

17.2: Determining the Solubility of Ionic Compounds

Learning Objectives

- To calculate the solubility of an ionic compound from its K_{sp} .

We begin our discussion of solubility and *complexation equilibria*—those associated with the formation of complex ions—by developing quantitative methods for describing dissolution and precipitation reactions of ionic compounds in aqueous solution. Just as with acid–base equilibria, we can describe the concentrations of ions in equilibrium with an ionic solid using an equilibrium constant expression.

The Solubility Product

When a slightly soluble ionic compound is added to water, some of it dissolves to form a solution, establishing an equilibrium between the pure solid and a solution of its ions. For the dissolution of calcium phosphate, one of the two main components of kidney stones, the equilibrium can be written as follows, with the solid salt on the left: As you will discover in [Section 17.4](#) and in more advanced chemistry courses, basic anions, such as S^{2-} , PO_4^{3-} , and CO_3^{2-} , react with water to produce OH^- and the corresponding protonated anion. Consequently, their calculated molarities, assuming no protonation in aqueous solution, are only approximate.



The equilibrium constant for the dissolution of a sparingly soluble salt is the solubility product (K_{sp}). The equilibrium constant expression for the dissolution of a sparingly soluble salt that includes the concentration of a pure solid, which is a constant, of the salt. Because the concentration of a pure solid such as $Ca_3(PO_4)_2$ is a constant, it does not appear explicitly in the equilibrium constant expression. (For more information on the equilibrium constant expression, see [Section 15.2](#).) The equilibrium constant expression for the dissolution of calcium phosphate is therefore

$$K = \frac{[Ca^{2+}]^3 [PO_4^{3-}]^2}{[Ca_3(PO_4)_2]} \quad (17.1.2a)$$

$$[Ca_3(PO_4)_2]K = K_{sp} = [Ca^{2+}]^3 [PO_4^{3-}]^2 \quad (17.1.2b)$$

At 25°C and pH 7.00, K_{sp} for calcium phosphate is 2.07×10^{-33} , indicating that the concentrations of Ca^{2+} and PO_4^{3-} ions in solution that are in equilibrium with solid calcium phosphate are very low. The values of K_{sp} for some common salts are listed in [Table 17.1.1](#); they show that the magnitude of K_{sp} varies dramatically for different compounds. Although K_{sp} is not a function of pH in [Equation 17.1](#), changes in pH can affect the solubility of a compound, as you will discover in [Section 17.4](#).

Note the Pattern

As with K , the concentration of a pure solid does not appear explicitly in K_{sp} .

Table 17.1.1 Solubility Products for Selected Ionic Substances at 25°C

Solid	Color	K_{sp}	Solid	Color	K_{sp}
Acetates			Iodides		
$Ca(O_2CCH_3)_2 \cdot 3H_2O$	white	4×10^{-3}	$Hg_2I_2^*$	yellow	5.2×10^{-29}
Bromides			PbI_2	yellow	9.8×10^{-9}
$AgBr$	off-white	5.35×10^{-13}	Oxalates		
$Hg_2Br_2^*$	yellow	6.40×10^{-23}	$Ag_2C_2O_4$	white	5.40×10^{-12}
Carbonates			$MgC_2O_4 \cdot 2H_2O$	white	4.83×10^{-6}

*These contain the Hg_2^{2+} ion.

Solid	Color	K_{sp}	Solid	Color	K_{sp}
CaCO_3	white	3.36×10^{-9}	PbC_2O_4	white	4.8×10^{-10}
PbCO_3	white	7.40×10^{-14}	Phosphates		
Chlorides			Ag_3PO_4	white	8.89×10^{-17}
AgCl	white	1.77×10^{-10}	$\text{Sr}_3(\text{PO}_4)_2$	white	4.0×10^{-28}
Hg_2Cl_2^*	white	1.43×10^{-18}	$\text{FePO}_4 \cdot 2\text{H}_2\text{O}$	pink	9.91×10^{-16}
PbCl_2	white	1.70×10^{-5}	Sulfates		
Chromates			Ag_2SO_4	white	1.20×10^{-5}
CaCrO_4	yellow	7.1×10^{-4}	BaSO_4	white	1.08×10^{-10}
PbCrO_4	yellow	2.8×10^{-13}	PbSO_4	white	2.53×10^{-8}
Fluorides			Sulfides		
BaF_2	white	1.84×10^{-7}	Ag_2S	black	6.3×10^{-50}
PbF_2	white	3.3×10^{-8}	CdS	yellow	8.0×10^{-27}
Hydroxides			PbS	black	8.0×10^{-28}
$\text{Ca}(\text{OH})_2$	white	5.02×10^{-6}	ZnS	white	1.6×10^{-24}
$\text{Cu}(\text{OH})_2$	pale blue	1×10^{-14}			
$\text{Mn}(\text{OH})_2$	light pink	1.9×10^{-13}			
$\text{Cr}(\text{OH})_3$	gray-green	6.3×10^{-31}			
$\text{Fe}(\text{OH})_3$	rust red	2.79×10^{-39}			

*These contain the Hg_2^{2+} ion.

