8.8 Buffer Systems

buffer an aqueous solution containing a conjugate acid–base pair that maintains a nearly constant pH when an acid or base is added

Much important chemistry, including almost all the chemistry of the natural world, occurs in aqueous solution. Living systems usually require a relatively narrow pH range for survival. Because of this, they depend on solutions that are resistant to changes in pH. Such a solution is called a **buffer**, or buffer solution: an aqueous solution that resists changes in pH upon the addition of an acid or base. A buffer contains a weak acid and its conjugate base (or a weak base and its conjugate acid) in an equilibrium mixture. Buffers may be prepared in a lab setting or they may be naturally occurring. Buffers can be created to maintain almost any pH value. In this section, we explore why buffers are important, their characteristics, and how they work.

Buffers in Action

Buffers may contain different acids, bases, and ions, but buffers all serve one purpose: to keep the pH of a solution constant when small quantities of acid or base are added to the solution.

Many biological processes include chemical reactions that must occur in a specific range of pH; buffers function to maintain this pH range. Gas exchange, muscle activity, and many other biological processes are sensitive to pH. Enzymes function at an optimal pH range. CAREER LINK

Although human blood contains many chemical components that help maintain the pH at a constant level, there is one particularly important conjugate acid–base pair: hydrogen carbonate, $HCO_3^-(aq)$, and carbonate, $CO_3^{2-}(aq)$. These ions produce a pH of 7.4 in healthy blood. Since our cells are so sensitive to pH, it is important that this pH be maintained. When reactions occur in our bodies, such as the formation of lactic acid, $HC_3H_5O_3(aq)$, by active muscles, these pH buffers must be capable of neutralizing the effects of this acid to maintain the blood's pH at 7.4. Even small changes in pH may have harmful effects on a person's health.

Buffers are also an important ingredient in many foods, drinks, and medicines. Buffers are useful to maintain the pH at a desired level. Sodium citrate, for example, is a common additive in products that contain fruit, such as jams and preserves (**Figure 1**). The citrate ion combines with the weak acids in fruit to form a buffer system that resists pH changes in the product. This helps to stabilize the properties of the product, including taste. Sodium citrate also acts as a preservative. Aspartame is a common artificial sweetener used in low-calorie soft drinks. The stability of the aspartame molecule is pH dependent. Aspartame decomposes readily unless its pH is maintained within a range of 4 to 5 by the additives in soft drinks. Some medicines use a buffer coating to reduce the risk of stomach upset from the acidic medicinal ingredient.

Figure 1 Buffers help to preserve foods containing fruits or vegetables.

The food industry uses weak acids to change the flavour of food and to improve its shelf life. Citric acid is one of the most commonly used acidic additives in the food industry. It is a natural organic (carboxylic) acid found in many fruits, vegetables, and dairy products. This acid is added to many foods, beverages, and candies. The sour taste of citric acid enhances fruit flavours like citrus and berry. The salt calcium citrate, which tastes sour and salty, contains the conjugate base of citric acid. The use of citric acid and calcium citrate deters bacterial growth and stabilizes pH. You will explore the use of buffers in the food industry in more detail in Section 8.9.

UNIT TASK BOOKMARK

For the Unit Task (page 582), you may ask a testable question regarding the buffering capacity of your consumer product.

Characteristics of a Buffer

A buffer keeps the pH of a solution within a certain narrow range. When an acid or base is added to this solution, the buffer prevents the pH from changing. Buffers always contain a mixture of a weak acid (or base) and its conjugate ion in solution. Living cells contain a phosphate buffer, which includes the conjugate acid–base pair of dihydrogen phosphate, $H_2PO_4^-(aq)$ (a weak acid), and hydrogen phosphate, $HPO_4^{2-}(aq)$ (its conjugate base). This buffer system maintains a neutral pH. Other combinations of an acid (or base) and its conjugate maintain the pH at different pH values.

