

# pH and Biological Systems

## pH Scale

- $\text{pH} = -\log[\text{H}^+]$
- Range: 0 (acidic) to 14 (basic)
- pH 7 is neutral
- Each unit = 10 $\times$  concentration change

## Buffer Systems

- Resist pH changes
- Blood pH: 7.35-7.45
- Bicarbonate buffer ( $\text{H}_2\text{CO}_3/\text{HCO}_3^-$ )
- Phosphate buffer in cells

## Henderson-Hasselbalch Equation

- $\text{pH} = \text{pK}_a + \log([\text{A}^-]/[\text{HA}])$
- Predicts buffer behavior
- Critical for enzyme function
- Used in drug design

## Enzyme pH Dependence

- Each enzyme has optimal pH
- Pepsin (stomach): pH 2
- Trypsin (intestine): pH 8
- pH affects protein charge state

# 1 pH Scale - Measure of Hydrogen Ion Concentration

## Definition and Concept

pH is defined as the negative logarithm of hydrogen ion ( $H^+$ ) concentration in a solution. This concept was introduced by Danish biochemist Søren Sørensen in 1909.

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

## Characteristics of pH Scale

- **Range:** 0 (strong acid) ~ 14 (strong base)
- **Neutral:** pH 7 (pure water)
- **Logarithmic scale:** Each pH unit change represents a 10-fold change in  $H^+$  concentration
- **Inverse relationship:** Lower pH means higher acidity

## Biological Significance

Living organisms function normally only within a very narrow pH range. Even small changes in pH can significantly affect protein structure, enzyme activity, membrane permeability, and more.

## Common Examples

- Gastric acid: pH 1.5-2.0
- Lemon juice: pH 2.0-2.5
- Blood: pH 7.35-7.45

## pH Scale Visualization



0

Strong Acid

7

Neutral

14

Strong Base

**pH 2:** Gastric acid, Lemon juice

**pH 4:** Wine, Tomato

**pH 7:** Pure water

**pH 7.4:** Blood

**pH 8:** Seawater

**pH 11:** Ammonia

- Baking soda solution: pH 8.5-9.0

## 2 Buffer Systems - pH Regulation Systems in Living Organisms

### What is a Buffer Solution?

A buffer solution is a solution that minimizes pH changes even when acids or bases are added. It consists of a mixture of a weak acid and its conjugate base (or a weak base and its conjugate acid).

### Mechanism of Action

- **When acid is added:** Base form ( $\text{A}^-$ ) neutralizes  $\text{H}^+$
- **When base is added:** Acid form ( $\text{HA}$ ) neutralizes  $\text{OH}^-$
- pH stabilization through equilibrium shift

### Major Biological Buffer Systems

#### 1. Bicarbonate Buffer System (Blood)



- Maintains blood pH at 7.35-7.45
- Regulated by lungs and kidneys
- Most important body fluid buffer system

#### 2. Phosphate Buffer System (Intracellular)



- Important in cytoplasm and urine
- $\text{pK}_a = 6.86$  (close to physiological pH)

#### Bicarbonate Buffer System

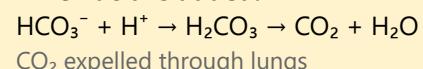


Carbonic anhydrase

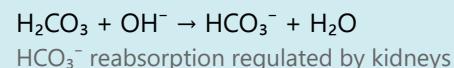


Normal blood pH range: 7.35 - 7.45

##### When acid is added:



##### When base is added:



### **3. Protein Buffer System**

- Amino acid residues such as histidine and cysteine
- Hemoglobin is a representative example

#### **Clinical Importance**

Acidosis or alkalosis can be life-threatening, and abnormalities in buffer systems are associated with various diseases.

