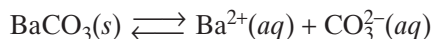


Calculating Solubilities

Once known, the solubility product constant can be used to determine the solubility of a sparingly soluble salt. Suppose you wish to know how many moles of barium carbonate, BaCO_3 , can be dissolved in 1 L of water at 25°C . From **Table 3**, K_{sp} for BaCO_3 has the numerical value 5.1×10^{-9} . The equilibrium equation is written as follows.



Given the value for K_{sp} , we can write the solubility equilibrium expression as follows.

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

Therefore, BaCO_3 dissolves until the product of the molar concentrations of Ba^{2+} ions and CO_3^{2-} ions equals 5.1×10^{-9} . The solubility equilibrium equation shows that Ba^{2+} ions and CO_3^{2-} ions enter the solution in equal numbers as the salt dissolves. Thus, they have the same concentration. Let $[\text{Ba}^{2+}] = x$. Then $[\text{CO}_3^{2-}] = x$ also.

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

$$(x)(x) = x^2 = 5.1 \times 10^{-9}$$

$$x = \sqrt{5.1 \times 10^{-9}}$$

The molar solubility of BaCO_3 is $7.1 \times 10^{-5} \text{ mol/L}$.

Thus, the solution concentration is $7.1 \times 10^{-5} \text{ M}$ for Ba^{2+} ions and $7.1 \times 10^{-5} \text{ M}$ for CO_3^{2-} ions.

SAMPLE PROBLEM C

Calculate the solubility of silver bromide, AgBr , in mol/L, using the K_{sp} value for this compound listed in Table 3.

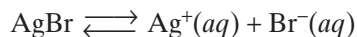
SOLUTION

1 ANALYZE

Given: $K_{sp} = 5.0 \times 10^{-13}$

Unknown: solubility of AgBr

2 PLAN



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

$[\text{Ag}^+] = [\text{Br}^-]$, so let $[\text{Ag}^+] = x$ and $[\text{Br}^-] = x$

3 COMPUTE

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

$$K_{sp} = x^2$$

$$x^2 = 5.0 \times 10^{-13}$$

$$x = \sqrt{5.0 \times 10^{-13}}$$

$$\text{Solubility of AgBr} = \sqrt{5.0 \times 10^{-13}} = 7.1 \times 10^{-7} \text{ mol/L}$$