Acidic Buffers

An acidic buffer includes a weak acid and its conjugate base. Let us consider an acidic buffer system in which ethanoic acid, $HC_2H_3O_2(aq)$, is the weak acid. Its conjugate base is the ethanoate ion, $C_2H_3O_2^-(aq)$. Imagine a titration in which a basic titrant, such as sodium hydroxide solution, is added to the ethanoic acid buffer. The pH is monitored throughout the titration as shown in **Figure 2**. You see that, in the highlighted region, the pH changes very little as $OH^-(aq)$ ions are added. This region, about halfway to the equivalence point, is called the "buffer zone." In the buffer zone there are approximately equal amounts of unreacted ethanoic acid, $HC_2H_3O_2(aq)$, and its conjugate base, $C_2H_3O_2^-(aq)$, in the receiving flask. This mixture of entities prevents large changes in pH as more hydroxide ions are added. The solution in the receiving flask at halfway to the equivalence point is a buffer: the pH of this solution changes very little as more base is added.

In contrast, at or near the equivalence point, the pH changes dramatically when even small amounts of hydroxide ions are added.

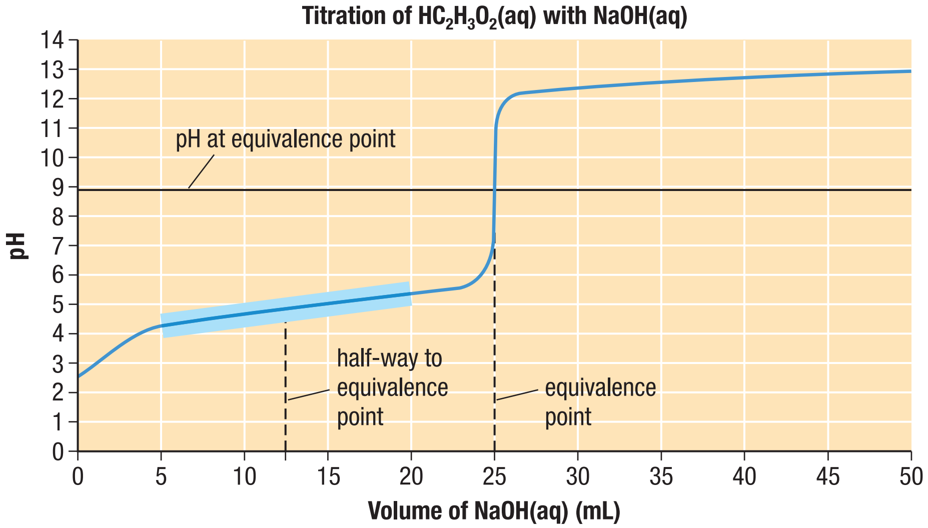

Figure 2 This graph shows the titration of ethanoic acid (a weak acid) with sodium hydroxide (a strong base). Note that, from the addition of 5 to 23 mL of base, the pH of the solution stays relatively constant.

Another example of an acidic buffer is a solution containing hydrofluoric acid, $HF(aq)$, and sodium fluoride, $NaF(aq)$. The conjugate base in this case is the fluoride ion, $F^-(aq)$. As you know, since hydrofluoric acid is a weak acid, a hydrofluoric acid solution has a relatively high concentration of non-ionized HF molecules and relatively low concentrations of $H^+(aq)$ and $F^-(aq)$ ions. A solution of hydrofluoric acid alone is not a buffer solution. To change this solution to a buffer, we must increase the concentration of $F^-(aq)$ ions until it is approximately equal to the concentration of non-ionized $HF(aq)$ molecules. This is done by adding a soluble salt containing fluoride ions, such as sodium fluoride, $NaF(s)$.

Buffer solutions are usually prepared by simply mixing a weak acid or base with a soluble salt containing the corresponding conjugate acid or base in approximately equal (molar) amounts.

By mixing appropriate acid–salt or base–salt combinations, we can prepare buffer solutions with virtually any pH. Acidic buffers (those based on a weak acid and a salt containing its conjugate base) generally have their buffering zone at a pH below 7.

Buffers with a higher pH can be made from a weak base and its conjugate acid. An example of a basic buffer is a solution containing ammonia, $NH_3(aq)$, and a salt. In this case, ammonia is the weak base component and the ammonium ion, $NH_4^+(aq)$, is the conjugate acid component. A soluble salt such as $NH_4Cl(s)$ could provide the additional $NH_4^+(aq)$ ions necessary to form the buffer.

How Do Buffers Work?

