Buffers

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1 Introduction

Buffers are solutions that can resist a change in pH when small amounts of acid or base are added to them. They are required in many biological systems such as the circulatory systems.

2 Theory

Acidic Buffers are made from a weak acid and its conjugate base:

$$
\begin{aligned}
\mathrm{HA} &\rightleftharpoons \mathrm{H^+ + A^-} \\
\mathrm{MA} &\longrightarrow \mathrm{M^+ + A^-}
\end{aligned}
$$

Upon adding hydroxide ions (OH^-), the concentration of hydroxonium ions (H^+) reduces, by L.C.P, the equilibrium will move to the right to resist this change. Because most of the acid has not dissociated, plenty of molecules can still break down to supply hydroxonium ions (H^+).

On adding hydroxonium ions (H^+), by L.C.P, the position of the equilibrium wants to move to the left to use up the extra hydroxonium ions (H^+). However we need the conjugate base (A^-) to react with the extra hydroxonium ions (H^+). The acid does not produce any conjugate base (A^-) to react with the extra hydroxonium ions (H^+) added. Therefore we need a secondary source of conjugate base (A^-), which comes from the salt.

Basic Buffers are made from a weak base and it conjugate acid:

$$
\begin{aligned}
\mathrm{B + H_2O} &\rightleftharpoons \mathrm{BH^+ + OH^-} \\
\mathrm{BHX} &\longrightarrow \mathrm{BH^+ + X^-}
\end{aligned}
$$

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The pH of Buffer Solutions:

$$K_a = \frac{[H^+][A^-]}{[HA]} \tag{1}$$

The value of K_a is a representation of how far the weak acid dissociates.With our knowledge of buffers, we can say that:

$$[HA]_{\text{equilibrium}} \simeq [HA]_{\text{initial}} \tag{2}$$

because very little of the acid dissociates. We can also say that:

$$[A^-]_{\text{equilibrium}} \simeq [A^-]_{\text{initial}} \tag{3}$$

because since there is a weak acid that does not dissociate significantly, only few extra conjugate base (A^-) at equilibrium are added. We can now say that:

$$K_a = \frac{[H^+][conc.base]}{[Acid]} \tag{4}$$

$$K_a \cdot [Acid] = [H^+][conc.base] \tag{5}$$

$$K_a \cdot \frac{[Acid]}{[Base]} = [H^+] \tag{6}$$

$$-log[H^+] = -log(K_a \cdot \frac{[Acid]}{[Base]}) \tag{7}$$

$$pH = -log(K_a) - log(\frac{[Acid]}{[Base]}) \tag{8}$$

$$pH = pK_a - log\frac{[Acid]}{[Base]} \tag{9}$$