

Determining K_a by the Half-Titration of a Weak Acid

A common analysis of a weak acid or a weak base is to conduct a titration with a base or acid of known molar concentration to help determine the equilibrium constant, K_a , for the weak acid or weak base. If this titration is conducted very carefully and very precisely, the results can lead to a valid approximation of an equilibrium constant. In this experiment, however, you will use a different technique to determine the K_a for a weak acid, acetic acid.

Your primary goal in this experiment is to calculate the K_a of acetic acid. The data that you will use to complete your calculations will come from the reaction of acetic acid with a solution of NaOH. Recall from your work with weak acid-strong base titrations that the point at which a reaction is half-titrated can be used to determine the pK_a of the weak acid. In this experiment, the half-titration point will exist when you have added half as many moles of $\text{HC}_2\text{H}_3\text{O}_2$ as moles of NaOH. Thus, OH^- will have reacted with half of the $\text{HC}_2\text{H}_3\text{O}_2$, leaving the solution with equal moles of $\text{HC}_2\text{H}_3\text{O}_2$ and $\text{C}_2\text{H}_3\text{O}_2^-$. At this point, according to the Henderson-Hasselbalch equation,

$$pH = pK_a + \log \frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}$$

if there are equal moles of $\text{HC}_2\text{H}_3\text{O}_2$ and $\text{C}_2\text{H}_3\text{O}_2^-$ at the half-titration point, then pK_a is equal to the pH value of the solution.

In this experiment, you may find it surprising that you do not need to keep close track of the volume of NaOH titrant added, as you would in most titrations. It is also unusual to conduct a titration without plotting or analyzing a conventional titration curve. This is the nature of a half-titration; it is only important to know when equal amounts of OH^- and $\text{HC}_2\text{H}_3\text{O}_2$ have been added.

OBJECTIVES

In this experiment, you will

- Conduct a reaction between solutions of a weak acid and sodium hydroxide.
- Determine the half-titration point of an acid-base reaction.
- Calculate the K_a and pK_a for the weak acid.

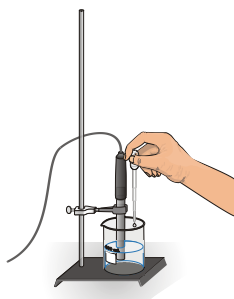


Figure 1

MATERIALS

LabQuest	1.00 M sodium hydroxide, NaOH, solution
LabQuest App	1.00 M acetic acid, HC ₂ H ₃ O ₂ , solution
Vernier pH Sensor	phenolphthalein indicator solution
50 mL buret	distilled water
buret clamp	magnetic stirrer and stirring bar
250 mL beaker	plastic Beral pipets
two ring stands	utility clamp

PROCEDURE

1. Obtain and wear goggles.
2. Use a buret clamp to connect a 50 mL buret to a ring stand. Rinse and fill the buret with 1.00 M acetic acid solution. *Handle the acetic acid with care. It can cause painful burns if it comes into contact with the skin.*
3. Transfer precisely 25.0 mL of the acetic acid solution to a 250 mL beaker.
4. Use a plastic Beral pipet to remove a small volume of the acetic acid from the 250 mL beaker. Draw enough acetic acid into the pipet so that the bulb is about 1/4 full. Carefully set aside the pipet of acid, to be used later.
5. Add 1–2 drops of phenolphthalein indicator solution to the beaker of acetic acid.
6. Connect the pH Sensor to LabQuest and choose New from the File menu. If you have an older sensor that does not auto-ID, manually set up the sensor.
7. Obtain about 50 mL of 1.00 M NaOH solution. **CAUTION:** *Sodium hydroxide solution is caustic. Avoid spilling it on your skin or clothing.*
8. Begin the half-titration.
 - a. Place the beaker of acetic acid on a magnetic stirrer and add a stirring bar.
 - b. Set up a ring stand and clamp to hold the pH Sensor in place (see Figure 1). Position the pH Sensor in the beaker so that the tip of the probe is completely immersed.
 - c. Gently stir the acetic acid solution.
 - d. Do not start data collection. Monitor the pH of the reaction mixture on LabQuest.
 - e. Use a new plastic Beral pipet to slowly add the 1.00 M NaOH solution, in ~1 mL increments, to the beaker of acetic acid solution (see Figure 1).
9. Conduct the titration carefully. As the reaction approaches the equivalence point, at about pH 6, add the NaOH solution drop by drop. When you reach the equivalence point, the pH will increase rapidly and the indicator will change color. If necessary, add another drop of NaOH, so that the reaction is slightly past the equivalence point. Remember that the pH will not increase rapidly beyond the equivalence point (pH ~10).
10. Add all of the acetic acid from the Beral pipet, which you removed in Step 4, to the beaker of reaction mixture. Check the pH readings and observe the indicator color. The mixture should be slightly acidic once again.

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- Carefully add NaOH, drop by drop, to the beaker of reaction mixture, until you reach the equivalence point as precisely as possible. A very slight pink color of the phenolphthalein indicator is visible. This is your half-titrated solution, because you have neutralized precisely 25.0 mL of the original 50.0 mL of acetic acid that you measured out into the buret.
- Transfer the remaining 25.0 mL of acetic acid from the buret to the 250 mL beaker of reaction mixture. Stir the solution in the beaker thoroughly. Read and record the pH of the solution in the beaker.
- When you have finished the testing, dispose of the reaction mixture as directed. Rinse the pH Sensor with distilled water in preparation for a second trial. Repeat the necessary steps to test a new sample of the acetic acid solution.

DATA TABLE

Titration Results	Trial 1	Trial 2
Equivalence point pH		
pH of half-titrated solution		

DATA ANALYSIS

- Calculate the pK_a and K_a using the results of your testing.
- Find the accepted values for the pK_a and K_a of acetic acid. How well do the accepted values compare with your calculated values? Explain.
- Explain why the pH at the half-titration point is equal to the pK_a in your experiment.
- Explain how this test could be done using only an indicator solution and no electronic means of measuring pH.