# LAB 6 Preparation of Buffer Solutions

## Learning Objectives

- **Employ the principles of the pH scale to create a buffer solution**
- **Calculate the pH of a buffer solution using the Henderson-Hasselbalch equation**

### INTRODUCTION

Buffers play a significant role in the world. For instance, natural soil contains buffers that help mitigate the consequences of outside factors, such as acid rain, in an ecosystem. Because of these buffers, the soil is able to maintain a relatively constant pH. If buffers were removed from the ecosystem and the pH were to drop or increase significantly, many of the organisms living in the soil would not be able to survive (Figure 1).

**Figure 1.** Natural rain has pH of approximately 5.6 and acid rain has a pH of approximately 3.5 - 4. Natural chemical buffers which exist in soil can help to mitigate the concerns created by acid rain.

Blood also contains natural buffers. Human blood has a pH of about 7.4. If this pH deviates more than ± 0.2, the natural systems would stop working, possibly resulting in death. Fortunately, the natural buffers in our blood are very effective and keep our blood pH between 7.35 and 7.45 pH at all times.

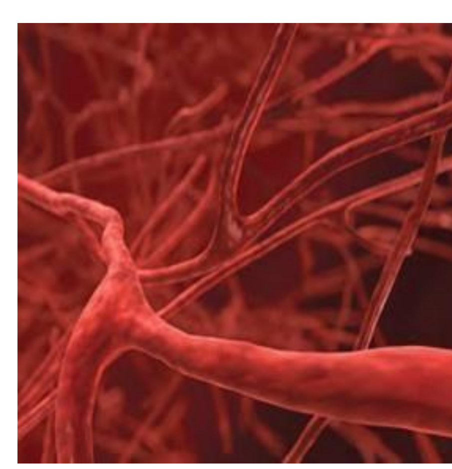

**Figure 2.** The pH of blood is maintained by the bicarbonate and carbonate buffer system.

### pH SCALE

The acidity of a solution is a measure of the hydrogen ion concentration, $[H^+]$ or $[H_3O^+]$. (Conversely, the basicity of a solution is a measure of the concentration of hydroxide ions $[OH^-]$.) Often these ion concentrations are very small, so in order to conveniently express the concentration of protons (or H+ ions) in an aqueous solution, the pH scale is used. pH is defined as the negative logarithm of the molar concentration of the hydronium (or hydrogen) ion. The equation looks like:

$$pH = -log[H_3O^+]$$

Thus, the **pH scale** is based on the hydronium ion concentration. It is most commonly indicated by a simple color or number scale. Red typically indicates acidic solutions and a low pH. On the other end of the spectrum, blue typically indicates basic solutions and a high pH (Figure 3). The most commonly used pH scale can be seen in Figure 3.

| pH: | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| | | Acidic | | | | | Neutral | | | | | Basic | | |

**Figure 3.** Note that many strong acids and bases do not have a pH that is indicated on this scale. For example, lead battery acid has a pH that is below one.

### WHAT IS A BUFFER?

A **buffer** can be thought of as a pH shock absorber. A buffer resists pH changes upon the addition of an acid or base, or when diluted. A buffer consists of a weak acid and its conjugate base in equilibrium (or vice-versa), so that the solution is able to react with either an added acid or base and maintain the equilibrium and thus the pH level.

# Preparation of a Buffer Solution

## CONJUGATE PAIRS

A conjugate pair is an acid-base couple that only differs by one proton. This definition is in accordance with the Brønsted-Lowery Theory, which states that an acid is a proton ($H^+$) donor and a base is a proton acceptor. When an acid gives up a proton, a species that can accept a proton, called a base, is formed from the acid. That base is called the conjugate base of the acid. For example, hydrogen chloride (HCl) and chloride ($Cl^-$) are conjugates. Hydrogen chloride (hydro is the acid - it can donate protons; it is protonated) and chloride is the base (it can accept protons; it is deprotonated). Another example is water ($H_2O$),and hydronium ($H_3O^+$). In this case, water is the base (it is deprotonated) and hydronium is the acid (it is protonated).

