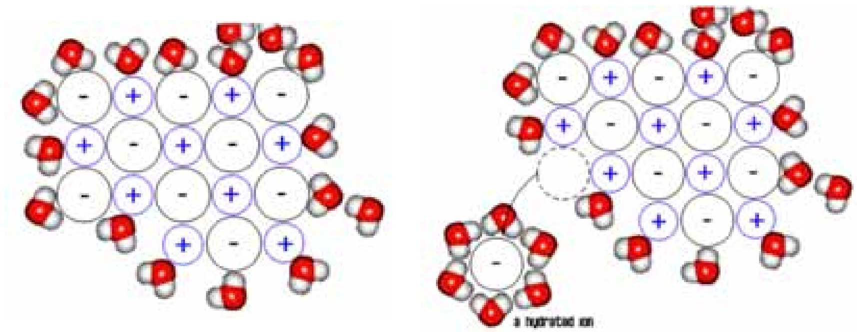
http://nobel.scas.bcit.ca/courses/wpmu/chem0011/unit-9/

## Chem 11 Review
## Preparation for Solubility Equilibrium

### CHEM 11 Review

The dissolution process.

http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/thermochem/solutionSalt.html






## Chem 11 Review
## Preparation for Solubility Equilibrium

### CHEM 11 Review

**1. Strong Electrolyte – Sodium chloride, NaCl**

- Forms from metal and nonmetal
  - eg: $NaCl\ (s) \rightarrow Na^+\ (aq) + Cl^-\ (aq)$
  - eg: $CaCl_2\ (s) \rightarrow Ca^{2+}\ (aq) + 2\ Cl^-\ (aq)$
  - eg: $NH_4NO_3\ (s) \rightarrow NH_4^+\ (aq) + NO^-_3\ (aq)$
- Totally ionizes in water.
- Forms electrolytic solutions (conducts electricity) with charged ions.
- Has ionic bonds.
- Commonly called “salts” which are crystalline solids (high melting points and boiling points) and dissolves well.

# Chem 11 Review
# Preparation for Solubility Equilibrium

## CHEM 11 Review

**2. Weak Electrolyte – Hydrofluoric acid, HF**

**Molecular Compounds – are weak or non-electrolytes when dissolved in water**

- Forms from nonmetal with nonmetal
  eg: HF (aq) ⇌ $H^+$ (aq) + $F^-$ (aq)
  eg: $H_2O$ (aq) ⇌ $H^+$ (aq) + $OH^-$ (aq)
- Partially ionizes in water.
- HF is a weak electrolyte (weakly conducts electricity).
- Has covalent bonds.
- Commonly called "molecules" and can be solid, liquids or gas.



# Chem 11 Review
# Preparation for Solubility Equilibrium

## CHEM 11 Review

**3. Nonelectrolyte – Methanol, $CH_3OH$**

- Does not ionize in water.
- $CH_3OH$ does not conduct electricity.
- Has covalent bonds.
- Upon dissolving, the methanol molecules remain intact.

http://highered.mcgraw-hill.com/olcweb/cgi/pluginpop.cgi?it=swf::525::530::/sites/dl/free/0072857684/322633/07_stg_wk_nonelelytes.swf::Strong Electrolyte, Weak Electrolyte, and Nonelectrolyte

## Chem 11 Review
## Preparation for Solubility Equilibrium

### CHEM 11 Review

State whether the following substance is expected to form ionic or molecular solutions.

$CuSO_4$ (s) ionic solution

$CH_4$ (g) molecular solution

$CrCl_3$ (s) ionic solution

$NaCH_3COO$ (s) ionic solution

**In this course, we are focus on the IONIC compounds.**



## Chem 11 Review
## Preparation for Solubility Equilibrium

### CHEM 11 Review

**Solubility of a substance:**

Describes the amount of the substance that will dissolve in a solvent like water at a given temperature.

A solution which has dissolved the **MAXIMUM** amount of substance at a given temperature is said to be a **SATURATED** solution.

## Chem 11 Review
## Preparation for Solubility Equilibrium

### CHEM 11 Review

It is expected that the student understands the concept of molarity and solution composition.

**Eg: What is the concentration of all the ions present in a saturated solution of $Ag_2CO_3$ having a concentration of $1.2 \times 10^{-4}$ M? Calculate the concentration of all the ions in the solution.**

When $Ag_2CO_3$ dissolves, the ions in solution are $Ag^+$ (aq) and $CO_3^{2-}$ (aq). The ratio of the ions is 2:1 $Ag^+$ : $CO_3^{2-}$.

$$Ag_2CO_3\,(s) \rightarrow 2\,Ag^+\,(aq) + CO_3^{2-}\,(aq)$$

$[Ag^+]$ = **$2.4 \times 10^{-4}$ M** $[CO_3^{2-}]$ = **$1.2 \times 10^{-4}$ M**

CHEM 0012 Lecture Notes 9

## Chem 11 Review
## Preparation for Solubility Equilibrium

### CHEM 11 Review

It is expected that the student understands the concept of dilution and how to calculate the dilution occurring when solutions are mixed.

