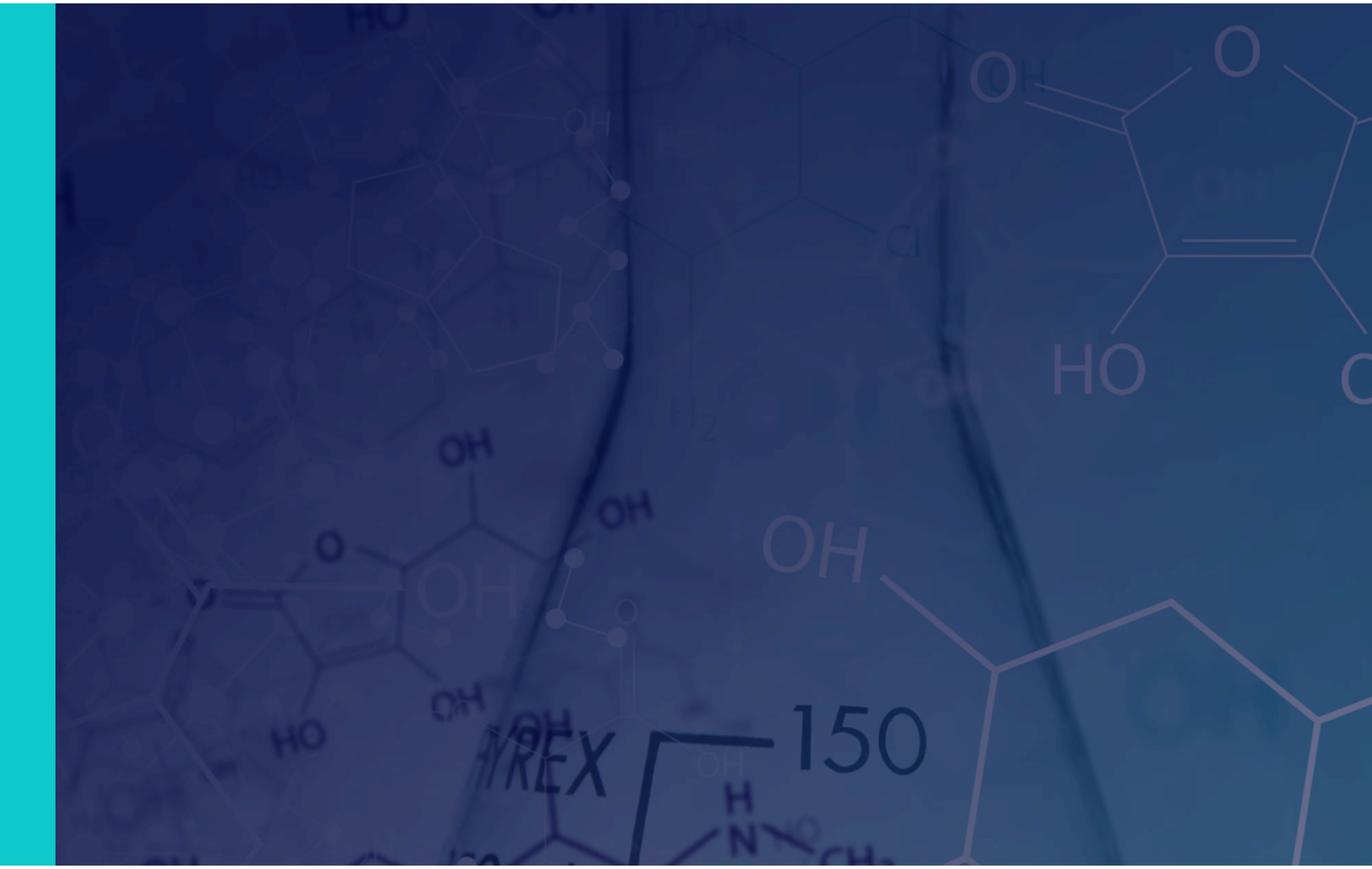

UWO

CHEM 1302

Winter 2025, Chapter 8 Notes

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8. Solubility of Ionic Compounds

8.1 Solubility Rules and Precipitation

8.1.1

Solubility Chart

The following table tells us if certain combinations of ions are soluble or insoluble!

WIZE CONCEPT

If a combination of ions is **soluble**, it means that **no precipitate will form.**
If a combination of ions is **insoluble**, it means a **precipitate will form!**

You can expect to see a question on your exam that asks you if a precipitate will form or not!

