SCH4U-C

Lesson 17

Solubility of Ionic Compounds

Introduction

In the last lesson, you reviewed solubility guidelines. These guidelines use the term "insoluble," but, in reality, every ionic substance will dissolve to some extent. The solubility of some compounds is so low that it appears as if they do not dissolve. Ionic compounds with low solubility will likely form precipitates in solutions.

One such ionic compound with low solubility is calcium oxalate. The forming of kidney stones, which are composed of calcium oxalate, is one of the human health issues that you will study in this unit.

Planning Your Study

You may find this time grid helpful in planning when and how you will work through this lesson.

Suggested Timing for This Lesson (Hours)	
Ionic Compounds and Their Solubility	½
Saturated Solutions and K_{sp} Expressions	½
Calculating Concentration and K_{sp}	¾
Predicting Precipitates	½
The Common Ion Effect	¾
Determining K_{sp} Experimentally	½
Key Questions	1

What You Will Learn

After completing this lesson, you will be able to

- describe and explain equilibrium in saturated solutions
- write K_{sp} expressions, and calculate molar solubility and K_{sp} values
- predict whether a precipitate will form during double displacement reactions

Ionic Compounds and Their Solubility

All ionic compounds will dissolve to some extent. Some are very soluble, while others have extremely low solubility. There is a range of solubility for ionic compounds. A set of solubility guidelines will help you to determine if an ionic substance will readily dissolve.

In this lesson, you'll need to refer to the solubility guidelines that you saw in Lesson 16:

Solubility guidelines for common ionic substances

Ionic substances that contain these positive ions are *soluble* in water:

- H^+
- NH_4^+
- Alkali metal ions

Ionic substances that contain these negative ions can be *either soluble or insoluble* in water:

- ClO_3^-: All chlorates are *soluble.*
- NO_3^-: All nitrates are *soluble.*
- CH_3COO^-: All acetates are *soluble.*
- F^-: All fluorides are *soluble* except for Mn^{2+}, Ba^{2+}, Ca^{2+}, Cu^{2+}, Fe^{2+}, Fe^{3+}, Pb^{2+}, Mg^{2+}, Ni^{2+}, Sr^{2+}, and Zn^{2+}.
- Cl^-: All chlorides are *soluble* except for Hg^{2+}, Ag^+, Cr^{3+}, Au^+, and Pb^{2+}.
- Br^-: All bromides are *soluble* except for Hg^{2+}, Ag^+, Au^+, and Pb^{2+}.
- I^-: All iodides are *soluble* except for Bi^{3+}, Cu^+, Au^+, Au^{3+}, Hg^{2+}, Ag^+, Pt^{3+}, and Pb^{2+}.
- SO_4^{2-}: All sulfates are *soluble* except for Sb^{3+}, Ba^{2+}, Ca^{2+}, Fe^{3+}, Pb^{2+}, Mg^{2+}, Ag^+, and Sr^{2+}.
- S_2: All sulfides are *insoluble* except for NH_4^+, H^+, alkali metal ions, Be^{2+}, Hg^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, and Ra^{2+}.
- CO_3^{2-}, PO_4^{3-}, and SO_3^{2-}: All carbonates, phosphates, and sulfites are *insoluble* except for NH_4^+, H^+, and alkali metal ions.
- OH^-: All hydroxides are *insoluble* except for NH_4^+, alkali metal ions, Ba^{2+}, Ca^{2+}, Sr^{2+}, Ra^{2+}, and Tl^+.

Consider the compounds $NaNO_3$, $Pb(NO_3)_2$, and $AgNO_3$. All will be soluble because, according to the guidelines, all nitrates are soluble. But if the metals were combined with chloride ions (that is, NaCl, $PbCl_2$, or AgCl), only NaCl will dissolve. According to the guidelines, most chlorides are soluble, but there are some exceptions, such as $PbCl_2$ and AgCl.

When ionic compounds dissolve in water, they dissociate (that is, separate) into their ions.

