

Buffered solutions

Introduction

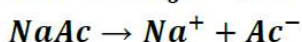
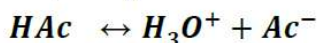
- Buffers are mixtures of compounds that resist changes in pH upon the addition of small quantities of acid or alkali.
- A buffer is composed of a weak acid (HA) and its salt (conjugate base A^-) or a weak base (B) and its conjugate acid BH^+ .



Buffer Equation

The pH of a buffer solution can be calculated by use of the *buffer equation*.

E.g. When sodium acetate ($NaAc$) is added to acetic acid (HAc), the salt and the acid have an ion in common.

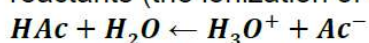


K_a for the weak acid is momentarily disturbed because the Ac^- supplied by the salt increases the $[Ac^-]$ term in the numerator:

$$K_a = \frac{[H_3O^+][Ac^- \uparrow]}{[HAc]}$$

Buffer Equation

To reestablish the constant K_a , $[H_3O^+]$ is instantaneously decreased, by shifting the equilibrium in the direction of the reactants (the ionization of acetic acid is repressed).



$$K_a = \frac{[H_3O^+ \downarrow][Ac^- \uparrow]}{[HAc]}$$

Since weak acid is slightly ionized, $[HAc] = [\text{total acid concentration}]$

Since the salt is completely ionized, $[Ac^-] = [\text{total salt concentration}]$

$$K_a = \frac{[H_3O^+][Salt]}{[acid]}$$

By rearranging the equation and using the logarithmic form:

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

Buffer Equation

The previous equation is known as the *buffer equation* or the *Henderson–Hasselbalch equation*.

For a weak acid (HA) and its salt (S):

$$pH = pKa + \log \frac{[S]}{[HA]}$$

For a weak Base (B) and its salt (S):

$$pH = pKa + \log \frac{[B]}{[S]}$$

Buffer Equation

Example

What is the pH of a buffer solution containing 0.1 M acetic acid and 0.1 M sodium acetate?

$$pH = pKa + \log \frac{[S]}{[HA]}$$

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

Buffer Equation

Example

How much sodium acetate (mol) should be added to 100 mL of 0.1 mol/L acetic acid solution to prepare a buffer of pH 5.2?

$$pH = pKa + \log \frac{[S]}{[HA]}$$

$$5.2 = 4.76 + \log \frac{[S]}{[0.1]}$$

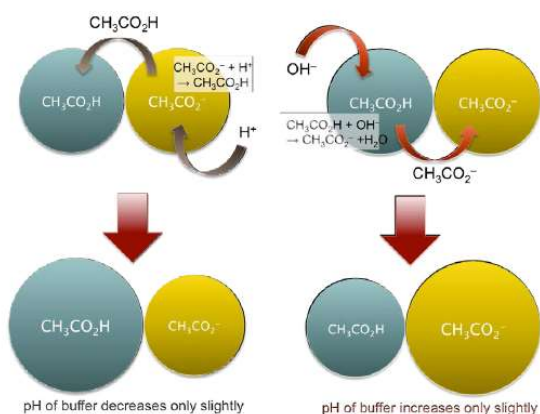
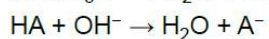
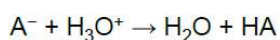
$$[S] = 2.754 \text{ mol/L}$$

$$n_{\text{sodium acetate}} = M \times v = 2.754 \times 0.01 = 0.02754 \text{ mol}$$

Buffer Capacity

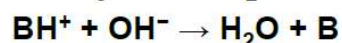
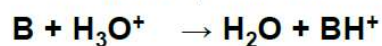


A weak acid and its conjugate base have buffering capacity because A⁻ ions remove the added H⁺ as undissociated weak acid, while HA remove the added OH⁻ ions as water:



Buffer Capacity

A weak base and its conjugate base have buffering capacity because B ions remove the added H^+ as undissociated weak acid, while BH^+ remove the added OH^- ions as water:



Buffer capacity is the quantity of strong acid or base that can be added to change the pH of one **liter** of buffer solution by one pH unit.

Buffer Capacity

Koppel and Spiro and Van Slyke devised an approximate equation for calculating buffer capacity (β):

$$\beta = \frac{\Delta B}{\Delta pH}$$

ΔB : number of moles of strong acid or base per liter of buffer.

ΔpH : change in pH.

When one of the buffer components is depleted completely, the solution lose its buffering capacity and can no longer resist the change in pH

Buffer Capacity

Koppel and Spiro and Van Slyke developed a more exact Equation for calculating buffer capacity:

$$\beta = 2.3C \frac{K_a [H_3O^+]}{(K_a + [H_3O^+])^2}$$

C is the total buffer concentration (the sum of the molar concentrations of the acid and the salt).

This equation allows the calculation of buffer capacity at any pH (even when no acid or base has been added to the buffer).

The equation shows that an increase in the concentration of the buffer components (C) results in a greater buffer capacity (β).

