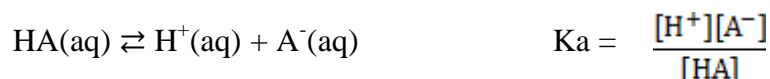


Lab 3: Introduction to Acids Base Chemistry

Part A

Experimental Determination of Acid Dissociation Constant, K_a

The molar concentration (M) of a solution of any acid (weak OR strong) is based on the moles per liter of acid before dissociation occurs. Hence, a 0.10 M solution of a strong acid, such as HCl contains 0.10 M H^+ because 0.10 M HCl dissociates completely to 0.1 M H^+ and 0.10 M Cl^- . However, H^+ concentration ($[H^+]$) of a weak acid is not equivalent to acid molarity, but instead is a function of the acid equilibrium constant for the dissociation reaction:



In this experiment, you will take pH measurements to determine $[H^+]$ ($pH = -\log [H^+]$) of a weak acid and use this information to experimentally determine the dissociation constant, K_a .

MATERIALS

computer
Vernier computer interface
LoggerPro
Vernier pH Sensor
100 mL beaker
2.00 M CH_3COOH

wash bottle
distilled water
100 mL volumetric flask
pipets
pipet bulb
Buffers for pH meter calibration

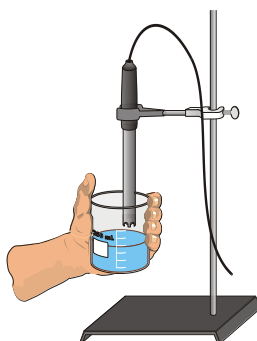


Figure 1

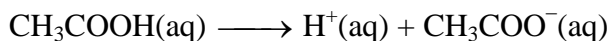
PROCEDURE

1. Obtain and wear safety goggles.
2. Put approximately 50 mL of distilled water into a 100 mL volumetric flask.
3. Use a pipet pump to pipet the required volume of 2.00 M acetic acid (calculated in Pre-Lab) to produce 100 mL of 0.200M acetic acid into the volumetric flask. Fill the flask with distilled water to the 100 mL mark. To prevent overshooting the mark, use a wash bottle filled with distilled water for the last few mL. Mix thoroughly.
4. Use a utility clamp to secure a pH Sensor to a ring stand as shown in Figure 1.

5. Connect the probe to the computer interface. Prepare the computer for data collection by opening the file "27 Acid Dissociation Ka" from the *Chemistry with Vernier* folder of *LoggerPro*. **CALIBRATE pH meter according to instructions given to you by your instructor.**
6. Determine the pH of your solution as follows:
 - Use about 40 mL of distilled water in a 100 mL beaker to rinse the pH Sensor.
 - Pour about 20 mL of your solution into a clean 100 mL beaker and use it to thoroughly rinse the sensor.
 - Repeat the previous step by rinsing with a second 20 mL portion of your solution.
 - Use the remaining 60 mL portion to determine pH. Swirl the solution vigorously. (Note: Readings may drift without proper swirling!) **Record the measured pH reading in your data table.**
 - When done, place the pH Sensor in distilled water, and **set aside your sample of 0.20M acetic acid to use in part B** of this lab (titration).
7. Repeat the procedure preparing a more dilute (0.020 M) solution of acetic acid for experimental determination of K_a . You do not need to save the 0.020 M solution for titration.

Data Treatment

1. Complete a separate ICE tables for each initial acid concentration (blank tables provided on next page)



Initial $[\text{CH}_3\text{COOH}]$ = the **known molarity** of the prepared solution.

2. Use a scientific calculator to determine **equilibrium** $[\text{H}^+]$ from the pH values for each solution:

$$\text{pH} = -\log [\text{H}^+]; \text{ thus } \mathbf{[H^+] = 10^{-\text{pH}} \text{ at equilibrium}}$$

3. Use the obtained value for $[\text{H}^+]$ at equilibrium to calculate all concentrations at equilibrium.
4. Substitute these calculated concentrations into the K_a expression and report your calculated K_a for acetic acid. Compare the values obtained with different initial concentrations of acid.

Data tables are provided, please show all calculations in your lab report.

Recorded Data:

Solution Prep

Sample	M of stock solution	M of Prepared solution desired	V of Prepared solution	Calculated Volume of Stock solution (to obtain desired M)
1	2.00M	0.200	100 mL	
2	2.00M	0.020	100 mL	

pH data

Sample	pH	[H ⁺]
1		
2		

I.C.E. tables

	[HA]	[H ⁺]	[A ⁻]
I			
C			
E (as function of x)			
E (inserting actual values)			

K_a =

	[HA]	[H ⁺]	[A ⁻]
I			
C			
E (as function of x)			
E (inserting actual values)			

K_a =

Post-lab questions

1.) What is the average experimental value for K_a of acetic acid based on your two results? Do your two values agree with each other? Is this what you expected? Why or why not?

2.) The literature value of K_a for acetic acid is 1.8×10^{-5} . Compare this to your experimental value. What sources of error can you identify (be specific – don't just say "measuring errors" only YOU can prevent measuring errors!)