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

$Pb(NO_3)_{2(s)} \rightarrow Pb^{2+}_{(aq)} + 2NO_3^-{}_{(aq)}$

In a solution, the water is the solvent that dissolves the solute (here, the ionic compound). In an unsaturated solution, the solvent can dissolve more solute. But in a saturated solution, the solvent has dissolved the maximum quantity of solute at that temperature. If you add more solute, the solute will remain undissolved at the bottom of the container. In a saturated solution, equilibrium is established. The equilibrium is between the ions that remain associated (the undissolved solute on the bottom), and the ions that have been dissolved in the solvent. As a result, equations that show the formation of saturated solutions are correctly written as:

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

$Pb(NO_3)_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2NO_3^-{}_{(aq)}$

These two ionic equations are examples of dynamic equilibrium, as indicated by the double-headed arrows. You will observe that the pile of solid on the bottom of the container, whether the table salt or the lead (II) nitrate, remains constant. If you could actually see the ions when the solution has reached equilibrium, you would see that ions are leaving the solid at the same rate that ions are returning to the solid. The forward rate is equal to the reverse rate. The pile of solid on the bottom does not change in size.

Support Questions

Be sure to try the Support Questions on your own before looking at the suggested answers provided.

1. Which of the following ionic compounds would have low solubility?

$Al(OH)_3$	$AgClO_3$	$BaSO_4$	$CaCO_3$	$CuSO_4$
CuS	PtI_3	NH_4OH	$AlPO_4$	$NaHCO_3$

Suggested answer

2. Write ionic equations for saturated solutions of each of the following ionic compounds:

a) Silver chromate $Ag_2CrO_{4(s)}$

b) Barium phosphate $Ba_3(PO_4)_{2(s)}$

c) Aluminum hydroxide $Al(OH)_{3(s)}$

d) Copper (II) iodate $Cu(IO_3)_{2(s)}$

e) Calcium oxalate $CaC_2O_{4(s)}$

Suggested answer

Saturated Solutions and K_{sp} Expressions

A saturated solution is formed when increasing amounts of an ionic compound are dissolved in water until no more of the solid will dissolve and equilibrium is reached. For instance, you can add a pinch of table salt to a glass of water and then stir the water to dissolve it, repeating the process over and over until the salt begins piling up at the bottom of the glass. At this point, the solution has become saturated, so no additional salt will dissolve.

An equilibrium expression for a saturated solution can be written following the same rules that were used in previous lessons. Solids are not included in these equilibrium expressions; only the aqueous or dissolved ions (aq) are included. The solid (concentration) does not form part of the equilibrium expression because its concentration does not vary.

The equilibrium constant for saturated solutions is known as the K_{sp} or **s**olubility **p**roduct constant. K_{sp} values, like K_{eq} values, have no units.

K_{sp} expressions are written from balanced ionic equations, as shown in the following examples:

$KOH_{(s)} \rightleftharpoons K^+_{(aq)} + OH^-_{(aq)}$

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

This equilibrium expression is formed in the same manner as all K expressions, [products]/[reactants]. In this case the [products] is omitted because it does not vary in concentration. You can see where the terminology "solubility product" originates for this type of equilibrium. The expression is simply a multiplication (product) of the concentrations.

$Al_2(SO_4)_{3(s)} \rightleftharpoons 2Al^{3+}_{(aq)} + 3SO_4^{2-}{}_{(aq)}$

$K_{sp} = [Al^{3+}]^2[SO_4^{2-}]^3$

Complete tables of K_{sp} values for various ionic compounds are included in chemistry textbooks and can be found on the Internet. The complete tables only list ionic compounds that have low solubility. A shortened K_{sp} table is shown on the next page. The table includes compounds that will be used in this course. The table specifies that the K_{sp} values have been measured at 25°C. As temperature increases, solubility generally increases. This causes the K_{sp} values to increase because there will be a higher concentration of ions in the solution. The smaller the K_{sp} values are, the less soluble the ionic compounds will be.