Solubility products are determined experimentally by directly measuring either the concentration of one of the component ions or the solubility of the compound in a given amount of water. However, whereas solubility is usually expressed in terms of m

mass of solute per 100 mL of solvent, K_{sp} , like K , is defined in terms of the *molar* concentrations of the component ions.

A kidney stone. Kidney stones form from sparingly soluble calcium salts and are largely composed of $\text{Ca}(\text{O}_2\text{CCO}_2) \cdot \text{H}_2\text{O}$ and $\text{Ca}_3(\text{PO}_4)_2$.

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Example 17.1.1

Calcium oxalate monohydrate [$\text{Ca}(\text{O}_2\text{CCO}_2) \cdot \text{H}_2\text{O}$, also written as $\text{CaC}_2\text{O}_4 \cdot \text{H}_2\text{O}$] is a sparingly soluble salt that is the other major component of kidney stones [along with $\text{Ca}_3(\text{PO}_4)_2$]. Its solubility in water at 25°C is 7.36×10^{-4} g/100 mL. Calculate its K_{sp} .

Given: solubility in g/100 mL

Asked for: K_{sp}

Strategy:

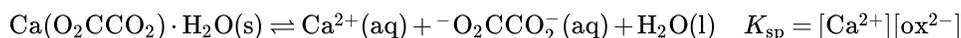
A Write the balanced dissolution equilibrium and the corresponding solubility product expression.

B Convert the solubility of the salt to moles per liter. From the balanced dissolution equilibrium, determine the equilibrium concentrations of the dissolved solute ions. Substitute these values into the solubility product expression to calculate K_{sp} .

Solution:



A We need to write the solubility product expression in terms of the concentrations of the component ions. For calcium oxalate monohydrate, the balanced dissolution equilibrium and the solubility product expression (abbreviating oxalate as ox^{2-}) are as follows:



Neither solid calcium oxalate monohydrate nor water appears in the solubility product expression because their concentrations are essentially constant.

B Next we need to determine $[\text{Ca}^{2+}]$ and $[\text{ox}^{2-}]$ at equilibrium. We can use the mass of calcium oxalate monohydrate that dissolves in 100 mL of water to calculate the number of moles that dissolve in 100 mL of water. From this we can determine the number of moles that dissolve in 1.00 L of water. For dilute solutions, the density of the solution is nearly the same as that of water, so dissolving the salt in 1.00 L of water gives essentially 1.00 L of solution. Because each 1 mol of dissolved calcium oxalate monohydrate dissociates to produce 1 mol of calcium ions and 1 mol of oxalate ions, we can obtain the equilibrium concentrations that must be inserted into the solubility product expression. The number of moles of calcium oxalate monohydrate that dissolve in 100 mL of water is as follows:

$$\frac{7.36 \times 10^{-4} \text{ g}}{146.1 \text{ g/mol}} = 5.04 \times 10^{-6} \text{ mol Ca}(\text{O}_2\text{CCO}_2) \cdot \text{H}_2\text{O}$$

The number of moles of calcium oxalate monohydrate that dissolve in 1.00 L of the saturated solution is as follows:

$$\left(\frac{5.04 \times 10^{-6} \text{ mol Ca}(\text{O}_2\text{CCO}_2) \cdot \text{H}_2\text{O}}{100 \text{ mL}} \right) \left(\frac{1000 \text{ mL}}{1.00 \text{ L}} \right) = 5.04 \times 10^{-5} \text{ mol/L} = 5.04 \times 10^{-5} \text{ M}$$

Because of the stoichiometry of the reaction, the concentration of Ca^{2+} and ox^{2-} ions are both $5.04 \times 10^{-5} \text{ M}$. Inserting these values into the solubility product expression,

$$K_{\text{sp}} = [\text{Ca}^{2+}][\text{ox}^{2-}] = (5.04 \times 10^{-5})(5.04 \times 10^{-5}) = 2.54 \times 10^{-9}$$

In our calculation, we have ignored the reaction of the weakly basic anion with water, which tends to make the actual solubility of many salts greater than the calculated value.

