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

He Wenxin何雯欣

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Student ID 518370910117

Section 3

Group 3

Members Qiu Jiahao 裘嘉豪

Feng Hanyang 冯瀚扬

Xu Jiaying 许佳滢

Telephone 15921193245

Email Address wenxin\_he@sjtu.edu.cn

University of Michigan-Shanghai Jiao Tong University Joint Institute(UM-SJTU JI)

# Ⅰ OBJECTIVES

* Study the strength of acids
* Learn the usage of pH meter
* Learn how buffers react
* Make a buffer solution of specific concentration and pH
* Identify a solution’s buffer capacity

# Ⅱ INTRODUCTION

The conditions must be controlled in some reactions to avoid sudden changes. A buffer can function to decrease an external impact. The human body uses the H2CO3/HCO3 system as a buffer to control the blood pH. Many bodily processes require a proper pH of which the range is very narrow. If the heart stops, lactic acid and CO2 cannot be removed from the blood, causing the blood pH to lower. Under this circumstance, hospital emergency staff can use the knowledge of buffers to help restore the blood pH by administering the fluids.

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**Figure 1** Hospitals treat patients with intravenous solution to help maintain the pH of blood.

# Ⅲ BACKGROUND

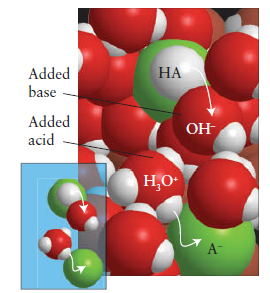
**A. The Properties of Buffers**

# Commonly, a buffer solution is composed of a conjugated acid-base pair of a weak acid which can neutralize OH- or H+ ions to resist changes when an acid or base component is added. A mixture of aqueous CH3COOH and CH3COO-, as an example, is a buffer which works under common-ion effect so that the components can stay almost the same when small amounts of OH- or H+ ions are added.

# Only when the added OH- or H+ are much less than the acid-base components can the buffer work. Take a buffer of C2H5COOH and C2H5COO- as example: (aq) ↔ (aq) + (aq)

# We can get .

# When a strong acid is added, increases and will react with H+ to form. Therefore, the concentration ratio of acid to base only increases a little bit, resulting in a small pH drop.



**Figure 2** A solution acts as a buffer if it contains a weak acid, HA, which donates protons when a base is added, and the conjugate base,, which accepts protons when an acid is added.

**B. Preparing a Buffer**

# We use the Henderson-Hasselbalch equation to calculate the pH of a buffer:

# pH=p+log. When preparing a Buffer, it is important to know the ratio of acid to base and the target pH. The two factors are dependent of each other when pis certain. Then by using [*buffer*]  [*acid*][*base*] we can get the concentration of the acid and base in the buffer.

# To calculate the volume of needed acid or base, we can use the equation where is the volume of the final solution, is the concentration of the final volume of the solution, and is the initial concentration of the stock solutions.

**C. Overview**

# We will learn how the effect of acid strength on its reaction with metals, how to use the pH meter and how to prepare a buffer with desired pH and concentration. We will also learn how the buffer helps to manage a change in or.

# In Part A, we will use pH meter and magnesium strip to determine the strength of acids.

In Part B, we will make calculations and prepare an acetic acid/sodium acetate buffer with a given total buffer concentration, a target pH and known concentration of acid-base pairs.

In Part C, we will discover the change in pH of the buffer accompanied with small amounts of strong acid or base added.

In Part D, we will determine the buffer capacity of the buffer we made in part B.

# Ⅳ EXPERIMENTAL PROCEDURES

**Part A pH Meter Calibration**

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.

**Part A.1 pH of Strong and Weak Acids**

|  |  |
| --- | --- |
| Chemicals used | Materials used |
| 0.50 M HCl  5.0 M HC2H3O2  Unknown CH3COOH solutions  Magnesium wire strips or shavings | pH Meter  Well plate |

1. Pipet 5.00 mL of 5.0 M CH3COOH 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-mLbeaker.

3. Pour 20-30 mL of the unknown solution CH3COOH 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 CH3COOH solution to the other well just to cover the Mg strip. Record your observations.

**Part B Designing a Buffer solution**

|  |  |
| --- | --- |
| Chemicals used | Materials used |
| 5.0M CH3COOH  2.00M CH3COONa | Graduated pipet and bulb  100- mL Volumetric flask  Plastic Pasteur pipet |

* Design each buffer:

1. You will be assigned to prepare CH3COOH / CH3COONa buffer solutions. A good first step is to look up the value of the Ka for acetic acid and then calculate its pKa. For example propanoic acid, Ka = 5 1.3 10at 25oC or pKa = 4.89. Make buffers with 5.0 M CH3COOH and 2.0M CH3COONa solutions. Calculate the appropriate volumes x & y to be used and make the buffers next.

