CHAPTER

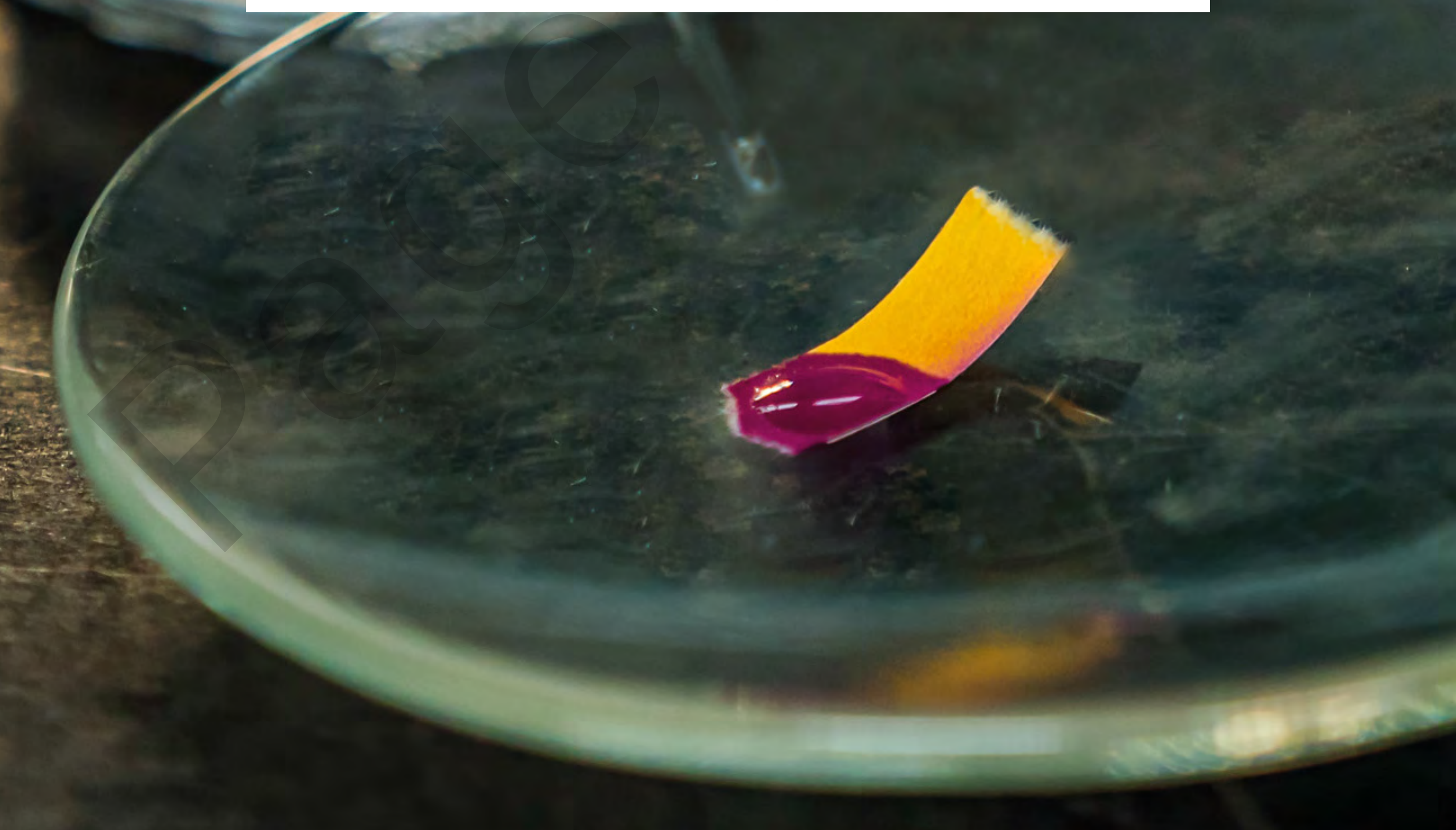
05 Buffers

Buffers are important in environmental and living systems. They have many uses. In medicines, buffers are used to maintain the stability and effectiveness of the medicine. In food, they preserve flavour and colour. Buffers are used during the fermentation of wine and in stages of textile dyeing. Buffers have a role in the maintenance and proper functioning of delicate chemical processes that can only occur within a narrow pH range. For example, there are buffer systems that ensure the pH of your blood is tightly maintained between pH 7.35 and pH 7.45.

In this chapter, you will study what a buffer is, how buffers are made and how they work. You will also examine the application of buffers as dynamic equilibrium systems and predict how they respond to pH changes according to Le Châtelier's principle.

Science understanding

- buffer solutions are conjugate in nature and resist changes in pH when small amounts of strong acid or base are added to the solution; buffering capacity can be explained qualitatively; Le Châtelier's principle can be applied to predict how buffers respond to the addition of hydrogen ions and hydroxide ions
- the characteristics of a system in dynamic equilibrium can be described and explained in terms of reaction rates and macroscopic properties



5.1 Introducing buffers

You might have come across buffers around your house, particularly if you have a swimming pool, spa or aquarium. You might also have used buffer solutions in your chemistry lessons to calibrate a pH meter (Figure 5.1.1). pH is probably the most commonly measured quantity in laboratories because pH affects many chemical and biochemical reactions. Buffers are used where it is important to maintain a solution at a known pH.

