

The Effect of Added Acid and Base on the pH of Buffer Solutions

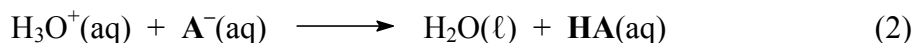
Introduction

Buffers are very important in many chemical and biological systems. For example, blood is a buffer solution that can resist changes in pH that result from metabolic processes. The normal pH range of human blood is 7.35 – 7.45. If the pH falls below 7.35, the condition is called acidosis. This can be caused by emphysema, diabetes or many other illnesses and can result in coma or death. If blood pH rises to 7.45, the condition is called alkalosis, which is also harmful.

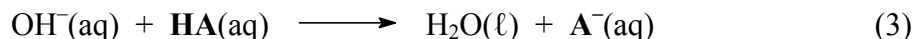
Buffers contain either a weak acid and its conjugate base or a weak base and its conjugate acid. A buffer solution has the ability to resist changes in pH when small amounts of acid or base are added. For example, consider a weak acid/conjugate base buffer. By rearranging the equilibrium constant expression, it is seen that the pH depends on the $[HA]/[A^-]$ ratio for a given conjugate acid/base pair (K_a is constant):

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} \Rightarrow [H_3O^+] = K_a \frac{[HA]}{[A^-]} \quad (1)$$

When a small amount of acid (H_3O^+) is added to the buffer, some of the conjugate base, A^- , is converted to HA.



The **ratio** of the two species changes, but as long as it is less than 10x, the pH will not change by more than 1 (because of the logarithmic relationship). If a little base (OH^-) is added, some of the HA is converted into A^- .



Again, the pH remains relatively constant for the same reason. It is not until much of the HA or A^- is consumed that the pH changes dramatically.

Two important characteristics of a buffer are its pH and its **buffer capacity**; which is the amount of acid or base the buffer can react with before giving a **significant** pH change. Buffer capacity depends on the total concentrations of the weak acid and its conjugate base in the solution.

Procedure

Making the Buffer, 0.10 M HCl and 0.10 M NaOH

1. Obtain three clean 250-mL Erlenmeyer flasks. Label one '**Buffer**'. Label the second flask '**0.100 M HCl**'. Label the third flask '**0.100 M NaOH**'.

2. To prepare 200.0 mL of buffer solution, add 20.0 mL of 1.00 M acetic acid solution to a 100-mL graduated cylinder and add enough distilled water to bring the volume to the 100.0 mL mark. Add to the 'Buffer' flask. Add 20.0 mL of 1.00 M sodium acetate solution to the 100-mL graduated cylinder, add enough distilled water to bring the volume to the 100.0 mL mark. Add to the 'Buffer' flask. Mix.
3. To prepare 200.0 mL of 0.100 M hydrochloric acid solution, add 20.0 mL of 1.00 M hydrochloric acid solution to a 100-mL graduated cylinder. Add enough distilled water to bring the volume to the 100.0 mL mark. Add to the '0.100 M HCl' flask. Add an additional 100.0 mL of distilled water to the flask. Swirl to mix thoroughly.
4. Prepare 200.0 mL of 0.100 M NaOH solution by repeating step 3 with 20.0 mL of 1.00 M NaOH.

Testing the Buffer and Distilled Water Against Added Acid and Base

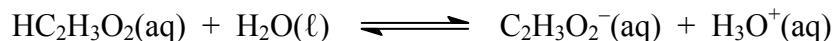
5. Obtain four clean 250-mL beakers. Pipet 50.00 mL of the buffer solution from step 2 into two **labeled** beakers. Pipet 50.00 mL of distilled water into two **labeled** beakers.
6. Add 4 drops of 0.03% methyl orange indicator solution and 4 drops of 0.1% malachite green indicator solution to one of the buffer and one of the distilled water beakers. Place the pH electrode in the buffer solution and measure the pH of solution with the pH meter. Note the color of the solution. **Keep the pH electrode in the solution throughout the following steps.**
7. Pipet 5.00 mL of 0.100 M HCl solution from step 3 to the buffer solution. Record the pH and note the color.
8. Using 5.00, 10.00 and 20.00 mL pipets, repeat step 7, but with the following 0.10 M HCl solution **additions**: 5.00 mL, 10.00 mL, 10.00 mL, 20.00 mL, 20.00 mL and 20.00 mL. You will have 8 data points: pH at 0.00, 5.00, 10.00, 20.00, 30.00, 50.00, 70.00 and 90.0 mL **total** added acid. Note the color each time an addition is made.
9. Rinse the pH electrode with distilled water. Repeat steps 6,7 and 8 in the distilled water beaker.
10. Add 4 drops of bromothymol blue indicator solution and 2 drops of 1% phenolphthalein indicator solution to the other buffer and distilled water beakers. Repeat steps 6, 7, 8 and 9 using 0.100 M NaOH from step 4 instead of HCl.

Testing the Buffer Against Dilution

11. Pipet 50.00 mL of the buffer solution into a **clean, dry** beaker. Record the pH. Pipet 50.00 mL of distilled water into the same beaker. Mix thoroughly and record the pH.

Question

In step 2, the **initial concentrations** of acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$, and acetate ion, $\text{C}_2\text{H}_3\text{O}_2^-$, are both 0.500 M. Calculate the pH of the solution given an equilibrium reaction of



and a K_a of 1.8×10^{-5} .

Data Treatment and Discussion

Include the following:

1. Plots of pH vs **total** mL of added HCl and **total** mL of added NaOH for **BOTH** buffer solution and distilled water (4 graphs).
2. Determine the **buffer** capacity (toward acid **and** base, in mL) by finding (and indicating on the plot) the point on the graph where the pH has changed by 1.00 unit from **its initial value**. Read this value to at least 0.1 mL by including gridlines (both directions) on the buffer plots.

Conclusion

Give the buffer capacity toward acid and toward base. Also address:

Considering the composition of your buffer (the amount of the weak acid and its conjugate base), do the **relative** buffer capacities toward acid and base make sense? Explain why or why not.

Does dilution have a **significant** effect on the pH of a buffer? Explain why or why not.