

## EXPERIMENT E2

### The Properties of Buffers: Resisting Change in a Turbulent World

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(Modified version of D. Van Dinh and University of Michigan General Chemistry Laboratory Manual)

#### I. OBJECTIVES

- Investigate strengths of acids
- Learn how to use pH meter
- Investigate how buffers work.
- Prepare a buffer solution with a target concentration and pH.
- Determine the buffer capacity of a solution.



#### II. INTRODUCTION

Many different systems require the control of the conditions so that radical changes do not occur when the system is stressed in some way. For example, our body temperatures remain stable in hot and cold days through perspiration and increased metabolism. In doing so, the effect of stress due to temperature changes is lessened. In everyday language, a buffer is something that lessens the impact of an external force. One of the most important examples of a system requiring controlled conditions through a buffer is our blood which is a  $\text{H}_2\text{CO}_3/\text{HCO}_3^-$  system. Human blood requires a pH between 7.35 and 7.45. If the pH of the blood drops below 6.9 or rises above 7.8, then death is likely. Even within the “safe” range, the optimum pH for many bodily processes is quite narrow; therefore slight disturbances to the pH can significantly impair normal processes. A disturbance that can severely alter the blood pH is cardiac arrest. When the heart stops, metabolic acidosis sets in. Lactic acid and  $\text{CO}_2$  cannot be removed from the blood and the pH is dramatically lowered. By applying their knowledge of buffers, hospital emergency room personnel are able to quickly administer the fluids necessary to help restore the victim’s blood to its normal state and stave off death (Figure 1). How do they know what concentration of fluids to use in order to restore the blood pH to its normal range?



Figure 1. A buffered intravenous drip

### III. BACKGROUND

#### A. The properties of Buffers

Chemists apply the concepts of a buffer to solutions with the ability to resist changes in pH. In many areas of research, chemists need an aqueous solution that resists a pH change when hydrogen ions from a strong acid, or hydroxide ions from a strong base, are added. An acid – base buffer is able to resist changes in its pH by containing an acidic component that can neutralize added  $\text{OH}^-$  ions and a basic component that can neutralize added  $\text{H}^+$  (or  $\text{H}_3\text{O}^+$ ) ions. Most commonly, the components of a buffer are the conjugate acid – base pair of a weak acid. For example, an aqueous mixture of weak acid such as acetic acid,  $\text{CH}_3\text{COOH}$ , and its conjugate base acetate,  $\text{CH}_3\text{COO}^-$ , (simply formed by a solution of sodium acetate added to the weak acid). Buffers work through a phenomenon known as the common – ion effect. The essential feature of a buffer is that it consists of high concentrations of the acidic ( $\text{CH}_3\text{COOH}$ ) and basic  $\text{CH}_3\text{COO}^-$ . This allows the relative concentrations of the buffer components to stay about the same when small amounts of  $\text{H}^+$  or  $\text{OH}^-$  ions are added to the buffer: the acidic component will neutralize any added base and the basic component will neutralize any added acid (Figure 2).

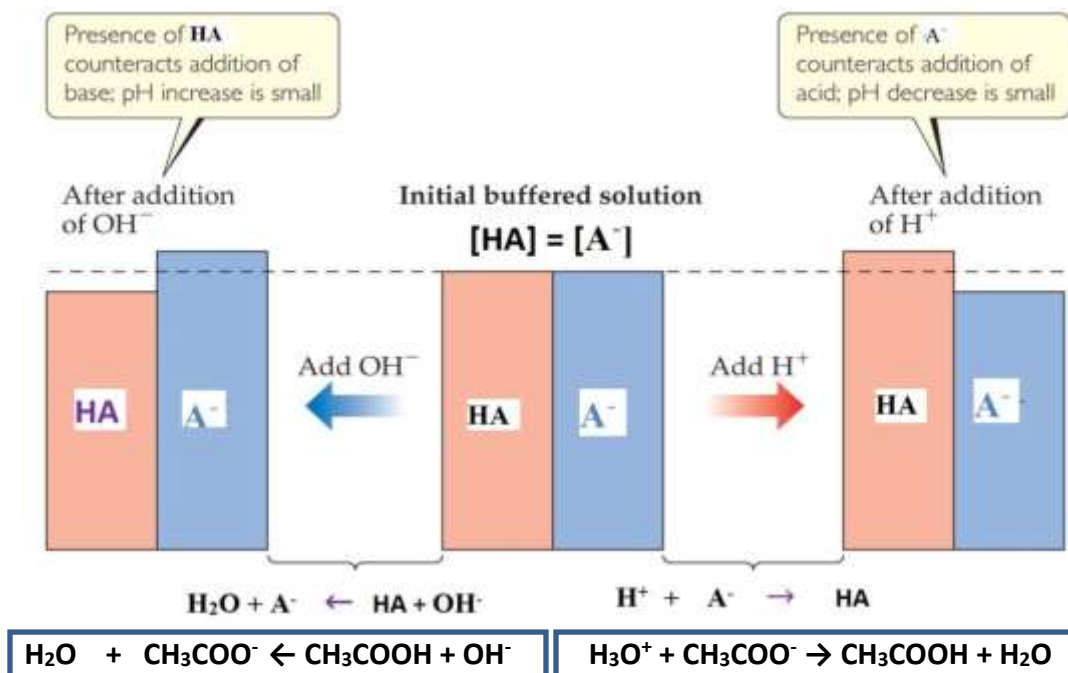
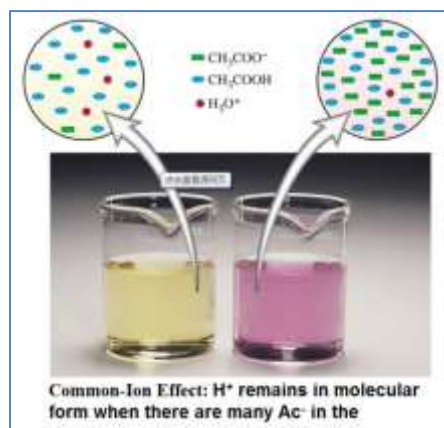