3.) What is the total % dissociation of HA for the 0.20 M sample? _____

What is the total % dissociation of HA for the 0.020 M sample? _____

4.) Which sample had a significant (albeit small) amount of H^+ dissociated, so [HA] at equilibrium was not identical to [HA] initial when expressed to the correct number of significant figures?

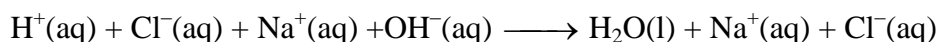
Lab 3 Part B: Titration of an Acid Using pH meter to Detect Endpoint

Acid-Base Titration

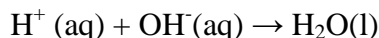
A titration is a process where one can monitor the volume of a solution required for a complete reaction with a specific amount of substance in another solution.

In this experiment, you will monitor the volume of 0.10M NaOH required to neutralize 10 mL of either HCl or acetic acid solutions (One group on each bench will be assigned to HCl and one to acetic acid, and you will share results). Given molarity and having established volume of NaOH required for neutralization, you can determine the moles required, and thus moles of acid present in the 10 mL sample of acid, allowing you to calculate the concentration of the acid.

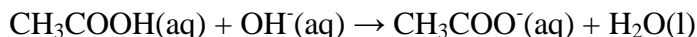
HCl is a strong acid and NaOH is a strong base. The titration reaction that occurs may be represented by the following complete ionic equation:



This equation can be simplified to a net-ionic equation as follows:



Since acetic acid (CH_3COOH) is a weak acid, very little of it dissociates in solution. The net-ionic equation that represents the titration of acetic acid with NaOH is:



In both cases, moles of NaOH added = moles of acid present in the 10mL sample, thus determining moles of NaOH required to complete neutralization will allow you to calculate the concentration of acid in the original sample.

In this experiment, you will use a computer to monitor pH as you titrate. Your data will produce a curve similar to the one shown in figure 1. Initially the acid sample will have a low pH, which increases gradually upon addition of NaOH. The region of most rapid pH change will be used to determine the equivalence point, where moles acid = moles NaOH added. The volume of NaOH used at the equivalence point will be used to determine the moles of NaOH required for neutralization. **Please save your data curve, all pH readings, AND the data curve and readings you obtain from your bench partners for further use later in this course!**

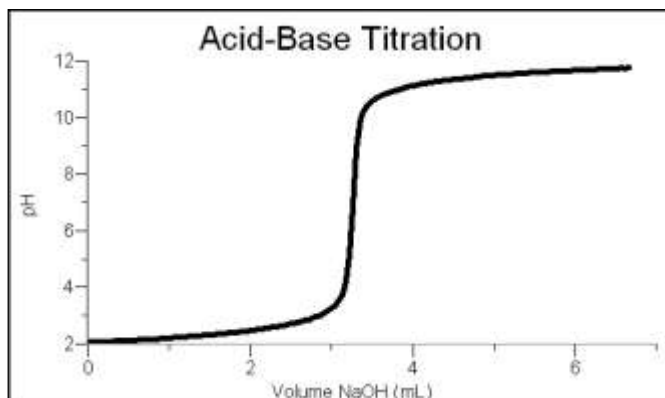


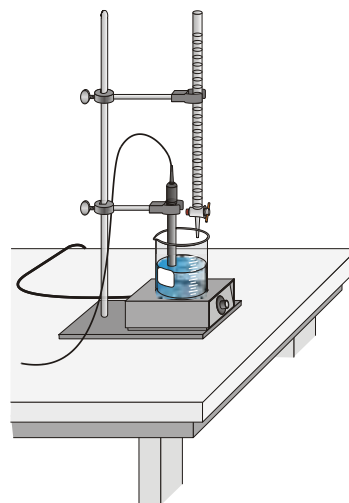
Figure 1

MATERIALS

computer	magnetic stirrer
Vernier computer interface	stirring bar
LoggerPro	wash bottle
Vernier pH Sensor	distilled water
acid solution, unknown concentration	ring stand
~0.1 M NaOH solution	1 utility clamp
pipet bulb or pump	250 mL beaker
50 mL buret	2nd utility clamp
10 mL pipet	2nd 250 mL beaker