CHEM 0011 – Unit 9 - Preparation of Solutions by Dilution
http://nobel.scas.bcit.ca/wiki/index.php/Solution_dilution

$M_1V_1 = M_2V_2$

where $M_1$ = concentration of the original solution
$V_1$ = volume of the original solution

where $M_2$ = concentration of the diluted solution
$V_2$ = volume of the diluted solution

Note: that the units on $M_1$ and $M_2$ must be the same and the units on $V_1$ and $V_2$ must be the same

CHEM 0012 Lecture Notes 10

# Chem 11 Review
# Preparation for Solubility Equilibrium

## CHEM 11 Review

Dilution Problem

$$M_1V_1 = M_2V_2$$

**If 5.0 mL of 0.020 M $Cl^-$ is added to 15.0 mL of 0.012 M $Br^-$, what is the molarity of $Cl^-$ and $Br^-$ ions?**

**Consider the $Cl^-$ ions:**

$M_1 = [Cl^-]_{final} = ?$
$M_2 = [Cl^-]_{original} = 0.020$ M
$V_1$ = Total volume of final solution = 20.0 mL
$V_2$ = Volume of original $Cl^-$ solution = 5.0 mL

$[Cl^-]_{final} = M_2V_2/V_1$ = (0.020 M)(5.0 mL) / 20.0 mL = 0.0050 M



# Chem 11 Review
# Preparation for Solubility Equilibrium

## CHEM 11 Review

Dilution Problem

$$M_1V_1 = M_2V_2$$

**If 5.0 mL of 0.020 M $Cl^-$ is added to 15.0 mL of 0.012 M $Br^-$, what is the molarity of $Cl^-$ and $Br^-$ ions?**

**Consider the $Br^-$ ions:**

$M_1 = [Br^-]_{final} = ?$
$M_2 = [Br^-]_{original} = 0.012$ M
$V_1$ = Total volume of final solution = 20.0 mL
$V_2$ = Volume of original $Br^-$ solution = 15.0 mL

$[Br^-]_{final} = M_2V_2/V_1$ = (0.012 M)(15.0 mL) / 20.0 mL = 0.0090 M

## Chem 11 Review
## Preparation for Solubility Equilibrium

### CHEM 11 Review

It is expected that the student understands the concept of writing chemically correct chemical formula when given cations and anions.

Chem 0011 – Unit 5 – Writing Chemical Formulae
http://nobel.scas.bcit.ca/wiki/index.php/Writing_chemical_formulae

Eg: Write the chemical formula for the compound formed between:

1. Na and Br: $NaBr$
2. K and I: $KI$
3. Ba and F: $BaF_2$
4. $K^+$ and nitrate ion: $KNO_3$
5. $Ca^{2+}$ and acetate ion: $Ca(C_2H_3O_2)_2$
6. ammonium ion and $Cl^-$ ion: $NH_4Cl$
7. $Fe^{3+}$ and carbonate ion: $Fe_2(CO_3)_3$
8. $Mg^{2+}$ and phosphate ion: $Mg_3(PO_4)_2$



## Chem 11 Review
## Preparation for Solubility Equilibrium

### CHEM 11 Review

It is expected that the student knows how to balance chemical equations.

Chem 0011 – Balancing Chemical Equations

http://nobel.scas.bcit.ca/wiki/index.php/Balancing_chemical_equations

## Chem 11 Review
## Preparation for Solubility Equilibrium

**CHEM 11 Review**

It is expected that the student understands calculations involving limiting reagent and solve stoichiometry-related reactions.

If 55.0 grams of $N_2$ (g) is reacted with 55.0 g $H_2$ (g), how many grams of ammonia gas, $NH_3$ (g), is formed?

Write the balanced chemical equation:

$$N_2\text{ (g)} + 3\ H_2\text{ (g)} \rightarrow 2\ NH_3\text{ (g)}$$

Convince yourself that:

1. $N_2$ (g) is the limiting reagent
2. $H_2$ (g) is in excess
3. The answer is 66.9 g $NH_3$ (g).



## Chem 11 Review
## Preparation for Solubility Equilibrium

**CHEM 11 Review**

It is expected that the student understands the concept of predicting solubility using the solubility table.

On a test or exam, a solubility table will be provided. The student is expected to know how to use it.

**Low solubility – does not dissolve in water.**

Predict whether the following is soluble in water:

| | |
|---|---|
| $PbI_2$ | low solubility |
| $MgSO_4$ | soluble |
| $CuCO_3$ | low solubility |
| LiCl | soluble |

## Chem 11 Review
## Preparation for Solubility Equilibrium

**CHEM 11 Review**

It is expected that the student understands the concept of predicting solubility using the solubility table.

Eg: Use the solubility table to predict whether the following double replacement reaction will occur? In other words, what happens when solutions of $Pb(NO_3)_2$ (aq) and KI (aq) are mixed? If a reaction occurs, write the balanced chemical equation.

Answer:

$$Pb(NO_3)_2\ (aq) + 2KI\ (aq) \rightarrow PbI_2\ (s) + 2KNO_3\ (aq)$$



## Chem 11 Review
## Preparation for Solubility Equilibrium

**CHEM 11 Review**

It is expected that the student understands the concept of predicting solubility using the solubility table.

Putting it all together!

Eg: What happens when 5.34 g of calcium chloride and 9.85 g of aluminum chloride are dissolved in 4.50 L of solution?

(a) Write the reaction if it occurs.
(b) Find the concentrations of all the ions in solution.

## Chem 11 Review
## Preparation for Solubility Equilibrium

### CHEM 11 Review

Eg: What happens when 5.34 g of calcium chloride and 9.85 g of aluminum chloride are dissolved in 4.50 L of solution?

(a) Write the reaction if it occurs.
(b) Find the concentrations of all the ions in solution.

Answer:
(a) No reaction occurs. All ions stay in solution.
(b) FW ($CaCl_2$) = 111 g/mole; FW($AlCl_3$) = 133.36 g/mole
$CaCl_2$ : 5.34 g/ 111 g/mole = 0.0481 moles $CaCl_2$ produces 0.0481 moles $Ca^{2+}$ ions
= 2(0.0481) moles of $Cl^-$ ions.
$AlCl_3$ : 9.85 g/ 133.36 g/mole = 0.0739 moles $AlCl_3$ produces 0.0739 moles $Al^{3+}$ ions.
= 3(0.0739) moles $Cl^-$ ions.