In this section, buffer solutions will be defined and you will be introduced to the preparation of buffer solutions.

FIGURE 5.1.1 A pH meter is calibrated by measuring a series of reference standards, known as pH buffers, that have known and accurate pH values.

BUFFER SOLUTIONS

Buffer solutions are able to resist a change in pH when small amounts of acid or base are added. This is important for processes that require stable and narrow pH ranges.

A buffer consists of a weak **conjugate acid–base pair**. This means that a buffer could be made of a:

- weak acid and its conjugate base
- weak base and its conjugate acid.

A **weak acid** donates hydrogen ions (protons) to a base to a limited extent. The acid's **conjugate base** contains one less hydrogen ion (proton) than the acid.

A **weak base** accepts hydrogen ions (protons) from acids to a limited extent. The base's **conjugate acid** contains one more hydrogen ion (proton) than the base.

The conjugate acid–base pair chosen for preparing a buffer helps determine the pH of the buffer.

> Buffer solutions can resist a change in pH when small amounts of acid or base are added.

BUFFER SOLUTIONS

Buffers can be made from a weak acid and one of its salts. For example, a mixture of ethanoic (acetic) acid and sodium ethanoate (acetate) solution produces the CH_3COOH/CH_3COO^- buffer. If the solution contains equal molar concentrations of ethanoic acid and sodium ethanoate salt solution, it would produce a buffer solution of pH 4.76.

It does not matter what the concentrations for the ethanoic acid and sodium ethanoate solutions are—as long as they are the same, the pH will always be 4.76. To change the pH of this buffer solution, you need to change the ratio of acid to salt.

The equilibrium system that acts as a buffer in this situation is:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

When making the buffer solution, the aim is to produce a solution that contains a high enough concentration of ethanoic acid molecules and ethanoate ions to enable the system to shift in either direction.

The ethanoic acid/sodium ethanoate buffer is prepared by dissolving approximately equal molar amounts of ethanoic acid (CH_3COOH) and sodium ethanoate (CH_3COONa) in water.

The sodium ethanoate dissociates completely in water, producing ethanoate ions and sodium ions:

$$CH_3COONa(s) \rightarrow CH_3COO^-(aq) + Na^+(aq)$$

At the same time, only a small proportion of ethanoic acid molecules will ionise, producing ethanoate (CH_3COO^-) and hydronium (H_3O^+) ions. The vast majority of the ethanoic acid molecules remain as molecules.

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

The resulting solution is an equilibrium mixture of CH_3COOH molecules, CH_3COO^- ions and H_3O^+ ions. Sodium ions are spectator ions and are not involved in the equilibrium:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq) \qquad K_a = 1.7\times10^{-5}$$

Therefore the buffer solution contains a significant amount of both the weak acid, CH_3COOH, and its conjugate base, CH_3COO^-. This equilibrium system can resist changes in pH when a small amount of an acid or base is added.

Buffers can be made from a weak base and one of its salts. For example, a mixture of ammonia solution and ammonium chloride solution produces the NH_3/NH_4^+ buffer. If equal molar concentrations of weak base and its salt are mixed, the buffer solution will have pH 9.25.

An NH_3/NH_4^+ buffer can be prepared by dissolving approximately equal molar amounts of ammonia solution, and an ammonium salt, such as ammonium chloride (NH_4Cl) in water.

The ammonium chloride dissociates completely in water, producing ammonium ions and chloride ions:

$$NH_4Cl(s) \rightarrow NH_4^+(aq) + Cl^-(aq)$$

At the same time, only a small proportion of ammonia molecules will ionise, producing ammonium ions (NH_4^+) and hydroxide ions (OH^-):

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

The resulting solution is an equilibrium mixture of NH_3 molecules, OH^- ions and NH_4^+ ions. Chloride ions are spectator ions and are not involved in the equilibrium. The equation for the equilibrium system is:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \qquad K_a = 5.6 \times10^{-10}$$

The buffer solution contains a significant amount of both the weak base, NH_3, and its conjugate acid, NH_4^+. This equilibrium system can resist changes in pH when a small amount of an acid or base is added.

5.1 Review

SUMMARY

- Buffer solutions resist changes in pH when small amounts of acid or base are added.
- A buffer consists of a conjugate acid–base pair (with a weak base or weak acid).
- Buffer solutions that are acidic are usually made by mixing approximately equal molar concentrations a weak acid and one of its salts.
- Buffer solutions that are alkaline are usually made by mixing approximately equal molar concentrations a weak base and one of its salts.

KEY QUESTIONS

1 Which one of the following substances could be used to maintain the pH range of water in an aquarium?

A a solution of a weak acid

B a solution of a strong acid

C a solution of a strong acid and its conjugate base

D a solution of a weak acid and its conjugate base

2 Write the equation for the equilibrium that exists in the ammonia/ammonium chloride buffer.

3 Write the equation for the equilibrium that exists in the hydrofluoric acid/fluoride buffer.

4 Write the equation for the equilibrium that exists in the hydrogencarbonate/carbonate buffer.

5 Classify the three buffer solutions identified in questions 2, 3 and 4 as either acidic or alkaline buffers (assume the components are present in roughly equimolar concentrations).

