

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+}
	$\text{C}_2\text{H}_3\text{O}_2^-$ 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+}

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

1) NaCl

No precipitate

2) AgCl

Yes

3) AgI

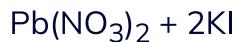
Yes

4) NaNO₃

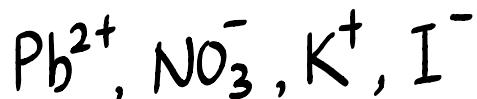
No

Example: Does a Precipitate Form?

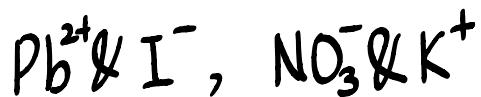
In the following example, do you think a precipitate is formed? If yes, what is the formula of the precipitate?



1) Write out all ions



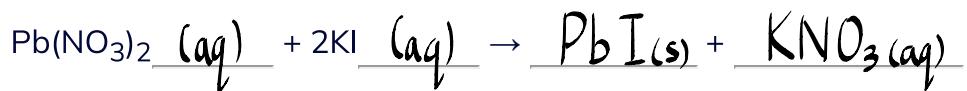
2) Consider combination



- 3) i) I^- is soluble, but Pb^{2+} is exception
 ii) anything with NO_3^- is soluble



Write the complete chemical reaction:



 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°C . Which salt is the most soluble in water?

- a) Na_2S Na^+ (any group 1 ion)
- b) CuS S have a charge of -2, Cu must have a charge 2^+
- c) SnS_2 Sn have a charge of 4^+
- d) Al_2S_3 Al^{3+} and S^{2-}

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

In order to have the highest $[S^{2-}]$,
that means the solution was most soluble
which is a).

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

- a) copper (II) perchlorate ClO_4^- with any compound is soluble
- b) ammonium sulphide NH_4^+ with any compound is soluble
- c) potassium hydroxide K^+ is a group I ions. This salt is soluble
- d) mercury(II) sulphate SO_4^{2-} is generally soluble but Hg^{2+} is exceptions, so this compound is insoluble.

Practice: Dissociation

What ions does Na_2CO_3 dissociate into in water?

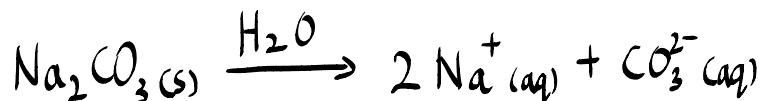
Na_2CO_3 only

$2 \text{Na}^+, \text{C}^{4+}, \text{O}_2^-$

$\text{Na}_2^+, \text{C}^{4+}$ and O_2^-

2Na^+ and CO_3^{2-}

Na^{+2} and CO_3^{2-}



Ions Formed:

Sodium ions (Na^+): There are spectator ions that do not participate in acid-base solution

Carbonate ions (CO_3^{2-}): This ion can act as a weak base, partially reacting with water in a hydrolysis reaction

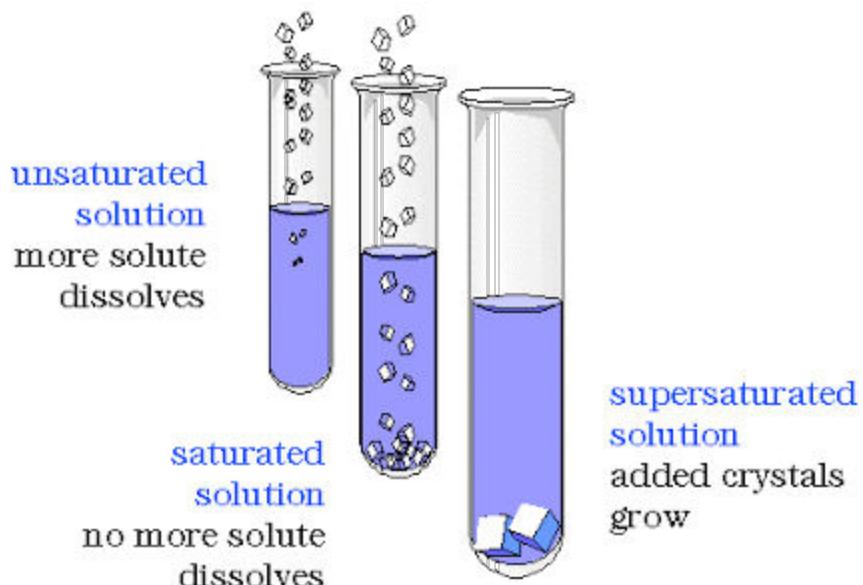
8.2

Solubility Products (K_{sp})

