Buffer Solution

2nd stage , 1st Semester

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Definition of buffer solution

Is a solution that **resist changes in pH** when small quantities of acid or alkali are added.

Uses of Buffer :

1/ in Biological Systems: pH of saliva typically ranges from 6.2 to 7.6
Blood maintains a pH range of 7.35 to 7.45

2/ in Cosmetics and Personal Hygiene Products:

Cosmetic and personal care products
Ointments and creams

Types of buffer solution

Acidic Buffer : Acidic buffers consist of a weak acid and its conjugate base (salt).

Example: Acetate buffer (CH_3COOH and CH_3COONa). Acetic acid (CH_3COOH) as a weak acid and sodium acetate (CH_3COONa) as its salt.

Note: The acetate buffer maintains a stable pH around 4.75, which is ideal for many biochemical processes and laboratory applications requiring an acidic environment.

Types of buffer solution

Basic Buffer: Basic buffers consist of a weak base and its conjugate acid (salt).

Example: Ammonium buffer or ammonia buffer (NH_4OH + NH_4Cl). Ammonium hydroxide (NH_4OH) as the weak base and ammonium chloride (NH_4Cl) as its salt.

Note: The ammonia buffer is well-known for maintaining a stable pH around 9.25, which is often required in biochemical and analytical applications.

Phosphate Buffer (Sorensen's Buffer Solution)

The phosphate buffer consists of Sodium dihydrogen phosphate (NaH_2PO_4) which acts as the acid component, and Disodium hydrogen phosphate (Na_2HPO_4) which acts as the conjugate base (salt) component.

pH Range: The phosphate buffer is effective across a broad pH range, generally around pH 6.8-7.4. By adjusting the ratio of NaH_2PO_4 to Na_2HPO_4, the pH can be precisely controlled.

Applications: This buffer is extensively used in biological and medical fields due to its compatibility with biological systems.

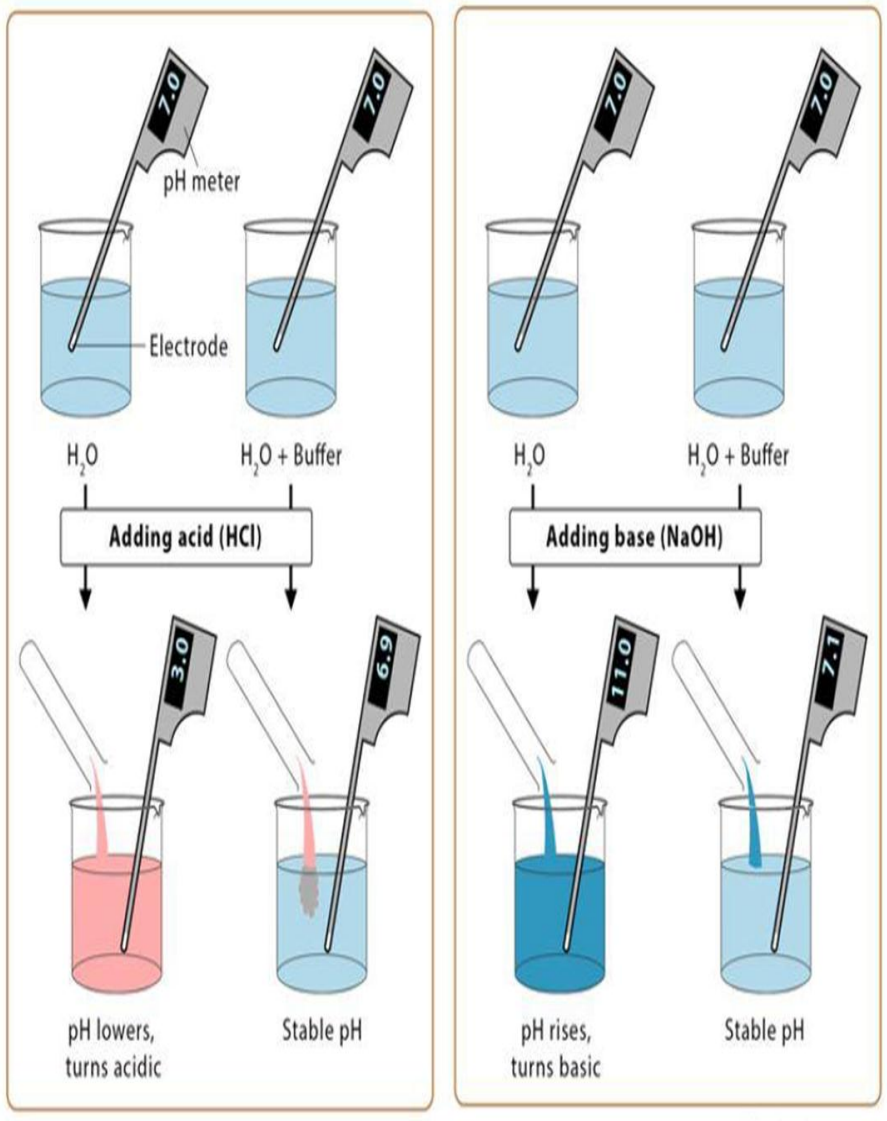
Differentiation between buffer system & non-buffer system:

