9.2: Buffers

Learning Outcomes

- Define buffer.
- Define buffer capacity.
- Describe how a buffer controls pH.
- Identify the components of a buffer solution.

Diabetes mellitus is a disorder of glucose metabolism in which insulin production by the pancreas is impaired. Since insulin helps glucose enter the cells, a decrease of this hormone means that glucose cannot be used in its normal fashion. When this happens, the body begins to break down fats, producing a decrease in blood pH. Chemical systems in the body can balance this pH shift for a while, but excessive acid production can create serious problems if not corrected by administering insulin to restore normal glucose use.

Buffers

If only 1.0 mL of 0.10 M hydrochloric acid is added to 1.0 L of pure water the pH drops drastically from 7.0 to 4.0. This is a 1000-fold increase in the acidity of the solution. For many purposes, it is desirable to have a solution which is capable of resisting such large changes in pH when relatively small amounts of acid or base are added to them. Such a solution is called a buffer. A **buffer** is a solution of a weak acid or a base and its salt. Both components must be present for the system to act as a buffer to resist changes in pH. The salt is the conjugate of the weak acid or of the weak base. It can be shown as the ion or with the counter ion. $\left(\text{COO}^- \text{ or } \text{COONa} \right)$

Some common buffer systems are listed in the table below. Note that the two components of the buffer system differ by only one hydrogen ion $\left(\text{H}^+ \right)$.

Table 9.2.1: Some Common Buffers

Buffer system	Buffer components	pH of buffer (equal molarities of both components)
Acetic acid/acetate ion	$\text{CH}_3\text{COOH}/\text{CH}_3\text{COO}^-$	4.74
Carbonic acid/hydrogen carbonate ion	$\text{H}_2\text{CO}_3/\text{HCO}_3^-$	6.38
Dihydrogen phosphate ion/hydrogen phosphate ion	$\text{H}_2\text{PO}_4^-/\text{HPO}_4^{2-}$	7.21
Ammonia/ammonium ion	$\text{NH}_3/\text{NH}_4^+$	9.25

One example of a buffer is a solution made of acetic acid (the weak acid) and sodium acetate (a conjugate of the acid). The pH of a buffer consisting of $0.50 \text{ M } \text{CH}_3\text{COOH}$ and $0.50 \text{ M } \text{CH}_3\text{COONa}$ is 4.74. If 10.0 mL of $1.0 \text{ M } \text{HCl}$ is added to 1.0 L of the buffer, the pH only decreases to 4.73. This ability to "soak up" the additional hydrogen ions from the HCl that was added is due to the reaction below.

$$\text{CH}_3\text{COO}^- \left(aq \right) + \text{H}^+ \left(aq \right) \rightleftharpoons \text{CH}_3\text{COOH} \left(aq \right) \tag{9.2.1}$$

Since both the acetate ion and the acetic acid were already present in the buffer, the only thing that changes is the ratio of one to the other. Small changes in that ratio have only very minor effects on the pH.

If 10.0 mL of $1.0 \text{ M } \text{NaOH}$ were added to another 1.0 L of the same buffer, the pH would only increase to 4.76. In this case, the buffer takes up the additional hydroxide ions.

$$\text{CH}_3\text{COOH} \left(aq \right) + \text{OH}^- \left(aq \right) \rightleftharpoons \text{CH}_3\text{COO}^- \left(aq \right) + \text{H}_2\text{O} \left(l \right) \tag{9.2.2}$$

Again the ratio of acetate ion to acetic acid changes only slightly, this time causing a very small increase in the pH. All buffers follow Le Chatelier's principle and respond to a stress on the system by responding to minimize the stress.

It is possible to add so much acid or base to a buffer that its ability to resist a significant change in pH is overwhelmed. The **buffer capacity** is the amount of acid or base that can be added to a buffer solution before a large change in pH occurs. The buffer capacity is exceeded when the number of moles of H^+ or OH^- that are added to the buffer exceeds the number of moles of the buffer components.

Human blood contains a buffer of carbonic acid $\left(H_2CO_3\right)$ and bicarbonate anion $\left(HCO_3^-\right)$ in order to maintain blood pH between 7.35 and 7.45, as a value higher than 7.8 (alkalosis) or lower than 6.8 (acidosis) can lead to death. In this buffer, hydronium and bicarbonate anion are in equilibrium with carbonic acid. The bicarbonate neutralizes excess acids in the blood while the carbonic acid neutralizes excess bases.

Furthermore, carbonic acid can decompose into CO_2 gas and water, resulting in a second equilibrium system between carbonic acid and water. Because CO_2 is an important component of the blood buffer, its regulation in the body, as well as that of O_2, is extremely important. The effect of this can be important when the human body is subjected to strenuous conditions.

$$CO_2 + H_2O \rightleftharpoons H_2CO_3 \rightleftharpoons HCO_3^- + H^+ \tag{9.2.3}$$

Contributors and Attributions

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