5.2 How buffers work

When 0.01 mol of hydrochloric acid is added to 1.0 L of water, the pH falls from 7.0 to 2.0, a drop of five pH units (Figure 5.2.1). The hydronium ion concentration in the solution has increased 100 000 times!

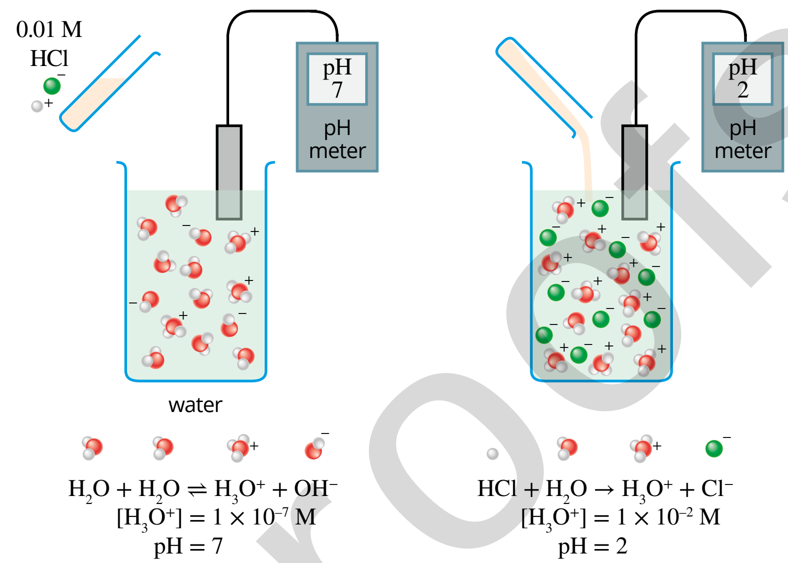


FIGURE 5.2.1 A significant change in pH occurs when a small amount of hydrochloric acid is added to 1.0 L of water.

When the same amount of hydrochloric acid is added to a particular buffer solution initially at pH 7, the pH might only drop to 6.9, a decrease by only 0.1 pH unit.

In this section, you will investigate how buffers work to maintain a stable pH range.

ADDING AN ACID TO A BUFFER

If a small amount of a strong acid, such as HCl, is added to an ethanoic acid/ethanoate buffer (CH_3COOH/CH_3COO^-), the pH will decrease to a small extent, but not as much as if the buffer were not present. For example, if you add HCl to a buffer solution, the extra H_3O^+ ions disturb the existing equilibrium of the buffer solution. **Le Châtelier's principle** tells you that the system will respond to oppose the change and restore equilibrium. Therefore, some of the additional H_3O^+ will react with CH_3COO^- ions present in the buffer. This produces more CH_3COOH molecules, decreasing the concentration of H_3O^+ ions in solution.

The equilibrium in a buffer system is represented in Figures 5.2.2. Note that when acid is added, the conjugate base (A^-) combines with any H_3O^+ that is added to maintain a stable pH.

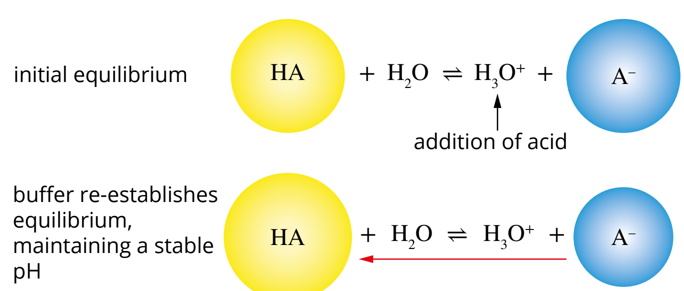


FIGURE 5.2.2 A constant pH is maintained by the equilibrium between a weak acid, HA, and its conjugate base, A^-, when acid is added because A^-(aq) combines with the added H_3O^+.

ADDING A BASE TO A BUFFER

When a strong base is added to the CH_3COOH/CH_3COO^- buffer, the OH^- ions from the base react with CH_3COOH molecules in the buffer. This generates more CH_3COO^- ions and consumes most of the extra OH^- ions:

$$CH_3COOH(aq) + OH^-(aq) \rightleftharpoons CH_3COO^-(aq) + H_2O(l)$$

The overall effect of adding a small amount of an acid or a base to a buffer solution on the final concentration of H_3O^+ or OH^- in the solution is small. Therefore, the overall change in pH is also very small.

The effect of adding an acid or base to a buffer system made from a weak acid with the general formula, HA, is shown in Figure 5.2.3.

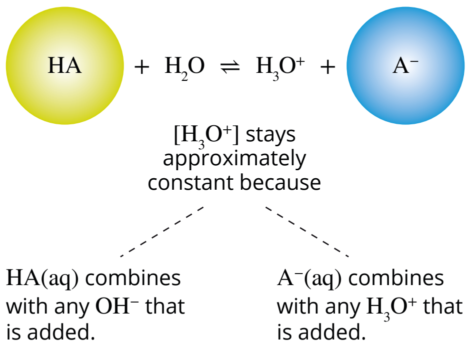


FIGURE 5.2.3 A constant pH is maintained by the equilibrium between a weak acid, HA, and its conjugate base, A^-.