8.2.1

Solubility and K_{sp}

- 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: $\text{AgCl}(s) \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq)$
 - Increasing temperature can cause the solubility to increase



K_{sp}= The solubility product constant

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

Write the K_{sp} expression for the above reaction in equilibrium:

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

! 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:

Silver Chloride

Let's say the K_{sp}= 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

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

2) Write out a K_{sp} expression and solve for x:

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

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

$$x = 1.34 \times 10^{-5} M = [Ag^+] = [Cl^-]$$

8.2.2

Express the K_{sp} for the following salts



8.2.3

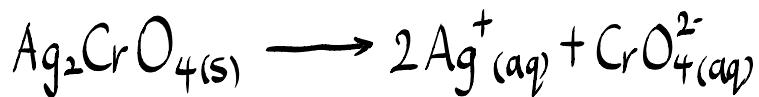
Estimate the molar solubility of Ag_2CrO_4 in pure water if K_{sp} for silver chromate is 1.1×10^{-12} .

$7.1 \times 10^{-7} \text{ M}$

$6.5 \times 10^{-5} \text{ M}$

$9.1 \times 10^{-5} \text{ M}$

$2.3 \times 10^{-4} \text{ M}$



$$K_{\text{sp}} = [\text{Ag}^+]^2 [\text{CrO}_4^{2-}]$$

Let solubility of Ag_2CrO_4 is x

$$K_{\text{sp}} = [2x]^2 [x]$$

$$4x^3 = 1.1 \times 10^{-12}$$

$$x = \sqrt[3]{\frac{1.1 \times 10^{-12}}{4}}$$

$$x \approx 6.5 \times 10^{-5} \text{ M}$$

∴ The molar solubility of Ag_2CrO_4 in pure water is $6.5 \times 10^{-5} \text{ M}$.

8.2.4

The solubility of copper(II) arsenate, $\text{Cu}_3(\text{AsO}_4)_2$, in pure water is 3.7×10^{-8} M. Calculate the value of K_{sp} for copper(II) arsenate from this data.

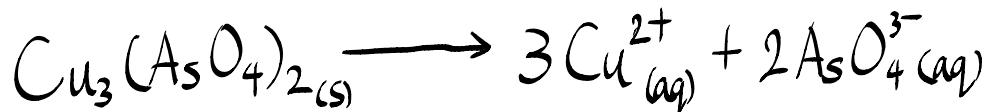
6.9×10^{-38}

4.2×10^{-37}

2.5×10^{-36}

3.7×10^{-36}

7.5×10^{-36}



let the molar solubility of $\text{Cu}_3(\text{AsO}_4)_2$ be $s = 3.7 \times 10^{-8}$ M

$$[\text{Cu}^{2+}] = 3 \times (3.7 \times 10^{-8}) = 1.1 \times 10^{-7} \text{ M}$$

$$[\text{AsO}_4^{3-}] = 2 \times (3.7 \times 10^{-8}) = 7.4 \times 10^{-8} \text{ M}$$

$$\begin{aligned} K_{\text{sp}} &= [\text{Cu}^{2+}]^3 [\text{AsO}_4^{3-}]^2 \\ &= (1.1 \times 10^{-7})^3 (7.4 \times 10^{-8})^2 \\ &= 7.51 \times 10^{-36} \end{aligned}$$

8.3

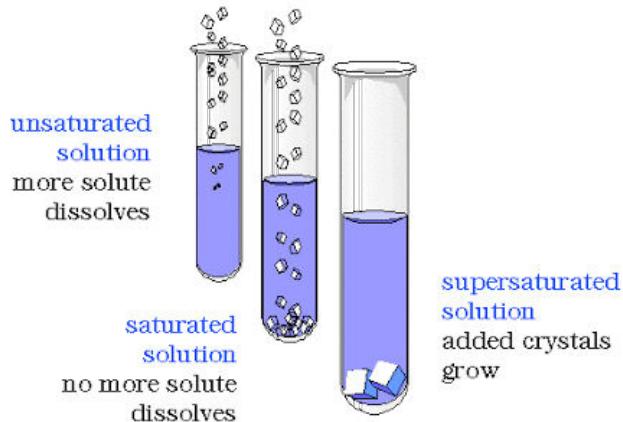
Does a Precipitate Form?

