## 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 $5.1 \times 10^{-9}$. The equilibrium equation is written as follows.

$$BaCO_3(s) \rightleftharpoons Ba^{2+}(aq) + CO_3^{2-}(aq)$$

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

$$K_{sp} = [Ba^{2+}][CO_3^{2-}] = 5.1 \times 10^{-9}$$

Therefore, $BaCO_3$ dissolves until the product of the molar concentrations of $Ba^{2+}$ ions and $CO_3^{2-}$ ions equals $5.1 \times 10^{-9}$. 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 $[Ba^{2+}] = x$. Then $[CO_3^{2-}] = x$ also.

$$[Ba^{2+}][CO_3^{2-}] = K_{sp} = 5.1 \times 10^{-9}$$

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

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

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

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

$$AgBr \rightleftharpoons Ag^+(aq) + Br^-(aq)$$

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

$[Ag^+] = [Br^-]$, so let $[Ag^+] = x$ and $[Br^-] = x$

**3 COMPUTE**

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

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

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

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

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