BUFFER CAPACITY

Buffer solutions have a working pH range and capacity that determines how much acid or base can be neutralised before the pH changes, and by how much the pH will change. It is the concentrations of the buffer, weak acid (HA) and its conjugate base (A^-) that influence how effective the buffer will be in resisting changes in pH. A buffer is most effective when more HA molecules and A^- are available to react with and neutralise the effect of the addition of a strong acid or base, so that the pH of the system only changes a small amount.

> The more A^- and HA molecules available, the less of an effect adding a strong acid or base will have on the pH of a system.

If HCl is added to an ethanoic acid/ethanoate buffer, the H_3O^+ ions from the HCl will react with CH_3COO^- ions to form CH_3COOH. As more HCl is added, a new stage is reached where most of the CH_3COO^- ions have reacted. At this point, the buffer system is no longer effective. Additional H_3O^+ ions can no longer be removed by reaction with CH_3COO^- ions. From this point onwards, there will be a sharp decrease in pH as more HCl is added. Figure 5.2.4 shows how the pH of a 1.0 L buffer solution (0.1 M CH_3COOH and 0.1 M CH_3COO^-) changes as HCl is added. A buffer can only resist changes in pH if some of each of its conjugate acid and base pair is present. If they are consumed, for example by the addition of a large amount of strong acid or base, the solution is no longer a buffer.

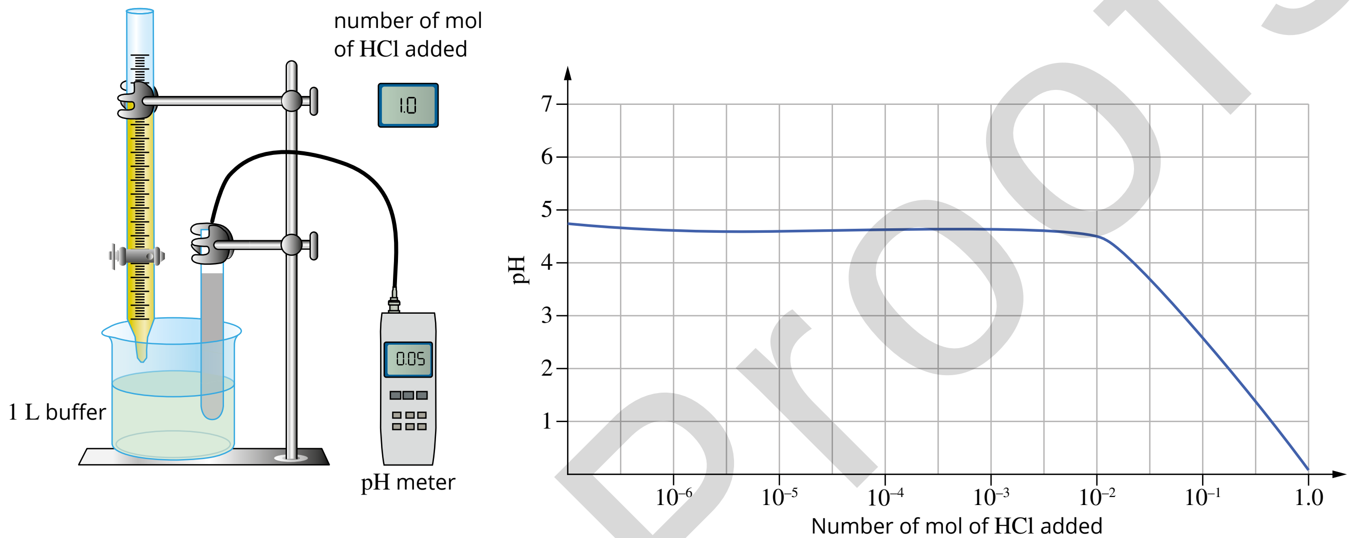
Buffer capacity is a measure of the effectiveness of a buffer solution at resisting a change in pH when either a strong acid or a strong base is added.

Buffer capacity is greatest when:

- there is a high concentration of the weak acid and its conjugate base
- the concentrations of the acid and its conjugate base are equal.

In the ethanoic acid/ethanoate buffer, a solution that has a high concentration of ethanoic acid can react with a greater amount of base. Similarly, a buffer that has a high concentration of ethanoate ions can react with a greater amount of added acid, H_3O^+. In general, for a buffer solution to be effective at maintaining pH, the ratio of the concentrations of the weak acid and its conjugate base should be within the range of 10:1 and 1:10.

Buffer capacity is a measure of the effectiveness of a buffer solution to resist a change in pH when either a strong acid or strong base is added.



Number of mol of HCl added	0	10^{-6}	10^{-5}	10^{-4}	10^{-3}	10^{-2}	10^{-1}	1.0
pH	4.74	4.7	4.7	4.7	4.7	4.6	2.7	0.05

FIGURE 5.2.4 The pH change of a 1.0 L buffer solution (0.1 mol L^{-1} CH_3COOH and 0.1 mol L^{-1} CH_3COO^-) when HCl is added

5.2 Review

SUMMARY

- When a buffer is present, the pH of a solution will change when a strong acid or base is added, but only to a small degree.
- Le Châtelier's principle can be used to predict the change in equilibrium of a buffer solution when an acid or a base is added.
- When acid is added to a buffer solution, the conjugate base combines with any H_3O^+ that is added to maintain a stable pH.
- When base is added to a buffer solution, the weak acid combines with any OH^- that is added to maintain a stable pH.
- Buffer capacity is a measure of the effectiveness of a buffer solution to resist a change in pH when either a strong acid or a strong base is added.
- Buffer capacity is greatest when:
  - there is a high concentration of the weak acid and its conjugate base
  - the concentrations of the acid and its conjugate base are equal.

