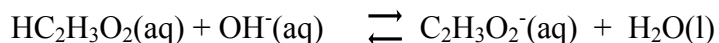


# Acid-Base Buffers

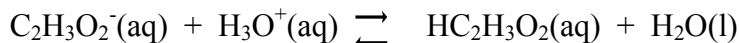
## Introduction

In this experiment we will focus on the topic of acid-base buffers. An acid-base buffer is a solution that resists pH change. Buffers are very important in chemistry since many reactions will only occur in certain pH ranges. This is especially true of many biological systems in which the pH must be maintained in very narrow ranges if the organism is to survive.

Buffers are solutions that simultaneously contain relatively large amounts of acid/base conjugate pairs. An example that you are already familiar with is the acetic acid/acetate ion conjugate pair. A solution containing both of these ions will be a buffer because the weak acid will react with added base to produce the conjugate base via:

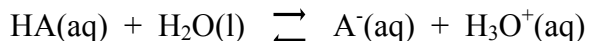


and the conjugate base present will react with added acid to produce the conjugate acid via:



In both cases the pH will change with the addition of acid or base, however the pH will change very little if the amounts of added base or acid is small relative to the concentration of the buffer conjugates already present in the solution.

Additionally, a buffer works best when the pH is about the same as the  $\text{pK}_a$  for the acid component of the buffer. To illustrate this, consider the reaction:



for which the  $K_a$  expression is:

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

If we take the  $-\log$  of both sides then we have,

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

or

$$\text{pH} = \text{pK}_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

Considering the 2<sup>nd</sup> term in the above equation, we see that in order for the pH change to be minimal, the contribution of the logarithm must be small. In fact, the logarithm will be zero if  $[A^-] = [HA]$  since the  $\log 1 = 0$ . Therefore, as strong acids or bases are added, we can expect a buffer solution to work best at stabilizing the pH when  $[A^-] = [HA]$ . If the pH is the same as the  $pK_a$ , it follows that  $[A^-] = [HA]$ .

The above equation can also be used to determine the conjugate acid-base concentrations required to make a buffer of specified pH. We can rearrange this equation to express the conjugate acid-base concentration ratio in terms of pH. We do this by subtracting  $pK_a$  from both sides of the equation then taking the antilog of both sides. Recall that the antilog function is  $10^x$ .

$$\frac{[A^-]}{[HA]} = 10^{(pH - pK_a)}$$

Given a target pH for the buffer and a desired concentration for either the conjugate acid or base, one can then find the concentration and thus a mass or volume required of the unspecified conjugate to complete the buffer solution.

Table 1 contains a list of useful  $pK_a$  values needed for this lab.

**Table 1.  $pK_a$  values for Acids used in the experiment.**

Name of Acid	Dissociation Reaction	$pK_a$
Acetic acid	$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$	4.74
Dihydrogen phosphate ion	$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{2-}(aq)$	7.20
Hydrogen carbonate ion	$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CO_3^{2-}(aq)$	10.33

In this experiment, you will prepare three buffers and study the effects of adding acid and base. For each of the buffers you will calculate the amounts of the conjugates required to prepare the buffer solutions. Then you will make small additions of acid and base to the buffer solutions and observe the pH changes that occur. You will graph these pH changes against volume and make comparisons to the previous experiments.

**As preparation for this experiment you should read sections 16.1 through 16.2 of Petrucci, 8<sup>th</sup> ed.**

## **Pre-Lab Preparation**

The calculations for this experiment are not trivial. For this reason you are required to prepare for this experiment by calculating the needed amounts of your reagents to make your buffer solutions at the assigned pH values. You should have been assigned a group number during the previous laboratory session. (You were asked to write it down in your lab manual immediately following the introduction of the Polyprotic System Experiment.) Table 2 identifies the assigned pH values by group numbers. If you do not complete the calculations before the laboratory session you may not have time to complete this experiment.