$[Ca^{2+}]$ = 0.0481 moles / 4.50 L = 0.0107 M
$[Al^{3+}]$ = 0.0739 moles / 4.50 L = 0.0164 M
$[Cl^-]$ = [2(0.0481) + 3 (0.0739)] / 4.50 L = 0.0706 M



## Chem 11 Review
## Preparation for Solubility Equilibrium

### CHEM 11 Review

(a) What happens when NaCl (aq) and $AgNO_3$ (aq) solutions are mixed?
(b) Find the concentrations of all the ions in solution when 2.72 g of NaCl and 3.75 g $AgNO_3$ are dissolved in 5.20 L solution.
(c) Write the balanced chemical equation.
(d) Write the complete ionic equation.
(e) Write the net ionic equation.
(f) Identify the spectator ions.

Answers:
(a) and (c) NaCl (aq) + $AgNO_3$ (aq) $\rightarrow$ $NaNO_3$ (aq) + AgCl (s)
(b) $[Na^+]$ = 0.00894 M, $[Cl^-]$ = 0.0047 M, $[Ag^+] \approx 0$ M, $[NO_3^-]$ = 0.00424 M
(d) $Na^+$ (aq) + $Cl^-$ (aq) + $Ag^+$ (aq) + $NO_3^-$ (aq) $\rightarrow$ $Na^+$ (aq) + $NO_3^-$ (aq) + AgCl (s)
(e) $Ag^+$ (aq) + $Cl^-$ (aq) $\rightarrow$ AgCl (s)
(f) $Na^+$ (aq) and $NO_3^-$ (aq) are spectator ions.


# Solubility Equilibrium

(Hebden Unit 3 – page 73-108)

## Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

**Solubility Equilibrium**

We will cover the following topics:

1. Define the term "solubility".
2. What is a saturated solution?
3. What is solubility equilibrium?
4. What is the meaning of $K_{sp}$ and what is the solubility product constant expression?
5. Calculate $K_{sp}$ from solubility data.
6. Compute molar solubility from $K_{sp}$.

## Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

1. Define the term "solubility".

Solubility is a measure of how much a solute can dissolve in a solvent at a given temperature.

In this course, we will focus on water as the solvent.



## Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

2. What is a saturated solution?

A solution which has dissolved the **MAXIMUM** amount of substance at a given temperature is said to be a **SATURATED** solution.

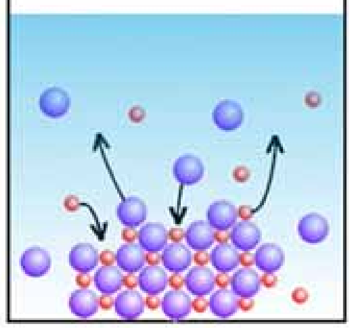
## Dynamic Equilibrium

Hebden – Unit 2 (page 37-69)

3. What is solubility equilibrium?

$$NaCl\ (s) \rightleftharpoons Na^{+}\ (aq) + Cl^{-}\ (aq)$$

A saturated sodium chloride solution



A saturated solution has an established equilibrium between dissolved and un-dissolved particles.

Rate of dissolution = Rate of crystallization



## Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

**Solubility Equilibrium**

Compounds of **low solubility** form **EQUILIBRIA** involving the solid and its aqueous ions.

For a general compound $A_xB_y$ (s) which is slightly soluble in water, the equilibrium between the solid and its soluble ions is established.

$$A_xB_y\ (s) \rightleftharpoons x\ A^{y+}\ (aq) + y\ B^{x-}\ (aq)$$

$$^{*}K_{eq} = [A^{y+}]^{x}\ [B^{x-}]^{y}$$

* $[A_xB_y]$ is not included in the equilibrium expression because $A_xB_y$ is a solid.

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

**Solubility Equilibrium**

For a general equilibrium

$$A_xB_y\text{ (s)} \rightleftharpoons x\,A^{y+}\text{ (aq)} + y\,B^{x-}\text{ (aq)}$$

where $A_xB_y$ (s) is a compound with low solubility.

$$K_{eq} = [A^{y+}]^x\,[B^{x-}]^y$$

Since the $K_{eq}$ is the product of the soluble ions, a special name is given to the $K_{eq}$, called

$K_{sp}$, Solubility Product Equilibrium Constant



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

4. What is the meaning of $K_{sp}$ and what is the solubility product constant expression?

Look at the Solubility Product Constant Reference Sheet

SOLUBILITY PRODUCT CONSTANTS AT 25°C

| Name | Formula | $K_{sp}$ |
|---|---|---|
| Silver bromide | AgBr | $5.4 \times 10^{-13}$ |
| Silver chloride | AgCl | $1.8 \times 10^{-10}$ |

What do these numbers mean?

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

**Solubility Product Constant Reference Sheet**

SOLUBILITY PRODUCT CONSTANTS AT 25°C

| Name | Formula | $K_{sp}$ |
|---|---|---|
| Silver bromide | AgBr | $5.4 \times 10^{-13}$ |
| Silver chloride | AgCl | $1.8 \times 10^{-10}$ |

The solubility constant equilibrium is:

$AgBr\ (s) \rightleftharpoons Ag^+\ (aq) + Br^-\ (aq)$ $K_{sp} = [Ag^+][Br^-]$

$AgCl\ (s) \rightleftharpoons Ag^+\ (aq) + Cl^-\ (aq)$ $K_{sp} = [Ag^+][Cl^-]$

$K_{sp}$ is the product of molar solubility of the ions in a saturated solution. This is the solubility product constant expression.



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

**Solubility Product Constant Reference Sheet**

SOLUBILITY PRODUCT CONSTANTS AT 25°C

| Name | Formula | $K_{sp}$ |
|---|---|---|
| Silver bromide | AgBr | $5.4 \times 10^{-13}$ |
| Silver chloride | AgCl | $1.8 \times 10^{-10}$ |

At 25°C, $K_{sp}$ (AgCl) > $K_{sp}$ (AgBr)

$[Ag^+][Cl^-] > [Ag^+][Br^-]$

This means that AgCl is more soluble than AgBr.

