Determination of K_a of Weak Acids

Theory: Introduction

Acids vary greatly in their strength—their ability to ionize or produce ions when dissolved in water. What factors determine the strength of an acid? In this experiment, the strength of acids will be measured by determining the equilibrium constants for their ionization reactions in water.

Concepts

- Weak acid
- Equilibrium constant
- Conjugate base
- Neutralization reaction

Experiment Overview

The purpose of this experiment is to measure the pK_a value for ionization of two unknown weak acids. Solutions containing equal molar amounts of the weak acids and their conjugative bases are prepared by "half-neutralization" of the acid. Their pH values are measured and used to calculate the pK_a value for the unknowns and thus determine their identities.

Pre-Lab Questions

Phosphoric acid is a triprotic acid (three ionizable hydrogens). The values of its stepwise ionization constants are $K_{a1} = 7.5 \text{ X } 10^{-3}$, $K_{a2} = 6.2 \text{ X } 10^{-8}$, and $K_{a3} = 4.2 \text{ X } 10^{-13}$.

1. Write the chemical equation for the first ionization reaction of phosphoric acid with water.
2. Write the equilibrium constant expression (K_{a1}) for the reaction.
3. What would be the pH of a solution when $[H_3PO_4] = [H_2PO_4^-]$? (Note: H+ion concentration is = to Ka1.
4. Phenolphthalein would not be an appropriate indicator to use to determine K_{a1} for phosphoric acid. Why not? Choose a suitable indicator from the following color chart.

	pH										
Indicator	1	2	3	4	5	6	7	8	9	10	11
Phenolphthalein	Colorless							Pink		Red	
Methyl Red	Red				Orange		Yellow				
Orange IV	Orange		Peach		Yellow						

Procedure

1. Label two weighing dishes #1 and #2.
2. Obtain an unknown weak acid and record the unknown letter in the Data Table.
3. Measure out a small quantity (0.15—0.20g) of the unknown into each weighing dish. *Note:* It is not necessary to know the exact mass of each sample.
4. Using a graduated cylinder, precisely measure 50.0mL of distilled water into a 150-mL beaker.
5. Transfer sample #1 to the water in the beaker and stir to dissolve.
6. Using a graduated cylinder, precisely transfer 25.0 mL of the acid solution in step 5 into an Erlenmeyer flask.
7. Add 3 drops of phenolphthalein to the acid solution in the Erlenmeyer flask.
8. Using a Beral-type pipet, add sodium hrdroxide solution dropwise to the flask. Gently swirl the flask while adding the sodium hydroxide solution.
9. Continue adding sodium hydroxide dropwise and swirling the solution until a faint pink color persists throughout the solution for at least 5 seconds. This is called the endpoint. *Note:* A pink color develops immediately when the base is added, but fades quickly once the solution is swirled. When nearing the

endpoint, the pink color begins to fade more slowly. Proceed cautiously when nearing the endpoint, so as not to "overshoot" it.

Note: At this point the solution in the beaker contains exactly one-half of the original amount of acid, essentially all of which is in the acid form, HA. The Erlenmeyer flask contains an equal amount of the conjugative base A^- obtained by neutralization.

10. Pour the contents of the Erlenmeyer flask back into the beaker. Pour the solution back and forth a few times to mix. *Note:* It is important to transfer the solution as completely as possible from the flask back into the beaker.
11. Using a pH meter, measure the pH of the resulting solution in the beaker, which contains equal molar amounts of the acid and its conjugate base. Record the pH in the Data Table.
12. Dispose of the beaker contents according to the teacher's instructions and rinse both the beaker and the Erlenmeyer flask with distilled water. Dry the beaker with a paper towel.
13. Repeat steps 4-12 using sample #2.
14. Repeat steps 1-13 for one of the remaining unknowns.

A	Sample #1				
	Sample #2				
B	Sample #1				
	Sample #2				
C	Sample #1				
	Sample #2				
D	Sample #1				
	Sample #2				
E	Sample #1				
	Sample #2				

Discussion

1. Average the pH readings for each trail (samples #1 and #2) to calculate the average pK_a value for the unknown weak acids and enter answers in the Data Table.

2. Comment on the precision (reproducibility) of the pK_a determinations. Describe sources of experimental error and their likely effect on the measured pK_a (pH) values.

3. The following table lists the identities of the unknowns in this experiment. Complete the table by calculation the pK_a value for each acid. *Note:* $pK_a = -\log K_a$.

Acid	**Formula**	K_a	pK_a
Potassium dihydrogen phosphate	KH_2PO_4	K_{a2} of $H_3PO_4 = 6.2 \times 10^{-8}$	
Potassium hydrogen sulfate	$KHSO_4$	K_{a2} of $H_2SO_4 = 1.0 \times 10^{-2}$	

Potassium hydrogen phthalate	$KHC_8H_4O_4$	K_{a2} of $H_2C_8H_4O_4 = 3.9 \times 10^{-6}$	
Potassium hydrogen tartrate	KHC_4H_4O6	K_{a2} of $H_2C_4H_4O_6 = 4.6 \times 10^{-5}$	
Acetylsalicylic Acid	2-$CH_3CO_2C_6H_4COOH$	$K_a = 3.2 \times 10^{-4}$	

4. Compare the experimental pK_a value for each unknown with the literature values reported in Question 3. Determine the probable identity of each unknown and enter the answers in the Data Table.

5. Write separate equations for each unknown potassium salt dissolving in water and for the ionization reaction of the weak acid anion that each of these salts contains. (See Equations 7 and 8)

6. Why was it not necessary to know the exact mass of each acid sample?

7. Why was it not necessary to know the exact concentration of the sodium hydroxide solution?

8. Why was it necessary to measure the exact volume of distilled water used to dissolve the acid, as well as the exact volume of solution transferred from the beaker to the Erlenmeyer flask?

Conclusion: