Dougherty Valley HS Chemistry - AP
Acid Base – Buffers

Worksheet #14

Name: **Period:** **Seat#:**

Required Sections: (Refer to R-15 for guidelines and requirements. Make note of any specific changes given by your teacher in class.)

Prelab: Prelab Questions, Purpose, Materials, Reagent Table, Procedures, and set up Data Tables before you get to class.
During Lab: Data section – Fill out your data table that is already set up from the prelab.
Post-lab: Post-Lab Questions, Post-Lab Two Pager done on separate Worksheet.

Background

A buffer is a mixture of a weak acid and its conjugate base, or a weak base and its conjugate acid. A buffer's function is to absorb acids (H^+ or H_3O^+ ions) or bases (OH^- ions) so that the pH of the system changes very, very little. In many systems, buffers are critical. Blood plasma, a natural example in humans, is a bicarbonate buffer that keeps the pH of blood between 7.2 and 7.6. By design, a buffer is an equilibrium system. For example, a buffer can be prepared with nitrous acid, HNO_2. The weak acid establishes an aqueous equilibrium as shown below.

$$\mathbf{HNO_2\ (aq) \leftrightarrow H^+\ (aq) + NO_2^-\ (aq)}$$

The equilibrium constant expression is shown below.

$$K_a = \frac{[H^+]\,[NO_2^-]}{[HNO_2]}$$

To prepare a buffer system with nitrous acid, a conjugate base is added, such as sodium nitrite ($NaNO_2$). The resulting system is a mixture of HNO_2 and NO_2^- ions. The nitrous acid molecule will neutralize hydroxide ions and the nitrite ion from the conjugate will neutralize hydrogen ions.

A variation of the equilibrium expression above, called the Henderson-Hasselbalch equation, is the best reference in preparing a buffer solution. For our nitrous acid/sodium nitrate buffer example, the Henderson-Hasselbalch equation is shown below.

$$pH = pK_a + \log \frac{[NO_2^-]}{[HNO_2]}$$

The pH range in which a buffer solution is effective is generally considered to be ±1 of the pK_a.

In this experiment, you will use the Henderson-Hasselbalch equation to determine the amount of acetic acid and sodium acetate needed to prepare two acidic buffer solutions. You will then prepare the buffers and test their buffer capacities by adding solutions of NaOH and HCl.

Figure 1

Objectives

In this experiment, you will

- Prepare and test two acid buffer solutions.
- Determine the buffer capacity of the prepared buffers.

Pre-Lab Questions

Use the Henderson-Hasselbalch equation to perform the following calculations. The K_a of acetic acid is 1.8×10^{-5}.

1) Buffer A: Calculate the mass of solid sodium acetate required to mix with 100.0 mL of 0.1 M acetic acid to prepare a pH 4 buffer. Record the mass in your data table.

2) Buffer B: Calculate the mass of solid sodium acetate required to mix with 100.0 mL of 1.0 M acetic acid to prepare a pH 4 buffer. Record the mass in your data table.

Materials

Chemicals

- 0.5 M sodium hydroxide, NaOH
- 0.5 M hydrochloric acid, HCl
- 0.1 M acetic acid, $HC_2H_3O_2$
- 1.0 M acetic acid, $HC_2H_3O_2$
- Solid sodium acetate, $NaC_2H_3O_2$
- Distilled H_2O

Equipment

- Computer with USB port, or a USB adaptor
- Logger Pro
- Vernier pH sensor
- Stir station and magnetic stirring bar.
- Electrode support
- 250mL beakers x3
- 100mL grad. cylinder
- 25mL grad. cylinder
- 50mL buret x2
- 50mL buret clamp x2
- Scale

Procedure

Part I Prepare and Test Buffer Solution A

1) Obtain and wear goggles.

2) Use your calculations from the Pre-Lab Exercise to prepare 100 mL of Buffer A. Weigh out the precise mass of sodium acetate and dissolve it in 100.0 mL of 0.1 M acetic acid solution.

3) Set up two burets, buret clamps, and Stir Station (see Figure 1). Rinse and fill one buret with 0.5 M NaOH solution. Rinse and fill the second buret with 0.5 M HCl solution.