Exercise

One crystalline form of calcium carbonate (CaCO_3) is the mineral sold as “calcite” in mineral and gem shops. The solubility of calcite in water is 0.67 mg/100 mL. Calculate its K_{sp} .

Answer: 4.5×10^{-9}

Note the Pattern

The reaction of weakly basic anions with H_2O tends to make the actual solubility of many salts higher than predicted.



A crystal of calcite (CaCO_3), illustrating the phenomenon of double refraction. When a transparent crystal of calcite is placed over a page, we see two images of the letters.

Calcite, a structural material for many organisms, is found in the teeth of sea urchins. The urchins create depressions in limestone that they can settle in by grinding the rock with their teeth. Limestone, however, also consists of calcite, so how can the urchins grind the rock without also grinding their teeth? Researchers have discovered that the teeth are shaped like needles and plates and contain magnesium. The concentration of magnesium increases toward the tip, which contributes to the hardness. Moreover, each tooth is composed of two blocks of the polycrystalline calcite matrix that are interleaved near the tip. This creates a corrugated surface that presumably increases grinding efficiency. Toolmakers are particularly interested in this approach to grinding.

Tabulated values of K_{sp} can also be used to estimate the solubility of a salt with a procedure that is essentially the reverse of the one used in Example 1. In this case, we treat the problem as a typical equilibrium problem and set up a table of initial concentrations, changes in concentration, and final concentrations as we did in [Chapter 15](#), remembering that the concentration of the pure solid is essentially constant.

Example 17.1.2

We saw that the K_{sp} for $\text{Ca}_3(\text{PO}_4)_2$ is 2.07×10^{-33} at 25°C . Calculate the aqueous solubility of $\text{Ca}_3(\text{PO}_4)_2$ in terms of the following:

1. the molarity of ions produced in solution
2. the mass of salt that dissolves in 100 mL of water at 25°C

Given: K_{sp}

Asked for: molar concentration and mass of salt that dissolves in 100 mL of water

Strategy:

A Write the balanced equilibrium equation for the dissolution reaction and construct a table showing the concentrations of the species produced in solution. Insert the appropriate values into the solubility product expression and calculate the molar solubility at 25°C .

B Calculate the mass of solute in 100 mL of solution from the molar solubility of the salt. Assume that the volume of the solution is the same as the volume of the solvent.

Solution:

1. **A** The dissolution equilibrium for $\text{Ca}_3(\text{PO}_4)_2$ (Equation 17.1) is shown in the following table. Because we are starting with distilled water, the initial concentration of both calcium and phosphate ions is zero. For every 1 mol of $\text{Ca}_3(\text{PO}_4)_2$ that dissolves, 3 mol of Ca^{2+} and 2 mol of PO_4^{3-} ions are produced in solution. If we let x equal the solubility of $\text{Ca}_3(\text{PO}_4)_2$ in moles per liter, then the change in $[\text{Ca}^{2+}]$ will be $+3x$, and the change in $[\text{PO}_4^{3-}]$ will be $+2x$. We can insert these values into the table.

$\text{Ca}_3(\text{PO}_4)_2(\text{s}) \rightleftharpoons 3\text{Ca}^{2+}(\text{aq}) + 2\text{PO}_4^{3-}(\text{aq})$			
	$\text{Ca}_3(\text{PO}_4)_2$	$[\text{Ca}^{2+}]$	$[\text{PO}_4^{3-}]$
initial	pure solid	0	0
change	—	$+3x$	$+2x$
final	pure solid	$3x$	$2x$

Although the *amount* of solid $\text{Ca}_3(\text{PO}_4)_2$ changes as some of it dissolves, its *molar concentration* does not change. We now insert the expressions for the equilibrium concentrations of the ions into the solubility product expression (Equation 17.1.2):

$$\begin{aligned}
 K_{sp} &= [\text{Ca}^{2+}]^3 [\text{PO}_4^{3-}]^2 = (3x)^3 (2x)^2 \\
 2.07 \times 10^{-33} &= 108x^5 \\
 1.92 \times 10^{-35} &= x^5 \\
 1.14 \times 10^{-7} \text{ M} &= x
 \end{aligned}$$

This is the molar solubility of calcium phosphate at 25°C . However, the molarity of the ions is $2x$ and $3x$, which means that $[\text{PO}_4^{3-}] = 2.28 \times 10^{-7}$ and $[\text{Ca}^{2+}] = 3.42 \times 10^{-7}$.

