Learning Objectives

- Quantitatively related \(K_{sp}\) to solubility

Considering the relation between solubility and \(K_{sp}\) is important when describing the solubility of slightly ionic compounds. However, this article discusses ionic compounds that are difficult to dissolve; they are considered "slightly soluble" or "almost insoluble." Solubility product constants \(K_{sp}\) are given to those solutes, and these constants can be used to find the molar solubility of the compounds that make the solute. This relationship also facilitates finding the \(K_{sp}\) of a slightly soluble solute from its solubility.

Introduction

Recall that the definition of solubility is the maximum possible concentration of a solute in a solution at a given temperature and pressure. We can determine the solubility product of a slightly soluble solid from that measure of its solubility at a given temperature and pressure, provided that the only significant reaction that occurs when the solid dissolves is its dissociation into solvated ions, that is, the only equilibrium involved is:

\[
\ce{M}_p\ce{X}_q(s) \rightleftharpoons p\ce{M^{m+}}(aq) + q\ce{X^{n−}}(aq)
\]

In this case, we calculate the solubility product by taking the solid's solubility expressed in units of moles per liter (mol/L), known as its molar solubility.

Example 1: Calculation of \(K_{sp}\) from Equilibrium Concentrations

We began the chapter with an informal discussion of how the mineral fluorite is formed. Fluorite, CaF₂, is a slightly soluble solid that dissolves according to the equation:

\[
\ce{CaF2}(s) \rightleftharpoons \ce{Ca^{2+}}(aq) + 2\ce{F}^−(aq)
\]

The concentration of Ca\(^{2+}\) in a saturated solution of CaF₂ is \(2.1 \times 10^{-4}\) M; therefore, that of F\(^{−}\) is \(4.2 \times 10^{-4}\) M, that is, twice the concentration of Ca\(^{2+}\). What is the solubility product of fluorite?

Solution

First, write out the \(K_{sp}\) expression, then substitute in concentrations and solve for \(K_{sp}\):

\[
\ce{CaF2}(s) \rightleftharpoons \ce{Ca^{2+}}(aq) + 2\ce{F}^−(aq)
\]

A saturated solution is a solution at equilibrium with the solid. Thus:

\[
\ce{K_{sp}} = \ce{[Ca^{2+}][F^{-}]^2} = (2.1 \times 10^{-4})(4.2 \times 10^{-4})^2 = 3.7 \times 10^{-11}
\]

As with other equilibrium constants, we do not include units with \(K_{sp}\).
In a saturated solution that is in contact with solid Mg(OH)$_2$, the concentration of Mg$^{2+}$ is $3.7 \times 10^{-5}$ M. What is the solubility product for Mg(OH)$_2$?

\[
\text{Answer}
\]

$2.0 \times 10^{-13}$

Example \(\PageIndex{2}\) : Determination of Molar Solubility from \(K_{sp}\)

The \(K_{sp}\) of copper(I) bromide, CuBr, is $6.3 \times 10^{-9}$. Calculate the molar solubility of copper bromide.

Solution

The solubility product constant of copper(I) bromide is $6.3 \times 10^{-9}$.

The reaction is:

\[
\text{CuBr}(s) \rightleftharpoons \text{Cu}^+(aq) + \text{Br}^-(aq)
\]

First, write out the solubility product equilibrium constant expression:

\[
K_{sp} = [\text{Cu}^+][\text{Br}^-]
\]

Create an ICE table (as introduced in the chapter on fundamental equilibrium concepts), leaving the CuBr column empty as it is a solid and does not contribute to the \(K_{sp}\):

<table>
<thead>
<tr>
<th></th>
<th>Cu$^+$</th>
<th>Br$^-$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial concentration (M)</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>Change (M)</td>
<td>x</td>
<td>x</td>
</tr>
<tr>
<td>Equilibrium concentration (M)</td>
<td>$0 + x = x$</td>
<td>$0 + x = x$</td>
</tr>
</tbody>
</table>

At equilibrium:

\[
K_{sp} = [\text{Cu}^+][\text{Br}^-] = (x)(x) = x^2
\]

\[
x = \sqrt{(6.3 \times 10^{-9})} = 7.9 \times 10^{-5}
\]

Therefore, the molar solubility of CuBr is $7.9 \times 10^{-5}$ M.

Summary

Solubility is defined as the maximum amount of solute that can be dissolved in a solvent at equilibrium. Equilibrium is the state at which the concentrations of products and reactant are constant after the reaction has taken place. The solubility product constant (\(K_{sp}\)) describes the equilibrium between a solid and its constituent ions in a solution. The value of the constant identifies the degree to which the compound can dissociate in water. The higher the \(K_{sp}\), the more
soluble the compound is. \( K_{sq} \) is defined in terms of activity rather than concentration because it is a measure of a concentration that depends on certain conditions such as temperature, pressure, and composition. It is influenced by surroundings. \( K_{sp} \) is used to describe the saturated solution of ionic compounds. (A saturated solution is in a state of equilibrium between the dissolved, dissociated, undissolved solid, and the ionic compound.)

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