Table 17.1: Solubility product constants for common ionic compounds

K_{sp} values measured at 25°C		
Nickel (II) hydroxide	$Ni(OH)_2$	1.6×10^{-16}
Calcium oxalate	CaC_2O_4	2.3×10^{-9}
Iron (III) hydroxide	$Fe(OH)_3$	2.6×10^{-39}
Silver iodide	AgI	8.5×10^{-17}
Lead (II) iodide	PbI_2	8.8×10^{-9}
Lead (II) sulfate	$PbSO_4$	1.8×10^{-8}
Calcium fluoride	CaF_2	3.5×10^{-11}
Calcium hydroxide	$Ca(OH)_2$	5.0×10^{-6}
Silver iodate	$AgIO_3$	3.2×10^{-8}
Silver chromate	Ag_2CrO_4	1.1×10^{-12}
Magnesium hydroxide	$Mg(OH)_2$	5.6×10^{-12}
Lead (II) chloride	$PbCl_2$	1.2×10^{-5}

Notice that all of these *K* values are very small, suggesting that at equilibrium, we will find very little of the product (dissolved) ions in the system. K_{sp} values are listed only for chemicals with very low solubilities. In most situations, these chemicals would be considered "insoluble" for general purposes.

Support Questions

3. Write K_{sp} expressions for each of the following ionic compounds. Use the balanced ionic equations that you wrote in the previous set of Support Questions.

a) Ag_2CrO_4

b) $Ba_3(PO_4)_2$

c) $Al(OH)_3$

d) $Cu(IO_3)_2$

e) CaC_2O_4

Suggested answer

4. How does the K_{sp} expression differ from most *K* expressions written earlier? Why is that?

Suggested answer

Calculating Concentration and K_{sp}

Molar solubility is the term that is used to describe the concentration of the solute in a saturated solution, which is the number of moles of solute dissolved in one litre of solution. Because the ionic compound has dissolved into ions in the solution, it is also possible to state the concentration of the ions as well.

For example, if 0.50 mol of KOH have been dissolved in 1.0 L of water, then the concentration of KOH in the solution will be 0.50 mol/L. To describe this solution, we write $[KOH_{(aq)}]$ = 0.50 mol/L. Note that this is describing the dissolved KOH, not any KOH that is lying on the bottom of the container. Even though we refer to the dissolved KOH, in fact that KOH exists in solution as dissociated ions, not KOH together (associated).

The balanced ionic equation is:

$KOH_{(s)} \rightleftharpoons K^+_{(aq)} + OH^-_{(aq)}$

The coefficients in the balanced equation are all equal to 1. The concentration of each individual ion will be the same as the overall compound's concentration:

$[K^+]$ = 0.50 mol/L

$[OH^-]$ = 0.50 mol/L

Consider a different ionic compound. If a solution of aluminum sulfate also has the same concentration, 0.50 mol/L, what are the concentrations of the individual ions? Again, you have to look at the balanced ionic equation:

$Al_2(SO_4)_{3(s)} \rightleftharpoons 2Al^{3+}_{(aq)} + 3SO_4{}^{2-}_{(aq)}$

The balanced chemical equation tells us that 1 mole of aluminum sulfate creates 2 moles of aluminum ion and 3 moles of sulfate ion. Therefore if we have 0.50 mol/L of dissolved aluminum sulfate, this will create:

$[Al^{3+}]$ = 1.00 mol/L (2 times 0.50 mol/L)

$[SO_4{}^{2-}]$ = 1.50 mol/L (3 times 0.50 mol/L)

You are now going to look at four different examples. Each one has a slightly different focus, but in each case, you will be following a similar strategy to the one in the previous two lessons. The following steps should be followed when you are asked to find concentration or K_{sp}.

The initial–change–equilibrium (ICE) strategy

1. Write the balanced ionic equation for the K_{sp} expression for the saturated solution given in the question.
2. Record the initial concentrations using an ICE chart.
3. Use the balanced equation and the ICE chart to determine the changes in concentrations of products (dissolved ions).
4. Substitute the equilibrium concentrations into the K_{sp} expression.
5. Solve for unknown values.