Figure 1. A buffer works by neutralizing small amounts of added acid or base

It is important to note that a buffer will only work when the amount of  $H^+$  and  $OH^-$  added is much smaller than the amounts of acid – base components of the buffer present. The amount of strong acid or strong base that can be added to a buffer depends on its buffer capacity. The added ions have little effect on the pH because one or the other buffer components consumes them. To help you better understand this concept, let's look at an example.

Consider what happens when small amounts of strong acid or base are added to a buffer containing high  $[C_2H_5COOH]$  and  $[C_2H_5COO^-]$ . As a weak acid, propanoic acid dissociates only slightly in water (Equation 1):



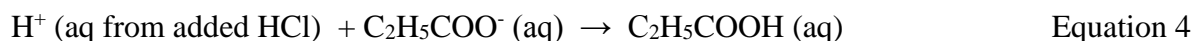
The expression for the acid dissociation constants is:

$$K_a = \frac{[H^+][C_2H_5COO^-]}{[C_2H_5COOH]} \quad \text{Equation 2}$$

Noting that  $K_a$  is a constant and solving for  $[H^+]$ , we see that the  $[H^+]$  of a solution is dependent on the conjugate acid – base pair concentration ratio:

$$[H^+] = K_a \times \frac{[C_2H_5COOH]}{[C_2H_5COO^-]} \quad \text{Equation 3}$$

From Equation 3, you can see that if the ratio of the acid to base increases, the  $[H^+]$  increases. Likewise, if the ratio of the acid to base decreases, the  $[H^+]$  decreases. When a small amount of strong acid is added, such as HCl, the increased amount of  $H^+$  reacts with a nearly stoichiometric amount of  $C_2H_5COO^-$  in the buffer to form more  $C_2H_5COOH$ . (Equation 4)



As a result, the  $[C_2H_5COO^-]$  goes down by the amount of  $[H^+]$  and the  $[C_2H_5COOH]$  goes up by the same amount. This increases the concentration ratio of the acid to base components of the buffer by a slight amount, resulting in a small pH drop. Adding a small amount of a strong base, such as NaOH, produces the opposite results (pH decreases with increased acidity or increases with increased basicity,  $pH = -\log [H^+]$ ).

## B. Preparing a Buffer

As stated in the introduction, many times emergency room personnel are required to administer a buffer solution that simulates the exact concentration and pH of normal human blood. How do they prepare a buffer with a specific concentration and particular pH? In order to calculate the expected pH for a buffer, the **Henderson–Hasselbalch** equation is used. The Henderson–

Hasselbalch equation is a mathematical expression useful for buffer calculations. Taking the negative logarithm of both sides of Equation 3 gives:

$$-\log[H^+] = -\log K_a - \log \frac{[C_2H_5COOH]}{[C_2H_5COO^-]} \quad \text{Equation 5}$$

Substituting definitions of  $pH$  and  $pK_a$  into Equation 5, we obtain:

$$pH = pK_a + \log \frac{[C_2H_5COO^-]}{[C_2H_5COOH]} \quad \text{Equation 6}$$

In general, the Henderson–Hasselbalch equation for any conjugate acid–base pair can be written as:

$$pH = pK_a + \log \frac{[base]}{[acid]} \quad \text{Equation 7}$$

In preparing a buffer, the correct [base]/[acid] ratio must be determined. Knowing the target pH, one can use the Henderson–Hasselbalch equation in order to calculate the correct concentrations of acid–base ratio to use in a buffer. The ratio gives some idea of the stability of the buffer. The most stable buffers have a ratio of 1:1. This ratio and the target pH are vital to know because it helps us to determine the conjugate acid–base pair to use when preparing a buffer.

