Determination of the Solubility Product of an Ionic Compound

Introduction

An equilibrium constant is a measure of how far a reaction proceeds to completion. A solubility product constant is an equilibrium constant that permits the calculation of the amount of a slightly soluble ionic compound that will dissolve in water.

Concepts

- Solubility product, K_{sp}
- Saturated solution
- Precipitation reaction

Background

The solubility product constant, K_{sp}, is a particular type of equilibrium constant. The equilibrium is formed when an ionic solid dissolves in water to form a saturated solution. The equilibrium exists between the aqueous ions and the undissolved solid. A saturated solution contains the maximum concentration of ions of the substance that can dissolve at the solution's temperature. The equilibrium equation showing the ionic solid lead chloride dissolving in water is:

$$PbCl_2(s) \rightleftarrows Pb^{2+}(aq) + 2Cl^-(aq)$$

The solubility product expression is:

$$K_{sp} = [Pb^{2+}][Cl^-]^2$$

where square brackets refer to molar concentrations of the ions. A knowledge of the K_{sp} of a salt is useful in determining the concentration of ions of the compound in a saturated solution. This allows us to control a solution so that precipitation of a compound will not occur, or to find the concentration needed to cause a precipitate to form.

The solubility product which will be determined in this experiment is that of the strong base, calcium hydroxide, $Ca(OH)_2$.

Experiment Overview

In this experiment, solutions of calcium ions and hydroxide ions are mixed to form a precipitate of calcium hydroxide. The solutions are each diluted and mixed until a precipitate no longer appears. The solubility product for calcium hydroxide is then calculated from the solution concentrations.

Pre-Lab Questions *(Use a separate sheet of paper to answer the following questions.)*

Before beginning work on this experiment, read the directions and answer the following questions.

The ionic compound silver chromate is not very soluble in water. It ionizes according to the following equation:

$$Ag_2CrO_4(s) \rightleftarrows 2Ag^+(aq) + CrO_4^{2-}(aq)$$

1. Write the solubility product expression for silver chromate.
2. If one has a solution of 0.10 M silver nitrate and it is diluted by a factor of two, what is the new concentration?
3. The dilution of 0.10 M silver nitrate by a factor of two is carried out five times. What is the concentration now?
4. The value for the K_{sp} of silver chromate is reported to be 1.1×10^{-12}. In a saturated solution of silver chromate, the silver ion concentration is found to be 2.5×10^{-4} M. What must the chromate ion concentration be? Show your work.

Materials

Chemicals

Calcium nitrate, $Ca(NO_3)_2$, 0.10 M, 5 mL	Water, distilled or deionized
Sodium hydroxide, NaOH, 0.10 M, 5 mL	

Equipment

Beral pipets with capillary tips, 5	Toothpicks, for stirring
Microplate, 96-well	

Safety Precautions

Calcium nitrate is slightly toxic by ingestion. Sodium hydroxide solution is a corrosive liquid. Avoid contact with skin and especially the eyes. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory.

Procedure

Part 1. Prepare a series of diluted calcium ion solutions

1. Arrange a microplate so that you have 12 wells across from left to right.
2. Put 5 drops of 0.10 M calcium nitrate in well #1 in the first row. *Note:* Hold the Beral capillary droppers vertically when dispensing the drops. Make sure no air bubbles are in the pipet tubes. Discard the first drop as it may contain an air bubble.
3. Place 5 drops of distilled or deionized water in each of the wells #2 through #12 in the first row.
4. Next, add 5 drops of 0.10 M calcium nitrate to well #2.
5. Use an empty Beral pipet to mix the solution thoroughly by drawing the solution into the pipet and then squirting it back several times. The solution in this well, #2, is now 0.050 M in Ca^{2+} ion.
6. Use your empty pipet to remove the solution from well #2 and put 5 drops of this solution into well #3.
7. Put the remaining solution back in well #2.
8. Mix the solution in well #3 as before.
9. Continue this serial dilution procedure, adding 5 drops of the previous solution to the 5 drops of distilled or deionized water in each well down the row until you fill the last one, #12.
10. Mix the solution in well #12, and discard 5 drops.
11. Determine the concentration of solution in each well. Verify that the concentration of calcium ions in well #12 is 4.9×10^{-5} M.

Part 2. Combine with sodium hydroxide

12. Place 5 drops of 0.10 M sodium hydroxide, NaOH, in each of the wells #1 through #12. When the sodium hydroxide is added to each well, the initial concentrations of the reactants are halved, as each solution dilutes the other.
13. Use an empty pipet to mix each of these combined solutions by drawing each solution up into the pipet and squirting it back into the well; or mix thoroughly with a toothpick. Now the concentration of Ca^{2+} ions in well #12 is 2.4×10^{-5} M.

Part 3. Observe the precipitates and calculate K_{sp}

14. Allow three or four minutes for the precipitates to form. Observe the pattern of precipitation.

15. At one point the concentration of both ions becomes too low to have any precipitate form. We will assume that the first well with no precipitate represents a saturated solution.

Part 4. Check using a dilution of sodium hydroxide

16. To check your results, repeat the procedure but use a serial dilution of the NaOH.
17. In a different row, put 5 drops of 0.10 M sodium hydroxide in well #1.
18. Put 5 drops of distilled or deionized water in wells #2 through #12.
19. Add 5 drops of the 0.10 M sodium hydroxide to well #2.
20. Use an empty Beral pipet to mix the solution by pulling the solution into the pipet and then squirting it back several times. The solution in this well, #2, is now 0.050 M in OH^- ion.
21. Continue the serial dilution to well #12, and then remove 5 drops from well #12.
22. Add 5 drops of 0.10 M $Ca(NO_3)_2$ to each of the wells, and mix each with an empty pipet or toothpick.
23. Again, determine the well where no more precipitate appears.

Disposal and Cleanup

Your teacher will provide disposal and cleanup instructions.

Data Table

Calcium Ion Serial Dilutions

First well with no precipitation ____________

Concentration of Ca^{2+} __________ mol/L

Concentration of OH^- __________ mol/L

K_{sp} $[Ca(OH)_2]$ = ______________

Hydroxide Ion Serial Dilutions

First well with no precipitation ____________

Concentration of Ca^{2+} __________ mol/L

Concentration of OH^- __________ mol/L

K_{sp} $[Ca(OH)_2]$ = ______________

Post-Lab Calculations and Analysis

(Use a separate sheet of paper to answer the following questions.)

1. Calculate the concentration of Ca^{2+} ions and OH^- ions in the first well of the calcium ion dilution series with no precipitate. Using these concentrations, determine the solubility product, K_{sp}, of calcium hydroxide. Enter these values in the Data Table under Calcium Ion Serial Dilutions.
2. Calculate the concentration of calcium and hydroxide ions in the first well of the hydroxide ion dilution series where there is no precipitate, and again calculate the value of K_{sp}. Enter these values in the Data Table under Hydroxide Ion Serial Dilutions.
3. How did the values obtained from the two trials compare with each other? Look up the accepted value for the solubility product of calcium hydroxide in a handbook and compare to your values.
4. Does this method give values that are too low or too high? Why?
5. What would make the method more accurate?
6. Would the results be better if the concentrations of the last well where precipitation occurred were averaged with the first well where there was no precipitate? Is there any justification for doing this? Try it!