Example

The concentration of a saturated solution (molar solubility) of PbI_2 is 1.3×10^{-3} mol/L. Determine the K_{sp} value of PbI_2.

Solution

Step 1:

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

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

The equation tells us that 1 mole of PbI_2 creates 1 mole of dissolved Pb^{2+} ion. The consequence of that is that $[Pb^{2+}_{(aq)}] = [PbI_{2(aq)}]$. If the molar solubility of lead iodide is given as 1.3×10^{-3} mol/L, this means that the concentration of the dissolved lead iodide in the saturated solution is $[PbI_{2(aq)}] = 1.3 \times 10^{-3}$ mol/L. We now can say that, in this saturated solution that is at equilibrium,

$[Pb^{2+}_{(aq)}]$ will be 1.3×10^{-3} mol/L

Note this value in our ICE table below.

Steps 2 and 3:

ICE	$[Pb^{2+}]$ in mol/L	$[I^{-}]$ in mol/L
Initial	0	0
Change	$+x$	$+2x$
Equilibrium	1.3×10^{-3}	2.6×10^{-3}

The initial concentration of the ions is recorded as 0. The ionic compound has just been added to the water and has yet to start to dissolve. When the maximum amount of this ionic compound has been dissolved, a saturated solution will have been created.

The x and the $2x$ are determined by coefficients in the balanced ionic equation. The x and $2x$ just reflect the relative concentrations of the lead and iodide ions in the solution as being 1:2 (mole ratio). We know from our previous discussion that the concentration of the lead ion will be 1.3×10^{-3} mol/L.

Steps 4 and 5:

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

$K_{sp} = (1.3 \times 10^{-3})(2.6 \times 10^{-3})^2$

$K_{sp} = 8.8 \times 10^{-9}$

The K_{sp} for PbI_2 is 8.8×10^{-9}.

Example

What is the molar solubility of nickel (II) hydroxide?

Solution

In this example, we will determine the molar solubility using the known K_{sp} value from our table. Molar solubility means $[Ni(OH)_{2(aq)}]$. The balanced equation below shows us that $[Ni(OH)_{2(aq)}]$ will be the same as $[Ni^{2+}{}_{(aq)}]$ (they are 1:1 in the equation). That ion concentration will be denoted as x on our ICE table.

Step 1:

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

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

Steps 2 and 3:

ICE	$[Ni^{2+}]$ in mol/L	$[OH^{-}]$ in mol/L
Initial	0	0
Change	$+ x$	$+ 2x$
Equilibrium	x	$2x$

Steps 4 and 5:

Since you are not given the K_{sp} value, look it up in Table 17.1. You will see that it is 1.6×10^{-16}.

$$K_{sp} = 1.6 \times 10^{-16}$$

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

$$1.6\times10^{-16} = 4x^3$$

$$x^3 = \frac{1.6\times10^{-16}}{4}$$

$$x^3 = 4.0\times10^{-17}$$

$$x = \sqrt[3]{4.0\times10^{-17}}$$

$$x = 3.4\times10^{-6}$$

Since x represented the concentration of the nickel ion, then we can conclude that the molar solubility of $Ni(OH)_2$ is 3.4×10^{-6} mol/L.

Example

In a saturated solution of silver carbonate, Ag_2CO_3, 0.0178 g of solute are dissolved in 500.mL of solution. Calculate the K_{sp} of silver carbonate.

Solution

Step 1:

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

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

Steps 2 and 3:

ICE	$[Ag^+]$ in mol/L	$[CO_3^{2-}]$ in mol/L
Initial	0	0
Change	$+ 2x$	$+ x$
Equilibrium	2.58×10^{-4}	1.29×10^{-4}

This problem becomes the same as the first example, where the molar solubility was given. They tell us that a *saturated* solution was created here when 0.0178 g was dissolved in 500. mL (0.500 L). Whatever concentration that creates must be the maximum that can be dissolved (the molar solubility). So determine what concentration has been created when this solution was created:

The molar mass of Ag_2CO_3 is 275.8 g/mol = M (using the periodic table).