2. **B** To find the mass of solute in 100 mL of solution, we assume that the density of this dilute solution is the same as the density of water because of the low solubility of the salt, so that 100 mL of water gives 100 mL of solution. We can then determine the amount of salt that dissolves in 100 mL of water:

$$\left(\frac{1.14 \times 10^{-7} \text{ mol}}{1 \text{ L}} \right) 100 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{310.18 \text{ g Ca}_3(\text{PO}_4)_2}{1 \text{ mol}} \right) = 3.54 \times 10^{-6} \text{ g Ca}_3(\text{PO}_4)_2$$

Exercise

The solubility product of silver carbonate (Ag_2CO_3) is 8.46×10^{-12} at 25°C . Calculate the following:

1. the molarity of a saturated solution
2. the mass of silver carbonate that will dissolve in 100 mL of water at this temperature

Answer

1. $1.28 \times 10^{-4} \text{ M}$
2. 3.54 mg

The Ion Product

The ion product (Q) is a quantity that has precisely the same form as the solubility product for the dissolution of a sparingly soluble salt, except that the concentrations used are not necessarily equilibrium concentrations. It is analogous to the reaction quotient (Q) discussed for gaseous equilibria in [Chapter 15](#). Whereas K_{sp} describes equilibrium concentrations, the ion product describes concentrations that are *not* necessarily equilibrium concentrations.

Note the Pattern

The ion product Q is analogous to the reaction quotient Q for gaseous equilibria.

As summarized in [Figure 17.1.1](#), there are three possible conditions for an aqueous solution of an ionic solid:

1. $Q < K_{sp}$. The solution is unsaturated, and more of the ionic solid, if available, will dissolve.
2. $Q = K_{sp}$. The solution is saturated and at equilibrium.
3. $Q > K_{sp}$. The solution is supersaturated, and ionic solid will precipitate.

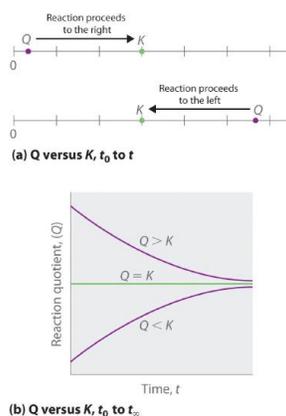


Figure 17.1.1 The Relationship between Q and K_{sp} . If Q is less than K_{sp} , the solution is unsaturated and more solid will dissolve until the system reaches equilibrium ($Q = K_{sp}$). If Q is greater than K_{sp} , the solution is supersaturated and solid will precipitate until $Q = K_{sp}$. If $Q = K_{sp}$, the rate of dissolution is equal to the rate of precipitation; the solution is saturated, and no net change in the amount of dissolved solid will occur.

The process of calculating the value of the ion product and comparing it with the magnitude of the solubility product is a straightforward way to determine whether a solution is unsaturated, saturated, or supersaturated. More important, the ion product tells chemists whether a precipitate will form when solutions of two soluble salts are mixed.

Example 17.1.3

We mentioned that barium sulfate is used in medical imaging of the gastrointestinal tract. Its solubility product is 1.08×10^{-10} at 25°C , so it is ideally suited for this purpose because of its low solubility when a “barium milkshake” is consumed by a patient. The pathway of the sparingly soluble salt can be easily monitored by x-rays. Will barium sulfate precipitate if 10.0 mL of 0.0020 M Na_2SO_4 is added to 100 mL of 3.2×10^{-4} M BaCl_2 ? Recall that NaCl is highly soluble in water.

Given: K_{sp} and volumes and concentrations of reactants

Asked for: whether precipitate will form

Strategy:

A Write the balanced equilibrium equation for the precipitation reaction and the expression for K_{sp} .

B Determine the concentrations of all ions in solution when the solutions are mixed and use them to calculate the ion product (Q).

C Compare the values of Q and K_{sp} to decide whether a precipitate will form.

