

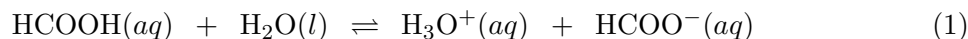
Lab Investigation 6 - How Well Can a Buffer Resist pH Change?

GUIDING QUESTION

How well can a buffer resist pH change?

INTRODUCTION

In dilute aqueous solutions, weak acids are slightly dissociated. They produce a small concentration of hydronium ion (H_3O^+) and an equal concentration of the conjugate base of the acid. Such dissociation reactions are equilibria, and equilibrium mathematics can be used to calculate concentrations of the species present in solution. Consider formic acid (CH_2O_2); its dissociation constant (K_a) is 1.7×10^{-4} . Incidentally, formic acid is what red ants inject when they bite. The concentration of H_3O^+ present in a 0.010-*M* solution of formic acid can be calculated from the equilibrium expression and a reaction table.



initial	0.010	0	0
Δ	$-x$	$+x$	$+x$
equilibrium	$0.010 - x$	x	x

$$K_a = 1.7 \times 10^{-4} = \frac{[\text{H}_3\text{O}^+][\text{HCOO}^-]}{[\text{HCOOH}]} = \frac{x^2}{0.010 - x} \quad (2)$$

Solving for x , one finds that the solution is 0.0012 *M* in H_3O^+ and HCOO^- .¹ Expressing $[\text{H}_3\text{O}^+]$ as a pH,

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log(0.0012) = 2.92. \quad (3)$$

Although a weak acid dissociates only slightly in water, the soluble salt of a weak acid (for example, sodium formate) is a strong electrolyte and dissociates completely.



The salt, if added to the weak acid solution, produces a large amount of formate ion in comparison to that produced by the acid dissociation. Adding formate ion to the equilibrium of equation

¹For this calculation, the quadratic formula was used. If one makes the simplifying assumption that x is small relative to $[\text{HCOOH}]$, the calculated value is 0.0013 *M*.

1 stresses the system by adding a product. According to Le Châtelier's principle, the equilibrium will shift to the left (toward reactants) and the concentration of H_3O^+ will decrease (and the pH will increase). The suppressed dissociation caused by adding an ion already present in the solution is called the common ion effect.

One can prepare a solution using a weak acid and its conjugate base (the common ion). The resulting solution will resist major changes in pH when an acid or base is added to the mixture. Such solutions are called buffer solutions.

Consider what would happen in a solution containing both formic acid and sodium formate when acid or base are added. Adding acid (a source of H_3O^+) stresses the system by adding a product. The equilibrium of equation 1, shifts toward the reactants, consuming formate ion and some of the added H_3O^+ . The result: a small decrease in the pH.

Adding a base causes the hydronium ion to neutralize the base. This stresses the system by removing a product. Some formic acid dissociates to replace the H_3O^+ and the equilibrium of equation 1 shifts toward the products. The result: a small increase in the pH.

The concentrations of the acid and its conjugate base in a buffer will determine how much additional acid or base can be added to the solution before its buffering ability is exhausted. This is called the **buffer capacity** of the solution. The higher the concentrations of acid and conjugate base, the larger the buffer capacity.

The preceding discussion will also apply if a buffer is prepared using a weak base and its conjugate acid. However, buffers cannot be made with strong acids or strong bases and their conjugates. No buffer capacity exists in such solutions because there is no equilibrium; everything has completely dissociated into ions.

The pH of a buffer solution may be calculated with the Henderson-Hasselbalch equation:

$$\text{pH} = \text{p}K_{\text{a}} + \log \left(\frac{[\text{base}]}{[\text{acid}]} \right) \text{ or } \text{pH} = \text{p}K_{\text{a}} + \log \left(\frac{\text{moles of base}}{\text{moles of acid}} \right). \quad (5)$$

The derivation of this equation follows from the general dissociation equilibrium expression for a weak acid, and includes the assumption that $[\text{H}_3\text{O}^+]$ is small relative to $[\text{HA}]$.



$$K_{\text{a}} = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]} \quad (7)$$

Solving for $[\text{H}_3\text{O}^+]$ gives:

$$[\text{H}_3\text{O}^+] = \frac{K_{\text{a}} \times [\text{HA}]}{[\text{A}^-]}. \quad (8)$$

Taking the negative logarithm of both sides puts the equation in terms of pH:

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log K_a - \log([\text{HA}]/[\text{A}^-]). \quad (9)$$

By definition, $\text{p}K_a = -\log K_a$ and $-\log([\text{HA}]/[\text{A}^-]) = \log([\text{A}^-]/[\text{HA}])$. Substituting these terms into equation 9:

$$\text{pH} = \text{p}K_a + \log([\text{A}^-]/[\text{HA}]). \quad (10)$$

Equation 5 shows that pH can be found using either concentrations of acid and base, or the number of moles of each. This follows from the fact that the volume term is the same for the acid and its conjugate base, and cancels in the calculation.

