

Lab 5A. Determination of the K_a of a Weak Acid and the K_b of a Weak Base from pH Measurements

Prelab Assignment

Before coming to lab:

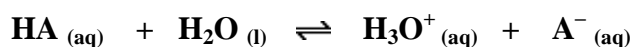
- Take note that this week's lab activity consists of two separate lab experiments, Lab 5A and Lab 5B, two separate lab handouts each with its own "abbreviated" lab write-up and Prelab questions.
- Read this lab handout thoroughly and then complete in your lab notebook the following sections of the report for this lab exercise: Title, a ruled data table—see the report guidelines on [page 3](#) for details. Ensure that the table of contents of your lab notebook is current.
- Answer the pre-lab questions that appear at the end of this lab exercise and hand in at the start of your lab period.
- Background information for this lab can be found in [Chapter 18](#) in your textbook (*Silberberg 5th ed*).

Purpose

In this experiment you will learn to calibrate and use a pH probe, and then use the probe to measure the pH of two different solutions, the weak acid, ethanoic acid (acetic acid, CH_3COOH), and the pH of sodium acetate (sodium ethanoate), NaCH_3COO . The pH of each solution will then be used to calculate the K_a (acid dissociation constant) of ethanoic acid and the K_b (base dissociation constant) of sodium acetate.

Introduction

Weak acids are incompletely dissociated in solution. The general equation for the undissociated weak acid, HA , in equilibrium with the hydronium ion, H_3O^+ , is:



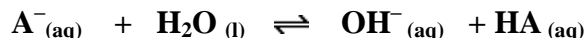
The **acid equilibrium constant** for the reaction, K_a , defined by the equation,

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

provides a quantitative measure of the degree to which an acid dissociates. A **small value for K_a** ($\ll 1$) indicates that the acid remains mostly as molecules in solution. On the other hand, a **large value for K_a** ($\gg 1$) indicates that the acid has dissociated to a greater extent. In fact, in the case of a strong acid such as HCl there is essentially no molecular acid remaining in solution—hence the K_a for HCl approaches infinity!

When a weak acid is labeled, for example, as 0.100 M, this refers to the initial molarity of the acid, before any of it dissociates. Thus the actual concentration of undissociated acid will always be less than the stated value. The amount that has dissociated can be determined by measuring the pH of the solution. Since **pH** is defined as $-\log [\text{H}^+]$ (Remember that H^+ is shorthand for the **hydronium ion**, H_3O^+), it is a simple step to determine the hydronium ion and conjugate base concentrations. And since there is a 1:1 stoichiometric relationship between the amount of hydronium ions formed and the amount of dissociated acid, one can readily calculate the concentration of undissociated acid present at equilibrium. For more detail on these calculations review [Sample Problem 18.7 on page 800](#) in your text.

Weak bases also dissociate incompletely. (It is actually the water that dissociates upon reaction with the base.) The general equation for the weak base, A^- , in equilibrium with the hydroxide ion, OH^- , is:



where **HA** is the conjugate acid. The **base dissociation constant** for the reaction, K_b , defined by the equation,

$$K_b = \frac{[OH^-][HA]}{[A^-]}$$

provides a quantitative measure of the degree to which a base dissociates. A **small value for K_b** indicates that the base remains mostly as the base form A^- in solution. On the other hand, a **large value for K_b** indicates that the base has dissociated to a greater extent.

Weak acids are related to their conjugate weak bases. For a conjugate acid-base pair, the K_a and K_b are related by the following equation:

$$K_a \cdot K_b = K_w = 1.00 \times 10^{-14} \quad (\text{at } 25^\circ\text{C})$$

In this experiment you will use a pH probe interfaced to a computer. The probe will need to be calibrated prior to use. This will be accomplished by sequentially putting the probe in each of two buffer solutions of known pH and setting the scale to read those particular pH values.