Converting 0.0178 g to moles: $n = \frac{m}{M}$ $\quad n = \frac{0.0178g}{275.8g/mol} = 0.0000645$ mol

$$[Ag_2CO_{3(aq)}] = \frac{n}{V} = \frac{0.0000645\,mol}{0.500L} = 1.29 \times 10^{-4} \text{ mol/L}$$

The balanced equation in Step 1 tells us that 1 mole of Ag_2CO_3 creates 1 mole of CO_3^{2-} ion. The CO_3^{2-} ion created will be denoted as *x* on our ICE table. Again, the 2*x* and the *x* just remind us that these ions are created in a mole ratio, 2:1. We know that if the $[Ag_2CO_{3(aq)}]$ is 1.29×10^{-4} mol/L, then $[CO_3^{2-}]$ will also be 1.29×10^{-4} mol/L. Study the values in the ICE table to follow what has been said.

The *x* and the 2*x* are determined by coefficients in the balanced ionic equation. The *x* is always equal to the molar solubility, in this case, 1.29×10^{-4} mol/L.

Steps 4 and 5:

$K_{sp} = (2.58 \times 10^{-4})^2(1.29 \times 10^{-4}) = 8.59 \times 10^{-12}$

The K_{sp} of silver carbonate is 8.59×10^{-12}.

Support Questions

Solve the following problems. Your solutions should be in the proper format and your final answers should have the correct units and number of significant figures.

5. A saturated solution of zinc sulfide has a concentration of 4.47×10^{-13} mol/L. Determine the K_{sp} of zinc sulfide. Suggested answer

6. Calculate the concentration of a saturated solution (molar solubility) of iron (III) hydroxide. Suggested answer

7. If 0.0161 g of strontium fluoride dissolve in 125 mL of water to form a saturated solution, determine the K_{sp} of strontium fluoride. Suggested answer

Predicting Precipitates

Figure 17.1: The Brook Bottom Calcareous precipitate formed as a result off the lime-rich water seeping through a mound of burnt lime that has filled the valley.

Source: http://upload.wikimedia.org/wikipedia/commons/0/03/Brook_Bottom_Calcareous_precipitate_-_geograph.org.uk_-_456623.jpg

Precipitates are solids that form when two solutions are mixed together and a double displacement reaction occurs. The ionic compounds that are found in the original solutions are soluble. When the solutions are mixed together, a combination of ions creates an ionic compound that has low solubility. The low solubility compound is the precipitate, which collects at the bottom of the test tube or reaction vessel. The precipitate forms because the two ions that make up the low-solubility compound exist in the solution at too high a concentration. They are too close to each other, so they join together and form a solid.

You can use the solubility guidelines to identify the potential precipitate, but does a precipitate always form? If the concentrations of the ions that form the precipitate are sufficiently low, then no precipitate will form. This means that the ions are not too close together in the mixture created.

Consider the following double displacement reaction and note the location of the two question marks:

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

This reaction goes to completion so the forward arrow $\rightarrow$ and not the double-headed arrow $\rightleftharpoons$ is used. The solubility guidelines indicate that PbI_2 has low solubility, so the equation is correctly completed as:

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

A very small amount of the lead iodide will remain dissolved as a saturated solution of PbI_2. Equilibrium is established between the solid precipitate lying on the bottom and the dissolved ions in the solution. That equilibrium is described:

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

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

From the tables, $K_{sp} = 8.8 \times 10^{-9}$

How do you know if a precipitate will form or not? A precipitate forms if the concentrations of the ions, when substituted into the K_{sp} expression, produce a value that is equal to or greater than the K_{sp}.

In Lesson 15, you used Q as a trial, to see if a reaction had reached equilibrium. You will now use Q again, to see if a precipitate has formed.