In the pre-lab exercise, you will prepare various acetate buffer solutions by the direct method and measure the pH of each solution. In the **direct method**, the conjugate acid and base are added together in solution to get the desired base to acid ratio. For example, acetic acid and sodium acetate will be combined in solution.

In the buffer investigation you will attempt to prepare an acetate buffer of pH 5.00 and determine the buffer capacity with respect to strong acid and strong base.

GOALS

As you complete this investigation you will:

- 1 Prepare a buffer with a pH of ~ 5.00 .
- 2 Verify the pH of the buffer.
- 3 Determine the buffer capacity with regard to strong acid and strong base.

MATERIALS AVAILABLE FOR USE

Volumetric flasks

Graduated cylinders

Pipettor and tips

Beakers

Erlenmeyer flasks

Buret

Funnel

Analytical balances

Vernier LabPro system

pH probe

Probe calibration solutions

6.0 and 1.0 *M* acetic acid ($\text{HC}_2\text{H}_3\text{O}_2$)

Sodium acetate ($\text{NaC}_2\text{H}_3\text{O}_2$)

0.50 and 0.050 *M* $\text{NaOH}(aq)$

0.50 and 0.050 *M* $\text{HCl}(aq)$

SAFETY PRECAUTIONS

CAUTION: Acetic acid, HCl, and NaOH are corrosive. They can attack the skin and cause permanent damage to the eyes. If one of these solutions splashes into your eyes, use the eyewash station immediately. Hold your eyes open and flush with water. If contact with skin or clothing occurs, flush the affected area with water. Have your lab partner notify your instructor about the spill.

WASTE DISPOSAL

All solutions can be flushed down the sink with water.

GETTING STARTED

- 1 What reagents will you use to obtain desired pH?
- 2 What amounts of reagents will you use to produce the buffer?
- 3 Review titration techniques² in preparation for determining the buffer capacity.

PROCEDURE

- 1 You need to prepare 100 mL of the buffer solution using the appropriate volume of 1.0 *M* acetic acid and the appropriate mass of sodium acetate. You will want the concentrations relatively dilute, so make the acid concentration 0.10 *M* or 0.010 *M*.
- 2 Check the pH of your buffer. It should be within ± 0.2 units.
- 3 Buffer capacity is defined as the number of moles of strong base or strong acid required to cause a one-unit change in pH of a buffer solution.
 - a Determine the buffer capacity for strong acid by titrating 25.0 mL of your buffer with 0.50 *M* HCl for the 0.1 *M* buffer or 0.050 *M* HCl for the 0.01 *M* buffer until the pH decreases by 1 unit. Report the capacity as moles acid per L buffer.
 - b Determine the buffer capacity for strong base by titrating 25.0 mL of your buffer with 0.50 *M* NaOH for the 0.1 *M* buffer or 0.050 *M* NaOH for the 0.01 *M* buffer until the pH increases by 1 unit. Report the capacity as moles base per L buffer.

²../videos/titration.html

INTERACTIVE POSTER SESSION

Once your group has completed your work, prepare a whiteboard that you can use to share and justify your ideas. See the handout³ provided for details on this process.

REPORT

Once you have completed your research, you will need to prepare an *investigation report* that consists of three sections. This report may require more than 2 pages with data tables. This report must be typed and any diagrams, figures, or tables should be embedded into the document.

Section 1: What concept were you investigating, and how does it relate to the guiding question? **How well can a buffer resist pH change?** Define and explain relevant terms, buffer, pH, and buffer capacity. See the getting started questions for suggestions as well as the introduction.

Section 2: Outline the calculations used to determine the amounts of materials you used in the buffer solution. Describe the procedure you used to prepare the buffer and to measure the buffer capacity.

Section 3: What is your **argument**? Present data for pH and buffer capacity. Show a calculation for prediction of buffer capacity. How did experimental pH and the buffer capacity of your prepared buffer compare to the calculated values? What would cause variations? What range of pH could you buffer with your proposed system? How did your system compare with other groups, particularly those that worked with a different buffer system?

This third section is where you not only present your data, but **use** the values you obtain as evidence in your reasoning. Statements like, “see data table for values” are not acceptable!

Remember: An argument is not just an answer to the question. It is a claim or conclusion supported by evidence with a rationale for why the evidence supports the claim or conclusion.

³../poster/manual.html