

K_a Acid Dissociation Constant

Minneapolis Community and Technical College

Principles of Chemistry II, C1152

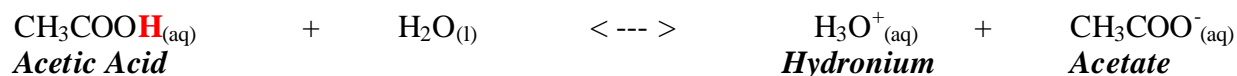
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I. Introduction

Acetic Acid

Monoprotic acetic acid, CH₃COOH is sometimes written as HCH₃COO, HC₂H₃O₂ or even HA to acknowledge the *one* ionizable hydrogen, attached to an oxygen atom, that is responsible for the acid's characteristics. The three hydrogen atoms attached to the carbon atom are strongly bonded and thus don't contribute significantly to acidic behavior.

Acetic acid is also a *weak* acid. When pure glacial acetic acid (containing no water) is dissolved in water, some acetic acid molecules react (the *forward* reaction) with water to form hydronium and acetate ions:



Simultaneously, acetate and hydronium ions recombine to form the original products (the *reverse* reaction). This equilibrium situation leaves most of the acetic acid molecules in their molecular form and only few as acetate ions. Thus, the solution contains relatively few ions and is a weak conductor.

The equilibrium constant for the reaction above, K_a, depends on the concentrations of product and reactant species as follows:

$$K_a = \frac{[\text{H}_3\text{O}^+] \times [\text{CH}_3\text{COO}^-]}{[\text{CH}_3\text{COOH}]} = 1.75 \times 10^{-5} \quad (@25^\circ\text{C}) \quad \text{Equation 1}$$

It will be the goal of this experiment to calculate the value of K_a in a variety of different situations and to compare it to the value above. The calculation will require mole and total volume information for all species used to calculate concentrations and lastly K_a.

Calculation of K_a

To calculate K_a, the equilibrium concentrations of H₃O⁺, CH₃COO⁻ and CH₃COOH must be known and substituted into the formula above (Equation 1).

1. [H₃O⁺]_{eq} is determined by first measuring the pH of the solution with the pH meter. The pH reading is then used to calculate [H₃O⁺]:

$$[\text{H}_3\text{O}^+] = 10^{-\text{pH}} \quad \text{Equation 2}$$

2. [CH₃COO⁻]_{eq} is determined indirectly by first calculating the *moles* of H₃O⁺ produced determined from step 1. Since the mole ratio of acetate to hydronium is 1:1, we have the following:

$$\text{moles}_{\text{CH}_3\text{COO}^-} = \text{moles}_{\text{H}_3\text{O}^+}$$

The concentration of acetate ion is next calculated by dividing the moles of acetate ion by the total volume of the solution.

Note: In some cases, there may be additional acetate ions present that have been added via other means (e.g. directly as NaCH₃COO salt). In such cases, the calculated acetate ion concentration *must include all sources of acetate ions* (i.e. the TOTAL moles of acetate ion).

3. $[\text{CH}_3\text{COOH}]_{\text{eq}}$ is calculated by first determining the number of moles of CH_3COO^- produced via dissociation of acetic acid (from step 2). This amount does NOT include acetate ions from other sources. Because CH_3COO^- is produced at the expense of consumed CH_3COOH and because they are in a 1:1 mole ratio, the number of equilibrium moles of CH_3COOH is the pre-equilibrium (initial) amount minus the amount converted into CH_3COO^- :

$$\begin{array}{rcccl} \textit{Equilibrium} & & \textit{Initial} & & \textit{Consumed} \\ \text{moles}_{\text{CH}_3\text{COOH}} & = & \text{moles}_{\text{CH}_3\text{COOH}} & - & \text{moles}_{\text{CH}_3\text{COO}^-} \end{array}$$

The equilibrium concentration of acetic acid is obtained by dividing the corresponding mole value by the total volume.

4. **K_a determination.** Substitute the values obtained from steps 1, 2 and 3 into Equation 1 above.

II. MSDS: Chemical Information

Acetic Acid (1 Molar)

EMERGENCY OVERVIEW

Appearance: colorless liquid.

Caution! May cause allergic skin reaction. May cause eye and skin irritation. May cause respiratory and digestive tract irritation.

Target Organs: Kidneys, teeth.

Potential Health Effects

Eye: May cause eye irritation and possible damage.

Skin: May cause severe skin irritation. May cause skin sensitization, an allergic reaction, which becomes evident upon re-exposure to this material.

Ingestion: May cause gastrointestinal irritation with nausea, vomiting and diarrhea.

Inhalation: Effects may be delayed. May cause respiratory tract irritation.

Chronic: Prolonged or repeated skin contact may cause dermatitis. Repeated inhalation may cause chronic bronchitis. Repeated exposure may cause erosion of teeth. May cause reproductive and fetal effects. Laboratory experiments have resulted in mutagenic effects.



III. Procedure


A. Solutions Preparation

...You will be working in pairs today.

** Download and print the handouts re. burette, volumetric flask, and volumetric pipette techniques available on the chemistry web site.*

Prepare each of the following 4 solutions using the stock 1.00 molar acetic acid stock solution provided in the laboratory.