Solution:

A The only slightly soluble salt that can be formed when these two solutions are mixed is BaSO_4 because NaCl is highly soluble. The equation for the precipitation of BaSO_4 is as follows:



The solubility product expression is as follows:

$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.08 \times 10^{-10}$$

B To solve this problem, we must first calculate the ion product— $Q = [\text{Ba}^{2+}][\text{SO}_4^{2-}]$ —using the concentrations of the ions that are present after the solutions are mixed and before any reaction occurs. The concentration of Ba^{2+} when the solutions are mixed is the total number of moles of Ba^{2+} in the original 100 mL of BaCl_2 solution divided by the final volume (100 mL + 10.0 mL = 110 mL):

$$\begin{aligned} \text{moles Ba}^{2+} &= 100 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{3.2 \times 10^{-4} \text{ mol}}{1 \text{ L}} \right) = 3.2 \times 10^{-5} \text{ mol Ba}^{2+} \\ [\text{Ba}^{2+}] &= \left(\frac{3.2 \times 10^{-5} \text{ mol Ba}^{2+}}{110 \text{ mL}} \right) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) = 2.9 \times 10^{-4} \text{ M Ba}^{2+} \end{aligned}$$

Similarly, the concentration of SO_4^{2-} after mixing is the total number of moles of SO_4^{2-} in the original 10.0 mL of Na_2SO_4 solution divided by the final volume (110 mL):

$$\begin{aligned} \text{moles SO}_4^{2-} &= 10.0 \text{ mL} \left(\frac{1 \text{ L}}{1000 \text{ mL}} \right) \left(\frac{0.0020 \text{ mol}}{1 \text{ L}} \right) = 2.0 \times 10^{-5} \text{ mol SO}_4^{2-} \\ [\text{SO}_4^{2-}] &= \left(\frac{2.0 \times 10^{-5} \text{ mol SO}_4^{2-}}{110 \text{ mL}} \right) \left(\frac{1000 \text{ mL}}{1 \text{ L}} \right) = 1.8 \times 10^{-4} \text{ M SO}_4^{2-} \end{aligned}$$

We can now calculate Q :

$$Q = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = (2.9 \times 10^{-4})(1.8 \times 10^{-4}) = 5.2 \times 10^{-8}$$

C We now compare Q with the K_{sp} . If $Q > K_{sp}$, then BaSO_4 will precipitate, but if $Q < K_{sp}$, it will not. Because $Q > K_{sp}$, we predict that BaSO_4 will precipitate when the two solutions are mixed. In fact, BaSO_4 will continue to precipitate until the system reaches equilibrium, which occurs when $[\text{Ba}^{2+}][\text{SO}_4^{2-}] = K_{sp} = 1.08 \times 10^{-10}$.

Exercise

The solubility product of calcium fluoride (CaF_2) is 3.45×10^{-11} . If 2.0 mL of a 0.10 M solution of NaF is added to 128 mL of a 2.0×10^{-5} M solution of $\text{Ca}(\text{NO}_3)_2$, will CaF_2 precipitate?

Answer: yes ($Q = 4.7 \times 10^{-11} > K_{sp}$)

The Common Ion Effect and Solubility

The solubility product expression tells us that the equilibrium concentrations of the cation and the anion are inversely related. That is, as the concentration of the anion increases, the maximum concentration of the cation needed for precipitation to occur decreases—and vice versa—so that K_{sp} is constant. Consequently, the solubility of an ionic compound depends on the concentrations of other salts that contain the same ions. This dependency is another example of the common ion effect discussed in [Section 16.6](#): adding a common cation or anion shifts a solubility equilibrium in the direction predicted by Le Chatelier's principle. As a result, *the solubility of any sparingly soluble salt is almost always decreased by the presence of a soluble salt that contains a common ion.* The exceptions generally involve the formation of complex ions, which is discussed in [Section 17.3](#).

Consider, for example, the effect of adding a soluble salt, such as CaCl_2 , to a saturated solution of calcium phosphate $[\text{Ca}_3(\text{PO}_4)_2]$. We have seen that the solubility of $\text{Ca}_3(\text{PO}_4)_2$ in water at 25°C is 1.14×10^{-7} M ($K_{sp} = 2.07 \times 10^{-33}$). Thus a saturated solution of $\text{Ca}_3(\text{PO}_4)_2$ in water contains $3 \times (1.14 \times 10^{-7} \text{ M}) = 3.42 \times 10^{-7}$ M Ca^{2+} and $2 \times (1.14 \times 10^{-7} \text{ M}) = 2.28 \times 10^{-7}$ M PO_4^{3-} , according to the stoichiometry shown in [Equation 17.1.1](#) (neglecting hydrolysis to form HPO_4^{2-} as described in [Chapter 16](#)). If CaCl_2 is added to a saturated solution of $\text{Ca}_3(\text{PO}_4)_2$, the Ca^{2+} ion concentration will increase such that $[\text{Ca}^{2+}] > 3.42 \times 10^{-7}$ M, making $Q > K_{sp}$. The only way the system can return to equilibrium is for the reaction in [Equation 17.1.1](#) to proceed to the left, resulting in precipitation of $\text{Ca}_3(\text{PO}_4)_2$. This will decrease the concentration of both Ca^{2+} and PO_4^{3-} until $Q = K_{sp}$.