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!

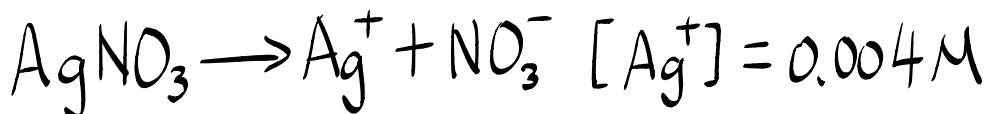
	What type of solution is this?	Will a precipitate form?
$Q = K_{sp}$	saturated solution	No
$Q > K_{sp}$	supersaturated solution	Yes
$Q < K_{sp}$	unsaturated solution	No

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? ($K_{\text{sp}}(\text{AgCl}) = 1.8 \times 10^{-10}$)

Solubility Of Ionic Compounds At SATP

		Anions						
		Cl^- , Br^- , I^-	S^{2-}	OH^-	SO_4^{2-}	CO_3^{2-} , PO_4^{3-}	$\text{C}_2\text{H}_3\text{O}_2^-$	NO_3^-
Cations	High solubility (aq) $\geq 0.1 \text{ mol/L}$	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) $< 0.1 \text{ mol/L}$	Ag^+ , Pb^{2+} , Tl^+ , Hg_2^{2+} , Cu^+	most	most	Ag^+ , Pb^{2+} , Ca^{2+} , Ba^{2+} , Sr^{2+} , Ra^{2+}	most	Ag^+	none



$$V_f = V_1 + V_2 = 500 + 300 = 800 \text{ mL} = 0.8 \text{ L}$$

$$[\text{Cl}^-] = 8.0 \times 10^{-6} \times \frac{500}{800} = 5.0 \times 10^{-6} \text{ M}$$

$$[\text{Ag}^+] = 0.004 \times \frac{300}{800} = 1.5 \times 10^{-3} \text{ M}$$

$$Q_{\text{sp}} = (1.5 \times 10^{-3})(5.0 \times 10^{-6}) \\ = 7.5 \times 10^{-9}$$

$Q_{\text{sp}} > K_{\text{sp}}$, supersaturated solution,
then precipitate will form.

8.3.3

Does a precipitate of PbBr_2 form if 50.0 mL of 0.0100 M $\text{Pb}(\text{NO}_3)_2$ is mixed with 50.0 mL of 0.100 M NaBr ? ($K_{\text{sp}} \text{ PbBr}_2 = 2.1 \times 10^{-6}$)



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.

$$V = V_1 + V_2 = 50 + 50 = 100 \text{ mL}$$

$$C_{\text{New}} = C_{\text{ini}} \times \frac{V_f}{V_i}$$

$$[\text{Pb}^{2+}] = (0.1) \times \frac{50}{100} = 0.05 \text{ M}$$

$$[\text{Br}^-] = (0.1) \times \frac{50}{100} = 0.05 \text{ M}$$

$$\begin{aligned} Q_{\text{sp}} &= [\text{Pb}^{2+}][\text{Br}^-]^2 \\ &= (0.05)(0.05)^2 = 1.25 \times 10^{-5} \end{aligned}$$

$$K_{\text{sp}} = 2.1 \times 10^{-6}$$

$Q_{\text{sp}} > K_{\text{sp}}$ the solution is supersaturated.

8.4

Factors that Affect Solubility

8.4.1

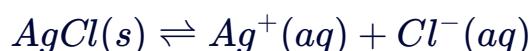
Factors that affect solubility:

Lower solubility:

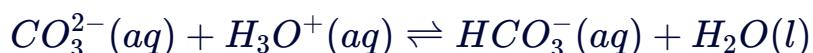
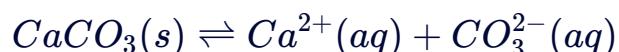
- **Common ion effect:** the solubility of a salt is reduced by the presence of another salt having a **common ion** (the ion that is common to both salts); Le Chatelier predicts that the equilibrium will shift to the left to reduce the disturbance.

