Experiment 5:

DETERMINATION OF THE SOLUBILITY PRODUCT CONSTANT FOR A SPARINGLY SOLUBLE SALT

OBJECTIVE

To become familiar with equilibria involving sparingly soluble substances by determining the value of the solubility product constant for a sparingly soluble salt.

DISCUSSION

Inorganic substances may be broadly classified into three different categories: acids, bases, and salts. According to the Brønsted-Lowry theory, acids are proton donors, and bases are proton acceptors. When an acid reacts with a base in aqueous solution, the products are a salt and water. With a few exceptions, nearly all common salts are strong electrolytes. The solubilities of salts span a broad spectrum, ranging from slightly or sparingly soluble to very soluble. This experiment is concerned with heterogeneous equilibria of slightly soluble salts. For a true equilibrium to exist between a solid and a solution, the solution must be saturated. Calcium hydroxide is slightly soluble salt, and in a saturated solution this equilibrium may be represented as follows:

$$Ca(OH)_2\,_{(s)} \rightleftharpoons Ca^{2+}\,_{(aq)} + 2OH^-\,_{(aq)} \tag{1}$$

The equilibrium constant for Equation [1] is

$$K_{sp} = [Ca^{2+}][OH^-]^2 \tag{2}$$

K_{sp} is called the solubility product constant. At a given temperature the value of K_{sp} is constant. The solubility product for a sparingly soluble salt can be easily calculated by determining the solubility of the substances in water. Suppose, for example, we determined that 7.4×10^{-2} g of $Ca(OH)_2$ dissolves in 100 mL of water. The molar solubility of this solution (that is, the molarity of the solution) is 0.010 *M*.

We see from Equation [1] that for each mole of $Ca(OH)_2$ that dissolves, one mole of Ca^{2+} and two moles of OH^- are formed. It follows, therefore, that

$$\text{solubility of } Ca(OH)_2 \text{ in moles/liter} = [Ca^{2+}] = ½\,[OH^-] = 0.010\ M$$

and

$$K_{sp} = [Ca^{2+}][OH^-]^2 = [0.010][2\times0.010]^2 = 4.0\times10^{-6}$$

In a saturated solution the product of the molar concentrations of Ca^{2+} and OH^- cannot exceed 4.0×10^{-6}. If the ion product $[Ca^{2+}][OH^-]^2$ exceeds 4.0×10^{-6}, precipitation of $Ca(OH)_2$ would occur until this product is reduced to the value of K_{sp}. Or if a solution of NaOH is added to a solution of $CaCl_2$, $Ca(OH)_2$ would precipitate if the ion product $[Ca^{2+}][OH^-]^2$ is greater than K_{sp}.

To determine the solubility product constant for a sparingly soluble substance, we need only to determine the concentration of one of the ions, because the concentration of the other ion is related to the first one's concentration by a simple stoichiometric relationship. Any method that we could use to accurately determine the concentration would be suitable. In this experiment, you will determine the solubility product constant K_{sp} for $Ca(OH)_2$. You will determine the concentration of hydroxide ion by titration with hydrochloric acid, according to:

$$H^+\ (aq)\ +\ OH^-\ (aq)\ \longrightarrow\ H_2O\ (l) \qquad [3]$$

The concentration of Ca^{2+} can be determined by noting that at equilibrium $[Ca^{2+}] = ½\,[OH^-]$.

PROCEDURE

1- You will be provided with a bottle containing about 2 g of powdered calcium hydroxide in 100 mL of distilled water. The bottle has been shaken well and set aside for more than a day. *Minimize shaking the bottle as possible*.
2- Rinse and fill the burette with standardized hydrochloric acid (0.100 *M*).
3- Filter the contents of the bottle, allowing the first 5 mL to run to waste and collecting the rest in a dry conical flask. The first few mL are rejected because they are less concentrated in solute than the rest. The filter paper absorbs solute until it attains equilibrium with the solution. Yet another equilibrium!

To minimize absorption of carbon dioxide, steps 4 and 5 should be done quickly (with due care!)

4- Rinse the pipette with the calcium hydroxide solution and transfer 25.0 mL to a conical flask (this needs not be dry).
5- Add two drops of phenolphthalein to the flask and titrate the solution until the pink color just disappears. Record your burette readings in your report sheet.
6- Repeat steps 4 and 5 for 2 or 3 solutions of calcium hydroxide.
7- Record the temperature.
8- Calculate the concentrations of OH^- and Ca^{2+} ions, and K_{sp} for $Ca(OH)_2$.

Name ………………………………………… ID ………… Sec …..

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REPORT SHEET

$Ca(OH)_2$–HCl Titration

		Trial	**1**	**2**	**3**	**4**
Burette readings (0.100 *M* HCl) / mL	Final					
	Initial					
Volume / mL						
Average volume / mL						

K_{sp} for $Ca(OH)_2$

Moles of OH^- = moles H^+ = ____________________________________ mol

$[OH^-]$ = _________________________________ *M*

$[Ca^{2+}]$ = __________________________ *M*

Solubility of $Ca(OH)_2$ = ________________ *M*

K_{sp} for $Ca(OH)_2 = [Ca^{2+}][OH^-]^2$ = ______________________________________

at _________ ^{o}C.