

## 8.8: Solubility Equilibria and the Solubility Product

Considering the relation between solubility and  $K_{sq}$  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 ( $K_{sq}$ ) 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  $K_{sq}$  of a slightly soluble solute from its solubility.

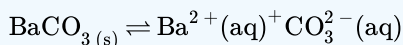
### Introduction

Solubility is the ability of a substance to dissolve. The two participants in the dissolution process are the solute and the solvent. The solute is the substance that is being dissolved, and the solvent is the substance that is doing the dissolving. For example, sugar is a solute and water is a solvent. 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 ( $K_{sq}$ ) 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. For example, the higher the  $K_{sq}$ , 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.  $K_{sq}$  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.)

#### ✓ Example 1: Barium Carbonate

Consider the compound barium carbonate  $\text{BaCO}_3$  (an ionic compound that is not very soluble):



#### Solution

First, write down the equilibrium constant expression:

$$K_c = \frac{[\text{Ba}^{2+}][\text{CO}_3^{2-}]}{[\text{BaCO}_3]}$$

The activity of solid  $\text{BaCO}_3$  is 1, and considering that the concentrations of these ions are small, the activities of the ions are approximated to their molar concentrations.  $K_{sq}$  is therefore equal to the product of the ion concentrations:

$$\begin{aligned} K_{sp} &= [\text{Ba}^{2+}][\text{CO}_3^{2-}] \\ &= 5.1 \times 10^{-9} \end{aligned}$$

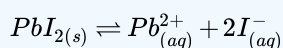
### How are $K_{sp}$ and Solubility Related?

The relation between solubility and the solubility product constants is that one can be used to derive the other. In other words, there is a relationship between the solute's molarity and the solubility of the ions because  $K_{sq}$  is the product of the solubility of each ion in moles per liter.

For example, to find the  $K_{sq}$  of a slightly soluble compound from its solubility, the solubility of each ion must be converted from mass per volume to moles per liter to find the molarity of each ion. These numbers can then be substituted into the  $K_{sq}$  formula, which is the product of the solubility of each ion. An example of this process is given below:

#### ✓ Example 2: Lead Iodide

Suppose the aqueous solubility for compound  $\text{PbI}_2$  is 0.54 grams/100 ml at 25 °C and calculate the  $K_{sq}$  of  $\text{PbI}_2$  at 25°C.



#### Solution

a. Convert 0.54 grams of  $PbI_2$  to moles:

$$0.54 \text{ grams} \times \frac{1 \text{ mol } PbI_2}{461.0 \text{ grams}} = 0.001171 \text{ mol } PbI_2$$

b. Convert ml to L:

$$\frac{100 \text{ mL}}{1000 \text{ L}} = 0.100 \text{ L}$$

c. Find the molarity:

$$\frac{0.001171 \text{ mol}}{0.100 \text{ L}} = 0.01171 \text{ M } PbI_2$$

d. Now find the molarity of each ion by using the stoichiometric ratio (remember there are two  $I^-$  ions for each  $Pb^{2+}$  ion):

$$[Pb^{2+}] = \frac{0.01171 \text{ M}}{1 \text{ L}} \times \frac{1 \text{ mol } Pb}{1 \text{ mol } PbI_2} \quad (8.8.1)$$

$$= 0.011714 \text{ M } Pb^{2+} \quad (8.8.2)$$

$$[I^-] = \frac{0.01171 \text{ M}}{1 \text{ L}} \times \frac{2 \text{ mol } I^-}{1 \text{ mol } PbI_2} \quad (8.8.3)$$

$$= 0.023427 \text{ M } I^- \quad (8.8.4)$$

e. Finally, plug in the molarity to find  $K_{sq}$ :

$$K_{sp} = [Pb^{2+}][I^-]^2 \quad (8.8.5)$$

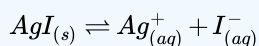
$$= (0.011714 \text{ M})(0.023427 \text{ M})^2 \quad (8.8.6)$$

$$= 6.4 \times 10^{-6} \quad (8.8.7)$$

\*This relation facilitates solving for the molar solubility of the ionic compounds when the  $K_{sq}$  is given to us. The process involves working backwards from  $K_{sq}$  to the molarity of the ionic compound.\*

### ✓ Example 3

Suppose the  $K_{sq}$  at 25 °C is  $8.5 \times 10^{-17}$  for the compound AgI. What is the molar solubility?



#### Solution

a. Let "g" represent the number of moles:

$$K_{sp} = [Ag^{2+}][I^-] \quad (8.8.8)$$

$$= g^2 \quad (8.8.9)$$

$$= 8.5 \times 10^{-17} \quad (8.8.10)$$

b. Solve for "g":

$$g^2 = 8.5 \times 10^{-17} \quad (8.8.11)$$

$$g = (8.5 \times 10^{-17})^{\frac{1}{2}} \quad (8.8.12)$$

$$= 9.0 \times 10^{-9} \quad (8.8.13)$$

The molar solubility of AgI is  $9.0 \times 10^{-9}$

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