For example, imagine you have to prepare 0.250 L of a  $pH = 5.19$  buffer solution of propanoic acid with a total concentration of 0.0500 M from stock solutions of 0.100 M propanoic acid and 0.100 M sodium propanate ( $C_2H_5COONa$ ). Propanoic acid ( $C_2H_5COOH$ ) has a  $pK_a$  of 4.89.

To determine the correct  $[C_2H_5COO^-]/[C_2H_5COOH]$  ratio, rearrange Equation 6 to obtain:

$$pH - pK_a = \log \frac{[C_2H_5COO^-]}{[C_2H_5COOH]} \quad \text{Equation 8}$$

$$5.19 - 4.89 = 0.30 = \log \frac{[C_2H_5COO^-]}{[C_2H_5COOH]} \quad \text{Equation 8a}$$

$$\text{Or } 10^{0.30} = 2.00 = \frac{[C_2H_5COO^-]}{[C_2H_5COOH]} \quad \text{Equation 8b}$$

$$[C_2H_5COO^-] = 2.00[C_2H_5COOH] \quad \text{Equation 8c}$$

This tells us that the amount of  $[C_2H_5COO^-]$  must be twice the amount of  $[C_2H_5COOH]$ . Next, we determine the actual concentrations of the weak acid and the weak base to use in the buffer to meet the required concentration. From Equation 9, we can see that we must have the two buffer component concentrations equal to the total buffer strength.

$$[buffer] = [acid] + [base] \quad \text{Equation 9}$$

Using our example:

$$0.0500M = [C_2H_5COOH] + 2.00[C_2H_5COOH] \quad \text{Equation 9a}$$

$$0.0500M = 3.00[C_2H_5COOH] \quad \text{Equation 9b}$$

$$0.0167M = [C_2H_5COOH] \quad \text{Equation 9c}$$

$$[C_2H_5COO^-] = 2.00[C_2H_5COOH] = 0.0333M \quad \text{Equation 9d}$$

The final step in preparing the buffer is to mix the correct volume amounts of the conjugate acid–base pair to give the desired concentration. Knowing the concentrations calculated from Equations 9c-d, the final volume of solution needed (0.250 L) and the initial concentration of the stock solutions (0.100 M) we can use the formula,  $M_1V_1 = M_2V_2$ , to calculate the exact volume of stock solutions needed.

To find the volume of  $C_2H_5COOH$  to use in our buffer we calculate:

$$V_1 = \frac{M_2V_2}{M_1} \quad \text{Equation 10}$$

$$V_1 = \frac{(0.0167M)(0.250L)}{0.100M} \quad \text{Equation 10a}$$

$$V_1 = 0.0418L \quad \text{Equation 10b}$$

Repeating these steps to calculate the volume of  $C_2H_5COO^-$  to use would give us 0.0833 L. Thus, we mix 41.8 mL of 0.100 M propanoic acid solution and 83.3 mL of 0.100 M sodium propanate solution in a 250.0- mL volumetric flask. Adding de-ionized  $H_2O$  while mixing will give a total solution with a volume of 250.0- mL, a pH of 5.19 and a total concentration of 0.0500 M.



## C. Overview

In this lab you will learn how acid strength affects reactions with metals and be able to measure pH using pH meter. Also you will learn how to prepare a buffer of a desired pH and concentration. You will observe how buffers stabilize the pH of a solution even after adding a strong acid or strong base.

Safety is part of science, please protect your safety and the safety of others.

### SAFETY WARNING:

Safety rules & chemical waste disposal guidelines must be followed in order to prevent personal injury, protect yourself, others & the environment. If you are unable to observe the rules then you are at risk of been dismissed from the lab.

### CAUTION:

Do not dump any of the reagents down the sink! Discard the waste in an appropriate waste container under the supervision of your instructor! Reduce waste by working diligently and do not repeat experimental trials without the approval of the instructor. Do not allow solutions to come in contact with your skin! Wear gloves & goggles.



In Part A you will learn the strengths of acid using pH digital meter and also magnesium strips.

In Part B, you will be acquainted with the procedures and calculations involved in designing and preparing a buffer. Each student will learn to design acetic acid/sodium acetate buffer with varying pH. Using the knowledge gained from the calculations, you will then prepare your own buffer, given a total buffer concentration, a target pH, and known concentrations of conjugate acid–base pairs. With a pH meter you will measure the initial pH of each buffer.