### 3 Henderson-Hasselbalch Equation - Quantitative Analysis of Buffer Solutions

#### Derivation of the Equation

The Henderson-Hasselbalch equation is derived from the dissociation equilibrium of a weak acid:

$$K_a = [H^+] [A^-] / [HA]$$

Taking -log of both sides

$$pH = pK_a + \log([A^-] / [HA])$$

#### Meaning of the Equation

- **pH**: Hydrogen ion concentration of the solution
- **pKa**: Acid dissociation constant (-logKa)
- **[A<sup>-</sup>]**: Concentration of conjugate base
- **[HA]**: Concentration of weak acid

#### Key Principles

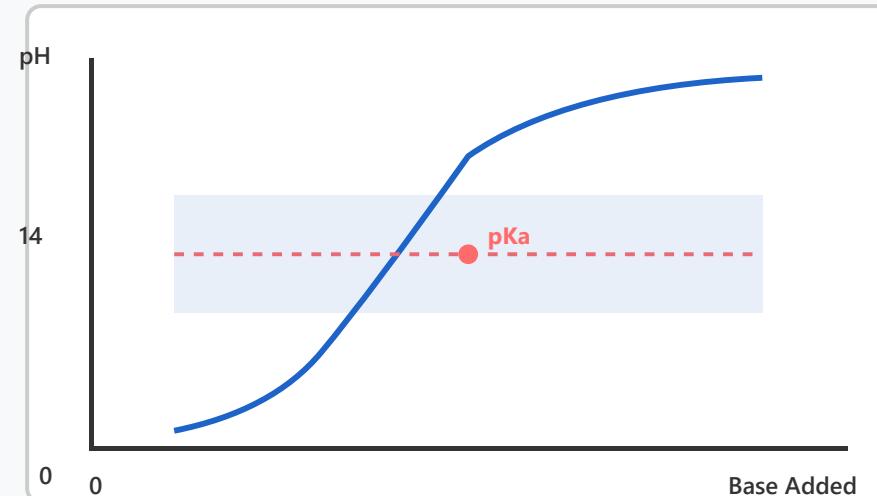
##### When pH = pKa:

- $[A^-] = [HA]$  (1:1 ratio)
- Maximum buffer capacity
- Equal concentrations of acid and base forms

**Effective buffer range:**  $pK_a \pm 1$

- Best buffering effect within this range

Henderson-Hasselbalch Titration Curve



##### Buffer Region:

pH change is gradual within  $pK_a \pm 1$  range

##### pH < pKa

$[HA] > [A^-]$   
Acid form predominates

##### pH > pKa

$[A^-] > [HA]$   
Base form predominates

- Effectively resists pH changes

## Biochemical Applications

- **Enzyme research:** Determining optimal pH conditions
- **Drug design:** Predicting bioavailability
- **Protein purification:** Maintaining stability
- **Diagnostics:** Blood gas analysis

## Practical Example

Acetate buffer ( $pK_a = 4.76$ ):

- At pH 4.76,  $\text{CH}_3\text{COOH}$  and  $\text{CH}_3\text{COO}^-$  are 1:1
- Effective buffering in pH 3.76-5.76 range
- Widely used in biochemical experiments

## 4 Enzyme pH Dependence - Relationship Between Enzyme Activity and pH

### Relationship Between Enzymes and pH

Every enzyme has a specific pH range at which it exhibits optimal activity. This reflects the enzyme's evolutionary adaptation and physiological function.

### Mechanisms by Which pH Affects Enzymes

#### 1. Changes in Ionization State

- Change in charge of amino acid residues in active site
- Proton acceptance/donation required for catalytic action
- pKa of His, Cys, Asp, Glu, etc. are important

#### 2. Changes in Protein Structure

- Changes in electrostatic interactions
- Changes in hydrogen bonding patterns
- Denaturation occurs at extreme pH