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

If $Q > K_{sp}$, a precipitate forms When Q is greater than the K_{sp} value, it simply means that the ions in question are too crowded together or too close together in the solution. The solvent (water) will be unable to prevent them from joining together and forming the solid precipitate. If $Q < K_{sp}$, there is no precipitate.

You need to know the concentration and volume of the two original solutions in order to determine the $[Pb^{2+}]$ and the $[I^{-}]$.

Imagine that you mixed 40.0 mL of 0.250 mol/L $Pb(NO_3)_{2(aq)}$ with 60.0 mL of 0.300 mol/L $KI_{(aq)}$.

Consider the lead nitrate solution before it was mixed with the other solution. The dissociation of $Pb(NO_3)_{2(aq)}$ is:

$Pb(NO_3)_{2(aq)} \rightarrow Pb^{2+}{}_{(aq)} + 2NO_3{}^{-}{}_{(aq)}$

We are told that in this lead nitrate solution, $[Pb(NO_3)_{2(aq)}] = 0.250$ mol/L. The dissociation equation shows us that 1 mole of $Pb(NO_3)_2$ yields 1 mole of the Pb^{2+} ion. Therefore, the $[Pb^{2+}{}_{(aq)}]$ must also be 0.250 mol/L *before* this solution is mixed with the second solution. Our thinking will be restricted to just the Pb^{2+} ion and the I^{-} ion (from the second solution), since it is these two ions that could potentially join together and form a precipitate. The NO_3– ion and the K^{+} ions are of no consequence since these always form soluble compounds.

Let's now look carefully at the second solution, and at the I^{-} ion it contains. The dissociation of KI is written as:

$KI_{(aq)} \rightarrow K^{+}{}_{(aq)} + I^{-}{}_{(aq)}$

We are told that in the KI solution, $[KI_{(aq)}]$ is 0.300 mol/L. Therefore the $[I^{-}{}_{(aq)}]$ will also be 0.300 mol/L.

Summarizing: We will mix two solutions together.

Solution 1 is 40 mL and has $[Pb^{2+}{}_{(aq)}] = 0.250$ mol/L

Solution 2 is 60 mL and has $[I^{-}{}_{(aq)}] = 0.300$ mol/L

When the two solutions are mixed together, the increase in volume (to 100 mL) causes a dilution to each ion concentration. The concentration of each ion must be calculated when the two solutions are combined.

Doing this for the lead ion: the volume changes from $V_1 = 40$ mL to $V_2 = 100$ mL. The concentration changes from $C_1 = 0.250$ mol/L to C_2. Since the number of moles of lead ion (C×V) does not change when it is diluted,

$$C_1V_1 = C_2V_2$$

$$C_2 = \frac{C_1V_1}{V_2}$$

$$[Pb^{2+}] = \frac{0.250 \text{ mol/L} \times 40.0 \text{ mL}}{100 \text{ mL}} = 0.100 \text{ mol/L}$$

and

$$[I^-] = \frac{0.300 \text{ mol/L} \times 60.0 \text{ mL}}{100 \text{ mL}} = 0.180 \text{ mol/L}$$

Substitute the concentrations into the Q expression to see if a precipitate forms.

$Q = (0.100)(0.180)^2 = 3.24 \times 10^{-3}$.

$Q > K_{sp}$, so a precipitate is formed.

The following is a summary of the strategy used to determine if a precipitate will form:

Step 1: Write out the balanced double displacement reaction and identify the potential precipitate, using solubility guidelines.

Step 2: Write out the equilibrium reaction for the precipitate and the Q expression.

Step 3: Look up the K_{sp} of the potential precipitate.

Step 4: Calculate the concentrations of the two ions after the two solutions have been mixed, and substitute the concentrations into the Q expression.

Step 5: Compare Q value and K_{sp} value.