## ACID-BASE BUFFERS MUST BE COMPOSED OF WEAK ACIDS AND BASES

Acids and bases fall into two general categories: weak or strong. Strong acids dissociate almost 100% to produce $H_3O^+$ ions and strong bases dissociate (or react with water) almost 100% to produce hydroxide ions ($OH^-$). In contrast, weak acids and bases dissociate only to a small extent, resulting in an equilibrium between a weak acid and its weak conjugate base, or vice-versa. Weak acids and their conjugate weak bases (or vice versa) must be used to create a buffer.

For example, suppose we mix a weak acid such as carbonic acid ($H_2CO_3$) with its conjugate base, the bicarbonate ion ($HCO_3^-$). This solution contains an acid and a base that are available to react with any added acids or bases. The following equilibrium us established in the solution:

$$H_2CO_3\,(aq) \rightleftharpoons H^+\,(aq) + HCO_3^-\,(aq)$$

When a hydrogen ion source ($H^+$) is added, the conjugate base, (the bicarbonate ion), will react as follows:

$$H^+\,(aq) + HCO_3^-\,(aq) \rightarrow H_2CO_3\,(aq)$$

As the added acid reacts with the bicarbonate ion, additional carbonic acid is formed, thus slightly changing the $H_2CO_3$ /$HCO_3^-$ in the solution.

If hydroxide ($OH^-$) is added, the weak acid, (sodium bicarbonate), will react as follows:

$$OH^-\,(aq) + H_2CO_3\,(aq) \rightarrow HCO_3^-\,(aq)$$

As the added base reacts with the carbonic acid, additional bicarbonate ion is formed, again slightly changing the $H_2CO_3$ /$HCO_3^-$ in the solution.

In both circumstances, the acid or base introduced does not significantly change the pH of the solution, due to the buffer interactions.

## BUFFER CAPACITY

The function of a buffer is to keep the pH from drastically changing. However, when all of the available acid or base in the buffer solution has been consumed, the pH will begin to significantly change. The buffer solution will not be able to maintain the pH as additional acid or base is added. This buffer has reached the end of its ability to absorb additional acid or base introduced; in other words, it has reached its **buffer capacity**. This is the maximum amount of strong acid or strong base that can be added before a significant change in pH will occur.

## HENDERSON-HASSELBALCH EQUATION

For any buffer solution, the pH can be determined from the acid and base concentrations using the ***Henderson-Hasselbalch equation***:

$$pH = pK_a + log\,\frac{[base]}{[acid]}$$

If a buffer solution is created with *equal concentrations* of a weak acid and its conjugate base, the log term of the equation will be equal to zero, and the pH of the buffer is equal to the $pK_a$ of the weak acid. In these calculations, the $K_a$ is the acid dissociation constant. Using this equation, we can prepare buffer solutions with the right acid/base ratio to maintain almost any pH. The **Henderson-Hasselbalch equation** describes a mathematical relationship between the pH, the $pK_a$ of a weak acid, and the concentrations of the weak acid and its conjugate base.

# Preparation of a Buffer Solution

## HENDERSON-HASSELBALCH EXAMPLE

Calculate the pH of a buffer solution which has been made by mixing 50.0 mL of a 0.25 M acetic acid ($CH_3COOH$) solution with 50 mL of a 0.15 M sodium acetate ($CH_3COONa$) solution.

**Note: The $K_a$ of acetic acid is 1.8 x $10^{-5}$**

## SOLUTION

First, we must determine the concentrations of the acid ($CH_3COOH$) and the base ($CH_3COONa$, or just $CH_3COO^-$) in the buffer. For this we will use a dilution calculation

$$M_cV_c = M_dV_d$$

$$[CH_3COOH]_d = \frac{[CH_3COOH]_cV_c}{V_d} = \frac{(0.25\ M)(0.050\ L)}{0.10\ L} = 0.125\ M\ CH_3COOH$$

$$[CH_3COO^-]_d = \frac{[CH_3COO^-]_cV_c}{V_d} = \frac{(0.15\ M)(0.050\ L)}{0.10\ L} = 0.075\ M\ CH_3COO^-$$

Using the Henderson-Hasselbalch equation, we can determine the pH of the solutions.

$$pH = pK_a + log\frac{[base]}{[acid]}$$

$$pH = -\log(1.8\ x\ 10^{-5}) + log\frac{0.075\ M}{0.125\ M} = 4.52$$