

Experiment 11: Preparing and Testing a Buffer

CH2250: Techniques in Laboratory Chemistry, Plymouth State University

Created by Jeremiah Duncan and Wavell Fogleman, Department of Atmospheric Science and Chemistry, Plymouth State University.

Suggested reading for background information: Section Ch 8.5-7, 9.1-5, *Exploring Chemical Analysis, 4th Edition*, Daniel C. Harris (2008).

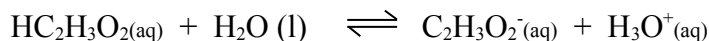
Introduction:

The preparation of buffer solutions is a common task in the lab, especially in biological sciences. A buffer is a solution that resists a change in pH, because it contains species in solution able to react with any added acid or base, according to the principles of equilibrium. You will study more about equilibrium and buffers in General Chemistry II; here we will limit ourselves to the practical matter of preparing a base at a desired pH. The Henderson-Hasselbalch equation, which can be easily derived from equilibrium equations, is used to find the pH of a buffer solution:

$$\text{pH} = \text{p}K_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$$

where $\text{p}K_a$ is an experimentally found constant for the acid HA, $[\text{HA}]$ is the concentration of the acid, and $[\text{A}^-]$ is the concentration of the conjugate base. Note that this is the acid version of the equation; a related equation exists for bases.

In this experiment, you will make a buffer using acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$) ($\text{p}K_a = 4.756$) and sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$), the sodium salt of the conjugate base. Note the relationship between the acid and its conjugate base in the equilibrium:



First, you will make 100 mL of a buffer with $\text{pH} = 5$ using 5 mL of a 0.3M acetic acid solution. Using the Henderson-Hasselbalch equation, the mass of sodium acetate needed to make the buffer is calculated as follows:

1. Calculate the concentration of acetic acid in the final 100 mL:

$$(5.00 \text{ mL} \times 0.300 \text{ M}) / (100.0 \text{ mL}) = 0.0150 \text{ M}$$

2. Plug the values into HH and solve for unknown:

$$5.00 = 4.756 + \log \frac{[\text{A}^-]}{0.0150 \text{ M}}$$

$$10^{0.244} = \frac{[\text{A}^-]}{0.0150 \text{ M}}$$

$$[\text{A}^-] = 0.0263 \text{ M}$$

3. Calculate the moles of A^- needed to obtain this concentration in 100 mL:

$$0.0263 \text{ M} \times 0.100 \text{ L} = 0.00263 \text{ moles}$$

4. Calculate the mass of NaA needed to obtain this number of moles:

$$0.00263 \text{ moles} \times 82.04 \text{ g/mol} = 0.216 \text{ g.}$$

Second, you will make 100 mL of a buffer also with $\text{pH} = 5$, but with a higher buffering capacity, using 5 mL of a 0.5 M acetic acid solution.



Although a buffer will resist a change in pH, eventually enough acid or base can be added to destroy it. The amount of acid or base needed to change the pH of a buffer is known as the "buffering capacity." You will measure the buffering capacity of the buffers you make, monitoring the titration with a pH meter.

Equipment: Read through the procedures and make a list of the equipment you will need.

Safety Considerations: Read through the procedures and note any safety considerations.

Pre-lab work: In addition to the normal pre-lab write up (Title, Date, Purpose, etc.), you must perform the needed calculations in the Procedure (Step B1) before coming to lab. Read through the Procedure and set up appropriate tables to record the data.

Procedure:

A. pH = 5.0 Buffer A

1. Add about 50 mL of distilled water to a 100 mL beaker.
2. Use a volumetric pipette to add 5.0 mL of 0.30 M acetic acid to the beaker.
3. Accurately weigh a small beaker, add about 0.4 g sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$), and weigh the beaker again. Record these masses in your notebook to all the sig figs possible.
4. Insert a previously calibrated pH meter into the beaker.
5. Add a little bit of the sodium acetate at a time, swirling the beaker to dissolve it (*do NOT stir with the pH probe!*) until the pH is 5.0. *In theory, this will take 0.216 g of sodium acetate.*
6. Weigh the beaker with the remaining sodium acetate.
7. Quantitatively transfer the buffer solution to a 100 mL volumetric flask.
8. Add distilled water up to the mark. Cap and invert the flask twice to mix.

B. pH = 5.0 Buffer B

1. Calculate the mass of sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$) that must be added to make 100 mL of an acetic acid/acetate buffer at pH = 5.0, given that you will use 5.0 mL of 0.50 M acetic acid. *You must perform this calculation as part of your pre-lab work. See the calculation in the Introduction for guidance.*
2. Make the buffer solution in a manner similar to that done in part A, starting with the mass of sodium acetate you calculated in B1 and using 0.50 M acetic acid. As you did in part A, weigh out a bit more sodium acetate than you will need.

C. Testing the Buffer Capacity

1. Calibrate a pH meter.
2. Using a volumetric pipette, transfer 25.0 mL of Buffer A into a 100 mL beaker.
3. Load a 50 mL buret with your Standardized NaOH solution (Experiment 7)
4. Use the pH meter to monitor the titration of the buffer until the pH changes 1 unit. Run two titrations: a quick one and a careful one. For the quick one, add NaOH in 1 mL increments. For the careful one, use your best judgment in adding titrant.
5. Repeat Steps 2-4 for Buffer B.



Analysis

1. Using the masses of sodium acetate and the beakers you weighed it in, determine the exact mass of sodium acetate you added to each buffer. Was the actual amount of sodium acetate needed different than the theoretical amounts? Why might this be?
2. Calculate the concentration of hydronium ion of the buffer solutions at their starting pH (5.0). Calculate the moles of hydronium ion present in 25.0 mL of the solutions at this pH.
3. A change of pH by one unit (i.e. from pH = 5.0 to 6.0) means a change in hydronium ion concentration by a factor of 10. Calculate the concentration of hydronium ion of the buffer solutions at their final pHs (6.0). Calculate the moles of hydronium ion present in 25.0 mL of the solutions at this pH.
4. Calculate the number of moles of NaOH that would theoretically be needed to change a solution from pH = 5.0 to pH = 6.0 (i.e., the difference in the number of moles in Questions 2 and 3).
5. Using your results from the careful titrations (Part C) and the concentration you previously calculated for your NaOH solution (Experiment 7), calculate the actual moles of titrant (NaOH) needed to change the pH by 1 unit for each buffer.

Conclusions

1. Compare your experimental results (Analysis Question 5) with the theoretical calculation (Analysis Question 4). Why are these numbers different?
2. Which buffer had a higher buffering capacity? Why?
3. If you wanted to make a buffer with an even higher buffering capacity, what change(s) would you make to the recipe?

Homework Problems

The following problems from your book must be completed in your lab notebook (see the Syllabus for other suggested problems): Ch 8: **11**, **21**; Ch 9: **8**, **14**, **17**