KEY QUESTIONS

1 A small amount of sodium hydroxide solution is added to the ammonia/ammonium chloride buffer. Which one of the following statements is true?

A The OH^- ions react with NH_4^+ ions and the pH decreases significantly.

B The OH^- ions react with NH_3 molecules and the pH decreases significantly.

C The OH^- ions react with NH_4^+ ions and the pH remains almost constant.

D The OH^- ions react with NH_3 molecules and the pH remains almost constant.

2 The chemical equation that describes the CH_3COOH/CH_3COO^- buffer system is:

$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$

Complete the following paragraph, choosing from the words and species listed:

right, left, small, large, H_3O^+, CH_3COOH, CH_3COO^-, OH^-, H_2O

The addition of a small amount of HCl to the ethanoic acid buffer system will disturb the equilibrium. As equilibrium is re-established, the added H_3O^+ reacts with ________. The position of the equilibrium in the equation given above shifts to the ________. Because the buffer contains a relatively large amount of CH_3COO^-, most of the added __________ is consumed and there is a _________ change in the concentration of H_3O^+.

When a small amount of base is added to the buffer system the OH^- reacts with _________. The excess OH^- is consumed without any large change in the concentration of _________.

3 The compositions of four solutions each containing 100 mL of a mixture of lactic acid (a weak acid), and sodium lactate (a weak base) are given below. Identify the solution with the greatest buffer capacity and explain your choice.

A 0.1 mol L^{-1} lactic acid and 0.1 mol L^{-1} sodium lactate

B 0.1 mol L^{-1} lactic acid and 0.01 mol L^{-1} sodium lactate

C 0.5 mol L^{-1} lactic acid and 0.5 mol L^{-1} sodium lactate

D 0.05 mol L^{-1} lactic acid and 0.25 mol L^{-1} sodium lactate

5.3 Applications of buffers

Buffers are used in a variety of fields, including analytical chemistry, biochemical systems, agriculture and industry.

In analytical chemistry they have applications in both qualitative and quantitative analysis. Buffers are used to control the concentration of ions such as phosphate, oxalate, borate and fluoride that can interfere with reactions. Section 5.1 noted the importance of the accurate measurement of pH and that buffer solutions of known pH values (usually pH 4, 7 and 10) are used to standardise pH meters.

This section investigates some specific buffers and their applications in biochemical systems, agriculture and the food and paper industries.

BIOCHEMICAL SYSTEMS

Biological fluids have characteristic pH values. Cell fluids, both inside and outside of all cells, have water as the solvent. Like all aqueous solutions, cell fluids contain the dissociation products of H_2O: hydrogen ions (H^+) and hydroxide ions (OH^-). Therefore, pH plays a very significant role in biochemical reactions. A small change of just 0.2 pH units can cause death, while certain enzymes are only activated at specific pH values.

Many reactions that occur in the human body involve acid–base reactions. Buffers provide a means of controlling the amount of H_3O^+ or OH^- ions present in cells and tissues and are vital. Without a buffer, the pH of body fluids could fluctuate from extremely basic to extremely acidic, making it impossible for biological systems, including those that involve enzymes to function.

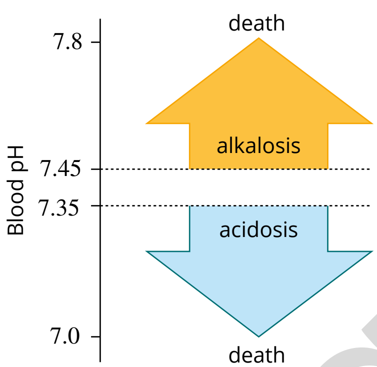


FIGURE 5.3.1 Physiological consequences of changes in blood pH

For example, blood is maintained within a narrow pH range of 7.35–7.45. Diseases such as pneumonia, emphysema and diabetes can cause pH to drop to potentially lethal levels, a condition called **acidosis**. On the other hand, hyperventilation, caused by rapid breathing, increases pH and causes **alkalosis**. Figure 5.3.1 shows a graphical representation of the effect on blood pH on the human body.

The presence of buffers maintains pH values within narrow limits in the body. Control of pH in the blood is achieved by different buffers. One important buffer is made up of carbonic acid (H_2CO_3) and the hydrogencarbonate ion (HCO_3^-):

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$$

How the carbonic acid buffer in blood works

Respiration in the cells produces carbon dioxide. The carbon dioxide dissolves and decreases blood pH according to the reaction:

$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$$

The buffers in the blood act to reduce the effect of additional acid (H_3O^+). According to Le Châtelier's principle, the system responds to oppose the change and restore equilibrium as some of the additional H_3O^+ produced by the carbon dixide will react with CH_3COO^- ions present in the buffer. This produces more CH_3COOH molecules, decreasing the overall concentration of H_3O^+ ions in solution. You can refer to the general equation showing the equilibrium in a buffer system in Figure 5.2.3.

