# Equilibrium

**Unit 13C:** Solubility Product Constants

### Knowledge & Understanding

* common ion effect

### Skills:

* calculate *K*sp values and concentrations from *K*sp values

###  Notes:

## Phases and Equilibrium

Solids and pure liquids do not appear in the equilibrium expression. This is because the moles of a pure solid or liquid per liter of that solid or liquid is really equivalent to the density, which is constant.

For example, in the chemical reaction:

CCl4 (g)   C (s) + 2 Cl2 (g)

The equilibrium expression would be:

Where does this come from?

Start with the usual products over reactants:

However, [C] is the concentration of carbon (in moles per liter) *in solid carbon*, which is just the density (a \_\_\_\_\_\_\_\_\_\_\_\_\_) divided by the molar mass (also a \_\_\_\_\_\_\_\_\_\_\_\_\_\_). A constant divided by a constant is just another constant, which we’ll call “*K*­carbon”. This means:



Now, applying the “a constant divided by a constant is just another constant” idea again, we simply adjust *K*c so that the new *K*c is the old *K*c divided by *K*carbon. Therefore, the equilibrium becomes:

## Solubility Product Constant (*K*sp)

Earlier in this course, we defined “insoluble” to mean a saturated solution would have a concentration of less than \_\_\_\_ (or less than 3 g/100 g solvent).

Actually, some amount of any ionic compound \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ in water, though the amount can be miniscule. For these “\_\_\_\_\_\_\_\_\_\_\_\_” compounds, when the compound is placed in water, the precipitate is in equilibrium with the solute.



Solubility Simulation: http://phet.colorado.edu/en/simulation/soluble-salts

The solubility product constant (Ksp) is an equilibrium constant for the process of the ionic solid \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_. (Recall that solubility varies with temperature; *K*sp values are usually reported at 25°C.)

For example, the chemical equation for the dissolution of Ag2SO4 is:

Ag2SO4 (s)   2 Ag+ (aq) + SO42− (aq)

As we saw earlier, solids do not appear in equilibrium constants, so the denominator is omitted. This gives the following equation:

For Ag­2SO4: *K*sp =

 *K*sp =

We can solve the above equation, because we know that one Ag2SO4 molecule dissolves into 2 Ag+ + 1 SO­42−. If we let [SO42−] = x, then [Ag+] = 2x, and:

Ksp = 1.2 × 10−5 =

In other words, a 0.0144 M solution of Ag2SO4 would be \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ (at 25°C).

## Common Ion Effect

*K*sp is a physical constant for a given compound. One useful corollary of this fact is the common ion effect.

As an illustrative example, suppose we start with 1.00 ℓ of a saturated solution of Ag2SO4. As we saw on the previous page, [Ag+] = 0.0288 M and [SO42−] = 0.0144 M. However, suppose we now add 1.00 mole of Na2SO4. This brings the total [SO42−] to 1.0144 M.

Notice that [Ag+] is substantially lower than the 0.0288 M that resulted when only Ag2SO4 was present.

Note that we should really have calculated [SO42−] as an I.C.E. problem. However, if we assume that *all* of the [Ag+] in solution was removed by precipitation, this would remove all of the 0.0144 M of SO42− that was initially in solution. Using 1.00 M instead of 1.0144 M for the [SO42−] gives [Ag+] = 3.46 × 10−3 *M* instead of 3.44 × 10−3 *M*. The difference is less than 1%, which means the simplification was reasonable.

The common ion effect is a common and powerful lab technique—a \_\_\_\_\_\_ concentration of a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ is added in order to precipitate virtually all of an ion of interest, such as removing toxic Hg2+ or Hg22+ ions from a solution before disposal.