Measuring pH using a digital pH meter.

Then in Part C, you will learn and test how the pH properties of each buffer changes with a small addition of strong acid and strong base and exchange data with other groups.

In Part D, you will titrate the buffer you made in Part B with a strong acid in order to determine its buffer capacity.

#### IV. EXPERIMENTAL PROCEDURES “confirm with the datasheet on the last page”

- Make sure you take photos of your favorite lab work for use in your final PPT presentation assigned by your TA about one of the experiments E1-E5.
- Always freshly wash & rinse glassware with deionized water before use.
- If 50-mL beakers are not available then add enough solutions to each beaker to make sure the pH meter electrode tip is completely immersed in solution.
- Each pipet must be rinsed with deionized water, use rubber bulb pump to force excess water out, then wipe with clean tissue. Do not use any stock solution to rinse the pipet since this may cause waste of chemicals.

##### **PART A: pH Meter Calibration (each group must do)**

- Accuracy of pH measurements depends on various factors including pH meter type, temperature, experimental procedure, concentration of chemicals, etc.
  - Make sure to conduct the entire experiment at minimum variation of room temperature hence the equilibrium constant and  $[H^+]$  depends on T (see Chapter 19 of VC211 course:  $\Delta G^0 = -RT \ln K$  where K &  $H^+$  are related in aqueous solution).
  - Minimize contamination of the pH meter electrode and when experiment is completed ensure the proper rinsing and storage of the electrode.
  - If pH meter electrode is not contaminated or misused then no need to recalibrate throughout the entire experiment of E2.
1. Using a clean beaker (50-mL) pour enough volume of the calibration stock buffer of pH = 4.003 (or as provided) to be able later to completely immerse the head of the pH meter electrode. The stock buffer pH may be different than 4.003 and also the stock buffer may be not the same buffer as that inside the storage bottle (where the electrode is normally stored).
  2. Turn on the pH meter, rinse electrode thoroughly with deionized water and carefully wipe the tip with clean dry tissue.
  3. Insert the electrode tip thoroughly into the 50-mL beaker buffer. Push the calibration button and wait (about 15 seconds) until the pH reaches 4.00 (in case another stock buffer is used then wait until the reading is very close to the stock solution pH).
  4. Carefully remove the electrode, rinse with deionized water, and wipe with clean dry tissue.
  5. Immediately stow away the electrode in its storage buffer bottle. Now the meter is ready to measure pH of all the samples for the entire experiment.

**Warning:** If you intentionally or ignorantly do not rinse the electrode with deionized water and wipe the tip, then you may contaminate the electrode and cause erroneous errors in pH measurement for the remaining samples. Under no circumstance the pH meter is left to dry out of its proper storage buffer bottle. In addition you may be subject to heavy grading penalty and be liable for the intentional damage of the meter. So please read procedures very carefully and follow the lab instructions.



### PART A.1: pH of Strong and Weak Acids

- Modified version of the original procedure from Univ. of Michigan, E1, Part B).
- Each group but alternating students should do three pH measurements & test acid strength using 2 Mg strips/group.

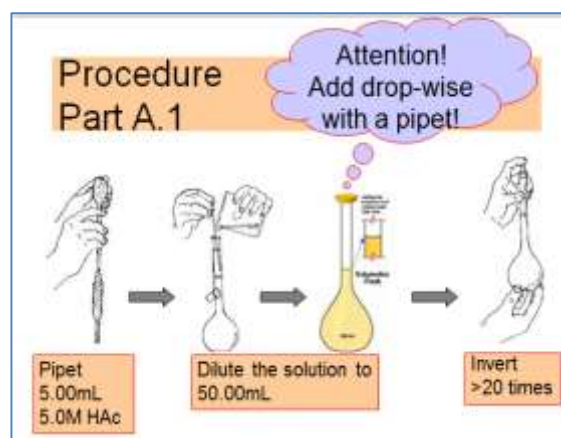
Chemicals used	Materials used
0.50 M HCl 5.0 M HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub> Unknown CH <sub>3</sub> COOH solutions Magnesium wire strips or shavings	pH Meter Well plate

1. Pipet 5.00 mL of 5.0 M CH<sub>3</sub>COOH solution into 50-mL volumetric flask. Dilute the solution by filling the volumetric flask with de-ionized water until the meniscus reach the mark. Then pour 20.0 mL of this solution into a 50-mL beaker.