Procedure (Perform in teams of 2 students)

1. Prepare a solution of your assigned acetic acid (ethanoic acid), CH_3COOH , concentration. Add about 50 mL of DI water to a 100-mL volumetric flask. **Q – Why is it important to have water already in the volumetric flask in this case?** Use a volumetric pipette to add the requisite amount of **2.00 M acetic acid** to the volumetric flask. **CAUTION: acetic acid can cause painful burns if it comes in contact with your skin or eyes.** Record the volume used. Consider in advance of lab how to calculate this volume. Fill the flask with DI water to the mark, cap and invert at least ten times to ensure the solution is thoroughly mixed.
2. Prepare a solution of your assigned sodium acetate (sodium ethanoate), $NaCH_3COO$, concentration. Weigh the required amount to the nearest mg. **Consider in advance of lab how to calculate this mass.** Record the exact amount used and recalculate the concentration if needed. Transfer the solid quantitatively to a clean 100-mL volumetric flask. Add DI water to about 2/3 full, then swirl until the entire solid is dissolved. Fill the flask with DI water to the mark, cap and invert at least ten times to ensure the solution is thoroughly mixed.
3. Obtain a pH probe from the lab cart. Clamp the probe on a ring stand so that you can move it up and down to insert it into solutions whose pH you want to measure.

CAUTION!!

- Be careful to handle the probe gently—the end where the potential of a solution is measured is **quite fragile**.
- As you are working through these steps you should also take care to **keep the probe moist at all times**.
- Before and after each time the probe is used to measure the pH of a solution, you will need to rinse the probe by squirting it with a small amount of DI water (catch the runoff in a large beaker) and then gently dry it with a *Kimwipe* tissue.

4. After attaching the probe to the “Go!Link”, start the Logger Pro software. In the “Chemistry with Vernier” folder open the file “Exp 27 Acid Dissociation K_a .”
5. The pH electrode may need to be calibrated. Two buffer solutions are available for this purpose. Test the electrode first by measuring the pH of each buffer. If the pH is within ± 0.05 there is no need to calibrate. If not, calibrate the electrode as follows:
 - a. Select the “Calibrate” command and then “Go!Link: pH 1” in the “Experiment” menu. Rinse and dry the electrode and insert the pH electrode into the first buffer solution. Click on “CalibrateNow”, wait a few seconds for the voltage to stabilize, enter the pH of that buffer solution in the “Value 1” box, and click “Keep.”
 - b. Rinse and dry the electrode and place it in the second buffer. Wait a few seconds, enter the pH of the second buffer, click “Keep” and then “Done.” You have now calibrated the electrode.
6. After rinsing the pH electrode with DI water, measure the pH of your assigned concentration of acetic acid solution. (Use your smallest beaker filled with about 2 cm of solution.)
7. After rinsing the pH electrode with DI water, measure the pH of your assigned concentration of sodium acetate solution.
8. Rinse the pH probe one last time with DI water and then gently dry it with a *Kimwipe* before returning the probe back to its original storage bottle. Make sure the tip of the pH probe is fully immersed in the protective solution (a pH 4-KCl buffer) it came in and return it to the lab cart.
9. Dispose of all chemicals in the waste containers under the hood.

Report Guidelines

As outlined below, your lab report for this lab will only consist of three sections: a **Data and Observations** section, an **Analysis of the Results** section and a **Conclusion**.

I. Data and Observations

- Cut and paste into your lab notebook [Table 1 on page 5](#) to record all data, observations and to summarize the results of your calculations.

II. Analysis of the Results

For the acetic acid solution, set up an ICE table with the following:

1. Use the pH to calculate the $[H_3O^+]$ and $[A^-]$.
2. Use the $[H_3O^+]$ and the initial concentration of the acid to calculate the concentration of the undissociated acid remaining at equilibrium, $[HA]$.
3. Calculate the K_a and the pK_a .

