COLLEGE OF PHARMACY
UNIVERSITY OF BASRAH

Buffer Solution

By
Dr. Mohammed Sattar
2018/2019

1

COLLEGE OF PHARMACY
UNIVERSITY OF BASRAH

Outlines

- Objectives
- Introduction
- Buffer equation
- Drugs as buffer
- pH indicators

2

COLLEGE OF PHARMACY
UNIVERSITY OF BASRAH

pH Indicators

- May be considered as weak acid or weak base that act like buffer and also exhibit color changes as their degree of dissociation varies with pH.

Ex. **Methyl Red**

- can give full alkaline color (yellow) at pH $\approx$6.
- And full acid color (red) at pH $\approx$ 4.

So, it is used for determination of pH of solutions(How?).

➢ **The dissociation of an acid indicator is given in simplified form as**

$$\underset{\substack{\text{Acid}_1\\\text{(Acid color)}}}{\text{HIn}} + \underset{\text{Base}_2}{\text{H}_2\text{O}} \rightleftharpoons \underset{\text{Acid}_2}{\text{H}_3\text{O}^+} + \underset{\substack{\text{Base}_1\\\text{(Alkaline color)}}}{\text{In}^-}$$

Then: $\dfrac{[\text{H}_3\text{O}^+][\text{In}^-]}{[\text{HIn}]} = K_{\text{In}}$

HIn = unionized form of indicator → *acid color*
In^- = ionized form → basic color
K_{In} = the indicator constant

11

COLLEGE OF PHARMACY
UNIVERSITY OF BASRAH

- **The color of indicator is a function of pH of the solution, how??**

➢ If acid is added to the solution of indicator, this causes increase in $[H_3O^+]$ shifting the reaction toward lift (common ion effect) forming more [HIn]= predominant → ***acid red color***.

➢ If base is added, causing decrease in $[H_3O^+]$ by reaction with acid, shifting the reaction toward right, forming more ionized form $[In^-]$ → ***Base Yellow color*** ***is predominante.***

Colorimetric methods for pH calculation

- **Applications**:

1. Less expensive compare to other methods
2. Used for non color aqueous solution
3. Used for non aqueous solutions

- **Limitations:**

1. Less accurate and less convenient.
2. They are acid or bases and if added to unbuffered solution causing significant change in pH.

13

Buffer capacity

- Or **buffer efficiency**, **buffer index** or **value**, represents *the magnitude of the resistance of a buffer to pH changes*. Or the ratio of the increment of strong base (or acid) to the small changes in pH brought about by this addition.

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

Where, Δ *is delta = a finite change*
ΔB = *the small increment in gram equivalent per liter of strong base added to the buffer solution to produce a pH change of* (ΔpH).

Then, according to this equation, addition of one gram equivalent of strong base (or acid) to one liter of the buffer solution results in the change of one pH unit.

14

How can you make an approximate calculation for buffer capacity?

- For ex. Acetate buffer containing (**0.1mole** each of acetic acid and sodium acetate in one liter of solution) to which **0.01mole** of NaOH is added

resulting in increase of salt conc. [$Na^+ Ac^-$] by (0.01 mole/liter), and [HAc] conc. decreases proportionally because each increment of base convert 0.01mole of acetic acid into 0.01mole of sodium acetate according to the following equation:

$$\underset{(0.1-0.01)}{\mathrm{HAc}} + \underset{(0.01)}{\mathrm{NaOH}} \rightleftharpoons \underset{(0.1+0.01)}{\mathrm{NaAc}} + H_2O$$

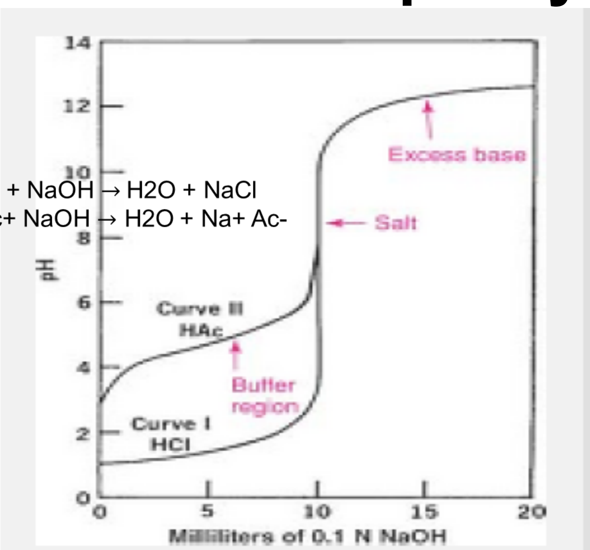
And

$$pH = pK_a + \log\frac{[\mathrm{Salt}] + [\mathrm{Base}]}{[\mathrm{Acid}] - [\mathrm{Base}]}$$