2. Pour 20.0 mL of 0.50 M HCl solution into another 50-mL beaker.

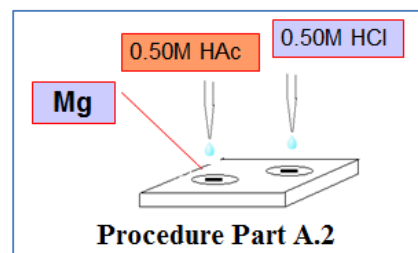
3. Pour 20-30 mL of the unknown solution CH<sub>3</sub>COOH into another third 50-mL beaker.

4. Using a calibrated pH meter, measure the pH of the solutions inside each of the three beakers. Record the measure pH.



### PART A.2: Reaction of Mg Metal with Addition of Strong and Weak Acids

5. Clean surface of two short Mg strips by lightly sand with a piece of abrasive paper and put each in the bottom of the 2 empty wells of a porcelain spot plate, add 5-10 drops 0.50 M HCl solution to one well & equal amount of CH<sub>3</sub>COOH solution to the other well just to cover the Mg strip. Record your observations.





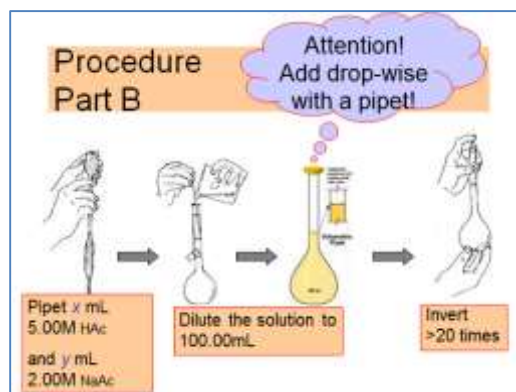
## Part B: Designing a Buffer Solution

- Each student of a group will design and prepare one pH of the buffer (four pH buffer solutions / group)

Chemical used	Materials used
5.0M CH <sub>3</sub> COOH 2.00M CH <sub>3</sub> COONa	Graduated pipet and bulb 100- mL Volumetric flask Plastic Pasteur pipet

### Design each buffer

- You will be assigned to prepare CH<sub>3</sub>COOH / CH<sub>3</sub>COONa buffer solutions. A good first step is to look up the value of the  $K_a$  for acetic acid and then calculate its  $pK_a$ . For example propanoic acid,  $K_a = 1.3 \times 10^{-5}$  at 25°C or  $pK_a = 4.89$ . Make buffers with 5.0 M CH<sub>3</sub>COOH and 2.0M CH<sub>3</sub>COONa solutions. Calculate the appropriate volumes x & y to be used and make the buffers next.
- Prepare 0.100L of the buffer in a 100-mL volumetric flask: Pipet (x) mL of 5.0M CH<sub>3</sub>COOH and (y) mL of 2.00M CH<sub>3</sub>COONa solution into a 100mL volumetric flask. Dilute the solution by filling the volumetric flask with de-ionized water until the meniscus reach the mark. **Be sure to save your buffers for Procedure Part D.**



**Note:** Only prepare the following 4 buffers, one per each student, 4 buffers per group. Each student in a group tests one type of buffer all the way to end of the experiment while the other students of the group test the remaining buffers.

- Follow the sample calculations next to design the buffer and record your data in the following table:

Buffer 1—0.100L of a pH=4.15 buffer with a total concentration of 0.100M

Buffer 2—0.100L of a pH=4.75 buffer with a total concentration of 0.100M

Buffer 3—0.100L of a pH=4.57 buffer with a total concentration of 0.100M

Buffer 4—0.100L of a pH=5.35 buffer with a total concentration of 0.100M

Buffer	Design pH	$c_b/c_a$	x mL (5.00M CH <sub>3</sub> COOH)	y mL (2.00M CH <sub>3</sub> COONa)
Example	4.00	0.178	1.7	0.75
1	4.15			
2	4.57			
3	4.75			
4	5.35			

### Sample calculations of x & y

Design the buffer  $\text{CH}_3\text{COOH} / \text{CH}_3\text{COONa}$  with desired  $\text{pH} = 4.0$ , where  $\text{CH}_3\text{COONa}$  is the conjugate base of the weak acid  $\text{CH}_3\text{COOH}$  with concentrations of  $C_b = [\text{CH}_3\text{COONa}]$  &  $C_a = [\text{CH}_3\text{COOH}]$ . Then in general, the **Henderson – Hasselbalch equation** for any conjugate acid–base pair can be written: ( $K_a = 1.85 \times 10^{-5}$ , or  $\text{p}K_a = 4.75$  at  $25^\circ\text{C}$ )

$$\text{pH} = \text{p}K_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\text{or } C_a/C_b = 10^{(\text{pH}-\text{p}K_a)} = 10^{(4.00 - 4.75)} = 0.178 = C_b/C_a$$

$$\text{or } C_b = 0.178C_a, \text{ but } C_a + C_b = 0.100 \text{ M}$$

Then  $C_a = 0.085\text{M}$  &  $C_b = 0.015\text{M}$ , but because of dilution in step 1:

$$5.00 \text{ M} \times V_a (\text{mL}) = (100 \text{ mL} \times 0.085 \text{ M}), \text{ then } V_a = 1.7 \text{ mL} = x$$

$$\text{Similarly: } 2.00 \text{ M} \times V_b (\text{mL}) = 100 \text{ mL} \times 0.015\text{M}, V_b = 0.75 \text{ mL} = y$$

