

Acid-Base Chemistry & Buffer Systems in Pharmacy

Introduction: Acids, Bases, and Buffer

- Acids, bases, and buffers are fundamental concepts in chemistry and are essential for many **biological, industrial, and laboratory** processes.
- They play a vital role in maintaining the **acid-base balance (pH)** of living organisms and various chemical reactions.

Definition of Acids, Bases, and Buffers

1. Acids:

- **Acids** are substances that donate hydrogen ions (H^+) or protons when dissolved in water.
- They typically have a sour taste and can be corrosive to metals.
- Acids can be classified into two main categories:

A. Strong Acids:

- These acids ionize completely in water, releasing all their hydrogen ions.
- **Examples** include hydrochloric acid (HCl), sulfuric acid (H_2SO_4), and nitric acid (HNO_3).

B. Weak Acids:

- These acids ionize only partially in water, releasing only a small fraction of their hydrogen ions.
- **Examples** include acetic acid (CH_3COOH), carbonic acid (H_2CO_3), and phosphoric acid (H_3PO_4).

2. Bases:

- Bases are substances that accept hydrogen ions (H^+) or donate hydroxide ions (OH^-) when dissolved in water.
- They typically have a bitter taste, feel slippery, and can be corrosive to organic materials.
- Bases can also be classified into **two** main categories:

A. Strong Bases:

- These bases ionize completely in water, releasing all their hydroxide ions.
- **Examples** include sodium hydroxide (NaOH), potassium hydroxide (KOH), and calcium hydroxide ($Ca(OH)_2$).

B. Weak Bases:

- These bases ionize only partially in water, releasing only a small fraction of their hydroxide ions.
- **Examples** include ammonia (NH_3), methylamine (CH_3NH_2), and pyridine (C_5H_5N).

3. Buffers:

- Buffers are solutions that resist significant changes in pH when small amounts of acid or base are added.
- They usually consist of a weak acid and its conjugate base, or a weak base and its conjugate acid.
- Buffers are essential for maintaining a stable pH in biological systems, industrial processes, and analytical chemistry.

Example: Bicarbonate Buffer System

- The bicarbonate buffer system is the major buffer system in human blood and helps maintain the blood pH within the normal range.

Components

- **Carbonic acid (H_2CO_3)** – Weak acid
- **Bicarbonate ion (HCO_3^-)** – Conjugate base

Mechanism of Action

- **When an Acid is Added**
 - The added hydrogen ions (H^+) combine with bicarbonate ions (HCO_3^-) to form carbonic acid (H_2CO_3).
 - This reaction neutralizes the added acid and minimizes the decrease in pH.
 - **Reaction:**



- **When a Base is Added**
 - Carbonic acid (H_2CO_3) releases hydrogen ions (H^+).
 - The released H^+ reacts with hydroxide ions (OH^-) to form water (H_2O).
 - This reaction neutralizes the added base and minimizes the increase in pH.
 - **Reaction:**



pH Scale & Its Significance

Basic Idea

- The pH scale is a measure of the acidity or alkalinity of a solution.
- It indicates the concentration of hydrogen ions (H^+) present in a solution.
- The pH scale usually ranges from 0 to 14.

pH Scale

- A pH value of 7 indicates a neutral solution.
- A pH value below 7 indicates an acidic solution.
- A pH value above 7 indicates an alkaline or basic solution.

Classification of pH

1. Acidic Solution

- pH value ranges from 0 to less than 7.
- Hydrogen ion concentration is high.
- Example: Hydrochloric acid solution.

2. Neutral Solution

- pH value is 7.
- Hydrogen ion and hydroxide ion concentrations are equal.
- Example: Pure water.

3. Basic Solution

- pH value ranges from more than 7 to 14.
- Hydroxide ion concentration is high.
- Example: Sodium hydroxide solution.

Significance of pH in Pharmacy

1. Drug Stability

- pH affects the chemical stability of medicines.
- Incorrect pH can cause drug degradation and loss of effectiveness.

2. Drug Solubility

- pH influences the solubility of drugs in pharmaceutical formulations.
- Proper pH selection improves drug dissolution.

3. Patient Comfort

- The pH of injections and eye preparations must be controlled to reduce irritation.
- Products with unsuitable pH may cause pain or tissue damage.

4. Preservation of Products

- pH affects the activity of preservatives used in pharmaceutical preparations.
- Proper pH helps prevent microbial growth.

5. Biological Importance

- The human body maintains specific pH ranges for normal functioning.
- Blood pH is maintained around 7.4 by natural buffer systems.

Buffer Equation

Basic Idea

- The buffer equation describes the relationship between the pH of a buffer solution and the concentration of its components.
- It is used to calculate and control the pH of buffer systems.

Henderson-Hasselbalch Equation

For an acidic buffer:

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{Salt}]}{[\text{Acid}]}\right)$$

➤ or

$$\text{pH} = \text{p}K_a + \log \left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$$

➤ Where:

- **pH** = measure of acidity of the solution
- **pKa** = negative logarithm of the acid dissociation constant (K_a)
- **[A⁻]** = concentration of the conjugate base (usually provided by the salt)
- **[HA]** = concentration of the weak acid

For a basic buffer

➤ The equation is often written as:

$$\text{pOH} = \text{p}K_b + \log \left(\frac{[\text{Salt}]}{[\text{Base}]}\right)$$

➤ Then calculate:

$$\text{pH} = 14 - \text{pOH}$$

Derivation (Brief)

➤ For a weak acid:

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

➤ Rearranging:

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

- Taking the negative logarithm of both sides:

$$pH = pK_a + \log \left(\frac{[A^-]}{[HA]} \right)$$

- This is the Henderson–Hasselbalch equation.

Significance

- Calculates the pH of buffer solutions.
- Helps in preparing buffers of a desired pH.
- Explains how buffers resist changes in pH when small amounts of acid or base are added.
- Most accurate when the concentrations of the weak acid and its conjugate base are relatively high and their ratio is between about 0.1 and 10.

Example

- A buffer contains 0.20 M acetic acid and 0.30 M sodium acetate. Given pK_a of acetic acid = 4.76:

$$\begin{aligned} pH &= 4.76 + \log \left(\frac{0.30}{0.20} \right) \\ &= 4.76 + \log(1.5) \\ &= 4.76 + 0.18 = 4.94 \end{aligned}$$

- **Answer:** The pH of the buffer is 4.94.

Calculation of pH for Buffer Solutions

Basic Idea

- The pH of a buffer solution can be calculated using the Henderson-Hasselbalch equation.
- The calculation depends on the type of buffer and the concentration of its components.

Calculation of pH for Buffer Solutions

- A buffer usually contains:
 - a weak acid and its conjugate base, or
 - a weak base and its conjugate acid.

1. Acidic buffer

- For a weak acid HA and its salt A^- :



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