AlAYEN IRAQI UNIVERSITY
Health and Medical Technologies
Anesthesia Department

General Chemistry
1st Year

Buffer Solution
Lec 7

Objectives

1. Buffer Solution and its Types
   Mechanism
2. Preparation of Buffer solutions
3. Handerson-Hasselbalch Equation
4. Solved Problems
5. pH Maintenance
6. Uses of Buffers

Buffer Solution

- **A buffer solution is an aqueous solution which consists of a mixture containing a weak acid and the conjugate base of the weak acid or a weak base and the conjugate acid of the weak base.**

Buffer Solution / Examples

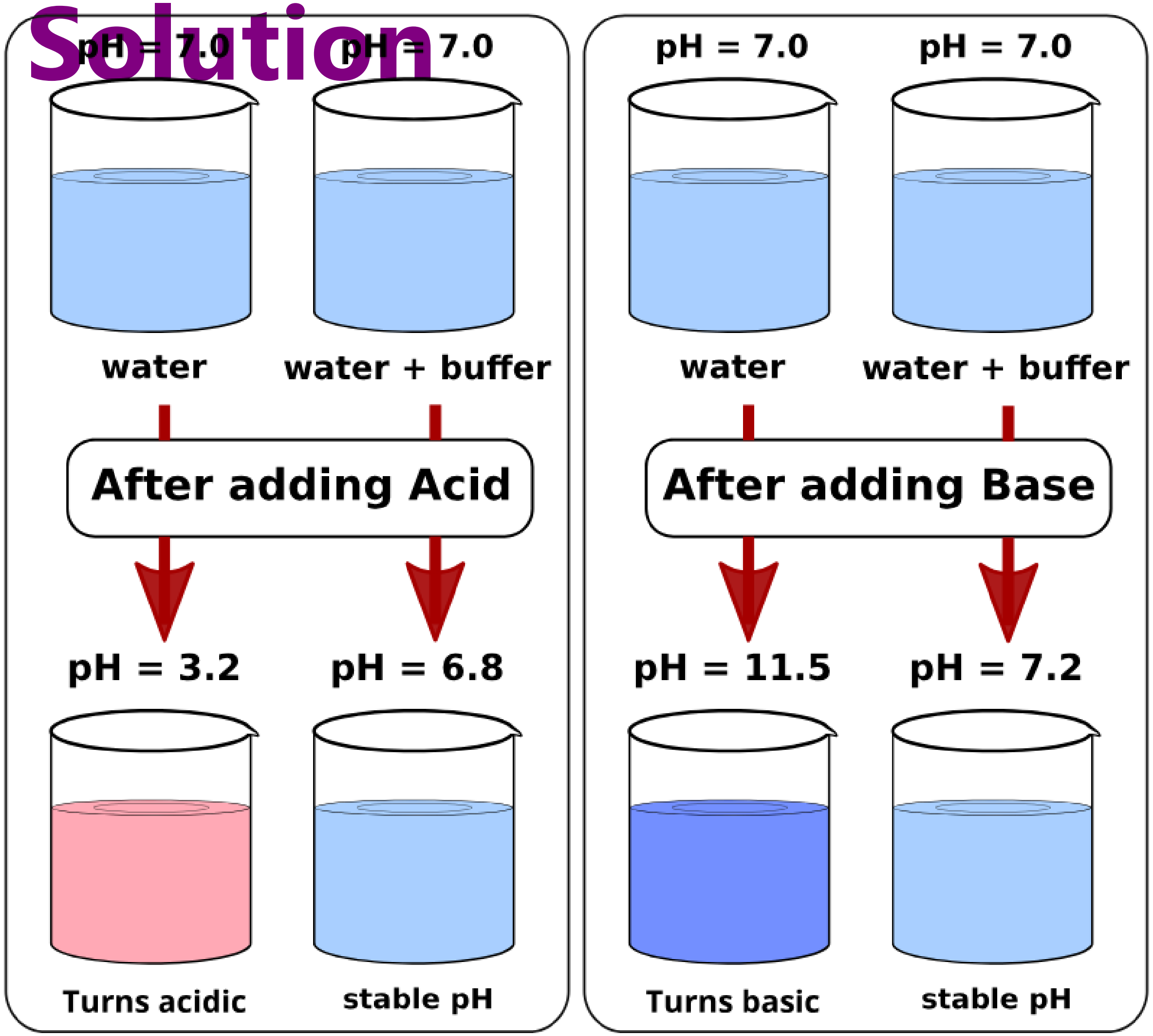
1- Acetic acid (CH_3COOH) and Sodium Acetate ($NaCH_3COO$)

2- Ammonia (NH_3) and
Ammonium Chloride (NH_4Cl)

3- Human blood buffer:
Carbonic acid (H_2CO_3) and
Sodium bicarbonate (HCO_3^-)

What Is the Role of Buffer Solution

- It resists a change in pH upon dilution or upon the addition of small amounts of acid or base to them.