### Preparing samples from each buffer solution

4. Share your results with your group in order to complete the previous table.
5. Each student, carefully pour one of three portions of 20.0 mL of the selected buffer, into a 50-mL graduated cylinder and then transferring each of the 20.0 mL portions into a 50-mL beaker. (Note: use three 50-mL beakers of same buffer per student and alternate buffers for the remaining students so each group will have prepared 12 samples, each in its own separate 50-mL beaker).
6. Using graduated cylinder the **group** measure two 20.0-mL portions of distilled  $\text{H}_2\text{O}$  and pour each into its own 50-mL beaker.
7. Label all beakers (3 beakers of the same buffer per student & two beakers of distilled water per group & and follow the procedure next).

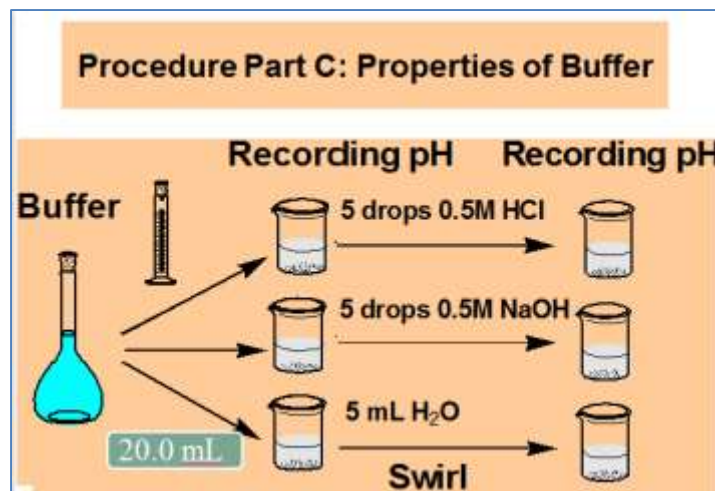
### Part C: Properties of a Buffer

- 3 beakers/buffer/student, 12beakers/group, plus 2 beakers of distilled water/group

Chemical used	Materials used
5.0M $\text{CH}_3\text{COOH}$	pH Meter, Funnel, Plastic Pasteur pipet
2.00M $\text{CH}_3\text{COONa}$	Graduated pipet and bulb
0.50 M HCl	50- mL Graduated cylinder
0.50 M NaOH	100- mL Volumetric flask, 100- mL Beaker (4)

1. Each **student** of a group uses a pH meter to measure and record the pH of one beaker of his own buffer prepared in step 5 of Part B.. Then add 5.0 mL of deionized  $\text{H}_2\text{O}$  to the same beaker and measure pH again. (two pH readings per student, 8 pH readings per group). Depending on time remaining for the experiment, instructor may ask you to skip measuring the pH before adding the 5.0 mL water in this step.

- Each **student**, of a group uses a clean disposable pipet to add 5 drops of 0.50 M HCl to one of the two beakers containing your prepared sample buffer of step 5 Part B, and 5 drops of 0.50 M NaOH to the other beaker. Swirl the solution, measure and record the resulting pH. (two pH readings per student, 8 pH readings per group)
- For the two beakers containing only distilled water in step 6 Part B, each **group** measures the pH of each beaker. Then with a clean, disposable pipet, add 5 drops of 0.50 M HCl to one of the beakers, and 5 drops of 0.50 M NaOH to the other beaker. Record data. Depending on time remaining for the experiment, instructor may ask you to skip measuring the pH before adding HCl or NaOH in this step.
- Tabulate all results.