Soluble Ionic Compounds	contain these ions	exceptions
	NH_4^+ group I cations: Li^+ Na^+ K^+ Rb^+ Cs^+	none
	Cl^- Br^- I^-	compounds with Ag^+, Hg_2^{2+}, and Pb^{2+}
	F^-	compounds with group 2 metal cations, Pb^{2+} and Fe^{3+}
	$C_2H_3O_2^-$ HCO_3^- NO_3^- ClO_3^-	none
	SO_4^{2-}	compounds with Ag^+, Ba^{2+}, Ca^{2+}, Hg_2^{2+}, Pb^{2+} and Sr^{2+}
Insoluble Ionic Compounds	contain these ions	exceptions
	CO_3^{2-} CrO_4^{2-} PO_4^{3-} S^{2-}	compounds with group 1 cations and NH_4^+
	OH^-	compounds with group 1 cations and Ba^{2+}

Decide whether the following salts would form a precipate or not:

1) NaCl

2) AgCl

3) AgI

4) $NaNO_3$

8.1.2

Example: Does a Precipitate Form?

In the following example, do you think a precipitate is formed? If yes, what is the formula of the precipitate?

$Pb(NO_3)_2 + 2KI$

Write the complete chemical reaction:

$Pb(NO_3)_2$________ + 2KI________ → __________ + ______________

WIZE TIP

In a chemical reaction, **(aq)** means that the substance is dissolved in aqueous solution (aka it is **soluble**),
whereas **(s)** means that the substance is in solid form and is **insoluble** in solution (aka **forms a precipitate**)

This reaction can be referred to as a **precipitation reaction.**

Example: Salt Solubility Practice

1 a) Saturated solutions of Na_2S, CuS, SnS_2, and Al_2S_3 are prepared in a lab at 25^{o}C. Which salt is the most soluble in water?

a) Na_2S

b) CuS

c) SnS_2

d) Al_2S_3

1 b) Which of the options above would produce the highest $[S^{2-}]$?

2) Which of the following salts has the lowest solubility?

a) copper (I) perchlorate

b) ammonium sulphide

c) potassium hydroxide

d) mercury(II) sulphate

8.1.4

Practice: Dissociation

What ions does Na_2CO_3 dissociate into in water?

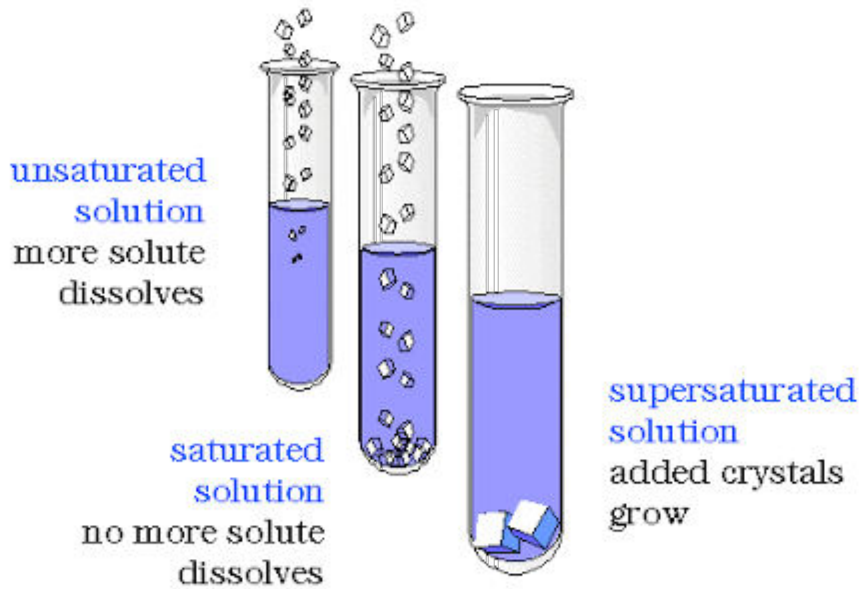
- Na_2CO_3 only
- 2 Na^+, C^{4+}, O_2^-
- Na_2^+,C^{4+} and O_2^-
- 2 Na^+ and CO_3^{2-}
- Na^{+2} and CO_3^{2-}

8.2 Solubility Products (Ksp)

8.2.1

Solubility and Ksp

- Recall that a **saturated solution** is one that has the maximum quantity of solute dissolved. It is at this point that an **equilibrium forms between the solid and the dissolved ions.**
  - Ex: $AgCl(s) \rightleftharpoons Ag^{+}(aq) + Cl^{-}(aq)$
  - Increasing temperature can cause the solubility to increase



Ksp= The solubility product constant

- It is the equilibrium constant (K) that is specifically for saturated solutions

Write the Ksp expression for the above reaction in equilibrium:

! WATCH OUT!

Remember that solids or liquids are never included in any K expression!

For these solids, we can write out ICE tables similar to how we did earlier:

Let's say the Ksp=1.8×10^{-10} for the AgCl(s) at 25°C and we start with 0.2M of AgCl(s). We could be asked to find the concentrations for the ions at equilibrium.