How does a buffer maintain its pH when an acid or a base is added? Suppose a buffer contains relatively large quantities of a weak acid, $HA(aq)$, and its conjugate base, $A^-(aq)$. The general equation for the ionization equilibrium is

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$$

When hydroxide ions are added to the solution, the non-ionized weak acid molecules react with the hydroxide ions, removing the hydroxide ions from the solution.

$$HA(aq) + OH^-(aq) \rightleftharpoons A^-(aq) + H_2O(aq)$$

The net result of this reaction is that, when a base is added, $OH^-(aq)$ ions are not allowed to accumulate but are replaced by $A^-(aq)$ ions (**Figure 3**).

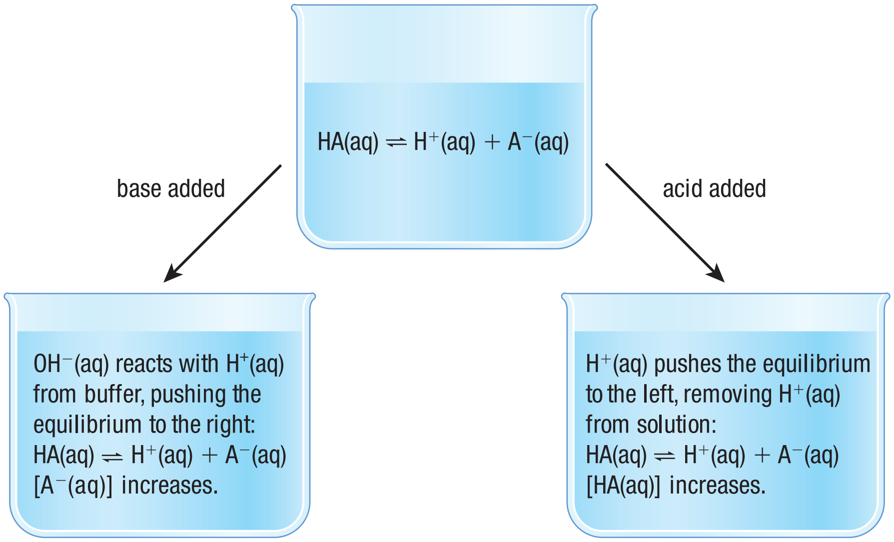

Figure 3 Le Châtelier's principle explains the changes that occur in an acidic buffer solution when a base or an acid is added.

When hydrogen ions are added to the solution, they push the ionization equilibrium to the left, removing the hydrogen ions from the solution and increasing the concentration of the acid molecules.

$$HA(aq) \rightleftharpoons H^+(aq) + A^-(aq)$$

The pOH (and hence the pH) of the solution thus remains relatively constant regardless of whether an acid or a base is added.

We can understand the stability of the pH under these conditions by examining the equilibrium equation for the ionization of a weak acid, HA(aq):

$$K_a = \frac{[H^+(aq)][A^-(aq)]}{[HA(aq)]}$$

Rearranging this equation to isolate $[H^+(aq)]$ on the left side, we get

$$[H^+(aq)] = K_a \frac{[HA(aq)]}{[A^-(aq)]}$$

This form of the equation tells us that the equilibrium concentration of hydrogen ions, $H^+(aq)$, (and thus the pH) is determined by the ratio $[HA(aq)]/[A^-(aq)]$. When $OH^-(aq)$ ions are added, HA(aq) is converted to $A^-(aq)$, and the $[HA(aq)]/[A^-(aq)]$ ratio decreases. However, if the amounts of HA(aq) and $A^-(aq)$ originally present are very large compared with the amount of $OH^-(aq)$ added (as is the case in buffer solutions), the change in the $[HA(aq)]/[A^-(aq)]$ ratio, and thus the change in pH, will be small.

The essence of buffering is that $[HA(aq)]$ and $[A^-(aq)]$ are large compared with the amount of $OH^-(aq)$ added. Thus, when a small quantity of $OH^-(aq)$ is added, the concentrations of HA(aq) and $A^-(aq)$ change only slightly. Under these conditions, the $[HA(aq)]/[A^-(aq)]$ ratio stays almost the same within the buffering zone. Thus $[H^+(aq)]$ (and pH) remain virtually constant.

Similarly, when hydrogen ions are added to a buffer, they react with $A^-(aq)$ ions to form the weak acid, HA(aq). There will be a net change of $A^-(aq)$ to HA(aq). However, if $[A^-(aq)]$ and $[HA(aq)]$ are large compared with the amount of $H^+(aq)$ added, little change in the pH will occur. **Figure 4** illustrates why a high initial concentration of the buffering entities is very important for them to function effectively.