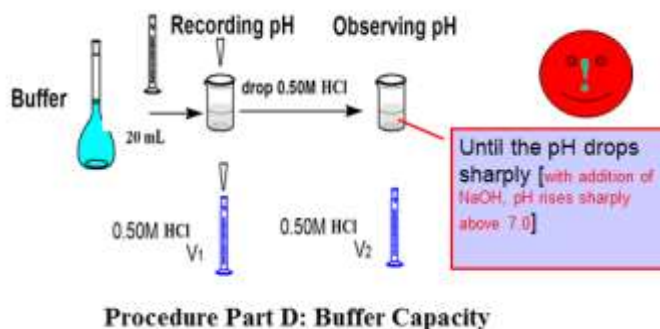


#### Part D: Determination of Buffer Capacity

- One buffer capacity per students, four buffer capacities per group

Chemicals used	Materials used
0.50M HCl or 0.50 M NaOH Buffer from Part B (50 mL)	pH Meter 50- mL Graduated cylinder 100- mL, 250- mL Beakers, Funnel 50- mL Buret and buret stand

- Fill a clean graduated pipette with 0.50M HCl solution (or use a 1-mL graduated pipet but not eye dropper). Record the initial volume  $V_1$ . Add the 0.50 M HCl solution drop-wise into the buffer with a pipet (students should use only the same buffer tested in Procedure Part C (step2) that contained 5 drops HCl), swirl the beaker after each addition. While monitoring the pH, record the volume of HCl solution to reach your buffer's capacity, record  $V_2$ . Do not exceed the buffer capacity since this may cause the need of excessive cleaning of the electrode later on.



- Make sure you add to the volume the 5 drops of HCl already inside the buffer sample from Procedure Part C (step 2). Then exchange data with other people who had a different buffer and prepare the data sheet information in a tabulated form.

## Common notes

- Buffer capacity is reached when pH starts to increase rapidly with continued addition of 0.50M HCl. This part is done by adding more drops of the HCl solution to the same beaker that you previously added 5 drops HCl. Students beware to account for the initial 5 drops volume (add volume of 5 drops from previous step Part C and the additional added volume to reach capacity).
- Each student should measure only 1 pH capacity end point, 4 end points per group. Students must use the 1mL graduated pipette to add all the drops in Part C & part D, then account for the volumes only for part D.
- Eye dropper is not the correct tool to add the drops for this experiment. An alternative method is to substitute the 0.50M HCl with the 0.50M NaOH but in this case you must use the corresponding beaker that you already added to it few drops of NaOH.
- Instructor/TA will give you direction if HCl or NaOH you should use to reach capacity.

## Desired data sheet information to tabulate

Discuss results & complete the post lab report for the experiment. All information in the report must be handwritten and include all the following data:

1. Buffer number.
2. Target pH, Calculated  $C_b / C_a$ .
3. Calculated  $x$  (to measure & add it to the 100mL volumetric flask).
4. Calculated  $y$  (to measure & add it to the same 100mL volumetric flask).
5. Actual pH after diluting the 100mL volumetric flask to mark (measure only once in one of the three beakers).
6. pH measured after adding 5drops 0.50M HCl to only one buffer beaker that contains 20mL buffer sample (note that total 5 beakers available : 3 buffer samples & 2 de-ionized water samples, 20 mL each sample beaker).
7. pH measured after adding 5 drops 0.50M NaOH to another beaker of the 3 buffer samples.
8. pH measured after adding 5mL de-ionized water to the third beaker of buffer sample.
9. Initial volume of 0.50M HCl before adding more drops to test buffer capacity (**Part C**).
10. Final volume after adding more drops of 0.50M HCl to capacity point.
11. Change in volume, end pH reading at capacity point

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Name:	Lab Instructor:
Date:	Lab Section:

### V. PRE-LABORATORY EXERCISES (PLE)

1. Define the term “common-ion effect” used in the BACKGROUND section. How do buffers work by taking advantage of the common-ion effect?
2. Give an example of a salt that could be used to make a buffer with  $\text{NH}_3$ .
3. If solution A has pH of 3.23 and solution B has a pH of 4.23, what is their relationship in terms of  $[\text{H}^+]$ ?
4. What is the relation between the concentration of buffer components and the buffer capacity? Explain.
5. For a buffer having a  $[\text{base}]/[\text{acid}]$  ratio of 1:1, what is the relationship between pH and  $pK_a$ ?
6. Using Equations 1 and 2 as models, write the acid dissociation reaction and  $K_a$  value for acetic and calculate its  $pK_a$  value (also record these values in your laboratory notebook so you will have them available during the experiment).

## E2: The Properties of Buffer: Resisting Change in a Turbulent World

Name:	Lab Instructor:
Date:	Lab Section:

### VI. RESULTS AND POST – LABORATORY QUESTIONS (PLQ)

#### Part B. Designing a Buffer

Attach a copy of your work showing the calculations of the volumes of acid and base needed to create the target buffer you were assigned.

