Experiment 6 Determination of a Solubility Product Constant

Introduction

The amount of a substance that will dissolve in a given amount of a solvent is termed its **solubility**. Although solubility is temperature dependent, the extent to which a substance dissolves also depends upon its chemical structure and the interaction with solvent molecules. Covalent or molecular compounds are generally not very soluble in water unless they are polar or contain sufficient polar functional groups to enable the molecule to associate with water molecules. In contrast, many ionic compounds dissolve quite readily in water due to the strength of ion-dipole forces between the ions in the salt and water molecules. However, not all ionic compounds dissolve to a great extent in water. If the strength of the ionic bond is too great for the ion-dipole forces between the ions and water to break the bond, the solubility of the ionic compound will be greatly decreased.

The dissolution process may be represented by a reversible chemical equation such as:

$$AgCl_{(s)} \rightleftharpoons Ag^{+}{}_{(aq)} + Cl^{-}{}_{(aq)}$$

For a sparingly soluble salt, this reaction represents and equilibrium process between the undissolved solid and its solvated ions. This equilibrium may be represented by an equilibrium constant, which is called the **solubility product constant**, K_{sp}. Since pure solids and liquids do not appear in equilibrium constant expressions, the solubility product consists of the product of the concentrations of the ions formed raised to the powers of their coefficients in the dissolution equation. For AgCl*(s)*, the K_{sp} expression is:

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

For the species Ag_2SO_4, which contains two silver ions per formula unit, the K_{sp} expression is:

$$K_{sp} = [Ag^{+}]^{2}[SO_4{}^{2-}]$$

K_{sp} values for various salts are tabulated in chemical reference books and can be used to calculate the solubilities of ionic substances. Conversely, the solubility of an ionic substance may be used to calculate the value of the K_{sp} for that substance.

In this experiment, you will be determining the solubility of calcium hydroxide, $Ca(OH)_2$. Like most metal hydroxides, calcium hydroxide is only sparingly soluble. The only common metal hydroxides that are readily soluble are the group 1A hydroxides such as NaOH and KOH. Calcium hydroxide dissolves via the following equation:

$$Ca(OH)_{2(s)} \rightleftharpoons Ca^{2+}{}_{(aq)} + 2\ OH^{-}{}_{(aq)}$$

When calcium hydroxide dissolves, it forms a basic solution (because it produces OH^{-} ion). You will prepare a saturated solution of calcium hydroxide and determine the concentration of hydroxide ion in that solution. The concentration of hydroxide ion in aqueous solution can be determined by titration with a standard acid solution. Since calcium hydroxide dissolves to form two OH^{-} ions for each Ca^{2+}, then:

$$[Ca^{2+}] = ½ [OH^{-}]$$

From these concentrations, you can calculate the value of the K_{sp} for calcium hydroxide.

Equipment

- 250-mL Erlenmeyer flask with stopper
- stirring rod
- burette
- 25-mL volumetric flask
- 125-mL Erlenmeyer flask
- Buchner funnel
- 250-mL filter flask
- piece of filter paper

Procedure

A. Preparation of a Saturated Calcium Hydroxide solution

1. Boil about 120 mL of deionized water in a 250-mL Erlenmeyer flask to remove dissolved carbon dioxide, then allow the water to cool to room temperature.
2. Add approximately 3 grams of $Ca(OH)_2$ to the flask. Stir the mixture thoroughly and stopper the flask to minimize the formation of $CaCO_3$ by reaction with CO_2 in the air. Allow the undissolved $Ca(OH)_2$ to settle to the bottom of the flask for 10-15 minutes.

(The above may already have been done by the Stockroom to save time).

B. Determination of $[OH^{-}]$ and K_{sp}

1. While the undissolved calcium hydroxide is settling out, rinse a burette with the standard 0.05 M HCl solution, which will be furnished in the lab. Discard the rinse then fill the burette with the standard solution. Record the initial volume and the actual concentration of the HCl solution on your data sheet.
2. Take the temperature of your saturated $Ca(OH)_2$ solution and record it on your data sheet.
3. Set up a Buchner funnel with a piece of filter paper onto a 250-mL filter flask. Attach the filter flask with a vacuum hose to the vacuum outlet. Turn on the vacuum. Carefully pour off a small amount of your saturated Ca(OH)2 solution onto the filter paper to wet it.
4. Then carefully filter about 60 mL of your solution through the filter paper. The liquid in the filter flask, called the filtrate, should be clear. If any cloudiness appears due to undissolved $Ca(OH)_2$, you will have to start the filtration over with a new piece of filter paper.
5. Pour a small portion of the clear filtrate (saturated solution) into a 25-mL volumetric flask, swirl, and discard.
6. Fill the flask up to the mark with the saturated solution.
7. Empty the saturated solution in the 25-mL volumetric flask into a clean 125-mL Erlenmeyer flask.
8. Rinse the volumetric flask with a small amount of deionized water.

9. Add 2-3 drops of methyl orange indicator to the solution and then titrate with the standard HCl solution to the endpoint. The endpoint is a color change from yellow to red (or orange).

10. Record the final volume of HCl in the burette.

11. Empty the titration flask down the sink, rinse well with deionized water, measure out a second 25 mL sample of $Ca(OH)_2$ and repeat the titration. If the results of the two titrations do not agree, repeat the titration a third time.

12. Calculate the $[OH^-]$ from the titration data. Then, calculate $[Ca^{2+}]$. Substitute these values into the Ksp expression for $Ca(OH)_2$ and calculate its value. Show all calculations.

Name: ____ CHEM 1412 Section: ________

Partner: ____ Date: ___/___/___

Report Sheet EXP6: Determination of a Solubility Product Constant

Concentration of standard HCl solution ____________________

Trial	1	2
Final burette reading	________________	________________
Initial burette reading	________________	________________
Volume of HCl added	________________	________________

Solution temperature ____________________

Concentration of OH^{-} ____________________

Calculations:

Concentration of Ca^{2+} ____________________

Calculations:

Value of K_{sp} ____________________

Calculations:

Name: ______________________________ Date: ___/___/___ CHEM1412 Section: ________

Experiment 6: Pre-Lab Exercise

An HCl solution has a concentration of 0.09714 M. 25.00 mL of this solution was then diluted to 250.00 mL in a volumetric flask. The diluted solution was then titrated against 10.00 mL of a saturated $Ca(OH)_2$ solution using methyl orange indicator to reach the endpoint.

1. What is the concentration of the diluted HCl solution?

2. If 6.82 mL of the diluted HCl solution was required to reach the endpoint, what is the concentration of OH^- in solution?

3. What is the concentration of Ca^{2+} in solution?

4. What is the K_{sp} expression for the dissolution of $Ca(OH)_2$?

5. Calculate the value of the K_{sp} for $Ca(OH)_2$?