Title: Determination of the Solubility Product of an Ionic Compound

Objective: To determine the solubility product of an ionic compound

Background:

When analyzing reactions, one factor is to analyze its rate. The rate of reaction is the speed at which a reaction takes place. Usually, most reactions are able to react one way, then react backwards. A reaction at equilibrium means that both the forward and reverse rates of reaction are the same. At equilibrium, neither the products or reactants have a tendency to want to react anymore. An equilibrium constant is a ratio between the amount of reactant and the amount of product at equilibrium. The equilibrium constant expression is:

$Kc = [products]^x / [reactants]^y$

Where Kc is the equilibrium constant, [products] is the concentration of products, [reactants] is the concentration of reactants, and x/y is the molar coefficients, respectively.

A type of equilibrium constant is the solubility product constant, Ksp. This equilibrium is established when an ionic solid dissociates in water to form a saturated solution. A saturated solution is a solution where no more of that particular ion can be dissolved under normal conditions. Similar to the equilibrium expression, the solubility product expression is:

$Ksp = [anion]^x * [cation]^y$

Note that there is no denominator because liquids and solids do not appear in the equilibrium expression. The square brackets refer to the molar concentrations of the ions.

Knowing the Ksp value of a salt is useful when determining what the concentration of ions of a compound in a saturated solution is. This allows us to control how much precipitate can be formed, or not form a precipitate at all.

To find the Ksp value, we can take the concentrations of ions of a molar solution and plug it into the equation to calculate for values.

Procedure p1 (dilute Ca^{2+} solutions)

1. Put 5 drops of 0.1 M $Ca(NO_3)_2$ in the first well. (hold the droppers vertically, make sure there are no air bubbles in pipette tubes.
2. Place 5 drops of distilled H_2O into the other wells (11)
3. Add 5 drops of 0.1 M $Ca(NO_3)_2$ to the second well.
4. Mix the solution with a pipette. The solution should be a 0.05 M Ca^{2+} solution.
5. Put 5 drops of that solution into a third well.
6. Repeat steps 4 to 6 by diluting it down the 12 wells. Determine the concentration of each solution. Make sure that the last well has a concentration of 4.9×10^{-5}

Procedure p2 (add NaOH)

7. Place 5 drops of 0.1 M NaOH into each well. When the NaOH is added, the initial concentrations will be halved, because of the constant concentration of Ca^{2+} and increasing volume.
8. Make sure each solution is mixed with a toothpick. Confirm that the Ca^{2+} concentration in the last well plate is 2.4×10^{-5}

Procedure p3 (precipitate and Ksp calculations)

9. Wait for precipitate to form. Observe the pattern of precipitation The first well that does not form a precipitate represents a saturated solution, as the ion concentration for both solutions will be too low to form any precipitate.

Procedure p4 (check the results)

10. Repeat the procedure from part 1, this time by diluting the NaOH instead of the $Ca(NO_3)_2$.
11. Observe, make sure the saturated solution is in the same well plate.