Phosphate buffer system

In the internal fluid of all cells another important buffer system is at work, the **phosphate buffer** system. This buffer system consists of dihydrogenphosphate ions ($H_2PO_4^-$) as a weak acid and hydrogenphosphate ions (HPO_4^{2-}) as the conjugate base of the weak acid. These two ions are in equilibrium with each other:

$$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{2-}(aq)$$

If additional hydrogen ions enter the cellular fluid, they are consumed in the reverse reaction. The HPO_4^{2-} reacts with the extra hydrogen ions and the equilibrium shifts to the left. If additional hydroxide ions enter the cellular fluid, they react with $H_2PO_4^-$, producing more H_3O^+ and HPO_4^{2-}. This shifts the equilibrium to the right to stabilise the pH. If the phosphate buffer in cell fluid were absent, sharp changes in pH of cell fluids would occur, causing cell death or different proteins and cell organelles to malfunction.

AGRICULTURE

The pH of the soil is very important. Soil pH affects all soil chemistry and nutrient reactions. It is usually the first consideration when testing and evaluating a soil and when maximising crop yield and choosing a fertiliser.

Most plants perform best in slightly acid soils at pH 6.0–7.0. Some plants, such as strawberries (Figure 5.3.2), blueberries and citrus grow better in more acid soils (pH 4.0–6.5). Other plants, such as plums, sage and sunflowers tolerate a slightly alkaline soil (pH 7.0–7.5). Small changes in a soil pH can have large effects on nutrient availability and plant growth. Soils are naturally buffered by the presence of carbonates, bicarbonates and phosphates salts as well as organic acids.

Soil acidification, a decrease in soil pH, is a natural process that occurs very slowly as soil is weathered. Unfortunately, this process is accelerated by agricultural practices such as the use of nitrogen fertilisers. The chemical and biological properties of soil change as it becomes more acidic. One chemical change is an increase in the solubility of aluminum (Al) and manganese (Mn), which can be toxic to plants.

FIGURE 5.3.2 Strawberry plants grow best in more acidic soils at pH 4.0–6.5.

When you measure the soil pH, you are measuring the concentration of hydrogen ions in the soil solution only. But hydrogen ions in soil are present both in the soil solution and adsorbed onto the soil surfaces. To counteract an increase in soil acidity, lime (calcium oxide) or limestone (calcium carbonate) can be added to the soil. By adding lime, hydrogen ions in the soil solution are neutralised. Hydrogen ions from the soil surfaces are then released into the soil solution. This buffering effect happens to maintain equilibrium and resist the change in pH. Well-buffered soils are slower to acidify. But once they are acidic, it takes more lime to increase pH. The Department of Agriculture and Food Western Australia (DAFWA) has identified that more than 40% of the agricultural land in south-west Western Australia consists of poorly buffered sandy soil types.

THE FOOD INDUSTRY

Buffers in food

The colour, flavor, texture and nutritional value of many foods can be altered by food additives. Some food additives help preserve the food. Food additives are classified by codes called **E numbers**. E numbers between E330 and E340 represent a group of antioxidants that can also act as buffers. Citric acid (E330) is added to control pH (Figure 5.3.3).

By keeping the acidity of a food at the appropriate level, flavour and appearance can be maintained. A stable pH is also critical to the preservation of many processed foods. Food additives that act as buffers usually consist of metal salts with a weak acid found naturally within the food to be preserved. For example, sodium citrate could be added to a food containing citric acid to create a buffer solution. Calcium and potassium citrate could also be used.

Citrate additives are widely used antioxidants and buffers with a range of applications. All citrate additives can reduce the chemical reaction that causes the discoloration of fruit and are often added to fruit products. Citrate additives include:

- E331, sodium citrate, a buffer mainly used in jams and jellies
- E332, potassium citrate, an antioxidant and buffer often found in cakes, biscuits and cheese
- E333, calcium citrate, an important buffer in carbonated drinks.

FIGURE 5.3.3 Citric acid, E330, is a food additive that acts as a buffer (acid regulator).

Potassium tartrate (E336) is obtained from grapes during the wine-making process. In addition to its buffering action, potassium tartrate helps bread to rise consistently.

Monopotassium phosphate (potassium dihydrogenphosphate) (E340), principally used as an antioxidant, also has buffering capabilities. It is used in products such as custard, milk powder and jelly mixes, and may be added to cooked meat. Monopotassium phosphate is an important ingredient in many sports drinks, acting as a buffer and providing the potassium electrolyte.

The presence of buffers maintains pH values within a narrow range in many foods. For example, control of pH in some jams can be achieved with citric acid. The formula for citric acid is shown in Figure 5.3.4.

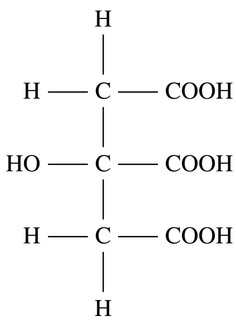


FIGURE 5.3.4 The structural formula of citric acid has three carboxylic acid groups (COOH) and can also be written as $H_6C_8O_7$. This shows that citric acid is triprotic, capable of releasing three hydrogen ions per molecule of citric acid.

