

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

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

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 (A^-) neutralizes H^+
- **When base is added:** Acid form (HA) neutralizes 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
- $pK_a = 6.86$ (close to physiological pH)

Bicarbonate Buffer System

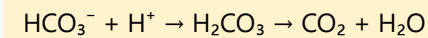


Carbonic anhydrase



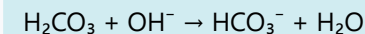
Normal blood pH range: 7.35 - 7.45

When acid is added:



CO_2 expelled through lungs

When base is added:



HCO_3^- reabsorption regulated by kidneys

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:

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

Taking -log of both sides

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

Meaning of the Equation

- **pH:** Hydrogen ion concentration of the solution
- **pKa:** Acid dissociation constant ($-\log K_a$)
- **[A⁻]:** Concentration of conjugate base
- **[HA]:** Concentration of weak acid

Key Principles

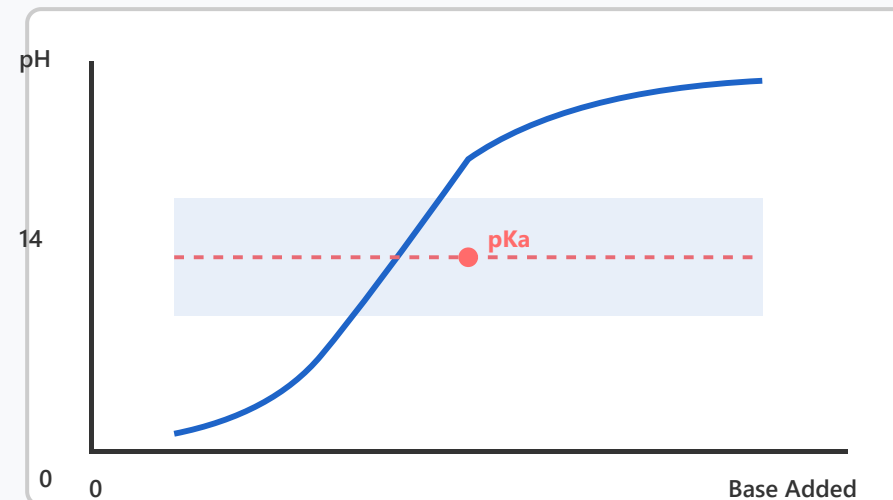
When pH = pKa:

- $[A^-] = [HA]$ (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, CH_3COOH and CH_3COO^- are 1:1
- Effective buffering in pH 3.76-5.76 range
- Widely used in biochemical experiments

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

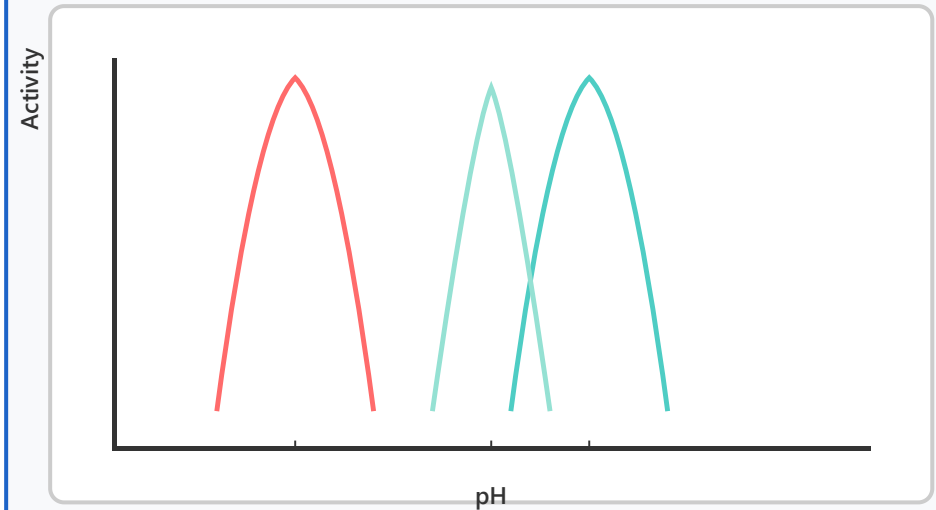
- Effect of substrate ionization state
- Efficiency of enzyme-substrate complex formation

Representative Enzyme Examples

Pepsin - pH 2.0

- Protein digestion in stomach

Enzyme Activity vs pH



— Pepsin (pH 2) — Amylase (pH 7) — Trypsin (pH 8)

Key Concepts:

- Each enzyme has maximum activity at a specific pH
- Activity decreases sharply away from optimal pH
- Enzyme's habitat matches its optimal pH

Physiological Significance:

pH changes activate different enzymes in each section of the digestive tract

- 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