Principle of Buffer action

•When a small quantity of an acid or base is added, the buffer behaves in such a way, to minimize the effect of that acid or base as much as possible so that the pH remains constant.

Bicarbonate buffer is present in human blood. It works to maintain the pH within the normal range (7.35 to 7.45).

Addition of Acid (H^+)

$$HCO_3^- + H^+ \rightleftharpoons H_2CO_3$$

When a small amount of H^+ is added, the equilibrium shifts forward and carbonic acid is formed. The concentration of H^+ remains constant.

Principle of Buffer action

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Addition of Base (OH^-)

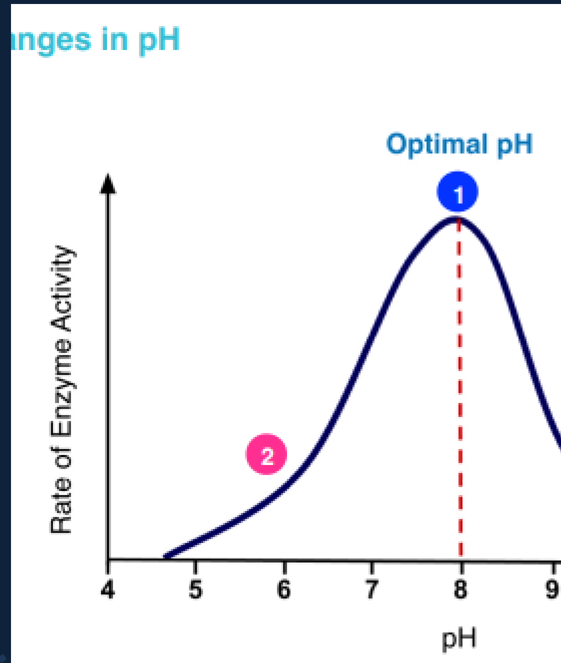
$$H_2CO_3 + OH^- \rightleftharpoons HCO_3^- + H_2O$$

When a small quantity of base (OH–) is added, the carbonic acid returns to its ionic form. Water, a neutral substance, is formed during this process, thus, the pH remains the same.

Buffer Capacity

β= millimoles /(ΔpH)

The number of millimoles of acid or base to be added to a liter of buffer solution to change the pH by one unit is the buffer capacity of the buffer.



Uses of Buffer Solution

Buffer solutions are used in food preservatives, drug delivery, and the activity of enzymes.

Types of Buffer Solutions

- **Acidic Buffers**

- The pH of these solutions is below seven.

- These solutions consist of a weak acid and a salt of a weak acid.

- An example of an acidic buffer solution is a mixture of sodium acetate and acetic acid (pH = 4.75).

Types of Buffer Solutions

- **Basic Buffers**

- The pH of these solutions is above seven.

- They contain a weak base and a salt of the weak base.

- An example of an alkaline buffer solution is a mixture of ammonium hydroxide and ammonium chloride (pH = 9.25).

Calculation of pH of a buffer solution

- **The pH of an *acid buffer* is given by Henderson-Hasselbalch equation.**

- **$pH = pK_a + \log([salt] / [weak\ acid])$**

- **$pH = pK_a + \log([conjugate\ base] / [weak\ acid])$**

- **where $pK_a = -\log K_a$ of the weak acid.**

Calculation of pH of a buffer solution

- If $[CH_3COOH]$ is equal to the 1.0 M and $[CH_3COONa]$ is 0.1 M, then find the pH of the buffer solution.

- pH = 4.74 + log 0.1/1

- = 4.74 + log 1/10

- = 4.74 + log 10^{-1}

- = 4.74 – 1

- pH = 3.74

Calculation of pH of a buffer solution

• The pH of a ***basic buffer*** is given by Henderson-Hasselbalch equation.

• $pOH = pK_b + \log ([salt] / [weak\ base])$

• $pOH = pK_b + \log ([conjugate\ acid] / [weak\ base])$

• $pH + pOH = 14$

• where $pK_b = -\log K_b$ of the weak acid.