Note the Pattern

The common ion effect usually decreases the solubility of a sparingly soluble salt.

Example 17.1.4

Calculate the solubility of calcium phosphate $[\text{Ca}_3(\text{PO}_4)_2]$ in 0.20 M CaCl_2 .

Given: concentration of CaCl_2 solution

Asked for: solubility of $\text{Ca}_3(\text{PO}_4)_2$ in CaCl_2 solution

Strategy:

A Write the balanced equilibrium equation for the dissolution of $\text{Ca}_3(\text{PO}_4)_2$. Tabulate the concentrations of all species produced in solution.

B Substitute the appropriate values into the expression for the solubility product and calculate the solubility of $\text{Ca}_3(\text{PO}_4)_2$.

Solution:

A The balanced equilibrium equation is given in the following table. If we let x equal the solubility of $\text{Ca}_3(\text{PO}_4)_2$ in moles per liter, then the change in $[\text{Ca}^{2+}]$ is once again $+3x$, and the change in $[\text{PO}_4^{3-}]$ is $+2x$. We can insert these values into the table.

$\text{Ca}_3(\text{PO}_4)_2(\text{s}) = 3\text{Ca}^{2+}(\text{aq}) + 2\text{PO}_4^{3-}(\text{aq})$			
	$\text{Ca}_3(\text{PO}_4)_2$	$[\text{Ca}^{2+}]$	$[\text{PO}_4^{3-}]$
initial	pure solid	0.20	0
change	—	$+3x$	$+2x$
final	pure solid	$0.20 + 3x$	$2x$

B The K_{sp} expression is as follows:

$$K_{\text{sp}} = [\text{Ca}^{2+}]^3[\text{PO}_4^{3-}]^2 = (0.20 + 3x)^3(2x)^2 = 2.07 \times 10^{-33}$$

Because $\text{Ca}_3(\text{PO}_4)_2$ is a sparingly soluble salt, we can reasonably expect that $x \ll 0.20$. Thus $(0.20 + 3x)$ M is approximately 0.20 M, which simplifies the K_{sp} expression as follows:

$$\begin{aligned} K_{\text{sp}} &= (0.20)^3(2x)^2 = 2.07 \times 10^{-33} \\ x^2 &= 6.5 \times 10^{-32} \\ x &= 2.5 \times 10^{-16} \text{ M} \end{aligned} \quad (17.2.1)$$

This value is the solubility of $\text{Ca}_3(\text{PO}_4)_2$ in 0.20 M CaCl_2 at 25°C. It is approximately nine orders of magnitude less than its solubility in pure water, as we would expect based on Le Chatelier's principle. With one exception, this example is identical to Example 2—here the initial $[\text{Ca}^{2+}]$ was 0.20 M rather than 0.

Exercise

Calculate the solubility of silver carbonate in a 0.25 M solution of sodium carbonate. The solubility of silver carbonate in pure water is 8.45×10^{-12} at 25°C.

Answer: 2.9×10^{-6} M (versus 1.3×10^{-4} M in pure water)

Summary

The equilibrium constant for a dissolution reaction, called the **solubility product** (K_{sp}), is a measure of the solubility of a compound. Whereas solubility is usually expressed in terms of *mass* of solute per 100 mL of solvent, K_{sp} is defined in terms of the *molar* concentrations of the component ions. In contrast, the **ion product** (Q) describes concentrations that are not necessarily equilibrium concentrations. Comparing Q and K_{sp} enables us to determine whether a precipitate will form when solutions of two soluble salts are mixed. Adding a common cation or common anion to a solution of a sparingly soluble salt shifts the solubility equilibrium in the direction predicted by Le Chatelier's principle. The solubility of the salt is almost always decreased by the presence of a common ion.

Key Takeaway

- The solubility product (K_{sp}) is used to calculate equilibrium concentrations of the ions in solution, whereas the ion product (Q) describes concentrations that are not necessarily at equilibrium.