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**.



- **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



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, $\text{PbCl}_2(s)$, and write out the reaction for when we have a saturated solution:



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

$\text{NaCl}(s)$ will be able to completely dissociate into ions:



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



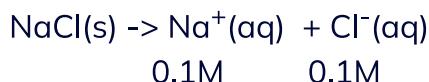
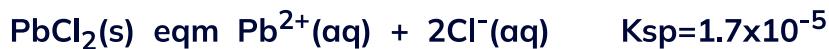
The added Cl^- ions will disrupt the equilibrium! With more products in our equation, according to Le Chatelier's Principle, the reaction will shift to the left!

Here, we call Cl^- the **common ion** (an ion that enters the solution through 2 different sources)

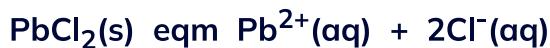
In general, do common ions increase or decrease the solubility of a salt? decrease

8.4.3

If we dissolve $\text{PbCl}_2(s)$ in solution containing 0.1M of NaCl, what is the equilibrium concentration of Pb^{2+} ?



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



I	0	0.1M
C	x	$+2x$
E	x	$0.1 + 2x$

What would the K_{sp} expression be? Can we solve for x ?

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

$$1.7 \times 10^{-5} = x [0.1 + 2x]^2$$

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~~ left

So the common ion caused an increase/decrease in the solubility of a salt: decrease

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? down

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:

$$0.1 + 2x \sim 0.1$$

Now we'll continue solving!

$$1.7 \times 10^{-5} = [x][0.1 + 2x]^2$$

Early we calculated $[\text{Pb}^{2+}]$ when no common ions were present. The answer was $1.6 \times 10^{-2} \text{ M}$.

Now, with common ions present the $[\text{Pb}^{2+}]$ was: $1.7 \times 10^{-3} \text{ M}$

The concentrations of Pb^{2+} changed because the common ion reduced the solubility of $\text{PbCl}_2(\text{s})$!

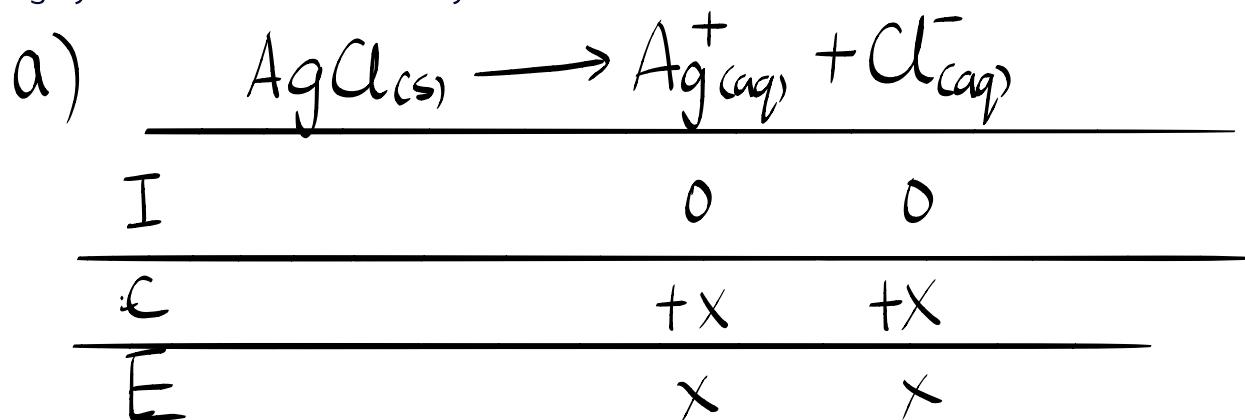
8.4.4

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

- 2.50 L pure water.
- 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)

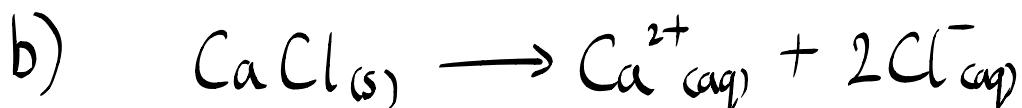
b) CaCl_2 highly soluble in water and fully dissociates.