For example:

- If 1 ml of **0.1N HCl** solution is added to 100ml of **pure water**, the pH is reduced from 7 to 3.

- When **strong acid** is added to 1M solution containing equal quantities of **acetic acid & sodium acetate** the pH change only by 0.09 units because the base $CH3COO^-$ ties up the H^+ ion according to the following equation:

- $CH3COO^- + H^+ \rightarrow CH3COOH + H_2O$



Mechanism of buffer action

Consider a buffer solution that consists of a weak acid and its salt such as the acetate buffer:

$CH_3COOH \rightleftharpoons H^+ + CH_3COO^-$ (incomplete dissociation)

$CH_3COONa \rightarrow Na^+ + CH_3COO^-$ (complete dissociation)

Addition of Acid (H^+):

When a strong acid (e.g. HCl) is added, it will be ***consumed*** by the conjugate base, forming the weak acid: $CH_3COONa + HCl \rightarrow CH_3COOH + NaCl$. This reaction consumes some of the added H^+ ions, minimizing the change in pH.

Addition of Base (OH^-):

When a strong base (e.g. NaOH) is added, it will be ***consumed*** by the weak acid, forming the conjugate base: $CH_3COOH + NaOH \rightarrow H_2O + CH_3COONa$. This reaction reduces the concentration of OH^- ions and again helps to stabilize the pH.

pH Measurement

pH value can be measured using:

1. **Electrometric method:** by using pH meter.

2. **Mathematical method:** by using the Handerson-Hasselbalch equation

- Handerson-Hasselbalch equation for a weak acid and its salt:

$$\mathbf{pH = pKa + \log \frac{[Salt]}{[Acid]}}$$

- Handerson-Hasselbalch equation for a weak base and its salt:

$$\mathbf{pH = pKw - pKa + \log \frac{[Base]}{[Salt]}}$$

pH Measurement

3. Colorimetric method:

It is less accurate and less convenient but less expensive than the electronic method. Colorimetric method involves the use of:

a. **Paper indicator:** a strip of paper that undergo color changes depending on the pH of the solution where the strip is immersed.

b. **liquid indicator:** they are weak acid or weak base that show color changes as their degree of dissociation varies with pH. The color of the indicator resembles a certain pH value.

- An examples of liquid indicators: methyl yellow, methyl red, bromothymol blue, thymol blue, and phenolphthalein.

