EXPERIMENT B5: ACID/BASE DISSOCIATION CONSTANT

Learning Outcomes

Upon completion of this lab, the student will be able to:

1) Estimate the dissociation constant for a weak acid and a weak base.
2) Relate the pH of a weak acid/base to the concentration of the weak acid/base

Introduction

A weak acid or a weak base only dissociates partially in an aqueous medium. For this discussion, assume that HA is a weak acid and B is a weak base. The dissociations of these and the expressions for the respective equilibrium constants are shown below:

$$HA_{(aq)} + H_2O_{(l)} \Leftrightarrow H_3O^+_{(aq)} + A^-_{(aq)} \qquad K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

$$B_{(aq)} + H_2O_{(l)} \Leftrightarrow BH^+_{(aq)} + OH^-_{(aq)} \qquad K_b = \frac{[BH^+][OH^-]}{[B]}$$

A simple method to determine the K_a of HA and the K_b of B is to determine the pH of known concentrations of these solutions. Assume that 0.1 M solutions of HA and B have been provided. The pH of the solutions can be used to determine the molarity of H_3O^+ in the respective solutions at equilibrium.

$$pH = -\log[H_3O^+]$$
$$[H_3O^+] = 10^{-pH}$$

The concentrations of H_3O^+ and OH^- in an aqueous solution are related to the autoionization of water which can be expressed by the ionic product of water (K_W).

$$K_W = 1.0 \times 10^{-14} = [H_3O^+] \times [OH^-]$$

Sample calculation of K_a:

Consider the calculation needed to determine the K_a of 0.100 M HA whose pH has been determined to be 2.87. An equilibrium table (ICE Table) will need to be set up as follows to perform the required calculations.

	$HA_{(aq)}$ +	$H_2O_{(l)}$ ⇔	$H_3O^+{}_{(aq)}$ +	$A^-{}_{(aq)}$
Initial concentration, M	0.100		~0	0
Change (amount reacted)	-x		+x	+x
Equilibrium concentration, M	0.100 – x		x	x

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{x^2}{0.100 - x}$$

Since "x" is the same as the $[H_3O^+]$ at equilibrium and the pH of the solution is known (pH is given to be 2.87), "x" is determined as follows:

$$x = [H_3O^+] = 10^{-pH} = 10^{-2.87} = 0.00135M$$

$$Therefore: K_a = \frac{(0.00135)^2}{(0.100 - 0.00135)} = 1.84 \times 10^{-5}$$

Using the data from the table in the textbook of acid dissociation constants, the acid in question can be inferred to be likely acetic acid.

<u>Sample calculation of K_b:</u>

Now, consider the calculation needed to determine the K_b of 0.100 M B whose pH has been determined to be 11.13. An equilibrium table will need to be set up as follows to perform the required calculations.

	$B_{(aq)}$ +	$H_2O_{(l)}$ ⇔	$BH^+{}_{(aq)}$ +	$OH^-{}_{(aq)}$
Initial concentration, M	0.100		0	~0
Change (amount reacted)	-x		+x	+x
Equilibrium concentration, M	0.100 – x		x	x

$$K_b = \frac{[BH^+][OH^-]}{[B]} = \frac{x^2}{0.100 - x}$$

Since "x" is the same as the [OH⁻] at equilibrium and the pH of the solution is known (pH is given to be 11.13), "x" is determined as follows:

$$pH = 11.13$$
$$[H_3O^+] = 10^{-11.13} = 7.42 \times 10^{-12}$$
$$[H_3O^+][OH^-] = 1.0 \times 10^{-14}$$
$$[OH^-] = \frac{1.0 \times 10^{-14}}{7.42 \times 10^{-12}} = 0.00135M$$

$$Therefore: K_b = \frac{(0.00135)^2}{(0.100 - 0.00135)} = 1.84 \times 10^{-5}$$

Using the data from the table of base dissociation constants, the base in question can be inferred to be likely ammonia.

An interesting observation in these instances is that, different concentrations of the same acid (or base) will result in different pH values. The equilibrium constant in each dilution will, of course, be the **same value**. This idea will also be explored in this experiment.

Experimental Design

Obtain approximately 20 mL of 0.100 M unknown acid. This solution will be diluted to obtain 0.0100 M and 0.00100 M concentrations of the same acid. The pH of all the three solutions will be measured using a pH meter. The pH values will be used to determine the K_a of the acid and the identity of the acid. The same experiment will be repeated using 20 ml of 0.100 M unknown base to determine the K_b of the base.

Reagents and Supplies

0.100 M unknown acid and 0.100 M unknown base

pH meter, Two 10-mL volumetric flask, Three 10-mL beakers

(See posted Material Safety Data Sheets)

Procedure

1. Read the instruction manual for the operation of a pH meter. The manual can be found in Appendix B. **Please read pages: 15, 16, 18, 21, 31, 22, and 28.**

2. The procedure described below in steps 6 and 8 is referred to as serial dilution. Diluting solutions is inherently prone to errors. In serial dilution, the errors are propagated into each subsequent dilution. Therefore great care must be taken in order to obtain the desired concentrations of the solutions. (Use a volumetric flask and a microburette for the purpose of dilution).

3. Obtain a pH meter from the stockroom. The instructor will demonstrate the proper use and calibration of the pH meter.

4. Obtain approximately 10 mL of an unknown weak acid of concentration 0.100 M.

5. Measure the pH of the acid in step 4.

6. Dilute the solution of 0.100 M weak acid to obtain a solution that is 0.0100 M. NOTE: The student must determine the exact volume needed for the dilution so that the desired final concentration is obtained (use a volumetric flask and a microburette for the purpose of dilution).

7. Measure the pH of the acid in step 6.

8. Dilute the solution of 0.0100 M weak acid prepared in step 6 to obtain a solution that is 0.00100 M.

9. Measure the pH of the acid in step 8.

10. Repeat steps 3 to 9 using 10 ml of an unknown weak base of concentration 0.100 M.

11. Use the calculated values of K_a and K_b to determine the identity of the weak acid and weak base, respectively.

Data Table

WEAK ACID

Concentration, M	pH
0.100	
0.0100	
0.00100	

Record clearly how the required dilutions were performed. Use the correct number of significant figures/decimal places reflecting the precision of the equipment used.

1. ________________ mL of 0.100 M weak acid was diluted to a final volume of ________________mL to obtain a solution whose molarity is ________________ M

2. ________________ mL of 0.0100 M weak acid was diluted to a final volume of ________________mL to obtain a solution whose molarity is ________________ M

WEAK BASE

Concentration, M	**pH**
0.100	
0.0100	
0.00100	

Record clearly how the required dilutions were performed. Use the correct number of significant figures/decimal places reflecting the precision of the equipment used.

1. ________________ mL of 0.100 M weak base was diluted to a final volume of ________________mL to obtain a solution whose molarity is ________________ M

2. ________________ mL of 0.0100 M weak base was diluted to a final volume of ________________mL to obtain a solution whose molarity is ________________ M

Data Analysis

Show the detailed calculations for the K_a of the weak acid for each concentration (as shown in the sample calculation section):

0.100 M

0.0100 M

0.00100 M

Average value of K_a = ____________________________________

Identity of the weak acid = ____________________________________

Show the detailed calculations for the K_b of the weak base for each concentration (as shown in the sample calculation section):

<u>0.100 M</u>

<u>0.0100 M</u>

<u>0.00100 M</u>

Average value of K_b = ____

Identity of the weak base = ____