**Table 2. pH of Buffer Solutions**

	<b>Acidic Buffer</b>	<b>Neutral Buffer</b>	<b>Basic Buffer</b>
<b>Group 1</b>	4.4	7.4	9.9
<b>Group 2</b>	4.5	7.5	10.0
<b>Group 3</b>	4.6	6.8	10.1
<b>Group 4</b>	4.7	6.9	10.2
<b>Group 5</b>	4.8	7.0	10.3
<b>Group 6</b>	4.9	7.1	10.4
<b>Group 7</b>	5.0	7.2	10.5
<b>Group 8</b>	5.1	7.3	10.6

You must have the calculation checked by the teaching assistant before you can begin the laboratory experiment

**Safety: Remember to always wear gloves when handling all acids and bases.  
WEAR YOUR GOGGLES!**

**You will be working in groups of three for this experiment.**

**Each student must collect data and submit a separate report.** The actual data analyses and the written reports must be done entirely independently of your lab partners or other students. Make sure that you avoid unauthorized collaboration and plagiarism. All suspected violations of the Code of Academic Conduct will be referred to Student Judicial Affairs.

### **Experimental Procedure**

#### **Part I. Preparing your Buffers – Correct first paragraph**

Working in groups of three, you will be preparing three 250 ml buffer solutions, an acidic, neutral and basic buffer. The acidic buffer will be prepared from a 6 M acetic acid solution and 2.5 M sodium acetate trihydrate. The neutral buffer is prepared from a 2.5 M sodium dihydrogen phosphate solution and solid sodium hydrogen phosphate heptahydrate. The basic buffer solution is prepared from solid sodium hydrogen carbonate and solid sodium carbonate monohydrate. After preparing the three solutions you will measure the pH level of the solution and adjust the levels by adding either strong acid or strong base as needed.

1. The TA will assign your group a number between 1 and 8. Each group is to prepare the 3 buffer solutions at the designated pH values given in Table 2. Have one group member prepare the 250 mL of the acidic buffer; another group member prepare 250 mL of the neutral buffer; and the third group member prepare 250 mL of the basic buffer at the designated pH values. You will need to share your 250 mL volumetric flasks, volumetric pipets, and graduated cylinders. You will also need to condition this glassware between measurements to avoid contamination.

#### Preparation of Acidic Buffer

2. Calculate and measure the volume of 2.5 M sodium acetate solution needed to prepare 250 mL of 0.2 M sodium acetate. Use the appropriate volumetric pipet to measure this volume. Transfer this volume to your 250 mL volumetric flask. Add about 100 mL of deionized water to the 250 mL volumetric flask and mix. Calculate the volume of 6 M stock acetic acid solution you need to make 250 mL of the buffer solution to the designated pH. You may use your graduated cylinder to measure this volume. Add this volume of 6 M acetic acid to your 250 mL volumetric flask that contains the sodium acetate solution. Now, add sufficient deionized water to the volumetric flask to bring the total volume to 250 mL. After you have mixed the buffer solution well, place the buffer solution into a clean and appropriately labeled 250 mL or 400 mL Erlenmeyer flask.

### Preparation of Neutral Buffer

3. Calculate and measure the volume of 2.5 M stock  $\text{NaH}_2\text{PO}_4$  solution needed to prepare 250 mL of 0.1 M  $\text{NaH}_2\text{PO}_4$ . Use the appropriate volumetric pipet to measure and transfer this volume to a clean 250 mL volumetric flask. Add about 100 mL of deionized water to the 250 mL volumetric flask and mix. Calculate the mass of solid  $\text{Na}_2\text{HPO}_4 \cdot 7 \text{H}_2\text{O}$  you need to make 250 mL of the buffer solution to the designated pH. Add this mass of  $\text{Na}_2\text{HPO}_4$  to your 250 mL volumetric flask that contains the diluted  $\text{NaH}_2\text{PO}_4$ . Now, add sufficient deionized water to the volumetric flask to bring the total volume to 250 mL. After you have mixed the buffer solution well, place the buffer solution into a clean and appropriately labeled 250 mL or 400 mL Erlenmeyer flask.