- **Solution #1**
 - Dispense approximately 50 mL of 1.00 CH₃COOH solution in a clean/dry 100mL beaker
- **Solution #2**
 - Obtain a 100 mL volumetric flask
 - Pipette 10.00 mL 1.00 M CH₃COOH into the volumetric flask.
 - Add distilled water to the volumetric flask up to the line. Use a wash bottle or eyedropper for the final additions.
 - Stopper and mix thoroughly
- **Solution #3: Buffer Solution A**
 - Use the supplied graduated cylinders to dispense 100.0 mL 1.00 M NaCH₃COO (sodium acetate) and 40.0 mL of 1.00 M CH₃COOH in a clean/dry 250 mL beaker.
 - Mix well using a clean glass stirring rod.
- **Solution #4: Buffer Solution B**
 - Take 10.00 mL of solution #3 and dilute with 90.0 mL of distilled water in a 150mL beaker.



B. pH probe calibration

Buffers are solutions constructed so as to maintain a constant pH. You will use two commercially available buffer solutions (pH = 2 and pH = 7) to calibrate your pH probe. The calibrated pH probe will then be used in part C to measure the pH values for the four test solutions.

1. Remove the pH probe from the storage solution by unscrewing the cap. (the cap remains attached to the probe but can be pushed up to the top).
2. Insert the pH probe in the vial and stir the solution by moving the pH probe in an up and down motion.
3. Immerse the pH probe in clean, distilled water when not in use and be sure to rinse with distilled water between measurements.
 - a. Click on “Experiment”
 - b. Click on “Calibrate” and then pH probe (menu pops up to right of calibrate)
 - c. Click on “Calibrate Now”
 - d. Enter the buffer’s pH value in the field next to “value 1” (either 2 or 7)
8. Continue stir in an up and down motion. When the voltage associated with “input 1” stabilizes click “Keep”
9. Rinse the pH probe with a small amount of distilled water.
10. Immerse the pH probe in the second buffer’s vial and stir in with an up and down motion.
11. Enter the second buffer’s pH value in the field next to the “value 2”
12. Continue to swirl the beaker and when the voltage next to “input 2” stabilizes click “Keep”
13. Click “DONE”
14. Store the pH probe in clean distilled water
15. Don’t shut down Logger Pro until you are finished making pH measurements as the calibration will be lost.

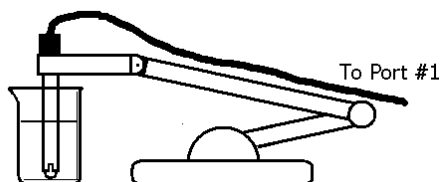


C. pH measurement

1. Pour approximately 20 mL of the solution to be tested into a clean, dry 50 mL beaker.
2. Immerse the pH probe in the solution and swirl the solution.
3. **Swirl the solution constantly** and record the pH reading in your notebook.
4. Rinse the beaker with a few mL of the next solution and dispose of the rinse solution.
5. Measure/record the pH. Repeat the procedure for the remaining solutions.
6. Discard all solutions in the sink EXCEPT solution #1. You will use this solution in part D.

D. Titration of a weak acid

1. Place 15.0 mL of solution #1 in a 100 mL beaker
2. Add an additional 15.0 mL of distilled water and stir.
3. Add a magnetic stir rod and place the beaker on an unheated stir plate
4. Begin stirring (~200)
5. Position the pH probe in the solution using the articulated arm
6. Rinse and fill a buret with 1.0 M NaOH solution.
7. Position the buret over the stirred beaker.
8. Add the NaOH in 0.2 mL increments. Wait ~5 seconds before recording the burette reading and the pH for each addition.
9. Continue adding base until the pH has leveled off at high pH (pH ~12).
10. Dispose of all solutions in the sink.



IV. Team Report

Page 1:

- **Upper right corner:** Student Names, lab section number and date of experiment
- **Titration Graph: (include labels for ...**
 - Buffering Region
 - * pK_a point (a.k.a. the “*half-equivalence point*” on pg. 622)
 - * equivalence point

*... Note that these points will not necessarily coincide with data points on your graph.

Pages 2 & 3: Work the complete I.C.E. problem for each of the 4 solutions.

Details should include:

- Measured pH
- Conversion to H₃O⁺ calculations
- Dilution calculations (Solutions 2 & 4)
- Balanced chemical equation
- Initial, Change and Equilibrium entries (beneath the equation)
- any important “math”
- Calculated value of K_a
- Δ% (use 1.75 × 10⁻⁵ as the known value)

Page 4:

Answers to the following questions:

1. Use your textbook to define the following buffer-related terms:
 - a. Buffering Region
 - b. endpoint
 - c. equivalence point (include how this is different from endpoint in your answer)
 - d. pK_a point (a.k.a. the “half-equivalence point” on pg. 622)
3. Identify an appropriate indicator for the titration using table 16.1 (pg. 624 in 1st ed.). Your indicator should change color in the vicinity of the equivalence point you identified above.
4. Calculate the number of moles of NaOH that have been added at the equivalence point and include this value on your graph. Compare this value to the number of moles of CH₃COOH that were originally present. They should be equal. Are they?
5. Calculate K_a and Δ% using the pK_a value you obtained from your graph. (K_a = 10^{-pK_a})