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

**Solubility Product Constant Reference Sheet**

Look up $K_{sp}$ for strontium fluoride, $SrF_2$. Write the solubility product constant expression for $SrF_2$.



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

**Solubility Product Constant Reference Sheet**

SOLUBILITY PRODUCT CONSTANTS AT 25°C

| Name | Formula | $K_{sp}$ |
|---|---|---|
| Strontium fluoride | $SrF_2$ | $4.3 \times 10^{-9}$ |

The solubility constant equilibrium is:

$$SrF_2\ (s) \rightleftharpoons Sr^{+2}\ (aq) + 2\ F^-\ (aq)$$

This is the solubility product constant expression.

$$K_{sp} = [Sr^{+2}][F^-]^2 = 4.3x10^{-9}$$

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## 5. Calculate $K_{sp}$ from solubility data.

One way to determine the value of $K_{sp}$ for a salt is to measure its **solubility** (i.e. How much of the salt is required to produce a saturated solution.) Solubility is expressed in grams/L.

**Molar solubility**: is the number of moles of salt dissolved in one liter of its saturated solution.

That is, the concentration of a dissolved solid present in a saturated solution, expressed in molarity. (moles/L)



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## 5. Calculate $K_{sp}$ from solubility data.

The solubility of AgBr in water was determined to be $1.38 \times 10^{-4}$ g/L at 25°C. Calculate the $K_{sp}$ for AgBr at that temperature.

Step 1: Convert $1.38 \times 10^{-4}$ g/L to moles/L to get molar solubility of AgBr.

$$1.3 \text{ x } 10^{-4} \text{ g/L} \cdot \frac{1}{187.77 \text{ g/mole AgBr}} = 7.34\text{x}10^{-7} \text{ moles/L}$$

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

5. Calculate $K_{sp}$ from solubility data.

The solubility of AgBr in water was determined to be 1.38 x $10^{-4}$ g/L at 25°C. Calculate the $K_{sp}$ for AgBr at that temperature.

Step 2: Set up ICE table.

Initially no ions in the solution.

| | AgBr (s) ⇌ | $Ag^+$ (aq) + | $Br^-$ (aq) |
|---|---|---|---|
| [I] | - | 0 | 0 |
| [C] | | 7.34 x $10^{-7}$ | 7.34 x $10^{-7}$ |
| [E] | | 7.34 x $10^{-7}$ | 7.34 x $10^{-7}$ |

Stoichio-metric ratio



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

5. Calculate $K_{sp}$ from solubility data.

Example 1: The solubility of AgBr in water was determined to be 1.38 x $10^{-4}$ g/L at 25°C. Calculate the $K_{sp}$ for AgBr at that temperature.

Step 3: Write the $K_{sp}$ expression.

$$K_{sp} = [Ag^+][Br^-] = (7.34 \times 10^{-7})(7.34 \times 10^{-7}) = 5.39 \times 10^{-13}$$

Look up the $K_{sp}$ of AgBr in the Solubility Product Constant Reference Sheet

SOLUBILITY PRODUCT CONSTANTS AT 25°C

| Name | Formula | $K_{sp}$ |
|---|---|---|
| Silver bromide | AgBr | 5.4 x $10^{-13}$ |

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## 5. Calculate $K_{sp}$ from solubility data.

Example 2: The molar solubility of silver chromate, $Ag_2CrO_4$, in water is $6.7 \times 10^{-5}$ moles/L at 25°C. What is the $K_{sp}$ of $Ag_2CrO_4$ at 25°C?

| | $Ag_2CrO_4$ (s) $\rightleftharpoons$ | 2 $Ag^+$ (aq) | + $CrO_4^{-2}$ (aq) |
|---|---|---|---|
| [I] | - | 0 | 0 |
| [C] | | $2(6.7 \times 10^{-5})$ * | $6.7 \times 10^{-5}$ |
| [E] | | $1.3 \times 10^{-4}$ | $6.7 \times 10^{-5}$ |

Stoichio-metric ratio

* For each mole of $Ag_2CrO_4$ dissolved, there are 2 moles $Ag^+$ and 1 mole $CrO_4^{-2}$.



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## 5. Calculate $K_{sp}$ from solubility data.

Example 2: The molar solubility of silver chromate, $Ag_2CrO_4$, in water is $6.7 \times 10^{-5}$ moles/L at 25°C. What is the $K_{sp}$ of $Ag_2CrO_4$ at 25°C?

Write the Ksp expression.

$$K_{sp} = [Ag^+]^2[CrO_4^{-2}] = (1.3 \times 10^{-4})^2 (6.7 \times 10^{-5})$$
$$= 1.1 \times 10^{-12}$$

SOLUBILITY PRODUCT CONSTANTS AT 25°C

| Name | Formula | $K_{sp}$ |
|---|---|---|
| Silver chromate | $Ag_2CrO_4$ | $1.1 \times 10^{-12}$ |

Look up the $K_{sp}$ of $Ag_2CrO_4$ in the Solubility Product Constant Reference Sheet

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## 6. Compute molar solubility from $K_{sp}$.

Example 1: What is the molar solubility of AgCl in pure water at 25°C?

| | $AgCl_{(s)}$ ⇌ | $Ag^+_{(aq)}$ + | $Cl^-_{(aq)}$ |
|---|---|---|---|
| [I] | - | 0 | 0 |
| [C] | | +x | +x |
| [E] | | x | x |

Stoichio-metric ratio

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

Molar solubility = $x = 1.3 \times 10^{-5}$ M



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## 6. Compute molar solubility from $K_{sp}$.