Calculation of pH of a buffer solution

- A buffered solution contains 0.25M NH_3 (K_b = 1.8×10^{-5}) and 0.40M NH_4Cl. Calculate the pH of this solution.

$$pOH = pK_b + \log [salt] / [base]$$

$$pOH = 4.75 + \log 0.4M/0.25M$$
$$pOH = 4.75 + 0.25$$
$$pOH = 4.95$$
$$pH = 14 – pOH$$
$$= 14 – 4.95$$

$$pH = 9.05$$

How is the buffer solution prepared using the Henderson-Hasselbalch equation?

•Henderson-Hasselbalch equation tells us how the pH of the buffer solution changes with the pKa of a weak acid. The best buffer solution is prepared when we take an equal amount of salt and acid. The pH is controlled by the pKa of the acid.

- For example, for the buffer of acetic acid and sodium acetate, the pH will be:
  - $[CH_3COOH] = [CH_3COONa]$
  - $pH = pK_a + \log [CH_3COONa] / [CH_3COOH]$
  - $pH = pK_a + \log (1)$
  - $pH = pK_a + 0$
  - $pH = pK_a$
- pK_a of the acid is 4.74. So the pH of this buffer is equal to the pK_a of the acid. Hence, the pH of the buffer is 4.74.

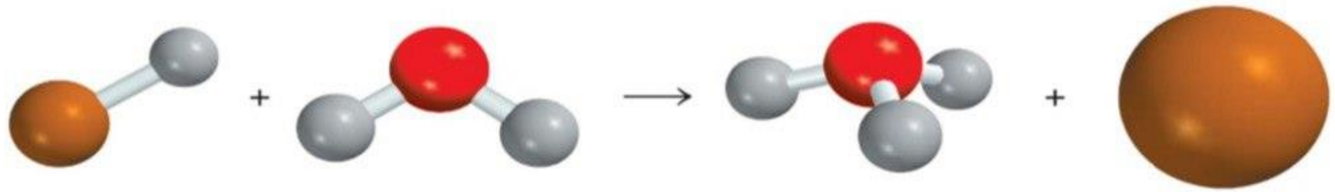
Real-life applications of buffer solutions

1. Photographic materials
2. Leathers
3. Dyes
4. Food preservation chemicals
5. Calibration of pH meters
6. Maintaining the pH of culture media for the growth of bacteria
7. The pH of soil, etc

Acid- Base Balance

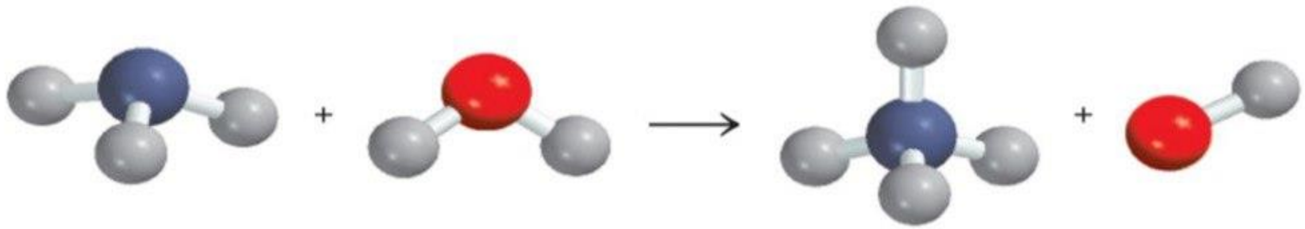
- Acids and Bases: Arrhenius Definition
- -An acid is a substance that take apart in water to yield H_3O^+ .
- -A base is a substance that take apart in water to yield OH^- .
- -This explains why all neutralization reactions between strong acids and bases and have similar heats of reaction:
- $H^+ (aq) + OH^- (aq)$? $H_2O(l)$ H 0 = -57 kJ/mol

Arrhenius acid is a substance that produces H^+ (H_3O^+) in water.



$$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$$

Arrhenius base is a substance that produces OH^- in water.



$$NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$$

Brønsted and Lowry proposed a more general definition…

Lewis Acids and Bases

- A Lewis acid is a substance that accepts a pair of electrons to form a covalent bond (ex; ccl4)

- A Lewis base is a substance that donates a pair of electrons to form a covalent bond (ex; NH3) a Lewis acid-base reaction is represented by the transfer of a pair of electrons from a base to an acid

Summary

- Acids are H+ donors.
- Bases are H+ acceptors, or give up OHin solution. Acids and bases can be:
- Strong – take apart completely in solution HCl, NaOH
- Weak – take apart only partially in solution Lactic acid, carbonic acid