Conceptual Problems

- Write an expression for K_{sp} for each salt.
 - AgI
 - CaF₂
 - PbCl₂
 - Ag₂CrO₄
- Some species are not represented in a solubility product expression. Why?
- Describe the differences between Q and K_{sp} .
- How can an ion product be used to determine whether a solution is saturated?
- When using K_{sp} to directly compare the solubilities of compounds, why is it important to compare only the K_{sp} values of salts that have the same stoichiometry?
- Describe the effect of a common ion on the solubility of a salt. Is this effect similar to the common ion effect found in buffers? Explain your answer.
- Explain why the presence of MgCl₂ decreases the molar solubility of the sparingly soluble salt MgCO₃.

Answers

- $K_{sp} = [Ag^+][I^-]$
 - $K_{sp} = [Ca^{2+}][F^-]^2$
 - $K_{sp} = [Pb^{2+}][Cl^-]^2$
 - $K_{sp} = [Ag^+]^2[CrO_4^{2-}]$
-
-
-
- For a 1:1 salt, the molar solubility is simply $\sqrt{K_{sp}}$; for a 2:1 salt, the molar solubility is $\sqrt[3]{K_{sp}/4}$. Consequently, the magnitudes of K_{sp} can be correlated with molar solubility *only* if the salts have the same stoichiometry.
-
- Because of the common ion effect. Adding a soluble Mg²⁺ salt increases [Mg²⁺] in solution, and Le Chatelier's principle predicts that this will shift the solubility equilibrium of MgCO₃ to the left, decreasing its solubility.

Numerical Problems

- Predict the molar solubility of each compound using the K_{sp} values given in [Table E3](#).
 - Cd(IO₃)₂
 - AgCN
 - HgI₂
- Predict the molar solubility of each compound using the K_{sp} values given.
 - Li₃PO₄: 2.37×10^{-11}
 - Ca(IO₃)₂: 6.47×10^{-6}
 - Y(IO₃)₃: 1.12×10^{-10}
- A student prepared 750 mL of a saturated solution of silver sulfate (Ag₂SO₄). How many grams of Ag₂SO₄ does the solution contain? $K_{sp} = 1.20 \times 10^{-5}$.
- Given the K_{sp} values in [Table 17.1.1](#) and [Table E3](#), predict the molar concentration of each species in a saturated aqueous solution.
 - silver bromide
 - lead oxalate
 - iron(II) carbonate
 - silver phosphate

5. copper(I) cyanide
5. Given the K_{sp} values in [Table 17.1.1](#) and [Table E3](#), predict the molar concentration of each species in a saturated aqueous solution.
 1. copper(I) chloride
 2. lanthanum(III) iodate
 3. magnesium phosphate
 4. silver chromate
 5. strontium sulfate
 6. Silicon dioxide, the most common binary compound of silicon and oxygen, constitutes approximately 60% of Earth's crust. Under certain conditions, this compound can react with water to form silicic acid, which can be written as either H_4SiO_4 or $Si(OH)_4$. Write a balanced chemical equation for the dissolution of SiO_2 in basic solution. Write an equilibrium constant expression for the reaction.
 7. The K_{sp} of $Mg(OH)_2$ is 5.61×10^{-12} . If you tried to dissolve 24.0 mg of $Mg(OH)_2$ in 250 mL of water and then filtered the solution and dried the remaining solid, what would you predict to be the mass of the undissolved solid? You discover that only 1.0 mg remains undissolved. Explain the difference between your expected value and the actual value.
 8. The K_{sp} of lithium carbonate is 8.15×10^{-4} . If 2.34 g of the salt is stirred with 500 mL of water and any undissolved solid is filtered from the solution and dried, what do you predict to be the mass of the solid? You discover that all of your sample dissolves. Explain the difference between your predicted value and the actual value.
 9. You have calculated that 24.6 mg of $BaSO_4$ will dissolve in 1.0 L of water at 25°C. After adding your calculated amount to 1.0 L of water and stirring for several hours, you notice that the solution contains undissolved solid. After carefully filtering the solution and drying the solid, you find that 22.1 mg did not dissolve. According to your measurements, what is the K_{sp} of barium sulfate?
 10. In a saturated silver chromate solution, the molar solubility of chromate is 6.54×10^{-5} . What is the K_{sp} ?
 11. A saturated lead(II) chloride solution has a chloride concentration of 3.24×10^{-2} mol/L. What is the K_{sp} ?
 12. From the solubility data given, calculate K_{sp} for each compound.
 1. AgI: 2.89×10^{-7} g/100 mL
 2. SrF_2 : 1.22×10^{-2} g/100 mL
 3. $Pb(OH)_2$: 78 mg/500 mL
 4. $BiAsO_4$: 14.4 mg/2.0 L
 13. From the solubility data given, calculate K_{sp} for each compound.
 1. $BaCO_3$: 10.0 mg/500 mL
 2. CaF_2 : 3.50 mg/200 mL
 3. $Mn(OH)_2$: 6.30×10^{-4} g/300 mL
 4. Ag_2S : 1.60×10^{-13} mg/100 mL
 14. Given the following solubilities, calculate K_{sp} for each compound.
 1. $BaCO_3$: 7.00×10^{-5} mol/L
 2. CaF_2 : 1.70 mg/100 mL
 3. $Pb(IO_3)_2$: 2.30 mg/100 mL
 4. SrC_2O_4 : 1.58×10^{-7} mol/L
 15. Given the following solubilities, calculate K_{sp} for each compound.
 1. Ag_2SO_4 : 4.2×10^{-1} g/100 mL
 2. $SrSO_4$: 1.5×10^{-3} g/100 mL
 3. CdC_2O_4 : 6.0×10^{-3} g/100 mL
 4. $Ba(IO_3)_2$: 3.96×10^{-2} g/100 mL
 16. The K_{sp} of the phosphate fertilizer $CaHPO_4 \cdot 2H_2O$ is 2.7×10^{-7} at 25°C. What is the molar concentration of a saturated solution? What mass of this compound will dissolve in 3.0 L of water at this temperature?