4) Use a graduated cylinder to measure out **40.0 mL** of the Buffer A solution into a 250 mL beaker and add **60 mL** of distilled water. Place the beaker on a Stir Station, beneath the buret of NaOH, be sure to add stirrer onto pH probe.

5) Connect a pH Sensor to Channel 1 of the Vernier computer interface. Connect the interface to the computer using the proper interface cable. Suspend the pH Sensor in the pH 4 buffer solution, as shown in Figure 1. Make sure that the sensor is not struck by the stirring bar.

6) Start the Logger *Pro* program on your computer. Open the file "19 Buffers" from the *Advanced Chemistry with Vernier* folder.

7) You are now ready to test Buffer A. You will slowly and carefully add 0.5 M NaOH solution to the buffer solution.

a. Take an initial pH reading of the buffer solution. Click ▶ Collect and monitor pH for 5–10 seconds. Once the displayed pH reading has stabilized, click Keep. In the edit box, type **0** (for 0 mL added). Click OK to continue. Record the initial pH value in your data table.

b. Add a small amount of the NaOH solution, up to 0.50 mL. When the pH stabilizes click Keep. Enter the current buret reading. Click OK to continue.

c. Continue adding the NaOH solution in small increments that raise the pH consistently and enter the buret reading after each increment. Your goal is to raise the pH of the buffer by 2 pH units.

d. When the pH of the buffer solution is 2 units greater than the initial reading, continue to add the NaOH solution in small increments until you have reached, and passed, the equivalence point of the titration.

e. Click ■ Stop.

8) Dispose of the reaction mixture as directed. Rinse the pH sensor with distilled water in preparation for the second titration.

9) Repeat Steps 7 and 8, using a fresh 10.0 mL sample of the Buffer A solution. For this second trial, titrate the buffer with 0.5 M HCl solution. Carefully add HCl in small increments until the pH of the solution has been lowered by 2 units. Record, in your data table, the volume of HCl that was used.

Part II Prepare and Test Buffer Solution B

10) Use your calculations from the Pre-Lab Exercise to prepare 100 mL of Buffer B. Weigh out the mass of sodium acetate and dissolve it in 100.0 mL of 1.0 M acetic acid solution. If necessary, refill the burets of NaOH and HCl solution. **Caution**: *Treat all laboratory chemicals with caution. Prudent laboratory practices should be observed*.

11) Use a graduated cylinder to measure out 10.0 mL of the Buffer B solution and add 15 mL of distilled water. Repeat the necessary steps to test Buffer B in a manner similar to the Part I trials. Print a copy of your graph of the titration using the NaOH solution. Record the volume of HCl that was used to lower the pH of Buffer B by 2 units.

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Data Table

1. Make your own data table! Remember, you need to make sure your data table has all required elements!

	Buffer A	Buffer B
Mass of $NaC_2H_3O_2$ used to prepare buffer (g)		
Volume of buffer prepared (mL)	100.0	100.0
Molar concentration of $HC_2H_3O_2$ in buffer (M)	0.1	1.0
Initial pH of buffer		
Volume of 0.5 M NaOH to raise pH by 2 units (mL)		
Volume of 0.5 M HCl to lower pH by 2 units (mL)		
Volume of 0.5 M NaOH at equivalence point (mL)		

2. Glue in a copy of your Logger Pro graph below your data table.

Post Lab Discussion Questions

1. Write reaction equations to explain how your acetic acid-acetate buffer reacts with an acid and how it reacts with a base.
2. Buffer capacity has a rather loose definition, yet it is an important property of buffers. A commonly seen definition of buffer capacity is: "The amount of H^+ or OH^- that can be neutralized before the pH changes to a significant degree." Use your data to determine the buffer capacity of Buffer A and Buffer B.
3. Say, for example, that you had prepared a Buffer C, in which you mixed 8.203 g of sodium acetate, $NaC_2H_3O_2$, with 100.0 mL of 1.0 M acetic acid.
    a. What would be the initial pH of Buffer C?
    b. If you add 5.0 mL of 0.5 M NaOH solution to 20.0 mL each of Buffer B and Buffer C, which buffer's pH would change less? Explain.