Example 2: Calculate the molar solubility of lead iodide in water from its $K_{sp}$ at 25°C? $K_{sp} = 8.5 \times 10^{-9}$

| | $PbI_{2(s)}$ ⇌ | $Pb^{+2}_{(aq)}$ + | $2\ I^-_{(aq)}$ |
|---|---|---|---|
| [I] | - | 0 | 0 |
| [C] | | +x | +2x |
| [E] | | x | 2x |

Stoichio-metric ratio

$K_{sp} = [Pb^{+2}][I^-]^2 = (x)(2x)^2 = 4x^3 = 8.5 \times 10^{-9}$

Molar solubility = $x = 1.3 \times 10^{-3}$ M

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## Solubility Equilibrium

We have covered the following topics:

1. Define the term "solubility".
2. What is a saturated solution?
3. What is solubility equilibrium?
4. What is the meaning of $K_{sp}$ and what is the solubility product constant expression?
5. Calculate $K_{sp}$ from solubility data.
6. Compute molar solubility from $K_{sp}$.



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## Solubility Equilibrium

We will cover the following topics:

1. Common Ion Effect
   - -Qualitatively using Le Chatelier's Principle
   - -Quantitatively by $K_{sp}$ calculation
2. Use $K_{sp}$ to predict whether a precipitate will form in a solution (i.e. Compare Q with $K_{sp}$)

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

1. Common Ion Effect

   Up to now, we have been studying the behaviour of low solubility salts dissolved in pure water.

   What if other ions are present in the water?

Recall, we calculate the solubility of $PbI_2$ (s)

Molar solubility $PbI_2$ in pure water $1.3 \times 10^{-3}$ M

Remember this number.



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

1. Common Ion Effect

Let's dissolve $PbI_2$ in a solution that already has sodium iodide, NaI, dissolved in it. How much of $PbI_2$ will dissolve?

To answer this question, think Le Chatelier's Principle. What does NaI do to the equilibrium?

$$PbI_2\ (s) \rightleftharpoons Pb^{+2}\ (aq) + 2\ I^-\ (aq)$$

## Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

1. Common Ion Effect

Let's dissolve $PbI_2$ in a solution that already has sodium iodide, NaI, dissolved in it. How much of $PbI_2$ will dissolve?

1. Ask yourself, what is the solubility property of NaI. Look it up!!

SOLUBILITY OF COMMON COMPOUNDS IN WATER

*The term soluble here means > 0.1 mol/L at 25°C*

| Negative Ions (Anions) | Positive Ions (Cations) | Solubility of Compounds |
|---|---|---|
| Chloride, $Cl^-$ or Bromide, $Br^-$ or Iodide, $I^-$ | All others | Soluble |
| | $Ag^+$, $Pb^{2+}$, $Cu^+$ | Low Solubility |

Is NaI soluble or insoluble in water?



## Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

1. Common Ion Effect

Let's dissolve $PbI_2$ in a solution that already has sodium iodide, NaI, dissolved in it. How much of $PbI_2$ will dissolve?

- Ask yourself, what is the solubility property of NaI. Look it up

Is NaI soluble or insoluble in water?

Answer: NaI is soluble!!

- What does it mean NaI is soluble?

It means that in solution,
$Na^+$ (aq) and $I^-$ (aq) ions are formed.

## Solubility Equilibrium
Hebden – Unit 3 (page 73-108)

1. Common Ion Effect

Let's dissolve $PbI_2$ in a solution that already has sodium iodide, NaI, dissolved in it. How much of $PbI_2$ will dissolve?

- Will dissolving $PbI_2$ in a solution full of $Na^+$ (aq) and $I^-$ (aq) affect the solubility of $PbI_2$? Yes!!

Solubility of $PbI_2$ is controlled by the equilibrium that has a $K_{sp}$ value equals to $8.9 \times 10^{-9}$. (from solubility constant table)

$$PbI_2\ (s) \rightleftharpoons Pb^{+2}\ (aq) + 2\ I^-\ (aq)$$

In the presence of NaI, the solution now has an extra source of $I^-$ (aq). This is equivalent to stressing the equilibrium by increasing the concentration of $I^-$ (aq). Le Chatelier's principle predicts that it will shift the equilibrium to the LEFT, causing LESS $PbI_2$ to dissolve.

## Solubility Equilibrium
Hebden – Unit 3 (page 73-108)

1. Common Ion Effect

Calculate the molar solubility of $PbI_2$ in 0.10 M NaI solution. Recall $K_{sp}$ of $PbI_2$ at 25°C? $K_{sp}$ = $8.5 \times 10^{-9}$

$I^-$ (aq) from The 0.10 M NaI.

| | $PbI_2$ (s) ⇌ | $Pb^{+2}$ (aq) + | $2\ I^-$ (aq) |
|---|---|---|---|
| [I] | - | 0 | 0.10 |
| [C] | | +x | +2x |
| [E] | | x | 0.10 +2x |

Stoichio-metric ratio

$K_{sp} = [Pb^{+2}][I^-]^2 = (x)(0.10 + 2x)^2 = 8.5 \times 10^{-9}$

Molar solubility = x = $8.5 \times 10^{-7}$ M

**Much smaller!! Recall that the molar solubiltiy of $PbI_2$ in pure water is $1.3 \times 10^{-3}$ M !!**

## Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

1. Common Ion Effect

Let's dissolve $PbI_2$ in a solution that already has sodium iodide, NaI, dissolved in it. How much of $PbI_2$ will dissolve?

In summary, dissolving $PbI_2$ in a solution of NaI will cause **LESS $PbI_2$ to dissolve** because of **common ion effect**.