$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = x \cdot x = x^2$$

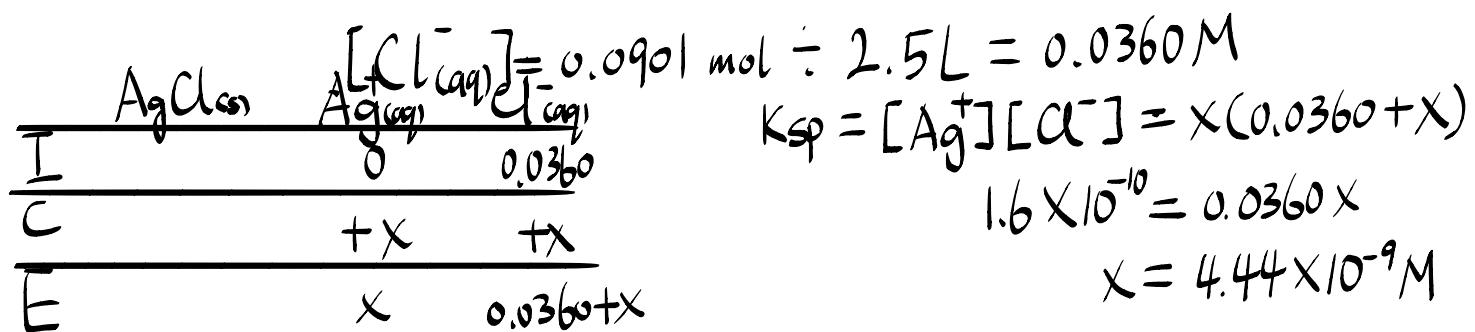
$$1.6 \times 10^{-10} = x^2$$

$$x = 1.26 \times 10^{-5} \text{ mol/L}$$



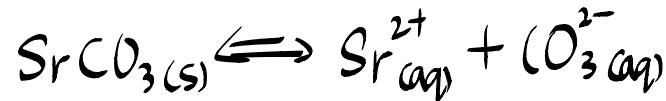
$$n_{\text{Cl}^{-}_{(aq)}} = (5.0 \text{ g}) \left(\frac{\text{mol}}{111 \text{ g}} \right) \left(\frac{2 \text{ mol Cl}^-}{1 \text{ mol CaCl}_2} \right)$$

$$= 0.0901 \text{ mol}$$



8.4.5

Find the molar solubility of SrCO_3 ($K_{\text{sp}} = 5.4 \times 10^{-10}$) in pure water and in 0.13 M $\text{Sr}(\text{NO}_3)_2$. Report your answers to two significant figures in scientific notation. Do not include units in your answer.



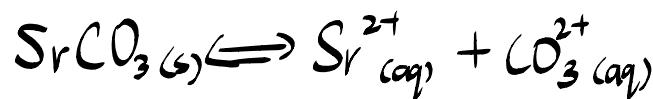
Let molar solubility of SrCO_3 be x

$$[\text{Sr}^{2+}] = s \quad [\text{CO}_3^{2-}] = s$$

$$K_{\text{sp}} = [\text{Sr}^{2+}][\text{CO}_3^{2-}]$$

$$5.4 \times 10^{-10} = s^2$$

$$s = 2.32 \times 10^{-5}$$



$$K_{\text{sp}} = [\text{Sr}^{2+}][\text{CO}_3^{2-}] = (0.13 + s)s$$

$$5.4 \times 10^{-10} = 0.13s$$

$$s = 4.2 \times 10^{-9}$$

\therefore The molar solubility of SrCO_3 in pure water is 2.32×10^{-5} , the molar solubility in 0.13 M $\text{Sr}(\text{NO}_3)_2$ is 4.2×10^{-9}

8.4.6

A student is preparing to test the molar solubility of PbHPO_4 (s) ($K_{\text{sp}}=3.5\times10^{-12}$) 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:

- water
- 0.1 M $\text{Pb}(\text{NO}_3)_2$ (aq) (which fully dissociates in water)
- 1.0 M HCl (aq)
- 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	<i>Lowest molar solubility of PbHPO_4 (s)</i>	\longrightarrow		<i>Highest molar solubility of PbHPO_4 (s)</i>
This beaker contains 100 mL of...	0.1 M $\text{Pb}(\text{NO}_3)_2$	1.0 M HCl	H_2O	0.001 M NaOH