### Preparation of Basic Buffer

4. Calculate and weigh out the grams of solid sodium hydrogen carbonate needed to prepare 250 mL of 0.1 M sodium hydrogen carbonate solution. Transfer this mass to a clean 250 mL volumetric flask and add about 100 mL of deionized water to the volumetric flask. Calculate the mass of solid sodium carbonate monohydrate required to make the basic buffer to the designated pH. Weigh out and add this mass of sodium carbonate monohydrate to your 250 mL volumetric flask that contains the diluted sodium hydrogen carbonate. Now, add sufficient deionized water to the volumetric flask to bring the total volume to 250 mL. After you have mixed the buffer solution well, place the buffer solution into a clean and appropriately labeled 250 mL or 400 mL Erlenmeyer flask.
5. Standardize the pH meter nearest your work area using the three buffer solutions at pH 4.00, pH 7.00, and pH 10.00 following the specific directions for use of the pH meters that are provided in the laboratory. Be sure to use your wash bottle to rinse off the pH electrode into a waste beaker before and after using it in each of the buffer solutions. Wipe the end of the electrode with a Kimwipe before placing it in a new solution. Always keep the pH electrode wet when not in use. For long-term storage place the electrode in a beaker of tap water or a container of pH 7.00 buffer solution.
6. For each of the 3 buffer solutions: Measure the pH. Using a disposable pipet, add either 6M HCl or 6M NaOH to adjust the pH until it is equal to the assigned pH. (See Table 2 above.) Stir the solution and record the pH to the nearest 0.02 pH unit. Rinse and wipe the electrode before placing it in a new solution.

## **Part II. Preparing your Reagents**

**\*Share your 0.2 M HCl & NaOH and your 0.2 M acetic acid with another group**

1. Have one group prepare the 250 mL of 0.2 M HCl from the 6 M HCl stock solution using a 250 mL volumetric flask. You may use your graduated cylinder to measure the 6 M HCl volume. Split the 0.2 M HCl solution in half by pouring it into two labeled 125 or 250 mL Erlenmeyer flasks. Share half of the solution with another group.
2. Have a second group prepare 250 mL of 0.2 M NaOH from the 6 M NaOH stock solution using a 250 mL volumetric flask. You may use your graduated cylinder to measure the 6 M NaOH volume. Split the 0.2 M NaOH solution in half by pouring it into two labeled 125 or 250 mL Erlenmeyer flasks. Share half of the solution with another group.
3. One of the two groups should prepare 250 mL of 0.2 M acetic acid from the 6 M acetic acid stock solution using a 250 mL volumetric flask. You may use your graduated cylinder to measure the 6 M acetic acid volume. Split the 0.2 M acetic acid solution in half by pouring it into two labeled 125 or 250 mL Erlenmeyer flasks. Share half of the solution with another group.
4. Each group needs to condition and fill two 50 mL burets, one with the 0.2 M HCl solution and the other buret with 0.2 M NaOH. Record the initial volume of HCl and NaOH to the nearest 0.02 mL. Label buret and flask

## **Part III. Addition of 0.2 M HCl**

In this part of the experiment each group will treat each of their three buffer solutions with 0.2 M HCl solution. After each addition you will measure the pH of the solution. You will then plot the pH vs. added volume of HCl to graphically observe the pH changes that occur. You will also explore the effect of adding 0.2 M HCl to 0.2 M acetic acid solution.

1. Place 50 mL of the acetic acid/acetate ion buffer in a clean and dry 250 mL beaker. Gently, place a stir bar into the 250 mL beaker containing buffer solution.
2. Set up the beaker containing the buffer, stir plate, electrode under the buret containing the 0.2 M HCl solution. Start gently rotating the stir bar.
3. Record the initial buret reading and pH meter reading. Add approximately 2 mL of HCl to the buffer. Record the buret reading to the nearest 0.02 mL. Stir the solution and measure the pH. Record the pH to the nearest 0.02 pH unit when the reading has stabilized.