1) Create an ICE Table

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

	AgCl(s)	⇌	$Ag^+(aq)$	$Cl^-(aq)$
I	0.2M	/	0	0
C	-x	/	x	x
E	0.2-x	/	x	x

2) Write out a Ksp expression and solve for x:

Ksp=[Ag+][Cl-]

$1.8\times10^{-10}=x^2$

$x=1.34\times10^{-5}$ M = [Ag+]=[Cl-]

8.2.2

Express the Ksp for the following salts

a) $NaOH_{(s)}$

b) MgF_2

c) Al_2S_3

8.2.3

Estimate the molar solubility of Ag_2CrO_4 in pure water if K_{sp} for silver chromate is 1.1×10^{-12}.

- 7.1×10^{-7} M
- 6.5×10^{-5} M
- 9.1×10^{-5} M
- 2.3×10^{-4} M

8.2.4

The solubility of copper(II) arsenate, $Cu_3(AsO_4)_2$, in pure water is 3.7×10^{-8} M. Calculate the value of K_{sp} for copper(II) arsenate from this data.

- 6.9×10^{-38}
- 4.2×10^{-37}
- 2.5×10^{-36}
- 3.7×10^{-36}
- 7.5×10^{-36}

8.3 Does a Precipitate Form?

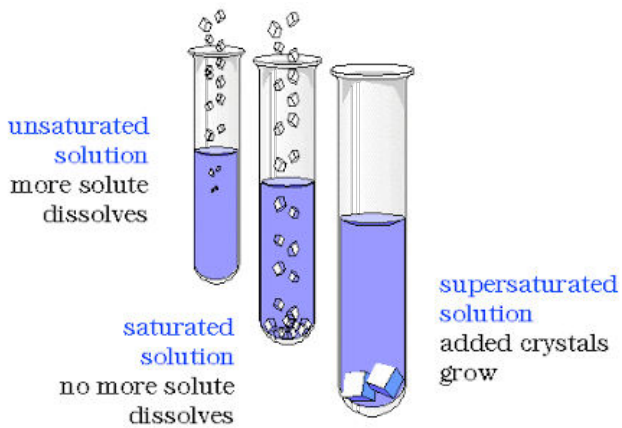
8.3.1

Q vs K to Predict Solubility

Recall:

- **There are three possible situations:**

1. **Q = K:** The system is at equilibrium.
2. **Q > K:** The [products] is too high compared to [reactants]. The equilibrium will shift left.
3. **Q < K:** The [products] is too low compared to [reactants]. The equilibrium will shift right.



When asked to predict solubility, we should be able to identify whether a solution is unsaturated, saturated, or supersaturated!

	What type of solution is this?	Will a precipitate form?
Q = Ksp		
Q > Ksp		
Q < Ksp		

8.3.2

If 500 mL of a 4.0×10^{-6} mol/L $CaCl_2$ solution is mixed with 300 mL of a 0.0040 mol/L $AgNO_3$ solution, will a precipitate form? ($K_{sp}(AgCl) = 1.8 \times 10^{-10}$)

Solubility Of Ionic Compounds At SATP

		Anions						
		Cl^-, Br^-, I^-	**S^{-2}**	**OH^-**	**SO_4^{2-}**	**CO_3^{2-}, PO_4^{3-}**	**$C_2H_3O_2^-$**	**NO_3^-**
Cations	High solubility (aq) ≥ 0.1 mol/L	most	Group 1, Group 2, NH_4^{+1}	Group 1, NH_4^{+1}, Sr^{2+}, Ba^{2+}, Tl^+	most	Group 1, NH_4^{+1}	most	all
	Low Solubility (s) <0.1 mol/L	Ag^+, Pb^{2+}, Tl^+, Hg_2^{2+}, Cu^+	most	most	Ag^+, Pb^{2+}, Ca^{2+}, Ba^{2+}, Sr^{2+}, Ra^{2+}	most	Ag^+	none

8.3.3

Does a precipitate of $PbBr_2$ form if 50.0 mL of 0.0100 M $Pb(NO_3)_2$ is mixed with 50.0 mL of 0.100 M NaBr? (K_{sp} $PbBr_2$ = 2.1×10^{-6})

- No, because this is a supersaturated solution.
- No, because equilibrium will shift towards products,.
- Yes, because this is a supersaturated solution.
- Yes, because equilibrium will shift towards products.

8.4 Factors that Affect Solubility

8.4.1

Factors that affect solubility:

Lower solubility:

- **Common ion effect:** the solubility of a salt is reduced by the presence of another salt having a **common ion** (the ion that is common to both salts); Le Chatelier predicts that the equilibrium will shift to the left to reduce the disturbance.

Higher solubility:

- **Complex ion formation:** reactions that use up one of the ions, to form a complex ion will increase the solubility of a salt. An ion made up of a metal ion bonded to one or more molecules or ions is called a **complex ion**.