Measuring Volume Using a Buret

1. Obtain and wear goggles.
2. Add 50 mL of distilled water to a 250 mL beaker. Use a pipet pump to pipet 10.0 mL of the acid solution into the distilled water in the 250 mL beaker. **CAUTION:** *Handle the acid solutions with care. They can cause burns if in contact with the skin.*
3. Place the beaker on a magnetic stirrer and add a stirring bar.
4. Use a utility clamp to suspend a pH Sensor on a ring stand as shown here. Position the pH Sensor in the acid solution and adjust its position so it will not be struck by the stirring bar. Turn on the magnetic stirrer, and adjust it to a medium stirring rate (with no splashing of solution).
5. Obtain approximately 60 mL of 0.10 M NaOH solution in a 250 mL beaker. Obtain a 50 mL buret and rinse the buret with a few mL of the 0.10 M NaOH solution. Use a utility clamp to attach the buret to the ring stand as shown here. Fill the buret a little above the 0.00 mL level of the buret with 0.10 M NaOH solution. Drain a small amount of NaOH solution into the beaker so it fills the buret tip *and* leaves the NaOH at the 0.00 mL level of the buret. Record the concentration of the NaOH solution (0.10M) in your data table. Dispose of the waste solution from this step as directed by your teacher. **CAUTION:** *Sodium hydroxide solution is caustic. Avoid spilling it on your skin or clothing.*
6. Connect the pH Sensor to the computer interface. Prepare the computer for data collection by opening the file “24a Acid-Base Titration” from the *Chemistry with Vernier* folder. Check to see that the pH value is between 2 and 3.
7. **Before adding NaOH titrant**, click and **monitor pH for 5-10 seconds**. Once the displayed pH reading has stabilized, click . In the edit box, type “0” (for 0 mL added). Press the ENTER key or click to store the first data pair for this experiment.
8. You are now ready to begin the titration. This process goes faster if one person manipulates and reads the buret while another person operates the computer and enters volumes.



- a. Add the next increment of NaOH titrant (enough to raise the pH about 0.15 units). When the pH stabilizes, again click . In the edit box, **type the current buret reading, to the nearest 0.01 mL.** Press ENTER or click . You have now saved the second data pair for the experiment.
 - b. Continue adding NaOH solution in increments that raise the pH by about 0.15 units and enter the buret reading after each increment. Proceed in this manner until the pH is 3.5.
 - c. When a pH value of approximately 3.5 is reached, change to a “one-drop increment.” Enter a new volume reading after each increment. Note: It is important that all increment volumes in this part of the titration be equal; that is, one-drop increments. Do not worry if your drops are not exactly the same volume each time, but don’t start adding larger increments until you reach a pH value of 10.
 - d. After a pH value of approximately 10 is reached, again add larger increments that raise the pH by about 0.15 pH units, and enter the buret level after each increment.
 - e. Continue adding NaOH solution until the pH value remains constant.
9. When you have finished collecting data, click . Dispose of the beaker contents as directed by your teacher.
 10. **Print copies of the table and the graph, and collect a copy of the table and graph from your bench partners. Include boths in your lab report.**
 11. If time permits, repeat the procedure.

Data Treatment

- 1.) MARK YOUR GRAPH! Label to indicate **which acid was titrated** (HCl or CH₃COOH). **Indicate the endpoint on the graph, and the volume and pH at the endpoint.** Do this for both your graph and the one you obtain from your “bench partners.” **Mark the values** for volume and pH at the equivalence point in the data chart as well as labeling the axis on the graph.
- 2.) Calculate moles of NaOH required to reach the equivalence point.
- 3.) Convert moles NaOH required to reach the equivalence point to moles of acid neutralized.
- 4.) Convert moles acid in a 10 mL sample to concentration of acid in 10 mL sample.
- 5.) Enter calculations in chart provided to include in your lab report

Data tables provided, please show all calculations in your lab report.

Data tables

	Volume of ____ M NaOH at equivalence point	Moles (mmoles) NaOH	Moles (mmoles) acid in 10 mL sample	Concentration of acid (M) in 10 mL sample
HCl trial 1				
HCl trial 2*				
CH ₃ COOH trial 1				
CH ₃ COOH trial 2				

Pre-lab questions FOR PART A

- 1.) Determine the volume, in mL, of 2.00 M CH₃COOH required to prepare 100 mL 0.200M CH₃COOH, and a separate 100mL 0.02 M CH₃COOH (show calculation):

Enter this information in the **Solution Prep** portion of your Data Tables.

- 2.) **Complete the ICE tables you will use for data treatment** (leaving the row for actual Equilibrium values blank since you'll need your data for that part), using [HA] initial = 0.200 for one table and [HA] initial = 0.0200 M for the other table. Make sure you complete the row calling for **E as a function of x**.
- 3.) Write the equilibrium constant expression, K_a , for the dissociation of acetic acid, CH₃COOH:
- a) in terms of [CH₃COOH], [H⁺], and [CH₃COO⁻]:
- b) as a function of x (using Equilibrium values f(x) in your ICE table (use the symbol M to represent molarity of acid):
- 4.) The measured pH for a 0.0200M solution of a certain acid, believed to be either Ascorbic OR Citric acid, is 2.46. Complete the ICE table below and calculate the K_a for this acid, then **use Appendix A in your text to determine the correct identity of the acid:**

	[HA]	[H ⁺]	[A ⁻]
I			
C			
E (as function of x)			
E values			

$K_a =$

Identity of the acid: _____

Pre-lab questions FOR PART B

- 1.) What is the **definition** of a strong acid compared to that of a weak acid?
- 2.) Write the NET IONIC equation for each of the titration reactions that you will observe in this experiment:

HCl titration:

CH₃COOH titration:
- 3.) A researcher takes a 10 mL sample of HCl from a bottle and titrates it with NaOH. It was found that 22.4 mL of 0.25 M NaOH were required to reach the equivalence point. What is the concentration of the HCl in original bottle?
- 4.) Would the calculation in question # 3 be different if the original acid was CH₃COOH? Why or why not?