4. Repeat step 3 until the pH of the buffer solution decreases by 1.5 pH units.
- 5. Save this solution for use in step 8 in Part IV.**
6. Repeat steps 1-4 using your dihydrogen phosphate ion/hydrogen phosphate ion buffer. Record your volumes and pH levels carefully. You do not need to save this solution.
7. Repeat steps 1-4 using your final buffer, hydrogen carbonate ion/carbonate ion solution. You may need to clean out your 250 mL beaker or use a 150 mL beaker. Record your volumes and pH levels carefully. You do not need to save this solution.
8. Place 50 mL of the 0.2 M acetic acid solution in a clean and dry 100 mL beaker. Gently, place a stir bar into the 100 mL beaker containing acetic acid solution. Repeat steps 2-4 using the 0.2 M acetic acid solution instead of a buffer solution. Record your volumes and pH levels carefully. You do not need to save this solution.

#### **Part IV. Addition of 0.2 M NaOH**

In this part of the experiment you will treat each of the three buffer solutions with 0.2 M NaOH solution. After each addition you will measure the pH of the solution. You will then plot the pH vs. added volume of NaOH to graphically observe the pH changes that occur. You will also explore the effect of adding 0.2 M NaOH to 0.2 M acetic acid solution.

1. Place a fresh 50 mL sample of the acetic acid/acetate ion buffer in a clean and dry 250 mL beaker. Gently, place a stir bar into the 250 mL beaker containing buffer solution. It is important that your beakers are clean and not contaminated with buffer/HCl mixture.
2. Set up the beaker containing the buffer, stir plate, electrode under the buret containing the 0.2 M NaOH solution. Start gently rotating the stir bar.
3. Add approximately 2 mL of NaOH to the buffer. Record the buret reading to the nearest 0.02 mL. Stir the solution and measure the pH. Record the pH to the nearest 0.02 pH unit when the reading has stabilized.
4. Repeat step 3 until the pH of the buffer solution increases by 1.5 pH units.
5. Repeat steps 1-4 using a fresh sample of your dihydrogen phosphate ion/hydrogen phosphate ion buffer. Record your volumes and pH levels carefully. You do not need to save this solution.

6. Repeat steps 1-4 using a fresh sample of your final buffer, hydrogen carbonate ion/carbonate ion solution. Record your volumes and pH levels carefully. You do not need to save this solution.
7. Place a fresh 50 mL sample of the 0.2 M acetic acid solution in a clean and dry 100 mL beaker. Gently place a stir bar into the 100 mL beaker containing acetic acid solution. Repeat steps 2-4 using the 0.2 M acetic acid solution instead of a buffer solution. Record your volumes and pH levels carefully. You do not have to save this solution.
8. Gently place a stir bar into the 250 mL beaker containing the **acetic acid/acetate ion buffer and HCl mixture reserved from step 5 of Part III**. Repeat steps 2-4 using this solution but adding NaOH until the pH increases by about 3.0 pH units. Record your volumes and pH levels carefully.

Clean-Up: All solutions can go down the drain with copious amounts of water. Be sure to rinse the electrode with deionized water before placing it in the storage solution. Note that it is critical to store the electrode in this solution at the end of the laboratory period.

