

Solubility Equilibrium

The Solubility Product Constant

Ionic compounds differ in their solubilities. Most salts of the alkali metals are soluble in water. There are many ionic compounds that are insoluble.

| Compounds | Solubility | Exceptions |
|---|--------------------|---|
| Salts of Group 1A metals and ammonia | Soluble | Some lithium compounds |
| Ethanoates, nitrates, chlorates, and perchlorates | Soluble | Few exceptions |
| Sulfates | Soluble | Compounds of Pb, Ag, Hg, Ba, Sr, and Ca |
| Chlorides, bromides, and iodides | Soluble | Compounds of Ag and some compounds of Hg and Pb |
| Sulfides and hydroxides | Most are insoluble | Alkali metal sulfides and hydroxides are soluble. Compounds of Ba, Sr, and Ca are slightly soluble. |
| Carbonates, phosphates, and sulfites | Insoluble | Compounds of the alkali metals and of ammonium ions |

*Table on page 561 (Table 18.1)

Solubility Equilibrium

Most insoluble salts will actually dissolve to some extent in water. They are partially soluble.

Example



A very small amount of silver chloride actually dissolves in the water.

We can also write this as an equilibrium expression

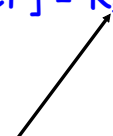
$$K_{eq} = \frac{[\text{Ag}^+] \times [\text{Cl}^-]}{[\text{AgCl}]}$$

As long as some undissolved $\text{AgCl}_{(s)}$ is present, the concentration of the AgCl is a constant.

The concentration of AgCl can then be combined with the equilibrium constant to form a new constant.

$$K_{eq} \times [\text{AgCl}] = [\text{Ag}^+] \times [\text{Cl}^-] = K_{sp}$$

solubility product
constant



K_{sp} equals the product of the concentrations of the ions each raised to a power similar to our K_{eq} equation.

Similar to K_{eq} , we get a value with no units.

The smaller the numerical value of K_{sp} , the lower the solubility of the compound.

| Solubility Product Constants (K_{sp}) at 25°C | | | | | |
|---|-----------------------|-------------------|-----------------------|---------------------------------|-----------------------|
| Salt | K_{sp} | Salt | K_{sp} | Salt | K_{sp} |
| Halides | | Sulfates | | Hydroxides | |
| AgCl | 1.8×10^{-10} | PbSO ₄ | 6.3×10^{-7} | Al(OH) ₃ | 3.0×10^{-34} |
| AgBr | 5.0×10^{-13} | BaSO ₄ | 1.1×10^{-10} | Zn(OH) ₂ | 3.0×10^{-16} |
| AgI | 8.3×10^{-17} | CaSO ₄ | 2.4×10^{-5} | Ca(OH) ₂ | 6.5×10^{-6} |
| PbCl ₂ | 1.7×10^{-5} | Sulfides | | Mg(OH) ₂ | 7.1×10^{-12} |
| PbBr ₂ | 2.1×10^{-6} | NiS | 4.0×10^{-20} | Fe(OH) ₂ | 7.9×10^{-16} |
| PbI ₂ | 7.9×10^{-9} | CuS | 8.0×10^{-37} | Carbonates | |
| PbF ₂ | 3.6×10^{-8} | Ag ₂ S | 8.0×10^{-51} | CaCO ₃ | 4.5×10^{-9} |
| CaF ₂ | 3.9×10^{-11} | ZnS | 3.0×10^{-23} | SrCO ₃ | 9.3×10^{-10} |
| Chromates | | FeS | 8.0×10^{-19} | ZnCO ₃ | 1.0×10^{-10} |
| PbCrO ₄ | 1.8×10^{-14} | CdS | 1.0×10^{-27} | Ag ₂ CO ₃ | 8.1×10^{-12} |
| Ag ₂ CrO ₄ | 1.2×10^{-12} | PbS | 3.0×10^{-28} | BaCO ₃ | 5.0×10^{-9} |

Example

What is the concentration of lead ions and chromate ions in a saturated lead chromate solution at 25°C? ($K_{sp} = 1.8 \times 10^{-14}$)

The Common Ion Effect

Looking at the saturated solution from the last example problem.



What do you think would happen if we were add some lead nitrate to the solution?



Immediately, the products would be greater than K_{sp} .

According to Le Chatelier's principle, what would happen to the equilibrium?



Lead chromate (PbCrO_4) will continue to precipitate from the solution until the product of $[\text{Pb}^{2+}]$ and $[\text{CrO}_4^{2-}]$ once again equals 1.8×10^{-14} .

In this example, lead is called a **common ion**.

A common ion is an ion that is found in both salts in a solution.

When a common ion addition lowers the solubility of an ionic compound it is called the **common ion effect**.

Example

What is the bromide-ion concentration of a 1.00-L saturated solution of AgBr to which 0.020 mol of AgNO_3 is added? The K_{sp} of silver bromide is 5.0×10^{-13} .

If the product of the concentrations of two ions in the mixture is greater than the solubility product (K_{sp}) of the compound formed from the ions, a precipitate will form.

If the product of the concentrations is less than K_{sp} , no precipitate will form.

Example

Suppose you have 0.50-L of 0.002M $\text{Ba}(\text{NO}_3)_3$ and mix it with 0.50 L of 0.008M Na_2SO_4 . Calculate if a precipitate will form.

Try problem on pages 562-565 #17-28

Solubility Equilibrium