Neutralization curves and buffer capacity:

- By considering the titration curves of strong and weak acids when they are mixed with increasing quantities of alkali, the reaction of an equivalent of an acid with an equivalent of the base called **neutralization**.

HCl + NaOH → H2O + NaCl
Hac+ NaOH → H2O + Na+ Ac-



Buffers in Pharma. & Biological systems

***In vivo* biologic buffer system**:

- **Blood** maintained the pH at **7.4** by so-called primary buffers in **the plasma** and the secondary buffers in **the erythrocytes**.
- **Plasma** contains (carbonic acid/bicarbonate, acid/alkali salts of phosphoric acid and plasma proteins).
- **Erythrocytes** contain two buffer systems (hemoglobin/oxy-hemoglobin and acid/alkali potassium salts of phosphoric acid).
- **It is usually life threatening for the pH of the blood to go below 6.9 or above 7.8, as in diabetic coma (pH as low as 6.8). Then, it is important for maintaining pH in a range of (7-7.8)**

17

✓ **Lacrimal fluid or tears,** have a great degree of buffer capacity allowing dilution of (1:15) with neutral D.w. before an alteration of pH is noticed (dilution value > buffer capacity).

- pH of tears ≈ 7.4 or in range (7-8) or slightly higher.

✓ **For urine**, the 24-hr. urine collection of a normal adult have the pH averaging about 6 units, or it may be as low as 4.5 or as high as 7.8.

- When the pH of urine is below normal value, hydrogen ions are excreted by kidney. Conversely when the pH is about 7.4, hydrogen ions are retained by the kidney action.

18

Influence of buffer capacity and pH on tissue irritation:

- Eye irritation my be resulted from the presence of free form of a drug at the physiologic pH, it is more often due to the acidity of the eye solution.
- Parenteral solution are usually not buffered or they are buffered to a low capacity so that, the buffer of the blood may readily bring them within the physiologic pH range.
- For oral administration, like aspirin is absorbed more rapidly in system buffered at low buffer capacity than in system containing no buffer or in highly buffered preparations. Gastric irritation is also affected by buffer capacity.

Stability versus optimum therapeutic response:

- Undissociated form of weakly acidic or basic drugs often has a higher therapeutic activity than the dissociated salt form (less lipophilic).
- For ophthalmic drugs (weakly basic alkaloid, pilocarpine, pk_b= 7.15), therapeutic response is increased as the pH of solution and hence the concentration of the undissociated base were increased.

✓ At low pH about 4, ionic form predominates, so low penetration or slow.

✓Tears with pH about 7.4, the drug found mainly in free base form.

Example 8-10

Mole Percent of Free Base

The pK_b of pilocarpine is 7.15 at 25°C. Compute the mole percent of free base present at 25°C and at a pH of 7.4. We have

$$\underset{\text{Pilocarpine base}}{C_{11}H_{16}N_2O_2} + H_2O \rightleftharpoons \underset{\text{Pilocarpine ion}}{C_{11}H_{16}N_2O_2} + OH^-$$

$$pH = pK_w - pK_b + \log\frac{[\text{Base}]}{[\text{Salt}]}$$

$$7.4 = 14.00 - 7.15 + \log\frac{[\text{Base}]}{[\text{Salt}]}$$

$$\log\frac{[\text{Base}]}{[\text{Salt}]} = 7.40 - 14.00 + 7.15 = 0.55$$

$$\frac{[\text{Base}]}{[\text{Salt}]} = \frac{3.56}{1}$$

$$\text{mole percent of base} = \frac{[\text{Base}]}{[\text{Salt}] + [\text{Base}]} \times 100$$

$$= [3.56/(1 + 3.56)] \times 100 = 78\%$$

The solutions of drugs can be buffered al a low buffer capacity and at pH that is a compromise between that of optimum stability and the pH for maximum therapeutic action.

Thanks for your attention