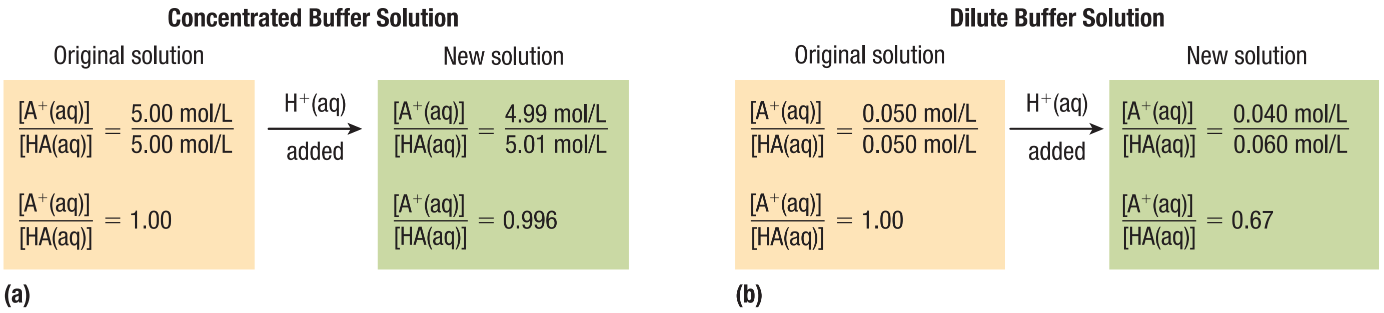

Figure 4 (a) When the initial concentrations of the buffering entities are high, the addition of a small amount of acid has little effect on their ratio. (b) When the initial concentrations are low, the addition of a small amount of acid has a much greater effect on their ratio.

Instructions for mixing buffer solutions can be found in many chemistry resources, including *The CRC Handbook of Chemistry and Physics*.

pH of Buffer Solutions

Buffers are simply solutions of weak acids or bases and their respective soluble salts. The pH calculations for buffers require exactly the same procedures used in studying acids, bases, and acid–base titrations. The problem-solving strategy was introduced in Section 8.4. *This is not a new type of problem.* In Tutorial 1, you will learn how to use this strategy to calculate the pH of buffer solutions.

Investigation **8.8.1**

Buffer Systems (page 572)
This observational study compares how pH changes in a buffer solution and in a control solution.

Tutorial 1 The Initial pH of Buffer Solutions

In this tutorial, you will learn how to calculate the pH of an acidic buffer and a basic buffer.

Sample Problem 1: Calculating the pH of an Acidic Buffer

A chemist in a leather-tanning facility is considering using a buffer containing 0.50 mol/L ethanoic acid, $HC_2H_3O_2(aq)$ ($K_a = 1.8 \times 10^{-5}$), and 0.50 mol/L sodium ethanoate, $NaC_2H_3O_2(aq)$. Calculate the pH of this solution.

Given: 0.50 mol/L ethanoic acid, $K_a = 1.8 \times 10^{-5}$; 0.50 mol/L sodium ethanoate

Required: pH

Solution:

Step 1. Write the formulas of all the entities in solution.

The entities in the solution are

$HC_2H_3O_2(aq)$,	$Na^+(aq)$,	$C_2H_3O_2^-(aq)$,	and	$H_2O(l)$
↑	↑	↑		↑
Weak acid	Neither acid nor base	Weak base (conjugate base of $HC_2H_3O_2$)		Very weak acid or base

Step 2. Write the reaction equation and the equilibrium constant equation for the equilibrium that determines the pH of the solution.

To decide which equilibrium reaction determines the pH of the solution, examine the list of entities in solution. The ethanoic acid ionization equilibrium, which involves both $HC_2H_3O_2(aq)$ and $C_2H_3O_2^-(aq)$, is the relevant reaction:

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

The equilibrium constant equation is

$$K_a = \frac{[H^+(aq)][C_2H_3O_2^-(aq)]}{[HC_2H_3O_2(aq)]}$$

$$1.8 \times 10^{-5} = \frac{[H^+(aq)][C_2H_3O_2^-(aq)]}{[HC_2H_3O_2(aq)]}$$

Step 3. Determine initial and equilibrium concentrations.