The citric acid buffer can be represented by the chemical equation:

$$H_6C_8O_7(aq) + H_2O(l) \rightleftharpoons H_5C_8O_7^-(aq) + H_3O^+(aq)$$

The citric acid buffer in a food such as jam acts to reduce the effect of additional acid (H_3O^+). According to Le Châtelier's principle, the system responds to oppose the change and restore equilibrium. In this case, the H_3O^+ produced by the fermentation of sugar in the jam will react with $H_5C_8O_7^-$ ions present in the buffer. This produces more citric acid molecules ($H_6C_8O_7$) and decreases the concentration of H_3O^+ ions, keeping the pH stable. Refer to the general equation in Figure 5.2.3 to see how the equilibrium in a buffer system works.

THE PAPER INDUSTRY

Nearly all industries use buffers in one process or the other. Major industries that employ the use of buffers include the paper, dye, ink, paint and drug industries.

Carbon dioxide is widely used as a buffer in the paper industry to adjust and stabilise pH. To buffer the papermaking system, to decrease calcium levels, and to increase dewatering which is needed to help with pulp brightness and in the production of acid free paper. Enzymes are often introduced into the paper bleaching process (Figure 5.3.5) to achieve a higher level of pulp brightness while using fewer bleaching chemicals. These enzymes are most effective when used in a narrow, neutral pH range.

FIGURE 5.3.5 (a) Unbleached paper pulp. (b) Bleached paper pulp. The paper bleaching process often uses enzymes to achieve a higher level of pulp brightness and to use less bleaching chemicals.

The production of acid free paper, is often buffered with the carbonic acid/hydrogencarbonate buffer to prevent the formation of additional acids that can occur during various stages of production.

$$CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)$$

$$H_2CO_3(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + H_3O^+(aq)$$

5.3 Review

SUMMARY

- Biological fluids have characteristic pH values.
- Biological buffer systems include the carbonic acid buffer system and the phosphate buffer system.
- Hydrogen ions in soil can be buffered by limestone (calcium carbonate).
- E numbers are used to classify food additives, with E330–E340 representing additives that can act as buffers.
- Carbon dioxide is widely used in the paper industry as a buffer.
- Le Châtelier's principle can be used to describe the response of a buffer system to external changes.

KEY QUESTIONS

1 List the buffers described in this section.

2 The carbonic acid buffer involving the weak acid/weak conjugate base pair H_2CO_3/HCO_3^- helps control the pH in the blood and in the manufacture of paper.

  a Write the equation for the equilibrium that exists in the H_2CO_3/HCO_3^- buffer.

  b Describe how the buffer works to stabilise the pH when acid is added.

3 The phosphate buffer system is described by the chemical equation: $H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{2-}(aq)$

Will the concentration of $H_2PO_4^-$ increase or decrease if the system is subjected to a more alkaline environment?

4 Food additives that act as buffers usually consist of metal salts with a weak acid found naturally within the food to be preserved. Write the equation for the sodium citrate buffer ($H_5C_8O_7^-/H_6C_8O_7$).

Chapter review

05

KEY TERMS

acidosis	conjugate acid	Le Châtelier's principle
alkalosis	conjugate acid–base pair	phosphate buffer
buffer capacity	conjugate base	weak acid
buffer solution	E number	weak base

Introducing buffers

1 Define a buffer solution.

2 A buffer consists of a weak conjugate acid–base pair. Define a conjugate acid–base pair.

3 Give two examples of buffer solutions.

4 Which of the two pairs of substances could be used to make a buffer system? Explain.
A: NaOH and H_2O
B: CH_3COONa and CH_3COOH

5 The phosphate buffer system ($H_2PO_4^-/HPO_4^{2-}$) is described by the chemical equation:
$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{2-}(aq)$
Identify the weak conjugate acid—base pair in the phosphate buffer system.

How buffers work

6 The phosphate buffer system is described by the chemical equation:
$H_2PO_4^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + HPO_4^{2-}(aq)$
To predict the effect of adding H_3O^+ to the $H_2PO_4^-/HPO_4^{2-}$ buffer system, complete the paragraph by filling in the blanks, using the words and species listed:
right, left, small, large, H_3O^+, $H_2PO_4^-$, HPO_4^{2-}, OH^-, H_2O
The addition of a small amount of HCl to the phosphate buffer system will disturb the equilibrium. As equilibrium is re-established, the added H_3O^+ reacts with ________. The position of the equilibrium in the equation given above shifts to the ________. Since the buffer contains a relatively large amount of HPO_4^{2-}, most of the added __________ is consumed and there is a _________ change in the concentration of H_3O^+.
When a small amount of base is added to the buffer system the OH^- reacts with _________. The excess OH^- is consumed without any large change in the concentration of _________.

7 Define buffer capacity.

8 Which one of the following statements is true?
A If a buffer is present, the pH of a solution will not change at all when a strong acid or base is added
B When base is added to a buffer solution, the weak acid combines with any OH^- that is added to lower the pH of the buffer solution.
C Buffer capacity is greatest when there is a low concentration of the weak acid and its conjugate base.
D Buffer capacity is greatest when the concentrations of the acid and its conjugate base are equal.

Applications of buffers

9 Name two buffers:
a essential in maintaining blood and cell pH
b used in the paper and food industry.

10 What are E numbers and which E numbers identify buffers used in the food industry?

11 Blood is buffered by the carbonic acid buffer, H_2CO_3/HCO_3^-. Write the equation for the buffer equilibrium. If the pH of blood increases, what happens to the concentrations of H_2CO_3, HCO_3^- and H^+?

