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## Acids, Bases and Buffers

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### INTRODUCTION

Electrolytes when dissolved in water, split up into two or more electrically charged particles. In 1887, Arrhenius explained this property of electrolytes, and hence, it is known as **Arrhenius theory of electrolyte dissociation**.

The charged particles, which are formed when the electrolyte is dissolved in water, are called ions, and this process is called ionization. The process of ionization is reversible. These ions may have positive or negative charge. The total number of positive charges is equal to the total number of negative charges in the solution. Since they are charged, they are responsible for carrying the electric current through the solution.

The extent of ionization is given by the ratio of the total number of dissociated molecules to the number of molecules dissolved. The degree of ionization can be increased by diluting the solution in case of weak electrolytes.

### ACIDS AND BASES

Brønsted and Lowry defined acids as substances which are able to donate protons, and bases as substances which accept protons.



In the above example, BH is an acid because it donates proton and  $\text{B}^-$  is an anion liberated by the deprotonation of the acid.  $\text{B}^-$  behaves like a base, so it is called **conjugate base**.

Acids can be classified into strong acids and weak acids.

1. **Strong acids** get dissociated almost completely, e.g., hydrochloric acid, sulphuric acid. This is because the conjugate bases of these acids are very weak (have less affinity for the proton).
2. **Weak acids** get dissociated partially, e.g., acetic acid, carbonic acid. This happens because the conjugate bases of these acids are strong (have greater affinity for proton).

Since the dissociation of the weak acids is partial, the equilibrium constant for the dissociation reaction of the weak acid (BH) can be written as follows:



By applying the law of mass action at equilibrium

$$K_a = \frac{[\text{B}^-][\text{H}^+]}{[\text{BH}]}$$

Ka is dissociation constant

The equilibrium constants for ionization reactions are commonly called ionization or dissociation constant ( $K_a$ ).

The above equation can be rearranged as:

$$[\text{H}^+] = \frac{K_a \times [\text{BH}]}{[\text{B}^-]}$$

By multiplying the above equation by  $-1$  and taking logarithm of both sides, the following expression can be derived:

$$-[\text{H}^+] = - \left[ \frac{K_a \times [\text{BH}]}{[\text{B}^-]} \right]$$

$$-\log_{10} [\text{H}^+] = -\log_{10} K_a - \log_{10} \frac{[\text{BH}]}{[\text{B}^-]}$$

$$-\log_{10} [\text{H}^+] \text{ is pH}$$

As the hydrogen ion concentration increases, the pH of the solution decreases. As the hydrogen ion concentration decreases, pH will increase. Hydrogen ion concentration and pH are reciprocally related.

$$\text{pH} = \text{p}K_a + \log_{10} \frac{[\text{B}^-]}{[\text{BH}]}$$

$$\text{pH} = \text{p}K_a + \log_{10} \frac{[\text{Conjugate base}]}{[\text{Acid}]}$$

This expression is called **Henderson-Hasselbalch equation**. If the conjugate base and acid concentration is the same, then **pH = pKa**. The value of pKa is lower for strong acids and higher for weak acids.

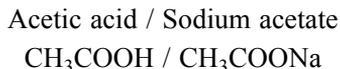
## BUFFERS

Buffers are solutions that **resist changes in the pH** of a when acid or alkali is added. The buffer system has an acid component and salt component. Buffers are prepared by mixing a weak acid and its salt with a strong base or by mixing weak base and its salt with a strong acid.

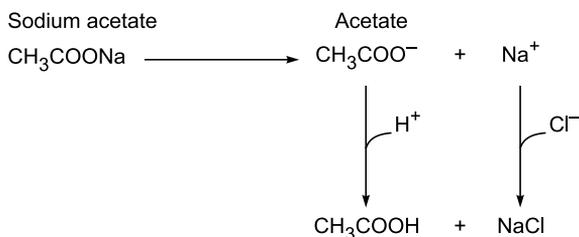
### Mechanism of Buffer Action

When a strong acid is added, protons are scavenged by the salt component of the buffer system. Similarly, when an alkali is added, acid component will get deprotonated and the alkali combines with the proton to form water.

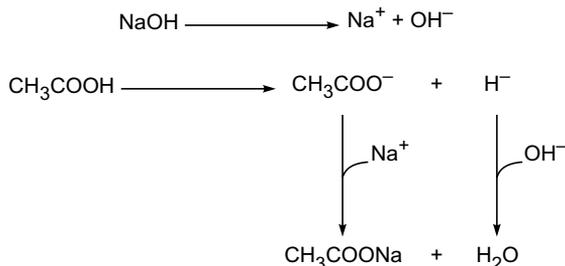
Acetate buffer is made up of acetic acid and sodium acetate.



- (a) When acid like HCl is added to the acetate buffer system,  $\text{CH}_3\text{COONa}$  (sodium acetate) gets converted to  $\text{CH}_3\text{COO}^-$  (acetate). Acetate combines with the protons released by the strong acids to give acetic acid, which is a weak acid, and the change in pH is very small.



- (b) When a strong alkali like NaOH is added,  $\text{OH}^-$  combines with the  $\text{H}^+$  released by the  $\text{CH}_3\text{COOH}$  to form water.  $\text{Na}^+$  combines with acetate to form sodium acetate.



According to Henderson-Hasselbalch equation, when the concentration of the salt component and the acid component of the buffer system is equal, pH of the solution equals pKa. At this pH, buffer can respond equally to both added acid as well as the alkali.

### Buffering Capacity

Capacity of the buffer to resist the change in the pH of a solution when an acid or alkali is added is called buffering capacity. It is estimated by calculating the **amount of acid or alkali required to change the pH of one litre of buffer by one unit.**

Buffering capacity depends upon the following factors:

**1. The concentration of the acid and base component of the buffer**

As the concentration of acid and base component of the buffer increases, buffering capacity of the buffer also increases.

**2. The pH of the Buffer**

Buffer can act best at  $\text{pH} = \text{pKa}$ , and its buffering range is about one pH unit above or below the pKa value.