Table 1 shows the ICE table. Note that, before reaction, the initial concentration of $HC_2H_3O_2(aq)$ is 0.50 mol/L and the initial concentration of $H^+(aq)$ is essentially 0 mol/L. Also, note that the initial concentration of $C_2H_3O_2^-(aq)$ is equal to the concentration of $NaC_2H_3O_2(aq)$.

Table 1 ICE Table for the Ethanoic Acid–Sodium Ethanoate Buffer Equilibrium (mol/L)

	$HC_2H_3O_2(aq)$	$\rightleftharpoons$	$H^+(aq)$	+	$C_2H_3O_2^-(aq)$
I	0.50		≈ 0		0.50
C	$-x$		$+x$		$+x$
E	$0.50 - x$		x		$0.50 + x$

Step 4. Substitute equilibrium concentrations from the ICE table into the equilibrium constant equation.

$$1.8 \times 10^{-5} = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$1.8 \times 10^{-5} = \frac{(x)(0.50 + x)}{0.50 - x}$$

Step 5. Use the hundred rule to determine if any simplifying assumptions may be made.

$$\text{Since } \frac{[\text{HC}_2\text{H}_3\text{O}_2(\text{aq})_{\text{initial}}]}{K_a} = \frac{0.50}{1.8 \times 10^{-5}}$$

$$\frac{[\text{HC}_2\text{H}_3\text{O}_2(\text{aq})_{\text{initial}}]}{K_a} = 2.8 \times 10^{4}$$

Since the $\frac{[\text{HC}_2\text{H}_3\text{O}_2(\text{aq})_{\text{initial}}]}{K_a}$ ratio is much greater than 100, you may make an assumption about the equilibrium concentration of $HC_2H_3O_2(aq)$:

$0.50 - x \approx 0.50$

By a similar calculation you may make an assumption about the equilibrium concentration of $C_2H_3O_2^-(aq)$:

$0.50 + x \approx 0.50$

Thus,

$$1.8 \times 10^{-5} \approx \frac{(x)(0.50)}{0.50}$$

$$x \approx 1.8 \times 10^{-5}$$

Step 6. Use the 5 % rule to check your result:

$$\frac{x}{[\text{HC}_2\text{H}_3\text{O}_2(\text{aq})_{\text{initial}}]} \times 100\ \% = \frac{1.8 \times 10^{-5}}{0.5} \times 100\ \%$$

$$\frac{x}{[\text{HC}_2\text{H}_3\text{O}_2(\text{aq})_{\text{initial}}]} \times 100\ \% = 0.0036\ \%$$

Since 0.0036 % < 5 %, the error in the result is acceptable.

A similar calculation shows that the assumption that $0.50 + x \approx 0.50$ is valid.

Step 7. Determine $[H^+(aq)]$ at equilibrium and calculate pH:

$$x \approx 1.8 \times 10^{-5}$$

$$[\text{H}^+(\text{aq})] \approx 1.8 \times 10^{-5}$$

$$\text{pH} \approx -\log(1.8 \times 10^{-5})$$

$$\text{pH} \approx 4.74$$

Statement: The pH of the ethanoic acid/sodium ethanoate buffer is 4.74.

Basic buffers are solutions of a weak base and a soluble salt of its conjugate acid in approximately equal concentrations. You can determine the pH of a basic buffer in much the same way as an acidic buffer.

Sample Problem 2: Calculating the pH of a Basic Buffer

Calculate the pH of a buffer made with 0.25 mol/L ammonia solution, $NH_3(aq)$ ($K_b = 1.8 \times 10^{-5}$), and 0.40 mol/L ammonium chloride solution, $NH_4Cl(aq)$.

Given: 0.25 mol/L ammonia solution, $K_b = 1.8 \times 10^{-5}$; 0.40 mol/L ammonium chloride solution

Required: pH

Solution:

Step 1. Write the formulas of all the entities in solution.

The entities in the solution are $\text{NH}_3(\text{aq}),\ \underbrace{\text{NH}_4^{\ +}(\text{aq}),\ \text{Cl}^-(\text{aq})}_{\text{From the dissolved NH}_4\text{Cl(aq)}}$, and $H_2O(l)$.

Step 2. Write the reaction equation and the equilibrium constant equation for the equilibrium that determines the pH of the solution.