Connecting the main ideas

12 Dihydrogenphosphate ($H_2PO_4^-$) and hydrogenphosphate (HPO_4^-) are important ions in intracellular fluids. They act as a buffer in the fluid inside cells. Using equations and Le Châtelier's principle, describe how these ions act to maintain a pH balance.

13 The concentration of carbonic acid (H_2CO_3) in blood is 1/20th of the concentration of the hydrogen carbonate ion (HCO_3^-), yet this blood buffer can buffer the pH of blood against bases as well as acids. Explain.

14 Distance running and other strenuous muscle activity can cause a buildup of lactic acid in the blood. You may have experienced the effect of lactic acid accumulation as a wobbly feeling in your leg muscles. How would the build-up of lactic acid in the blood affect blood pH? Using Le Châtelier's principle describe how the buffers in blood works to regulate this pH change.

15 Cardiac arrest occurs when the heart fails to effectively pump and circulate blood around the body. Cell reactions still continue.
a Explain the effect of cardiac arrest on blood pH.
b Doctors often injected sodium hydrogencarbonate solution ($NaHCO_3$) into the heart muscle. Explain how this injection counteracts the pH effect.

16 Citric acid, E330, is a food additive that acts as a buffer and helps to preserve food.

- **a** Write an equilibrium equation for the citric acid buffer. (Hint: citric acid can be written as $H_6C_8O_7$)
- **b** Identify the conjugate base of citric acid.
- **c** Describe how this buffer behaves to stabilise pH when acid is added.

17 Lakes with lake beds and surrounding rock and soil rich in limestone ($CaCO_3$) have a capacity to be able to buffer themselves against appreciable changes in pH. Using equations, explain how the water in these lakes behave as a buffer to neutralise the effects of acid rain. For this question assume the acid rain is mainly a dilute solution of sulfurous acid, H_2SO_3.

18 Figure 5.4.1 shows how HCO_3^- ions act as a buffer to prevent the oceans from becoming too acidic or alkaline.

- **a** If sea water is slightly alkaline, would you expect the concentration of dissolved CO_2 to be higher or lower than in pure water? Explain.
- **b** Describe and explain the impact of increasing atmospheric carbon dioxide concentration on coral reefs.

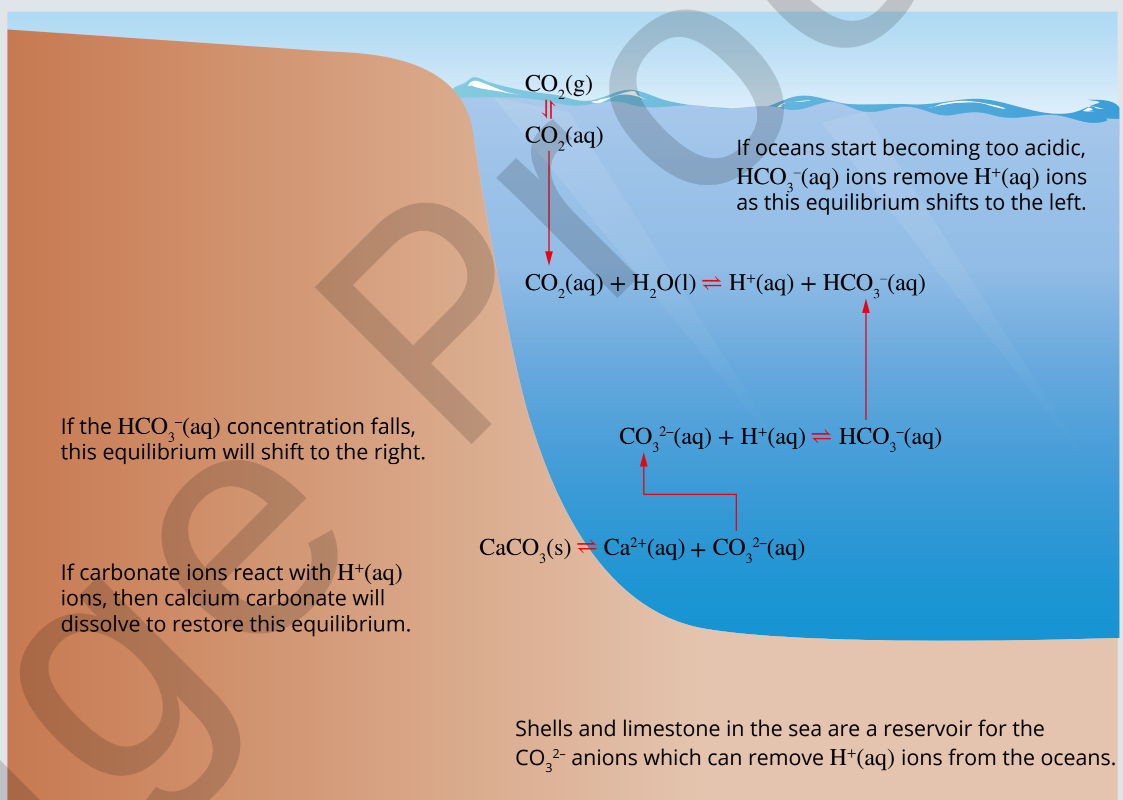


FIGURE 5.4.1 Buffering the oceans