Support Questions

8. How is Q the same as K_{sp}? Why do we evaluate the Q expression and what does it tell us?

Suggested answer

9. Solve the following problems. Your solutions should be in the proper format and your final answers should have the correct units and number of significant figures.

a) Will a precipitate form when 20.0 mL of $Pb(NO_3)_{2(aq)}$ with a concentration of 2.50×10^{-4} mol/L is reacted with 45.0 mL of $Na_2SO_{4(aq)}$ with a concentration of 1.25×10^{-3} mol/L?

b) Will a precipitate form when 60.0 mL of $NaF_{(aq)}$ with a concentration of 8.50×10^{-5} mol/L is reacted with 85.0 mL of $CaCl_{2(aq)}$ with a concentration of 1.25×10^{-4} mol/L?

Suggested answer

The Common Ion Effect

Figure 17.2: Barbara Askins, NASA chemist

Source: http://upload.wikimedia.org/wikipedia/commons/thumb/0/0c/Barbara_Askins%2C_Chemist_-_GPN-2004-00022.jpg/474px-Barbara_Askins%2C_Chemist_-_GPN-2004-00022.jpg

Whenever chemists prepare solutions, they use distilled water, instead of tap water, as the solvent. Tap water contains ions. Sometimes the ions in tap water are identical to the ions in the ionic compound being dissolved. Ions like these that are found in both the solute and the solvent are known as common ions. How does the presence of a common ion, such as chlorine, affect the molar solubility of an ionic compound? The common ion effect describes what occurs in a solution of two dissolved solutes that contain the same ion or ions.

Consider the molar solubility of AgCl in distilled water. The K_{sp} for AgCl is 1.8×10^{-10}.

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

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

ICE	$[Ag^{+}]$ in mol/L	$[Cl^{-}]$ in mol/L
Initial	0	0
Change	$+x$	$+x$
Equilibrium	x	x

$$1.8 \times 10^{-10} = (x)(x)$$

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

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

The molar solubility of silver chloride is 1.3×10^{-5} mol/L, when the solution is prepared using distilled water.

But what if tap water, having a concentration of chloride ions of 1.0×10^{-4} (0.00010 mol/L), was used? Would the molar solubility be the same? Would the same amount of silver chloride dissolve before reaching saturation?

ICE	$[Ag^+]$ in mol/L	$[Cl^-]$ in mol/L
Initial	0	0.00010
Change	$+x$	$+x$
Equilibrium	x	$0.00010 + x$
Revised equilibrium	x	0.00010

The K_{sp} expression for this system is

$K_{sp} = [Ag^+_{(aq)}][Cl^-_{(aq)}]$. Using the equilibrium concentrations shown in the ICE table (before the revision), we can write

$1.8 \times 10^{-10} = x(0.00010 + x)$

We could solve this equation for x, but because it is a quadratic equation, we should consider whether we can make a simplifying assumption. The K for this system is very small, so we can expect that the dissolved silver concentration, x, will be very small. We would expect that x will be much smaller than 0.00010 in value. If that is so, then $0.00010 + x$ can be assumed to be just 0.00010. This simplifies our equation to:

$1.8 \times 10^{-10} = x(0.00010)$

The “100 rule” seen earlier will also confirm this simplification. Applying the 100 rule:

$$\frac{[original]}{K_{sp}} = \frac{0.00010 \text{ mol/L}}{1.8 \times 10^{-10}} = 55555.6 > 100$$

$1.8 \times 10^{-10} = x(0.00010)$

$$x = \frac{1.8 \times 10^{-10}}{0.00010} = 1.8 \times 10^{-6}$$

Since x represents the concentration of silver ion that was created by the addition of the AgCl, the concentration of AgCl that must have been added must be the same value. The molar solubility of silver chloride in tap water is 1.8×10^{-6} mol/L, which is lower than it is in distilled water.