Which set of data is yours? Buffer 1, Buffer 2, Buffer 3, or Buffer 4

	Buffer 1	Buffer 2	Buffer 3	Buffer 4
[base]/[acid] ratio				
Average pH of buffer				
Average pH of DI-H <sub>2</sub> O				
pH of buffer after adding HCl				
pH of buffer after adding NaOH				
pH of DI - H <sub>2</sub> O after adding HCl				
pH of DI - H <sub>2</sub> O after adding NaOH				

#### QUESTIONS FROM PART B

1. Verify that the volumes of the acid and base you were asked to use will result in the correct total concentration of the buffer.
2. Calculate the [base]/[acid] ratio used for your buffer.

3. Closely examine the data from Part B, looking at all four buffer solutions, do you see a relationship between the  $pK_a$  value of the acid, the desired or target pH of the buffer, and the  $[\text{base}]/[\text{acid}]$  ratio? Explain.
4. How does the relationship in question 3 help you decide which conjugate acid–base pair to choose when designing a buffer?
5. After performing Part C, did you notice a difference between the buffer and the DI–water when the strong acid/base was added? Did the buffer “do its job”? Explain.

#### Part D. Determination of Buffer Capacity

Which set of data is yours? Buffer 1, Buffer 2, Buffer 3, or Buffer 4.

	Buffer 1	Buffer 2	Buffer3	Buffer 4
Conjugate acid–base pair used				
Moles of acid in buffer				
Total volume of NaOH to reach buffer capacity				
Total moles of NaOH to reach buffer capacity				

How does the total moles of NaOH needed to reach buffer capacity relate to the number of moles of acid in the buffer? Explain.



# SAMPLE DATASHEET FOR A LAB SECTION (FOR REFERENCE ONLY, STUDENTS TO OMIT THIS PAGE)

VC211 DATASHEET FOR EXPERIMENT: E2 PROPERTIES OF BUFFERS																		PAGE 1		
SECTION:			TA:				LAB ROOM:													
BUFFER			#1	#2	#3	#4	Start with (x) mL 5.0 M HAc & (y) mL 2.0 M AcNa. TA will compare "pH Design"													
pH Design			4.15	4.37	4.75	5.35	to measured "pH" of each buffer made by a student													
PROCEDURE PART →			A	A1	A1	A1	A2	A2	C	C	C	C	C	C	C	C	C	D		
EFFORTS			← GROUP EFFORTS →						← INDIVIDUAL EFFORTS →											
GRP	NAME	ID	CALIB. BUFFER	HCl	HAc	?HAc	Mg	Mg	Design	H <sub>2</sub> O /	H <sub>2</sub> O /	H <sub>2</sub> O /	H <sub>2</sub> O /	← BUFFER pH: H <sub>2</sub> O, NaOH, & HCl →						
#	Chinese		STOCK	0.50M	0.50M	?M	0.50M	0.50M	Buffer	H <sub>2</sub> O	H <sub>2</sub> O	HCl	NaOH	INIT.	H <sub>2</sub> O	NaOH	HCl	TOT. mL		
			pH	pH1	pH2	pH3	HCl	HAc	pH	pH4	pH5	pH6	pH7	pH8	pH9	pH10	pH11	pH12	V(HCl)	
1									4.15											
1									4.57											
1									4.75											
1									5.35											
2									4.15											
2									4.57											
2									4.75											
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4									4.75											
4									5.35											
4									4.15											
4									4.57											

## SUMMARY OF E2 Properties of Buffers

- PART A:** Measure pH of 0.50M HCl & 0.50M HAc (diluted from 5.0M), then test their strength in dissolving Mg strip. (3 pH measurements (pH1-pH3) & 2 Mg strips tested /group)
- PART B:** Design 4 buffers with different pH using same stock solutions of HAc & AcNa diluted to 100 mL each, but using different volumes that will be calculated. (1buffer/student, 4buffers/group)
- PART C:** Test the initial pH of each buffer to compare with target pH. Then add 5 drops of 0.50M HCl to one 20mL buffer sample then add 5 drops of 0.50M NaOH to another 20mL sample of same buffer. Add 5ml de-ionized water to the 3<sup>rd</sup> beaker of 20mL sample of same buffer. Measure pH to conclude that pH remains about constant. Measure pH for the remaining 2 water beakers of 20mL distilled water before and after using 5drops of 0.5M HCl for one beaker & 5drops 0.50M NaOH for the other beaker. (pH4-pH9) (3beakers/buffer/student, 12beakers/group, plus 2beakers of distilled water/group)
- PART D:** Test the total capacity of a buffer using the previous sample from Part C when 5drops of 0.50M HCl were added (not NaOH) by also adding more drops until pH starts dropping rapidly. Make sure you account for the initial drops of 0.50M HCl added at start of Part C, to find volume added to capacity. (pH10-pH13) (1 buffer sample/student, 4buffer samples/group)