1. Qualitatively, this is as predicted by Le Chatelier's Principle. The extra $I^-$ (aq) from NaI will cause the solubility equilibrium of $PbI_2$ to shift to the left, forming more $PbI_2$ (s), or suppressing the dissolution of $PbI_2$.
2. Quantitatively, this can be calculated using the solubility product constant of $PbI_2$.
   - Molar solubility of $PbI_2$ in pure water is $1.3 \times 10^{-3}$ M.
   - Molar solubility of $PbI_2$ in a 0.10 M NaI solution is $8.5 \times 10^{-7}$ M.



## Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

2. We can use $K_{sp}$ to determine whether a precipitate will form in solution.

This is done by computing Q, the ion product, and compare it to the $K_{sp}$.

- Precipitate will form if $Q > K_{sp}$
- No precipitate will form if:
  - $Q = K_{sp}$ (at saturation point)
  - $Q < K_{sp}$ (an unsaturated solution)

## Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

2. We can use $K_{sp}$ to determine whether a precipitate will form in solution.

Suppose we wish to prepare 0.500 L of a solution containing 0.0075 mole of **NaCl** and 0.075 mole of **$Pb(NO_3)_2$**. Will a precipitate form?

This question draws on the following knowledge:

1. Solubility rules
2. Predict double-replacement reactions
3. Write chemical formulae
4. Balance chemical reaction
5. Write $K_{sp}$ equilibrium constant expression



## Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

2. We can use $K_{sp}$ to determine whether a precipitate will form in solution.

Suppose we wish to prepare 0.500 L of a solution containing 0.0075 mole of NaCl and 0.075 mole of $Pb(NO_3)_2$. Will a precipitate form?

Write the balanced equation:

$$2\ NaCl\ (aq) + Pb(NO_3)_2\ (aq) \rightarrow PbCl_2\ (s) + 2\ NaNO_3\ (aq)$$

0.0075 mole $Na^+$
0.0075 mole $Cl^-$

0.075 mole $Pb^{2+}$
0.150 mole $NO_3^-$

$K_{sp} = 1.2 \times 10^{-5}$

Is there enough $Pb^{2+}$ and $Cl^-$ ions in solution to precipitate $PbCl_2$?

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

2. We can use $K_{sp}$ to determine whether a precipitate will form in solution.

Suppose we wish to prepare 0.500 L of a solution containing 0.0075 mole of NaCl and 0.075 mole of $Pb(NO_3)_2$.
(a) Will a precipitate form?
(b) Calculate the concentration of the ions after the reaction.

(a) $2\ NaCl\ (aq) + Pb(NO_3)_2\ (aq) \rightarrow PbCl_2\ (s) + 2\ NaNO_3\ (aq)$

0.0075 mole $Na^+$
0.0075 mole $Cl^-$

0.075 mole $Pb^{2+}$
0.150 mole $NO_3^-$

**$K_{sp}$ = 1.2 x $10^{-5}$**

$Na^+$ and $NO_3^-$ are spectator ions.

$[Pb^{2+}] = 0.075 / 0.500 = 0.15\ M$
$[Cl^-] = 0.0075 / 0.500 = 0.015\ M$

$PbCl_2\ (s) \rightleftharpoons Pb^{2+}\ (aq) + 2\ Cl^-\ (aq)$

$Q = [Pb^{2+}]\,[Cl^-]^2 = (0.15)(0.015)^2 = 3.4 \times 10^{-5}$

**$Q > K_{sp}$, precipitate will form!**



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

2. We can use $K_{sp}$ to determine whether a precipitate will form in solution.

Suppose we wish to prepare 0.500 L of a solution containing 0.0075 mole of NaCl and 0.075 mole of $Pb(NO_3)_2$.
(a) Will a precipitate form?
(b) Calculate the concentration of the ions after the reaction.

(b) $2\ NaCl\ (aq) + Pb(NO_3)_2\ (aq) \rightarrow PbCl_2\ (s) + 2\ NaNO_3\ (aq)$

0.0075 mole NaCl
This is the limiting reagent. → 

0.075 mole $Pb(NO_3)_2$
Using up 0.00375 mole $Pb(NO_3)_2$ →

Producing 0.00375 mole $PbCl_2$ →

Producing 0.0075 mole $NaNO_3$

Left over is
$0.075 - 0.00375 = 0.07125$ mole $Pb(NO_3)_2$ → 0.1425 M $Pb(NO_3)_2$ → 0.1425 M $Pb^{2+}$

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

2. We can use $K_{sp}$ to determine whether a precipitate will form in solution.

Suppose we wish to prepare 0.500 L of a solution containing 0.0075 mole of NaCl and 0.075 mole of $Pb(NO_3)_2$.
(a) Will a precipitate form?
(b) Calculate the concentration of the ions after the reaction.

(b) $Na^+$ and $Cl^-$ are spectator ions, $[Na^+] = 0.015$ M and $[NO_3^-] = 0.30$ M

| | $PbCl_2$ (s) $\rightarrow$ | $Pb^{2+}$ (aq) | + 2 $Cl^-$ (aq) |
|---|---|---|---|
| I | | 0.1425 | 0 |
| C | | x | 2x |
| E | | 0.1425 + x | 2x |

$K_{sp} = 1.2x10^{-5} = [Pb^{2+}][Cl^-]^2 = (0.1425 + x)(2x)^2 \approx (0.1425)(4x^2)$

$$x = \sqrt{\frac{1.2x10^{-5}}{(4)(0.1425)}} = 4.6x10^{-3}$$

$[Pb^{2+}] = 0.1425+0.0046 = 0.147M$

$[Cl^-] = 2(0.0046) = 0.0092$ M

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

2. We can use $K_{sp}$ to determine whether a precipitate will form in solution.

What possible precipitate might form by mixing 50.0 mL of $1.0 \times 10^{-4}$ M NaCl with 50.0 mL of $1.0 \times 10^{-6}$ M $AgNO_3$? Will a precipitate form?