For the acetate ion, set up an ICE table with the following:

4. Use the pH to calculate the $[OH^-]$ and $[HA]$.
5. Use the $[OH^-]$ and the initial concentration of the base to calculate the concentration of the undissociated base remaining at equilibrium, $[A^-]$.
6. Calculate the K_b and the pK_b .

Analysis Questions:

7. Look up the accepted value of the K_a and pK_a and compare your experimental results with the accepted values. Yes, it would be appropriate to calculate percent errors and comment on them. What are the possible sources of error?
8. Compare your value of K_a to that of at least 3 other groups that had different initial concentrations. What effect if any does the initial acetic acid concentration have on the K_a ? Explain.
9. Look up the accepted value of the K_b and pK_b , and compare your experimental results with the accepted values. Calculate percent errors and comment on them. What are the possible sources of error?
10. Compare your value of K_b to that of at least 3 other groups that had a different initial concentration. What effect if any does the initial sodium acetate concentration have on the K_b ? Explain.
11. Summarize all your data and calculations on the provided Results table.
12. Calculate the K_b for the acetate ion from your experimentally determined K_a for the acetic acid. How does it compare with the K_b you found experimentally?

III. Conclusion

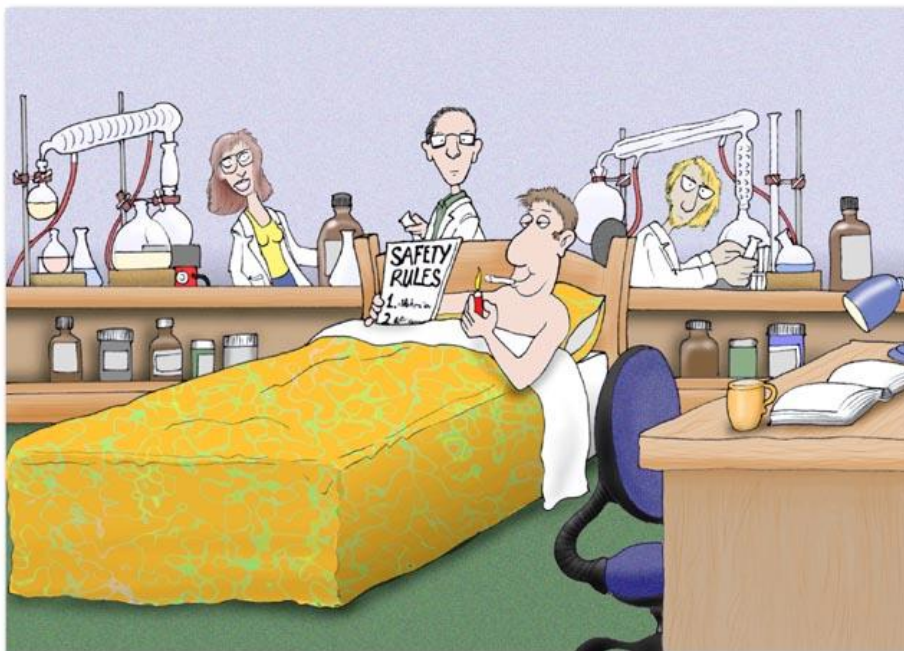
- Write a brief conclusion using bullets.

Acknowledgement: This lab is adapted from similar labs produced by my colleagues at GRCC.

