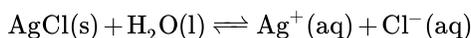


## 10.7: Solubility Equilibria

In [Chapter 5](#) we learned about a class of reactions that involved the formation of a solid that was “insoluble” in water, and precipitated from the solution. In these “precipitation reactions”, one ionic salt was described as “insoluble”, driving the reaction towards the formation of products. Silver chloride is a classic example of this. If you mix silver nitrate (almost *all* nitrate salts are “soluble” in water) with sodium chloride, a copious white precipitate of silver chloride formed and the silver nitrate was deemed “insoluble”.

Nonetheless, if you took the clear solution from above the silver chloride precipitate and did a chemical analysis, there will be sodium ions, nitrate ions, and *traces* of chloride ions and silver ions. The concentrations of silver and chloride ions would be about  $1.67 \times 10^{-5}$  M, far below the concentrations we typically work with, hence we say that silver chloride is “insoluble in water”. That, of course, is not true. Solubility is an equilibrium in which ions leave the solid surface and go into solution at the same time that ions are re-deposited on the solid surface. For silver chloride, we could write the equilibrium expression as:



In order to write the expression for the equilibrium constant for this solubility reaction, we need to recall the rules stated in [Section 10.2](#) of this chapter; Rule #4 states, “Reactants or products that are present as solids or liquids or the solvent, all have an activity value of 1, and so they do not affect the value of the equilibrium expression.” Because silver chloride is a solid, and water is the solvent, the expression for the equilibrium constant is simply,

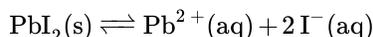
$$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

Note that we have denoted the equilibrium constant as  $K_{sp}$ , where “sp” refers to solubility equilibrium, or “solubility product” (the product of the concentrations of the ions). We can calculate the value of  $K_{sp}$  for silver chloride from the analytical data that we cited above; an aqueous solution above solid silver chloride has a concentration of silver and chloride ions of  $1.67 \times 10^{-5}$  M, at 25° C. Because the concentrations of silver and chloride ions are both  $1.67 \times 10^{-5}$  M, the value of  $K_{sp}$  under these conditions must be:

$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = (1.67 \times 10^{-5})^2 = 2.79 \times 10^{-10}$$

This is very small, considering that  $K_{sp}$  for sodium chloride is about 29!

For a salt such as  $\text{PbI}_2$  chemical analysis tells us that the lead concentration in a saturated solution (the maximum equilibrium solubility under a specified set of conditions, such as temperature, pressure, etc.) is about  $1.30 \times 10^{-3}$  M. In order to calculate  $K_{sp}$  for lead (II) iodide, you must first write the chemical equation and then the equilibrium expression for  $K_{sp}$  and then simply substitute for the ionic concentrations.



As you do this, remember that there are *two* iodide ions for every lead ion, therefore the concentrations for lead (II) and iodide are  $1.30 \times 10^{-3}$  M and  $2.60 \times 10^{-3}$  M, respectively.

$$K_{sp} = [\text{Pb}^{2+}][\text{I}^-]^2 = (1.30 \times 10^{-3})(2.60 \times 10^{-3})^2 = 8.79 \times 10^{-9}$$

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