Yes, there could be a possibility that a precipitate will form because AgCl has low solubility.

$NaCl\ (aq) + AgNO_3\ (aq) \rightarrow AgCl\ (s) + NaNO_3\ (aq)$

Whether AgCl will precipitate out will depend on the concentrations of the $Ag^+$ and $Cl^-$ ions.

Think dilution when calculating the $[Ag^+]$ and $[Cl^-]$ !!

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

2. We can use $K_{sp}$ to determine whether a precipitate will form in solution.

What possible precipitate might form by mixing 50.0 mL of $1.0 \times 10^{-4}$ M NaCl with 50.0 mL of $1.0 \times 10^{-6}$ M $AgNO_3$? Will a precipitate form?

$NaCl\ (aq) + AgNO_3\ (aq) \rightarrow AgCl\ (s) + NaNO_3\ (aq)$

moles of $Ag^+$ = (0.0500 L) ($1.0 \times 10^{-6}$ M) = $5.0 \times 10^{-8}$ moles
moles of $Cl^-$ = (0.0500 L) ($1.0 \times 10^{-4}$ M) = $5.0 \times 10^{-6}$ moles

$[Ag^+] = [NO_3^-] = 5.0 \times 10^{-8}$ moles / 0.1000 L = $5.0 \times 10^{-7}$ M
$[Na^+] = [Cl^-] = 5.0 \times 10^{-6}$ moles / 0.1000 L = $5.0 \times 10^{-5}$ M

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

2. We can use $K_{sp}$ to determine whether a precipitate will form in solution.

What possible precipitate might form by mixing 50.0 mL of $1.0 \times 10^{-4}$ M NaCl with 50.0 mL of $1.0 \times 10^{-6}$ M $AgNO_3$? Will a precipitate form?

$NaCl\ (aq) + AgNO_3\ (aq) \rightarrow AgCl\ (s) + NaNO_3\ (aq)$

$[Ag^+] = 5.0 \times 10^{-8}$ moles / 0.1000 L = $5.0 \times 10^{-7}$ M
$[Cl^-] = 5.0 \times 10^{-6}$ moles / 0.1000 L = $5.0 \times 10^{-5}$ M

$AgCl\ (s) \rightleftharpoons Ag^+\ (aq) + Cl^-\ (aq)$

$K_{sp} = 1.8 \times 10^{-10} = [Ag^+][Cl^-]$
$Q = (5.0 \times 10^{-7})(5.0 \times 10^{-5}) = 2.5 \times 10^{-11}$

$Q < K_{sp}$, therefore, no precipitate forms.

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

**Solubility Equilibrium**

We have covered the following topics:

1. Common Ion Effect
   - -Qualitatively using Le Chatelier's Principle
   - -Quantitatively by $K_{sp}$ calculation
2. Use $K_{sp}$ to predict whether a precipitate will form in a solution (i.e. Compare Q with $K_{sp}$)



# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

Solubility Problems

0.24 g $CaSO_4$ dissolve in 100.0 mL of water at 25°C to form a saturated solution. What is the $K_{sp}$ for $CaSO_4$ at 25°C.

# Solubility Equilibrium
Hebden – Unit 3 (page 73-108)

## Solubility Problems

0.24 g $CaSO_4$ dissolve in 100.0 mL of water at 25°C to form a saturated solution. What is the $K_{sp}$ for $CaSO_4$ at 25°C.

$CaSO_4(s) \rightleftharpoons Ca^{2+}(aq) + SO_4^{2-}(aq)$

$$[CaSO_4] = \frac{(0.24\ g\ CaSO_4)\left(\frac{mol\ CaSO_4}{136.14\ g\ CaSO_4}\right)}{0.1000\ L} = 1.76x10^{-2}\ M = [Ca^{2+}] = [SO_4^{\ 2-}]$$

$$K_{sp} = [Ca^{2+}][SO_4^{\ 2-}] = (1.76x10^{-2})^2 = 3.1x10^{-4}$$

# Solubility Equilibrium
Hebden – Unit 3 (page 73-108)

## Solubility Problems:

1. Calculate the molar solubility of $PbCl_2$ in pure water.

2. The molar solubility of $Ag_2SO_4$ in pure water is $1.2x10^{-5}$ M. Calculate the $K_{sp}$.

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## Solubility Problems:

1. Calculate the molar solubility of $PbCl_2$ in pure water.

$PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$

$$K_{sp} = 1.2x10^{-5} = [Pb^{2+}][Cl^-]^2 = s(2s)^2 = 4s^3$$

$$s = \sqrt[3]{\frac{1.2x10^{-5}}{4}} = 1.4x10^{-2}\,M$$

2. The molar solubility of $Ag_2SO_4$ in pure water is $1.2x10^{-5}$ M. Calculate the $K_{sp}$.

$Ag_2SO_4(s) \rightleftharpoons 2Ag^+(aq) + SO_4^{2-}(aq)$

$$K_{sp} = [Ag^+]^2[SO_4^{2-}] = [2(1.2x10^{-5})]^2(1.2x10^{-5}) = 6.9x10^{-15}$$

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## Solubility Problems:

A 200.0 mL sample of a solution that is $1.5x10^{-3}$ M in $Cu(NO_3)_2$ is mixed with a 250.0 mL sample of a solution that is 0.20 M in $NH_3$. After the solution reaches equilibrium, what concentration of $Cu^{2+}(aq)$ remains?