17. The K_{sp} of zinc carbonate monohydrate is 5.5×10^{-11} at 25°C . What is the molar concentration of a saturated solution? What mass of this compound will dissolve in 2.0 L of water at this temperature?
18. Silver nitrate eye drops were formerly administered to newborn infants to guard against eye infections contracted during birth. Although silver nitrate is highly water soluble, silver sulfate has a K_{sp} of 1.20×10^{-5} at 25°C . If you add 25.0 mL of 0.015 M AgNO_3 to 150 mL of 2.8×10^{-3} M Na_2SO_4 , will you get a precipitate? If so, what will its mass be?
19. Use the data in [Table E3](#) to predict whether precipitation will occur when each pair of solutions is mixed.
- 150 mL of 0.142 M $\text{Ba}(\text{NO}_3)_2$ with 200 mL of 0.089 M NaF
 - 250 mL of 0.079 M K_2CrO_4 with 175 mL of 0.087 M CaCl_2
 - 300 mL of 0.109 M MgCl_2 with 230 mL of 0.073 M $\text{Na}_2(\text{C}_2\text{O}_4)$
20. What is the maximum volume of 0.048 M $\text{Pb}(\text{NO}_3)_2$ that can be added to 250 mL of 0.10 M NaSCN before precipitation occurs? $K_{sp} = 2.0 \times 10^{-5}$ for $\text{Pb}(\text{SCN})_2$.
21. Given 300 mL of a solution that is 0.056 M in lithium nitrate, what mass of solid sodium carbonate can be added before precipitation occurs (assuming that the volume of solution does not change after adding the solid)? $K_{sp} = 8.15 \times 10^{-4}$ for Li_2CO_3 .
22. Given the information in the following table, calculate the molar solubility of each sparingly soluble salt in 0.95 M MgCl_2 .

Saturated Solution	K_{sp}
$\text{MgCO}_3 \cdot 3\text{H}_2\text{O}$	2.4×10^{-6}
$\text{Mg}(\text{OH})_2$	5.6×10^{-12}
$\text{Mg}_3(\text{PO}_4)_2$	1.04×10^{-24}

Answers

- 1.184×10^{-3} M
 - 7.73×10^{-9} M
 - 1.9×10^{-10} M
-
- 3.37 g
-
- 4.15×10^{-4} M
 - 7.26×10^{-4} M
 - 6.26×10^{-6} M
 - 6.54×10^{-5} M
 - 5.86×10^{-4} M
-
- 22.4 mg; a secondary reaction occurs, where OH^- from the dissociation of the salt reacts with H^+ from the dissociation of water. This reaction causes further dissociation of the salt (Le Chatelier's principle).
-
- 1.2×10^{-10}
-
- 1.70×10^{-5}
-
-
-
- 8.8×10^{-6}
 - 6.7×10^{-9}

3. 9.0×10^{-8}

4. 2.16×10^{-9}

16.

17. 7.4×10^{-6} M; 2.1 mg

18.

19. Precipitation will occur in all cases.

20.

21. 8.27 g

22.

Contributors

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