Since the chloride ion is such a weak base and water is a weak acid or base, the important equilibrium is

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

The equilibrium constant equation is

$$K_b = \frac{[NH_4^+(aq)][OH^-(aq)]}{[NH_3(aq)]}$$

$$1.8 \times 10^{-5} = \frac{[NH_4^+(aq)][OH^-(aq)]}{[NH_3(aq)]}$$

Step 3. Determine initial and equilibrium concentrations (**Table 2**).

Table 2 ICE Table for the Ionization of Ammonia, $NH_3(aq)$ (mol/L)

	$NH_3(aq)$	$+$	$H_2O(l)$	$\rightleftharpoons$	$NH_4^+(aq)$	$+$	$OH^-(aq)$
I	0.25		—		0.40		≈ 0
C	$-x$		—		$+x$		$+x$
E	$0.25 - x$		—		$0.40 + x$		x

Step 4. Substitute equilibrium concentrations from the ICE table into the equilibrium constant equation:

$$1.8 \times 10^{-5} = \frac{[NH_4^+(aq)][OH^-(aq)]}{[NH_3(aq)]}$$

$$1.8 \times 10^{-5} = \frac{(0.40 + x)(x)}{0.25 - x}$$

Step 5. Use the hundred rule to determine if any simplifying assumptions may be made.

Assume that $0.25 - x \approx 0.025$ and that $0.40 + x = 0.40$.

$$1.8 \times 10^{-5} \approx \frac{(0.40)(x)}{0.25}$$

$$x \approx 1.1 \times 10^{-5}$$

Step 6. Use the 5 % rule to check your result.

The approximation is valid by the 5 % rule. (Check for yourself.)

Step 7. Determine $[OH^-(aq)]$ at equilibrium and calculate pH.

$$x \approx 1.1 \times 10^{-5}$$

$$[OH^-(aq)] \approx 1.1 \times 10^{-5}$$

$$pOH \approx -\log(1.1 \times 10^{-5})$$

$$pOH \approx 4.96$$

$$\begin{aligned} \text{pH} &\approx 14.00 - \text{pOH} \\ &\approx 14.00 - 4.96 \\ \text{pH} &\approx 9.04 \end{aligned}$$

Statement: The pH of the buffer with 0.25 mol/L $NH_3(aq)$ and 0.40 mol/L $NH_4Cl(aq)$ is 9.04.

Practice

1. A buffer solution contains 2.250 mol/L ethanoic acid, $HC_2H_3O_2(aq)$ ($K_a = 1.8 \times 10^{-5}$), and 2.250 mol/L sodium ethanoate, $NaC_2H_3O_2(aq)$. Calculate the pH of this buffer solution. K/U [ans: 4.74]
2. A buffer contains 0.10 mol/L hydrofluoric acid, HF(aq) ($K_a = 6.6 \times 10^{-4}$), and 0.10 mol/L fluoride ions, $F^-(aq)$. Calculate the pH of this buffer solution. K/U [ans: 3.18]
3. Calculate the pH of a buffer solution containing 0.15 mol/L ammonia $NH_3(aq)$ ($K_b = 1.8 \times 10^{-5}$), and 0.04 mol/L ammonium chloride, $NH_4Cl(aq)$. K/U [ans: 8.83]

Buffers and Equilibria

We can better understand buffers in action if we consider their Brønsted–Lowry equations. For example, when sodium hydroxide is added to a buffer solution of ethanoic acid–ethanoate ions, the following equilibrium occurs:

$$HC_2H_3O_2(aq) + OH^-(aq) \rightleftharpoons C_2H_3O_2^-(aq) + H_2O(l)$$

The hydroxide ions from sodium hydroxide convert ethanoic acid to ethanoate ions. This reduces the ratio of ethanoic acid to ethanoate ions, which results in a slight increase in pH. If there were no buffer present, there would be a large increase in pH. This ethanoic acid–ethanoate buffer can also buffer a strong acid in a similar way, except that added hydrogen ions (from the strong acid) encourage the formation of ethanoic acid:

$$C_2H_3O_2^-(aq) + H^+(aq) \rightleftharpoons HC_2H_3O_2(aq)$$

The ratio of ethanoic acid to ethanoate ions will increase and the pH will decrease slightly. The pH decrease would be much greater if there were no buffer present.