$Cu^{2+}(aq) + 4NH_3(aq) \rightleftharpoons Cu(NH_3)_4^{2+}(aq)$  $K = 1.7x10^{13}$

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## Solubility Problems:

A 200.0 mL sample of a solution that is $1.5x10^{-3}$ M in $Cu(NO_3)_2$ is mixed with a 250.0 mL sample of a solution that is 0.20 M in $NH_3$. After the solution reaches equilibrium, what concentration of $Cu^{2+}(aq)$ remains?

$$Cu^{2+}(aq) + 4NH_3(aq) \rightleftharpoons Cu(NH_3)_4^{2+}(aq) \quad K = 1.7x10^{13}$$

$$[Cu^{2+}]_0 = \frac{(200.0\ mL)(1.5x10^{-3}\ M)}{200.0\ mL + 250.0\ mL} = 6.7x10^{-4}\ M$$

$$[NH_3]_0 = \frac{(250.0\ mL)(0.20\ M)}{200.0\ mL + 250.0\ mL} = 0.11\ M$$

Since K >> 1 let the reaction go to completion and then let it approach equilibrium


# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## Solubility Problems:

A 200.0 mL sample of a solution that is $1.5x10^{-3}$ M in $Cu(NO_3)_2$ is mixed with a 250.0 mL sample of a solution that is 0.20 M in $NH_3$. After the solution reaches equilibrium, what concentration of $Cu^{2+}(aq)$ remains?

$$Cu^{2+}(aq) + 4NH_3(aq) \rightleftharpoons Cu(NH_3)_4^{2+}(aq) \quad K = 1.7x10^{13}$$


| | $Cu^{2+}$ | $4NH_3$ | $Cu(NH_3)_4^{2+}$ |
|---|---|---|---|
| I | $6.7x10^{-4}$ | 0.11 | 0 |
| R | $-6.7x10^{-4}$ | $-4(6.7x10^{-4})$ | $6.7x10^{-4}$ |
| C | x | 4x | -x |
| E | x | 0.11 + 4x | $6.7x10^{-4}$-x |

$$K = 1.7x10^{13} = \frac{6.7x10^{-4} - x}{(x)(0.11 + 4x)^4} \approx \frac{6.7x10^{-4}}{(x)(0.11)^4} \rightarrow x = 2.7x10^{-13}\ M\ assumption\ is\ OK$$

$[Cu^{2+}] = 2.7x10^{-13}$ M $\quad$ $[NH_3] = 0.11M + 4(2.7x10^{-13}\ M) = 0.11$ M

$[NH_3] = 6.7x10^{-4}$ M – $2.7x10^{-13}$ M = $6.7x10^{-4}$ M

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## Solubility Problems:

1. What is the molar solubility of $CaF_2$ in 0.100 M NaF. $K_{sp} = 1.5x10^{-10}$

2. A 30.0 mL sample of a solution that is 0.0500 M in $Pb(NO_3)_2$ is mixed with a 70.0 mL sample of a solution that is 0.00500 M in NaBr. Will a precipitate form?

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

## Solubility Problems:

1. What is the molar solubility of $CaF_2$ in 0.100 M NaF.
   $K_{sp} = 1.5x10^{-10}$

   $CaF_2(s) \rightleftharpoons Ca^{2+}(aq) + 2F^-(aq)$   $K = 1.5x10^{-10} = [Ca^{2+}][F^-]^2$

   | | | |
   |---|---|---|
   | I | 0 | 0.100 |
   | C | s | 2s |
   | E | s | 0.100+2s |

   $K = 1.5x10^{-10} = s(0.100+2s)^2 \approx s(0.100)^2$

   $s = 1.5x10^{-8}$ M  assumption is OK

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

Solubility Problems:

2. A 30.0 mL sample of a solution that is 0.0500 M in $Pb(NO_3)_2$ is mixed with a 70.0 mL sample of a solution that is 0.00500 M in NaBr. Will a precipitate form?

Must consider $PbBr_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Br^-(aq)$

$$[Pb^{2+}]_0 = \frac{(30.0\ mL)(0.0500M)}{100.0\ mL} = 0.0150\ M$$

$$[Br^-]_0 = \frac{(70.0\ mL)(0.00500M)}{100.0\ mL} = 0.00350\ M$$

$Q = [Pb^{2+}][Br^-]^2 = (0.0150)(0.00350)^2 = 1.8x10^{-7} < Ksp = 6.6x10^{-6}$

Since $Q < K_{sp}$ no precipitate forms

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

Solubility Problems

What is the solubility of AgCl in:

a) Pure water

b) $4.00x10^{-3}$ M $CaCl_2$

c) $4.00x10^{-3}$ M $Ca(NO_3)_2$

d) $4.00x10^{-3}$ M $CaBr_2$

# Solubility Equilibrium

Hebden – Unit 3 (page 73-108)

Solubility Problems

$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$ $K = 1.8 \times 10^{-10} = [Ag^+][Cl^-]$

Let s be the solubility of AgCl

What is the solubility of AgCl in:

a) Pure water

$1.8 \times 10^{-10} = (s)(s) \rightarrow s = 1.3 \times 10^{-5}$ M

b) $4.00 \times 10^{-3}$ M $CaCl_2$

$1.8 \times 10^{-10} = (s)(8.00 \times 10^{-3} + s) \approx (s)(8.00 \times 10^{-3}) \rightarrow s = 2.3 \times 10^{-8}$ M

c) $4.00 \times 10^{-3}$ M $Ca(NO_3)_2$ same as pure water

$1.8 \times 10^{-10} = (s)(s) \rightarrow s = 1.3 \times 10^{-5}$ M

d) $4.00 \times 10^{-3}$ M $CaBr_2$

$1.8 \times 10^{-10} = (s - 8.00 \times 10^{-3})(s)$ let $s = x + 8.00 \times 10^{-3}$

$1.8 \times 10^{-10} = (x)(x + 8.00 \times 10^{-3}) \approx (x)(8.00 \times 10^{-3}) \rightarrow x = 2.3 \times 10^{-8}$ M $\rightarrow s = 8.00 \times 10^{-3}$ M