2. Prepare 0.100L of the buffer in a 100-mL volumetric flask: Pipet (x) mL of 5.0M CH3COOH and (y) mL of 2.00M CH3COONa solution into a 100mL volumetric flask. Dilute the solution by filling the volumetric flask with de-ionized water until the meniscus reach the mark.

3. 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 | cb/ca | x mL (5.00M CH3COOH) | y mL (2.00M CH3COONa) |
| Example | 4.00 | 0.178 | 1.7 | 0.75 |
| 1 | 4.15 |  |  |  |
| 2 | 4.57 |  |  |  |
| 3 | 4.75 |  |  |  |
| 4 | 5.35 |  |  |  |

* Preparing samples from each buffer solutions:

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 groupmeasure two 20.0-mL portions of distilled H2O 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**

|  |  |
| --- | --- |
| Chemicals used | Materials used |
| 5.0M CH3COOH  2.00M CH3COONa  0.50 M HCl  0.50 M NaOH | pH Meter, Funnel, Plastic Pasteur pipet  Graduated pipet and bulb  50- mL Graduated cylinder  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 H2O 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.

2. 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)

*3.* 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.

4. Tabulate all results.

**Part D Determination of Buffer Capacity**

|  |  |
| --- | --- |
| 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 |

1. 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 V1. 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 V2. Do not exceed the buffer capacity since this may cause the need of excessive cleaning of the electrode later on.

2. 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.

# Ⅴ CALCULATION/ANALYSIS/DATA PROCESSING

**Part A pH Meter Calibration**



As is shown in figure 3, when the reading is 4.00 and the “A” on the upper-left corner stops shining, the meter is ready to measure.

**Figure 3** calibration state

**Part A.1 pH of Strong and Weak Acids**

The pH of 0.5M HCl, 0.5M HAc, and unknown concentration of HAc are shown in Table 1

|  |  |  |  |
| --- | --- | --- | --- |
| Chemical | 0.5M HCl | 0.5M HAc | ?M HAc |
| Temperature(℃) | 22.0 | 22.0 | 22.0 |
| pH | 0.67 | 2.42 | 2.52 |

**Table 1** pH measurement of Part A.1

Theoretically, the pH of 0.5M HCl should be:

pH=-log[H+]= -log(0.5)= 0.3

To calculate the theoretical pH of 0.5M HAc under 22℃, we have to find out the [H+], so we set up an ICE table:

|  |  |  |  |
| --- | --- | --- | --- |
|  | HAc | H+ | Ac- |
| Initial | 0.5 | 0 | 0 |
| Change | -x | x | x |
| Equilibrium | 0.5-x | x | x |

**Table 2** The ICE table for HAC

We use the equation lnKa=lnKa1 to calculate Ka1 under 22℃:

We come up with the equation:Ka1=, and we get x=2.83

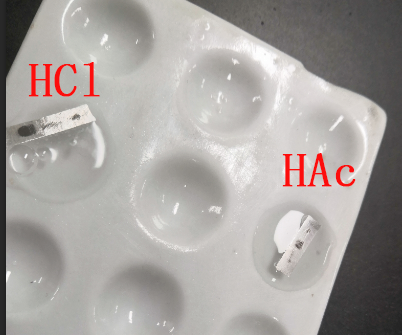
Therefore, pH= -log[H+]= -log(2.83)=2.55

The error of the pH of 0.5M HCl is:

The error of the pH of 0.5M HAC is

**Part A.1 pH of Strong and Weak Acids**

From Figure 4 we can see that the Mg strip acts quickly with HCl, as there exist bubbles soon after HCl is added. However, the reaction between Mg and HAc is very slow.



**Figure 4** Mg reacts with HCl and HAc with equal concentration and amount.

**Part B Designing a Buffer Solution**

Using the Henderson-Hasselbalch equation, we can get .

Suppose Va=x(mL) and Vb=y(mL)

We can get the system: 0.005x+0.002y=0.1M=0.01(mol)

We solve the equation set for each buffer and get 4 sets of values listed in Table3:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Buffer | Design pH | cb/ca | x mL (5.00M CH3COOH) | y mL (2.00M CH3COONa) |
| 1 | 4.15 | 0.251 | 1.6 | 1.00 |
| 2 | 4.57 | 0.661 | 1.2 | 1.99 |
| 3 | 4.75 | 1 | 1.0 | 2.50 |
| 4 | 5.35 | 3.981 | 0.4 | 4.00 |

**Table 3** Results of Calculation

**Part C Properties of a Buffer**

We took 20mL of buffer solution each time. Another 20mL deionized water was experimented on through the same procedure as a control experiment. The deionized water used in the experiment has an average pH of 5.04 I individually tested Buffer 4. The rest of the data is from my other group members.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | Buffer 1 | Buffer 2 | Buffer 3 | Buffer 4 | De-ionized water |
| Initial pH | 4.07 | 4.55 | 4.71 | 5.22 | 5.04 |
| pH after adding 5mL water | 4.01 | 4.54 | 4.69 | 5.22 | 5.04 |
| pH after adding 5 drops of NaOH | 4.17 | 4.66 | 4.78 | 5.40 | 11.22 |
| pH after adding 5 drops of HCl | 3.81 | 4.37 | 4.54 | 5.09 | 2.67 |

**Table 4** Experiment data for Part C

We then calculate the change of pH and get the following table.