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

$$Ag^{+}(aq) + 2CN^{-}(aq) \rightleftharpoons Ag(CN)_2^{-}(aq)$$

- **Acid-base neutralization:** if the anion of the salt is the conjugate base of a weak acid, the solubility of the salt will increase in an acidic solution

$$CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + CO_3^{2-}(aq)$$

$$CO_3^{2-}(aq) + H_3O^{+}(aq) \rightleftharpoons HCO_3^{-}(aq) + H_2O(l)$$

8.4.2

Common-Ion Effect Introduction

By now we are familiar with the equation below where we have a solid dissolving and can form an equilibrium.
Now let's consider adding common ions to this equilibrium....

Let's take the solid we've seen, $PbCl_2(s)$, and write out the reaction for when we have a saturated solution:

$PbCl_2(s)$ eqm $Pb^{2+}(aq)$ + $2Cl^-(aq)$ $K_{sp}=1.7\times10^{-5}$

Now let's say we wanted to add NaCl(s) to the same solution. How will this solid dissociate?

NaCl(s) will be able to completely dissociate into ions:
NaCl(s) -> Na+ + Cl-

So now we will have additional Na+ and Cl- ions in solution. How do you think these would affect this equilibrium?

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

The added Cl- ions will disrupt the equilibrium! With more products in our equation, according to Le Chatelier's Principle, the reaction will shift to the _______!
Here, we call Cl- the **common ion** (an ion that enters the solution through 2 different sources)

In general, do common ions increase or decrease the solubility of a salt? ____________

8.4.3

If we dissolve $PbCl_2(s)$ in solution containing 0.1M of NaCl, what is the equilibrium concentration of Pb^{2+}?

$PbCl_2(s)$ eqm $Pb^{2+}(aq)$ + $2Cl^-(aq)$ Ksp=1.7×10^{-5}

$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$
0.1M 0.1M

Now let's do an ICE table for our main salt equation:

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

What would the Ksp expression be? Can we solve for x?

At this point, let's think if we can make a **simplification** :)
We said that adding a common ion caused an eqm shift to the left/right? ______
So the common ion caused an increase/decrease in the solubility of a salt: __________
If this is the case, then do you think x (the molar solubility) for the salt would go down or up when a common ion is present? ________

With a common ion present, **because the solubility of the salt is decreased, x becomes very small and we can ignore it when it is added to the concentration of ions that is already present. *Assume the concentration of common ions is entirely due to the other solution!**

Here we have the Cl- ions with an equilibrium concentration of 0.1 + 2x. 0.1 is a lot bigger than x and x is very tiny so:

0.1 + 2x ~0.1

Now we'll continue solving!

$1.7x10^{-5}=[x][0.1 + 2x]^2$

Early we calculated [Pb2+] when no common ions were present. The answer was $1.6x10^{-2}$M.

Now, with common ions present the [Pb2+] was: ____________

The concentrations of Pb^{2+} changed because the common ion reduced the solubility of $PbCl_2(s)$!

8.4.4

Calculate the molar solubility of AgCl ($K_{sp} = 1.60 \times 10^{-10}$) in:

a) 2.50 L pure water.

b) 2.50 L solution containing 5.00 g of dissolved $CaCl_2$. Molar mass of $CaCl_2$ is 111.0 g/mol. Note that $CaCl_2$ fully dissociates in water.

a)

b) $CaCl_2$ highly soluble in water and fully dissociates.

8.4.5

Find the molar solubility of $SrCO_3$ (Ksp = 5.4×10^{-10}) in pure water and in 0.13 M $Sr(NO_3)_2$. Report your answers to two significant figures in scientific notation. Do not include units in your answer.

8.4.6

A student is preparing to test the molar solubility of $PbHPO_4$ (s) (K_{sp}=3.5×10^{-12}) in different aqueous solutions. When this salt dissociates, it produces Pb^{2+} (aq) and the amphiprotic ion HPO_4^{2-} (aq).They prepare separate beakers with 100 mL of each of the following solutions to test:

a. water
b. 0.1 M $Pb(NO_3)_2$ (aq) (which fully dissociates in water)
c. 1.0 M HCℓ (aq)
d. 0.001 M NaOH (aq)

Unfortunately, the student forgot to label which beaker is which! They continue with their experiment and ranked each solution in terms of the molar solubility of $PbHPO_4$ (s) in that solution. Use their data to complete the bottom row of the table below, and determine which solution is in which beaker.

	 **Beaker 1**	 **Beaker 2**	 **Beaker 3**	 **Beaker 4**
Relative molar solubility in this beaker	*Lowest molar solubility of $PbHPO_4(s)$*	$\longrightarrow$		*Highest molar solubility of $PbHPO_4(s)$*
This beaker contains 100 mL of…				