Support Questions

10. Why should solutions, both in laboratory and hospital settings, always be prepared with distilled water?

Suggested answer

11. Solve the following problems. Your solutions should be in the proper format and your final answers should have the correct units and number of significant figures.

a) What is the maximum concentration of a silver iodate solution prepared with 1.00 L of distilled water? (**Hint:** This is another way of asking, "What is the solubility?")

b) What is the maximum concentration of a silver iodate solution prepared with 1.00 L of 0.014 mol/L calcium iodate solution? (Hint: This is another way of asking, "What is the molar solubility?") Suggested answer

12. What mass of $Ca(OH)_2$ will dissolve in 1.0 L of 0.030 mol/L solution of $NaOH_{(aq)}$? (**Hint:** The common ion is OH. Find the solubility and change to grams.) Suggested answer

Determining K_{sp} Experimentally

You will now read about an experiment, but you will not be asked to do the experiment yourself. You will be calculating the K_{sp} of silver chromate, using a single-displacement reaction between a saturated solution of silver chromate and copper metal. The procedure you will be following is listed below.

Materials:

- a one-litre graduated cylinder
- a bottle of saturated silver chromate solution
- a piece of copper, such as a thick copper wire

Procedure:

Step 1: Using a graduated cylinder, carefully measure 1.00 L of saturated silver chromate solution from the bottle and pour it into a clean, dry beaker.

Step 2: Clean the piece of copper with steel wool and then measure its mass using the electronic balance, and record the mass in the data table.

Step 3: Place the piece of copper in the beaker of silver chromate solution and allow it to react over night.

Step 4: Clean the dark residue off the piece of copper using steel wool, and then remeasure the piece of copper's mass.

Data table:

If you had done this experiment, this is what you would have observed.

Mass of copper before reaction (g)	Mass of copper after reaction (g)
1.250	1.240

Support Questions

13. Answer the following questions based on the experiment you have just read.

a) What observation would you make that would show that the solution of silver chromate in the bottle was saturated at the beginning of the experiment?

b) Write the balanced chemical equation for this reaction.

c) Why was the copper cleaned before starting the experiment?

d) What was the dark residue found on the piece of copper, at the end of the reaction?

e) Why did the mass of copper decrease?

f) What mass of copper reacted? Convert this to moles. Referring to the balanced equation, how many moles of silver chromate must have reacted as well?

g) Was all of the copper consumed in the reaction? What was the limiting reactant? Explain what "limiting reactant" means.

h) What is the accepted K_{sp} for silver chromate?

Suggested answer

In the Key Questions, you will continue this analysis and determine the experimental value of K_{sp} for silver chromate.

Key Questions

Now work on your Key Questions in the online submission tool. You may continue to work at this task over several sessions, but be sure to save your work each time. When you have answered all the unit's Key Questions, submit your work to the ILC.

57. In order to answer the following questions, you need to use information from this lesson that you were given in "Determining K_{sp} Experimentally" and in your answers to the Support Questions.

- **a)** Calculate the K_{sp} of Ag_2CrO_4. Organize your answer so that your solution is logical. Make sure that your K_{sp} has the correct number of significant figures. **(7 marks)**
- **b)** Calculate your percentage error, using the following equation. (You first used this equation in Lesson 11.) **(2 marks)**

$$\%\ error = \frac{(experimental - theoretical)}{theoretical} \times 100\%$$

- **c)** Describe two sources of error that could have been present while doing the experiment "Determining K_{sp} Experimentally." **(2 marks)**

58. Solve the following problems. Your solutions should be in the proper format and your final answers should have the correct units and number of significant figures.

- **a)** What mass of $Mg(OH)_2$ will dissolve in 1.0 L of 0.050 mol/L solution of $MgSO_{4(aq)}$? **(9 marks)**
- **b)** Will a precipitate form when 20.0 mL of $Pb(NO_3)_{2(aq)}$ with a concentration of 2.50×10^{-4} mol/L is mixed with 45.0 mL of $CaCl_{2(aq)}$ with a concentration of 1.25×10^{-3} mol/L? **(8 marks)**

Now go on to Lesson 18. Send your answers to the Key Questions to the ILC when you have completed Unit 5 (Lessons 17 to 20).