### Data Analysis

#### **Part I.**

1. What was your assigned pH value of your acidic buffer?
2. What volume of 2.5 M sodium acetate was needed to make the 250 mL acetic acid/acetate ion buffer at that pH? Show your calculations.
3. What was the concentration of the acetic acid in the 250 mL acetic acid/ acetate ion buffer solution?
4. What volume of the 6 M acetic acid solution was needed to prepare the 250 mL acetic acid/acetate ion buffer solution?
5. What was your assigned pH value of your neutral buffer?
6. What volume of the 2.5 M stock sodium dihydrogen phosphate was needed to prepare 250 mL of the neutral buffer that was 0.1 M in sodium dihydrogen phosphate? Show your calculations.
7. What was the concentration of the sodium hydrogen phosphate in the dihydrogen phosphate ion/hydrogen phosphate ion buffer solution? Show your calculations.
8. What mass of sodium hydrogen phosphate heptahydrate was needed to make the 250 mL dihydrogen phosphate ion/hydrogen phosphate ion buffer at that pH? Show your calculations.

9. What was your assigned pH value of your basic buffer?
10. What mass of sodium hydrogen carbonate was needed to make the buffer solution 0.1 M in sodium hydrogen carbonate? Show your calculations.
11. What was the concentration of the sodium carbonate in the hydrogen carbonate ion/ carbonate ion buffer solution? Show your calculations.
12. What mass of sodium carbonate monohydrate was required to make the 250 mL hydrogen carbonate ion/carbonate ion buffer at that pH? Show your calculations.
13. Provide reasons as to why the measured pH level is different from the calculated value.

## Part II

14. What volume of the 6 M HCl stock solution is needed to prepare 100 mL of 0.2 M HCl? Show your calculation.
15. What volume of the 6 M NaOH stock solution is needed to prepare 100 mL of 0.2 M NaOH?

## Part III

16. Using graph paper or a spreadsheet program, such as Excel make the following graphs.
  - a. Plot pH vs. added volume of HCl for all 3 buffer solutions and the 0.2 M acetic acid solution. You should have four separate graphs when you are finished. Label each graph appropriately.
  - b. Plot pH vs. added volume of NaOH for all 3 buffer solutions and the 0.2 M acetic acid solution. You should have four separate graphs when you are finished. Label each graph appropriately.
  - c. Plot pH vs. added volume of NaOH for the acetic acid/acetate ion buffer and HCl mixture used in step 8 of Part IV. Label the graph appropriately.
17. Let's begin comparing corresponding graphs. First, take the two graphs of the acidic acid buffer. Line up these two graphs along the pH axis (y-axis) and the volume axis (x-axis). One graph will be on top of the other. Flip the top graph 180°, keeping the pH axis aligned. After the flip, the volume axis will be lined up end-to-end. You should now have a curve that looks like an "S" laying on its side and one of the graphs will be face down. Hopefully your graph paper is "see through." What is the pH range over which the buffer effectively neutralizes the added acid and base and maintains a reasonably constant pH? This is referred to as the buffer range.

18. Compare the acidic buffer graph constructed in question 15 to the graph you made in question 14c. What are the differences, if any? How do the buffer ranges compare?
19. Repeat the procedure described in question 15 for the graphs involving the neutral buffer. What is the buffer range for the neutral buffer?
20. Repeat the procedure described in question 15 for graphs involving the basic buffer. What is the buffer range for the basic buffer?
21. Considering the ranges of pH of each buffer, write an equation in terms pH and  $pK_a$  that defines buffer range.
22. Buffer capacity is defined as the amount of acid or base that can be added to a buffer before any substantial change in pH. When is the buffer capacity at its maximum?
23. Repeat the procedure described in question 15 for the graphs involving the 0.2 M acetic acid solution. How do these graphs compare to the graphs of the acetic acid/acetate ion buffer? For example, compare the slopes of the curve of each graph at corresponding points. At corresponding pH values, compare how much HCl or NaOH is added before  $\Delta pH = 1$ .
24. Consider the titration curve you plotted for the "Titration of a Weak Acid" experiment. How does the titration curve compare to the graph involving the acetic acid/acetate ion buffer in this experiment? Does the titration curve include a buffer region? If so, where is the buffer region? If not, why not?

**Conclusion** After reflecting on the nature of buffer solutions, and their effectiveness over pH different ranges compose a few paragraph summary of this experiment.