Buffer Capacity

At a hydrogen ion concentration of 1.75×10^{-5} , what is the capacity of a buffer containing 0.10 mole each of acetic acid and sodium acetate per liter of solution? ($K_a = 1.75 \times 10^{-5}$)

$$C = [\text{Acid}] + [\text{Salt}] = 0.1 + 0.1 = 0.20 \text{ mole/liter}$$

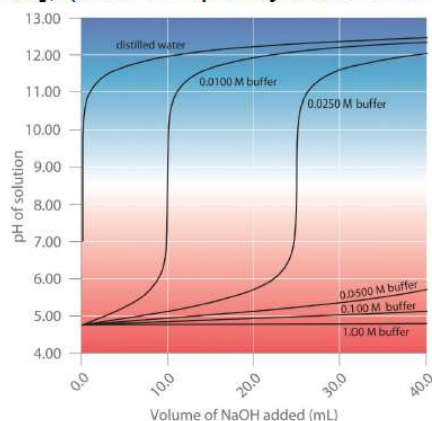
$$\beta = 2.3C \frac{K_a [H_3O^+]}{(K_a + [H_3O^+])^2}$$

$$\beta = 2.3 \times 0.2 \times \frac{(1.75 \times 10^{-5})(1.75 \times 10^{-5})}{[(1.75 \times 10^{-5}) + (1.75 \times 10^{-5})]^2} = 0.115 \text{ mol/L}$$

Buffer Capacity

The buffer capacity depends on:

- the value of the ratio $[\text{Salt}]/[\text{Acid}]$, (buffer capacity increases as the ratio approaches 1)
- the magnitude of the individual concentrations of the buffer components (buffer capacity increases as the salt and acid concentrations are increased).

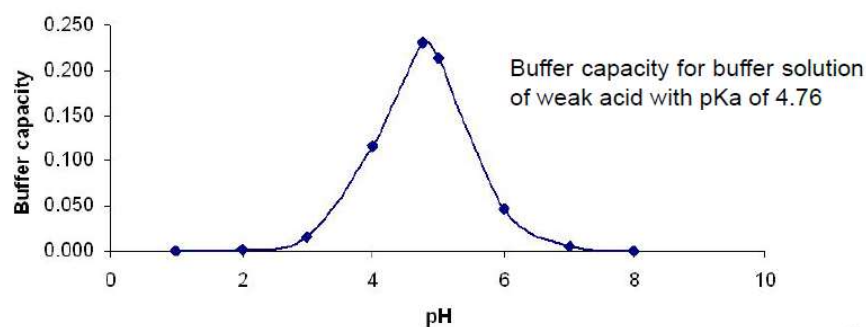


Buffer Capacity

The maximum buffer capacity occurs where $\text{pH} = \text{pK}_a$, or, in equivalent terms, where $[\text{H}_3\text{O}^+] = K_a$.

$$\beta_{\text{max}} = 0.576 C$$

Where C is the total buffer concentration



Buffer Capacity

Example

What is the maximum buffer capacity of an acetate buffer with a total concentration of 0.020 mole/liter?

$$\beta_{max} = 0.576 \text{ C}$$

$$\beta_{max} = 0.576 \times 0.02 = 0.012 \text{ Mol/liter}$$

Buffer in Biological Systems

Some body fluids have natural buffer capacity:

1. pH of tears is 7-8 with higher buffer capacity so that a reasonably wide pH range of medicines can be tolerated.
2. pH of blood is maintained at approximately 7.4 by buffer component in the plasma (bicarbonate and phosphate buffers) and erythrocytes (hemoglobin and phosphate buffers).



Hb: hemoglobin, O₂Hb: oxyhemoglobin

Pharmaceutical Buffers

Buffer solutions are widely used to adjust pH of aqueous pharmaceutical solutions to ensure:

- Prevention of tissue irritation
- Optimum therapeutic effect
- Maximum drug stability
- Maximum drug solubility

Pharmaceutical Buffers

Tissue Irritation Prevention

- Solutions to be applied to delicate tissues (e.g. eye) or administered parenterally are liable to cause irritation if their pH is greatly different from the normal pH of the relevant body fluid.
- If there is a large pH difference between the solution and body fluid, tissue irritation will be minimal if:
 - The volume and buffer capacity of the solution is low
 - The volume and buffer capacity of the physiologic fluid is high.

Pharmaceutical Buffers

Maximum Therapeutic Effect vs Stability

- The undissociated form of a weakly acidic or basic drug often has a higher therapeutic activity than that of the dissociated salt form because they can penetrate body membranes readily due to their lipid solubility.
- The pH for maximum stability of a drug for ophthalmic use may be far below that of the optimum physiologic effect.
- Under such conditions, the solution of the drug can be buffered at a low buffer capacity and at a pH that is between that of optimum stability and that for maximum therapeutic action.
- When the solution is instilled in the eye, the tears bring the pH to about 7.4, converting the drug to the physiologically active form

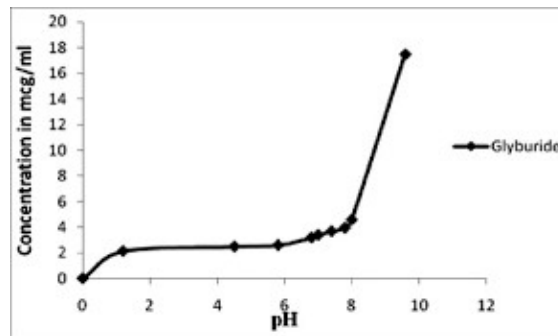
Pharmaceutical Buffers

Maximum Solubility

- The pH of the solution can affect the solubility of the drug.
- At a low pH, a base is predominantly in the ionic form, which is usually very soluble in aqueous media.
- As the pH is raised, more undissociated base is formed, which has poor water solubility, leading to precipitation of this form from solution.
- Therefore, the solution should be buffered at a sufficiently low pH so that the concentration of the free base is less than its solubility.

Pharmaceutical Buffers

pH Solubility profile



Pharmaceutical Buffers

• pH Stability profile

