# Titration of Buffered and Unbuffered Solutions

**Description**
One important property of weak acids and weak bases is their ability to form buffers. A buffer is the combination of a weak acid and a salt of the weak acid, or a weak base and a salt of the weak base. Acetic acid and sodium acetate are an example of this kind of buffer pair. Buffers resist changes in pH upon the addition of small amounts of $H^+$ or $OH^-$ ions. The dissociation equation for acetic acid contains both of the buffer components, $HC_2H_3O_2$ and $C_2H_3O_2^-$:

$$HC_2H_3O_{2(aq)} \leftrightarrows H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

When a small amount of an HCl solution is added to the buffer solution, most of the $H^+$ ions are removed when they react with acetate ions:

$$H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)} \rightarrow HC_2H_3O_{2(aq)}$$

When a solution of NaOH is added to the buffer, most of the $OH^-$ ions are removed when they react with acetic acid molecules:

$$OH^-_{(aq)} + HC_2H_3O_{2(aq)} \rightarrow H_2O_{(\ell)} + C_2H_3O_2^-{}_{(aq)}$$

In this lab, you will be "virtually titrating" an acid buffered by a salt containing its conjugate base, as well as titrating the unbuffered acid. You will graph the titration curves produced by the two titrations and compare the results.

**Materials**
Carnegie Mellon University's "Virtual Laboratory" Software
Lab Book
Graph paper

**Procedure**
1. Open the Virtual Lab Applet
   http://www.sciencegeek.net/VirtualLabs/VLab.html
2. Open the Lab
   File → Load Homework → Allan → Buffers
3. First, we will titrate the unbuffered solution.
   a) In a 250 mL Erlenmeyer flask (tools → glassware → Erlenmeyers), prepare 2.0 mL of 0.1 M acetic acid using the stock 1 M $CH_3COOH$ solution and distilled water. In your calculations section, show the method you used to determine the amounts of each needed to achieve the proper dilution. Add 0.01 mL of phenolphthalein to your beaker when the solution is prepared.
   b) Fill a 50 mL buret (tools → glassware → 50 mL buret) with 45 mL of 0.1 M NaOH stock solution.
   c) Add NaOH to your acetic acid solution, 0.15 mL at a time. In a table in your lab book, record the pH after each addition of base. Also, record the step in the process where the indicator undergoes its color change. Stop the titration when a total volume of 3.0 mL of NaOH has been added.

4. Second, we will titrate a buffered solution.
   a) In a 250 mL beaker, prepare 2.0 mL of a buffered solution that is 0.1 M acetic acid and 0.1 M in $NaCH_3COO$ as well. In your calculations section, show the method you used to determine the amounts of each needed to achieve the proper dilution. Add 0.05 mL of phenolphthalein to your beaker when the solution is prepared.
   b) Add NaOH to your buffered solution, 0.15 mL at a time. In a table in your lab book, record the pH after each addition of base. Also, record the step in the process where

the indicator undergoes its color change. Stop the titration when a total volume of 3.0 mL of NaOH has been added.

## Data

This section should include appropriate tables in which you have recorded "Volume of NaOH added" and "Solution pH."

## Calculations and Graphs

1. Graph each of the titrations, on GRAPH paper. Volume of base added should be the independent variable. The solution pH should be the dependent variable. Be certain to label the axes on your graph, as well as marking an appropriate scale on each. Graph both titrations on THE SAME GRAPH to make comparisons easier.
2. On each graph, label the point at which you first noticed a change in indicator color.
3. In association with your titration graph for the first titration (unbuffered), show the law of mass action for the dissociation of acetic acid, and calculate the THEORETICAL value of the pH for a 0.1 M solution of acetic acid ($K_a = 1.8 \times 10^{-5}$) at 25 °C.
4. In association with your titration for the buffered solution, show the Henderson-Hasselbalch equation for the buffer system, and calculate the THEORETICAL value of the pH for the buffered solution prior to the addition of base.
5. If your lab book does not have "graph paper pages" then attach your graphs with staples or tape in this section of your lab write-up. You may also use Excel or a similar application to generate a graph. Be certain to label the axes appropriately.

## Conclusion

At a minimum, address the following questions…

1. In what way were the graphs of your two titrations different? What does this say about the effect of a buffer in a solution?
2. What was the relationship between the ending pH of two titrations? What does this say about the limitations of buffered solutions?

## Discussion of Theory

Address the "WHY" of the observations made in your conclusion. In other words, provide a cogent explanation for the observations made in the lab.

## Error Analysis

1. How close was the agreement between the software pH and your calculated initial pH for each solution? Does the software algorithm for calculating pH appear to be accurate?

## Questions for Discussion

1. Describe the usefulness of buffers in commercial food products such as the lemonade which uses citric acid but buffers with a citrate salt, usually sodium citrate.
2. Aspirin is also commonly buffered. What is the acid in aspirin, and what is used to complete the "buffer system?"
3. Why is the 1:1 ratio of acid to conjugate base (or vice-versa) referred to as a "best buffer?"