Examples of calculating the pH of a buffer solution using Handerson-Hasselbalch equation

Example 1. Calculate the pH of a buffer solution that is **0.050 M** in benzoic acid and **0.150 M** in sodium benzoate, **pka = 4.2**.

$$\mathbf{pH = pka + \log \frac{[salt]}{[acid]}}$$

$$pH = 4.2 + \log \frac{0.150}{0.050}$$

$$pH = 4.2 + \log 3$$

$$pH = 4.2 + 0.48 = 4.68$$

Example 2. A buffer is prepared containing **1.00 M** acetic acid and **1.00 M** sodium acetate. What is its pH? the **pKa = 4.76.**

$$\mathbf{pH = pKa + \log \frac{[salt]}{[acid]}}$$

pH = 4.76 + log 1/1

pH = 4.76 + log 1

pH = 4.76 + 0 = 4.76

Example 3: A buffer is prepared containing **0.800 molar** acetic acid and **1.00 molar** sodium acetate. What is its PH ? **pKa = 4.76.**

$$\mathbf{pH = pka + \log \frac{[salt]}{[acid]}}$$

$$pH = 4.76 + \log \frac{1}{0.8}$$

$$pH = 4.76 + \log 1.25$$

$$\mathbf{pH = 4.76 + 0.09 = 4.85}$$

Example 4: A buffer is prepared containing **1.00 molar** acetic acid and **0.800 molar** sodium acetate. What is its pH ? the pKa of acetic acid = 4.76

$$\mathbf{pH = pka + \log \frac{[salt]}{[acid]}}$$

$$pH = 4.76 + \log \frac{0.8}{1}$$

$$pH = 4.76 + \log 0.8$$

$$\mathbf{pH = 4.76 - 0.09 = 4.67}$$

Buffer capacity

Refers to the ability of a buffer solution to resist changes in pH when strong acids or bases are added. The buffer capacity is numerically expressed as the minimum concentration of strong acid or base required to cause a one-unit change in pH.

To calculate buffer capacity, we use the following formula:

$$\beta = \Delta B/\Delta pH$$

Where:

β is a buffer capacity (which is unitless) .

ΔB is the amount of strong acid or strong base added (the concentration in a normality 'N').

ΔpH is the pH difference between the initial buffer's pH and the pH after the addition of acid or base to the buffer.

Examples of Buffer Capacity

Example 1: 0.5 ml of 0.1 M NaOH is added to 40 ml of buffer solution having pH 3.75 and the pH will be 3.78 after the addition, calculate the buffer capacity?

Solution: $\beta = \Delta B / \Delta pH$

First, we find ΔB using the following formula: $C1V1 = C2V2$

Since, 0.1 M NaOH = 0.1 N NaOH

because the Eq. wt = M.wt

$$0.1\ N \times 0.5\ ml = C2 \times (40+0.5)\ ml$$

$\Delta B = C2 = 0.0012$ N is the normality of NaOH in the final solution.

$\Delta pH = 3.78 - 3.75 = 0.03$

$\beta = 0.0012 / 0.03 = 0.04$

Thus, the buffer capacity of solution is **0.04**

Example 2: We are given 60 mL of a sodium phosphate buffer with a pH of 7.39. We then add to it 150 mL of 0.2 M HCl. This addition gives the buffer solution a new pH of 7.03. What is the capacity of our sodium phosphate buffer?

Solution 1: $\beta = \Delta B / \Delta pH$

0.2 M HCl = 0.2 N HCl

because the Eq. wt = M.wt

$C1V1 = C2V2$

$0.2\ N \times 150\ ml = C2 \times (60+150)\ ml$

$\Delta B = C2 = 0.04$ N is the normality of HCl in the final solution

$\Delta pH = 7.03 - 7.39 = 0.36$

$\beta = 0.14 / 0.36 = 0.39$

Thus, the buffer capacity of solution is **0.39**

Note 1:

1. Addition of **strong acid** to **acidic buffer solution** lead to **increase** the acid concentration and **decrease** the salt concentration. **pH = pka + log**

$$\frac{[\text{salt} - \text{strong acid}]}{[\text{acid} + \text{strong acid}]}$$

2. Addition of **strong base** to **acidic buffer solution** lead to **decrtease** the acid concentration and **increase** the salt concentration. **pH = pka + log**

$$\frac{[\text{salt} + \text{strong base}]}{[\text{acid} - \text{strong base}]}$$

3. Addition of **strong base** to **basic buffer solution** lead to **increase** the base concentration and **decrease** the salt concentration.
4. Addition of **strong acid** to **basic buffer solution** lead to **decrease** the base concentration and **increase** the salt concentration

Note 2:

1. If a buffer has a higher concentration of its components (e.g. the weak acid and its conjugate base or a weak base and its conjugate acid), its capacity will be higher.