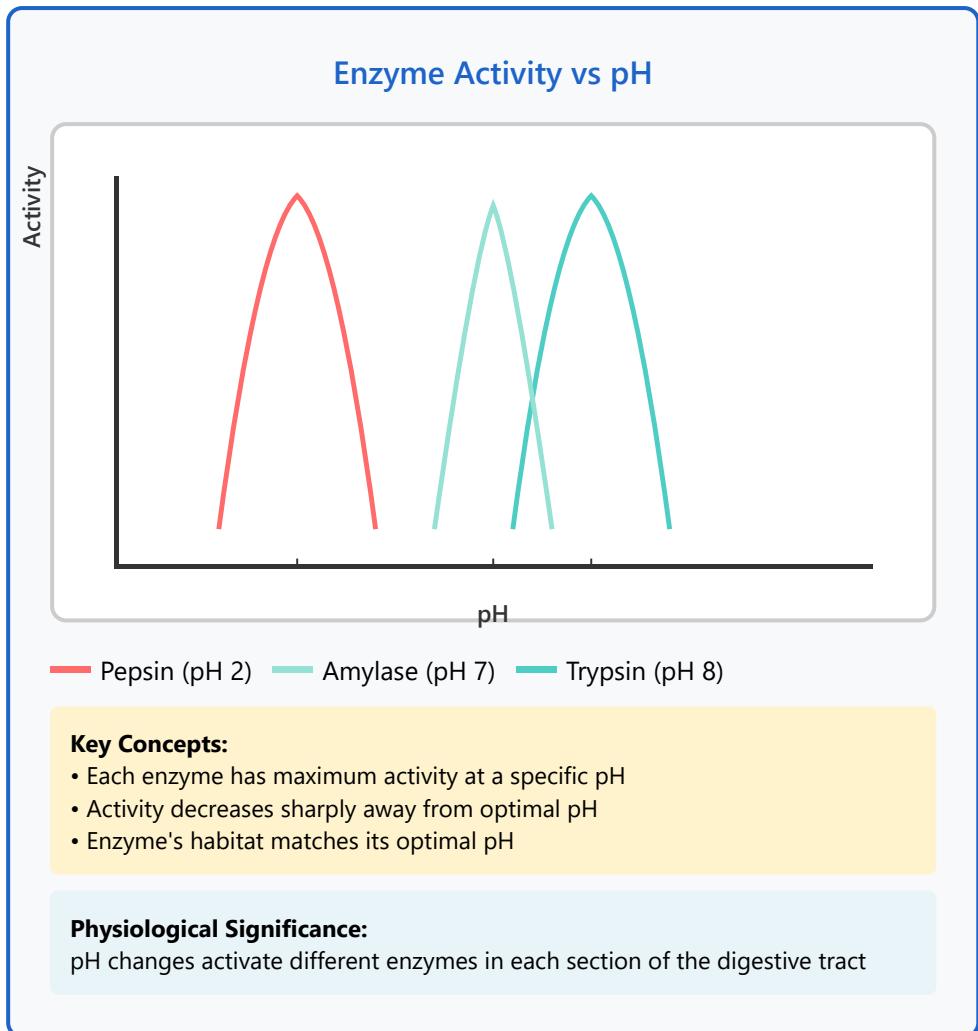
#### 3. Substrate Binding Affinity

- Effect of substrate ionization state
- Efficiency of enzyme-substrate complex formation

### Representative Enzyme Examples

#### Pepsin - pH 2.0

- Protein digestion in stomach



- Optimized for strongly acidic environment
- Rich in acidic amino acid residues

### Trypsin - pH 8.0

- Protein digestion in small intestine
- Active in slightly alkaline environment
- Secreted by pancreas

### Amylase - pH 7.0

- Starch-degrading enzyme
- Present in saliva and pancreas
- Optimal at neutral pH

## Clinical and Industrial Applications

- **Enzyme therapeutics:** Requires maintenance of appropriate pH
- **Food industry:** Control enzyme reactions by pH adjustment
- **Diagnostic kits:** Testing performed at optimal pH
- **Biotechnology:** Research on improving enzyme stability