WEB LINK

We can summarize the characteristics of buffers as follows:

- Buffers contain relatively large concentrations of a weak acid and its conjugate base. They can involve a weak acid, HA(aq), and the conjugate base, $A^-(aq)$, or a weak base, B(aq), and the conjugate acid, $BH^+(aq)$.
- When an acid is added to a buffered solution (of either type), the hydrogen ions react essentially to completion with the weak base present:

$$H^+(aq) + A^-(aq) \rightleftharpoons HA(aq) \quad \text{or} \quad H^+(aq) + B(aq) \rightleftharpoons BH^+(aq)$$

- When a base is added to a buffer, the hydroxide ions react essentially to completion with the weak acid present:

$$OH^-(aq) + HA(aq) \rightleftharpoons A^-(aq) + H_2O(l) \quad \text{or}$$

$$OH^-(aq) + BH^+(aq) \rightleftharpoons B(aq) + H_2O(l)$$

The pH of a buffer is determined by the ratio of the concentrations of the weak acid and weak base.

$$\frac{[HA(aq)]}{[A^-(aq)]} \quad \text{or} \quad \frac{[B(aq)]}{[BH^+(aq)]}$$

Investigation 8.8.2

The Effectiveness of an Antacid (page 573)

This investigation involves a back-titration procedure to determine how much acid is neutralized by a crushed antacid tablet. You may compare several brands of tablets.

As long as this ratio remains virtually constant, the pH will remain virtually constant. This will be the case as long as the amounts of the buffering entities (HA(aq) and A^-(aq) or B(aq) and BH^+(aq)) present are large compared with the amounts of H^+(aq) or OH^-(aq) added.

When solving problems involving the addition of a strong acid or base to a buffer, deal first with the stoichiometry of the resulting reaction. A strong acid will dissociate completely and react with the chemical entities present in solution. After the stoichiometric calculations are completed, then consider the equilibrium calculations (**Figure 5**).

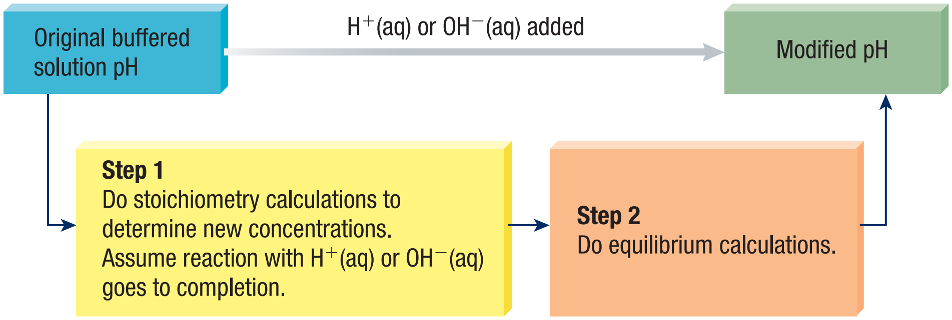

Figure 5 The procedure for solving calculations involving buffers

The Capacity of a Buffer

The **buffering capacity** of a buffer solution represents the amount of added hydrogen ions or hydroxide ions that the buffer can absorb without a significant change in pH (**Figure 6**). A buffer with a large capacity contains large concentrations of buffering components and so can absorb a relatively large amount of hydrogen ions or hydroxide ions and show little pH change. The pH of a buffered solution is determined by the [HA(aq)]/[A^-(aq)] ratio. The capacity of a buffered solution is determined by the magnitudes of [HA(aq)] and [A^-(aq)].

buffering capacity the ability of a buffer to resist changes in pH by reacting with added hydrogen ions or hydroxide ions