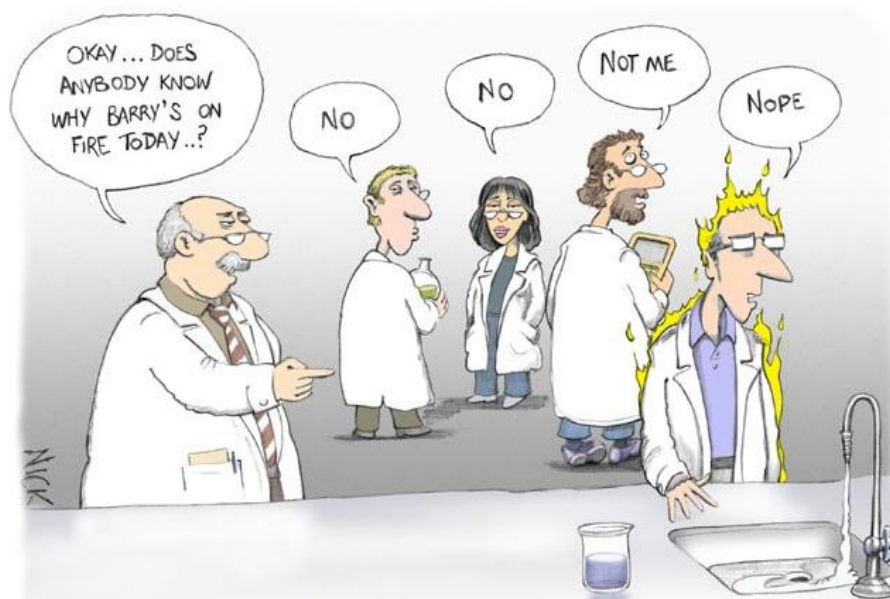
Table 1. Summary table for the results for Lab 5A, K_a of Acetic Acid and K_b of sodium acetate from pH.

Acetic Acid		Sodium Acetate	
Assigned Concentration		Assigned Concentration	
Volume of 2.00 M acetic acid		Mass of sodium acetate	
Measured pH		Measured pH	
Equilibrium equation		Equilibrium equation	
K_a Expression		K_b Expression	
$[\text{CH}_3\text{COOH}]_{\text{init}}$		$[\text{CH}_3\text{COO}^-]_{\text{init}}$	
$[\text{H}_3\text{O}^+]_{\text{eq}}$		$[\text{OH}^-]_{\text{eq}}$	
$[\text{CH}_3\text{COO}^-]_{\text{eq}}$		$[\text{CH}_3\text{COOH}]_{\text{eq}}$	
$[\text{CH}_3\text{COOH}]_{\text{eq}}$		$[\text{CH}_3\text{COO}^-]_{\text{eq}}$	
K_a		K_b	
$\text{p}K_a$		$\text{p}K_b$	
K_a values from three other groups		K_b values from three other groups	

This page is reserved for cartoons and therapeutic doodling...



Advances in Science, 37. Iconoclastic researcher Dr Alvin Henderson demonstrates that it IS in fact possible to become too familiar with the safety rules.



Studies reveal that at the first hint of OSH paperwork, 82.3% of scientists are able to voluntarily jettison their short-term memories.

Labs 5A. Prelab Questions **K_a and K_b from pH**

Name _____

Date _____ Section _____ Group No. _____

Instructions: Complete the following five questions and hand in at the start of your lab period. Show your work with units and correct significant figures for all questions. ***CIRCLE YOUR ANSWERS!***

1. A 1.5×10^{-3} M solution of a weak acid, HA , has a pH of 5.12. Complete the equilibrium equation below and then calculate the K_a for the acid. Show your work using an "ICE" table. Circle your answer.



2. Calculate the hydronium ion concentration and pH of 0.15 M HA , a weak acid, with $K_a = 1.8 \times 10^{-5}$.

3. Write the *net ionic equation* for the basic hydrolysis reaction for sodium cyanide, $NaCN$.

4. How would you expect the pH of 0.10 M $NaCN$ to compare to 0.10 M KCN ? Circle your choice and explain your reasoning below.

(a.) They have the same pH (b.) $pH\ NaCN > pH\ KCN$ (c.) $pH\ NaCN < pH\ KCN$

Explanation:

Turn the page for more pleasure....

5. A 0.10 M solution of NaX has a pH of 8.11. Complete the equilibrium equation below and then calculate the K_b for the salt, NaX. Show your work using an “ICE” table. Circle your answer.