For example: a buffer with [weak acid] = 0.50 M and [conjugate base] = 0.50 M would have a higher capacity than a buffer with [weak acid] = 0.050 M and [conjugate base] = 0.050 M.

2. If a buffer has a greater concentration of the weak acid than that of the conjugate base, then it will have a higher capacity for added base. For example, in an acetate buffer with 0.1 M CH_3COOH (weak acid) and 0.01 M CH_3COO^- (conjugate base), the higher weak acid concentration increases the buffer's capacity to resist changes in pH when a base is added.

3. Likewise, if a buffer has a greater concentration of the conjugate base than that of the weak acid, then it will have a higher tolerance for added acid.

1. Addition of **strong acid** to **acidic buffer solution** lead to increase the acid concentration and decrease the salt concentration. **pH = pka + log** $\frac{\text{[salt − strong acid]}}{\text{[acid + strong acid]}}$

Example: An acidic buffer is prepared containing **0.1 M** acetic acid and **0.1 M** sodium acetate. 0.01 M of HCl has been added. Calculate the pH before and after HCl added? **pKa = 4.76.**

The pH before HCl added:

$$pH = pka + \log \frac{[salt]}{[acid]}$$

$$pH = 4.76 + \log \frac{0.1}{0.1}$$

$$pH = 4.76 + \log 1$$

$$pH = 4.76 + 0 = \mathbf{4.76}$$

• The pH after HCl added:

$$\mathbf{pH = pka + \log \frac{[salt - strong\ acid]}{[acid + strong\ acid]}}$$

$$pH = pka + \log \frac{[0.1 - 0.01]}{[0.1 + 0.01]}$$

$$pH = 4.76 + \log \frac{0.09}{0.11}$$

$$pH = 4.76 + \log 0.82$$

$$pH = 4.76 - 0.09 = \mathbf{4.67}$$

2. Addition of **strong base** to **acidic buffer solution** lead to decrease the acid concentration and increase the salt concentration. **pH = pka + log** $\frac{[\text{salt+ strong base}]}{[\text{acid− strong base}]}$

Example: An acidic buffer is prepared containing **0.10 M** acetic acid and **0.10 M** sodium acetate. 0.01 M of NaOH has been added. Calculate the pH before and after NaOH added? **pKa = 4.76.**

The pH before NaOH added:

$$pH = pka + \log \frac{[salt]}{[acid]}$$

$$pH = 4.76 + \log \frac{0.1}{0.1}$$

$$pH = 4.76 + \log 1$$

$$pH = 4.76 + 0 = \mathbf{4.76}$$

The pH after NaOH added:

$$pH = pka + \log \frac{[\text{salt+ strong base}]}{[\text{acid− strong base}]}$$

$$pH = pka + \log \frac{[0.1+0.01]}{[0.1-0.01]}$$

$$pH = 4.76 + \log \frac{0.11}{0.09}$$

$$pH = 4.76 + \log 1.22$$

$$pH = 4.76 + 0.09 = \mathbf{4.85}$$

Example: An acidic buffer is prepared containing **0.05 M** acetic acid and **0.05 M** sodium acetate. 0.01 M of NaOH has been added. Calculate the pH after NaOH added? **pKa = 4.76**

pH = pka + log ([salt+ strong base]) / ([acid− strong base])

pH = pka + log ([0.05+0.01]) / ([0.05−0.01])

pH = 4.76 + log 0.06/0.04

pH = 4.76 + log 1.5

pH = 4.76 + 0.17= 4.93

Thank You