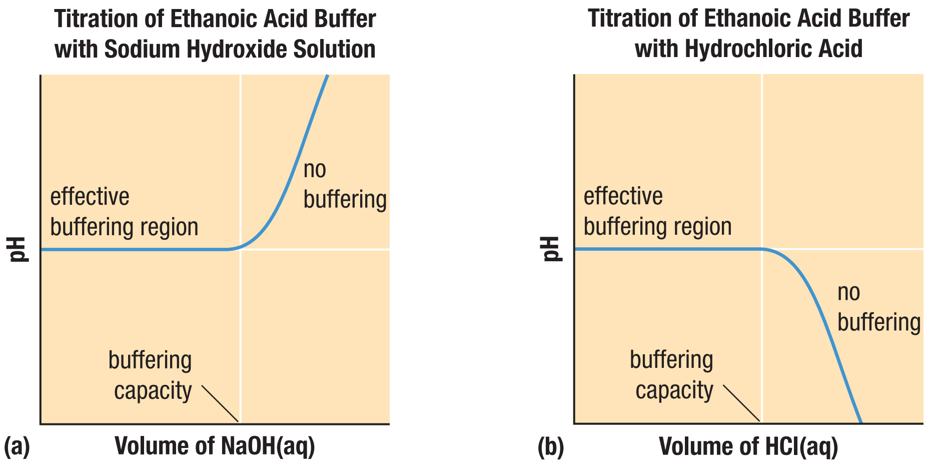

Figure 6 (a) Ethanoic acid buffer with a strong base added (b) Ethanoic acid buffer with a strong acid added. The pH changes quickly once all of the available buffer is depleted.

8.8 Review

Summary

- Buffers contain a weak acid, HA(aq), and a salt of its conjugate base, A^-(aq), or a weak base, B(aq), and a salt of its conjugate acid, BH^+(aq).
- Buffers play an important role in many biological and industrial processes.
- When an acid or a base is added to a buffer, the system resists change in pH by removing OH^-(aq) or H^+(aq) ions from solution until its buffering capacity is exceeded.
- The buffering capacity of a solution containing HA(aq) and A^-(aq) depends on the ratio of [HA(aq)] to [A^-(aq)] (**Figure 7**). Buffering is most efficient when the ratio is close to 1.

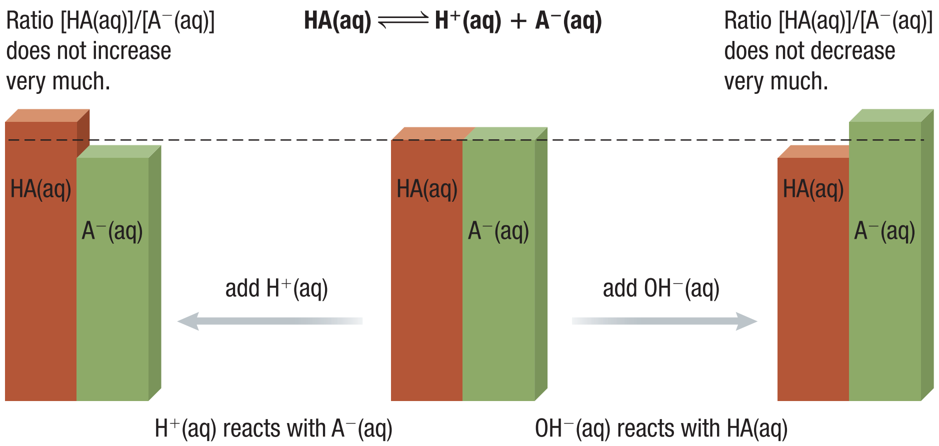

Figure 7 The effect of adding a strong acid or base to a buffer system

Questions

1. Describe the function of a buffer. K/U
2. What compounds make up a buffer? K/U
3. Write the equilibrium reaction equation for each of the following weak acids or bases. Name the conjugate base or acid in each case. K/U
   (a) phosphoric acid, H_3PO_4(aq)
   (b) carbonic acid, H_2CO_3(aq)
   (c) hydrogen sulfite ion, HSO_3^-(aq)
   (d) hypochlorite ion, ClO^-(aq)
4. For each acid or base listed in Question 3, suggest a salt that could be added to form a buffer. K/U
5. How will the equilibrium shift if a small amount of strong acid is added to a basic buffer solution? Write an example of this reaction equation. K/U
6. Explain what occurs in the buffer zone. K/U
7. A buffer contains 2.63 mol/L ethanoic acid, $HC_2H_3O_2$(aq) ($K_a = 1.8 \times 10^{-5}$), and 0.50 mol/L sodium ethanoate, $NaC_2H_3O_2$(aq). Calculate the pH of this solution. K/U
8. Calculate the pH of a buffer with 0.15 mol/L ammonia solution, NH_3(aq) ($K_b = 1.8 \times 10^{-5}$), and 0.04 mol/L ammonium chloride, NH_4Cl(aq). K/U
9. Research two buffer systems in the human body. Present your findings in a format of your choice. Your report should
   - explain why these buffers are important
   - include the equilibrium reaction equations
   - describe how each buffer responds to the addition of hydrogen or hydroxide ions

   T/I C A

WEB LINK