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | Buffer 1 | Buffer 2 | Buffer 3 | Buffer 4 | De-ionized water |
| Error of Initial pH | -1.9% | -0.004% | -0.008% | -2.4% | / |
| Change of pH after adding 5mL water | -0.06 | -0.01 | -0.02 | ±0 | ±0 |
| Change of pH after adding 5 drops of NaOH | +0.10 | +0.11 | +0.07 | +0.18 | +6.15 |
| Change of pH after adding 5 drops of HCl | -0.26 | -0.18 | -0.17 | -0.13 | -2.40 |

**Table 5** Data analyzation for Part C

From the data we can know that compared to pure de-ionized water, the pH change is very small in buffer solution when strong acid or base is added.

Theoretically, when 5mL de-ionized water is added, the pH of each buffer should become higher. However, the experimental data shows the opposite.

Also, from buffer 1 to 4, the concentration of acid grows smaller and the concentration of conjugate base grows bigger. Therefore, from buffer 1 to 4, the pH change should become bigger when a strong acid is added and become smaller when a strong base is added. From the analyzation of experiment data, buffer 1, 2, 4 show the right tendency yet buffer 3 behaves differently when NaOH is added.(the pH change should be between +0.11 and +0.18)

**Part D Determination of the Capacity of Buffer**

We first calculate the volume of 5 drops of HCl: the volume of 18 drops of HCl is 0.70mL; therefore, the volume of 5 drops of HCl is

|  |  |  |  |
| --- | --- | --- | --- |
| Buffer | Change in Vol. of HCl(mL) | Total consumption of HCl(mL) | pH at end point |
| 1 | 0.98 | 1.17 | 1.92 |
| 2 | 1.37 | 1.56 | 1.88 |
| 3 | 1.64 | 1.83 | 1.92 |
| 4 | 2.61 | 2.80 | 1.99 |

**Table 6** Experiment data for Part D

Then we analyze the data:

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Buffer | Initial pH | Moles of base in buffer(mol) | Consumption of HCl(mol) |  |
| 1 | 4.07 | 2 | 5.85 | -2.15 |
| 2 | 4.55 | 4 | 7.8 | -2.67 |
| 3 | 4.71 | 5 | 9.15 | -2.79 |
| 4 | 5.22 | 8 | 1.4 | -3.23 |

**Table7** Data analyzation for Part D

In line with theoretical property, buffer 4 shows the biggest capacity when HCl is added.

**Ⅵ DISCUSSION**

In Part A, the pH of 0.5M HCl measured is much higher than theoretical value, while the pH of 0.5M HAc measured is lower than theoretical value. The possible reasons are as follows:

1. There is stock buffer of pH4.0 left on the electrode when measuring the pH of HCl.

2. There is 0.5M HCl left on the electrode when measuring the pH of HAc.

In Part C, all the pH of designed buffer are lower than theoretical values. The possible causes are as follows:

1. The actual environment temperature is 22°C, which is lower than the standard room temperature (25°C). This causes the pKa of HAc to be smaller than 4.75, resulting in a smaller pH.

2. When pipetting NaAc, there is some solution staying in the pipet.

3. There is previous (more acidic) solution left on the electrode when measuring the pH of the next buffer.

When it comes to buffer 3 behaving differently when NaOH is added compared to other buffers, we believe there must be some mistakes in operating that causes the amount of NaOH added is smaller than required.

In Part D, because buffer 4 has the most moles of base, it needs more HCl to reach its end point. The data is in line with the theoretical values. However, we achieved the volume of added HCl by counting the number of drops and then multiplying it by the volume per drop. The right way is to directly read the initial and final readings of the pipet followed by taking the volume of the initial 5 drops into account.

**Ⅶ CONCLUSIONS ＆ RECOMMENDATIONS**

In Part A, we’ve learned that strong acid has a lower pH than weak acid with the same concentration. Strong acid also acts with Mg much faster than weak acid.

In Part B, we’ve learned how to prepare a buffer solution with desired pH.

In Part C, we experimented with buffers of different pH and tested their properties when strong acid or base is added.

In Part D, we’ve collected data and compared the capacity of buffers with different pH.

We strongly recommend you to divide labor properly during the experiment so that your work will be more efficient. Also, we recommend you to read the readings carefully and patiently to increase the accuracy of the data.

**Ⅷ REFERENCES**

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