وزارة التعليم العالي والبحث العلمي
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مختبر الصيدلة الفيزياويه I / المرحلة الثانية

Buffer solutions

Buffers: are compounds or mixtures of compounds that, by their presence in solution, resist changes in pH upon the addition of small quantities of acid or alkali.

- **buffer action :** The resistance to a change in pH .
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- **What is a Buffer?**
- A combination of a weak acid and its conjugate base (i.e., its salt) or
- a weak base and its conjugate acid
- Consider a buffer solution that includes of a weak acid and its salt such as the acetate buffer:

- $CH_3COOH \leftrightarrow H_3O^+ + CH_3COO^-$ (incomplete dissociation)

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- $CH_3COOK \rightarrow K^+ + CH_3COO^-$ (complete dissociation)

- **How can you differentiate between buffer system & non-buffer system?**

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

- When strong acid is added to 0.01 M solution containing equal quantities of acetic acid & sodium acetate the pH change only by 0.09 units because the base AC^- ties up the H^+ ion according to the following equation

- $\mathbf{AC^- + H_3O \rightarrow HAC + H_2O}$

- To illustrate the way that buffer resist pH change ,let's take acetate buffer as example:

$HAC + H_2O \leftrightarrow AC^- + H_3O^+$

$NaAC \rightarrow AC^- + Na^+$

- if strong acid added $\rightarrow$ $H_3O^+ \rightarrow$ shifts the equation to the left so ties up the H_3O^+ ion.

- If strong base added → OH^- → shifts the equation to the right so ties up OH^- ion .

- When a strong base, such as KOH is added, the following occurs:

$$KOH \rightarrow OH^- + K^+$$
$$CH_3COOH \leftrightarrow H_3O^+ + CH_3COO^- \text{ (shifts to the right)}$$
$$CH_3COOK \rightarrow K^+ + CH_3COO^-$$

The added OH^- ions react with the $H3O^+$ ions to form H_2O

The decrease in $[H_3O^+]$ causes a shift to the right and more CH_3COO^- is formed.

Buffer equations:-

Weak acid & it's salt	*Weak Base & it's salt*
$PH = Pka + \log \frac{salt}{acid}$	$pH = pK_w - pK_b + \log \frac{[base]}{[salt]}$

pH Indicators

The pH of the buffer solution can be measured by:

1- **Colorimetric method**:

a)chemical indicator

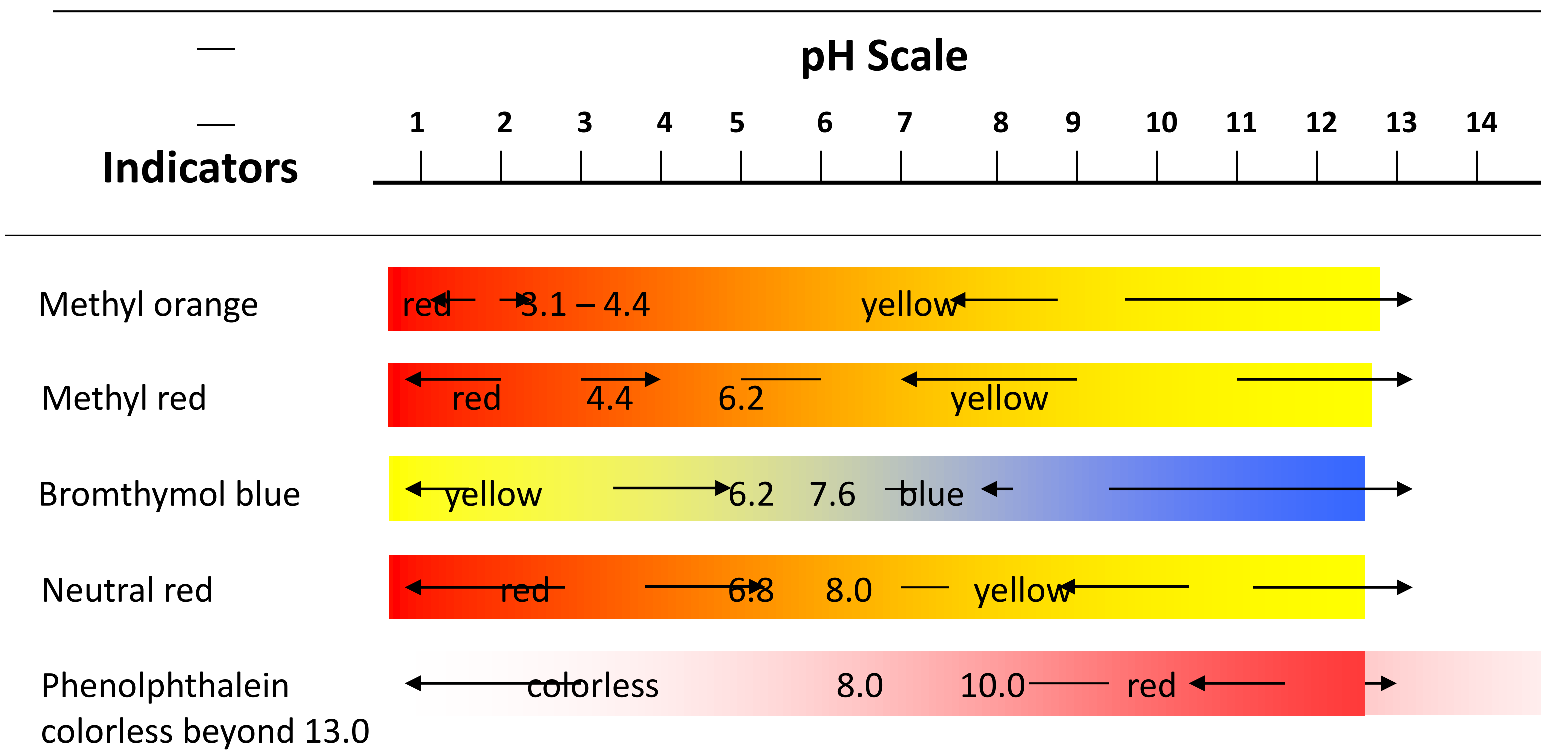
b) paper indicators

may be considered as weak acids or weak bases that act like buffers and also exhibit color changes as their degree of dissociation varies with pH.

— **For example**, methyl red shows its full alkaline color, yellow, at a pH of about 6 and its full acid color, red, at about pH 3.

2- **Electrometric method (pH meter).**

Range and Color Changes of Some Common Acid-Base Indicators



Bromthymol blue indicator would be used in titrating a strong acid with a strong base.
Phenolpthalein indicator would be used in titrating a weak acid with a strong base.
Methyl orange indicator would be used in titrating a strong acid with a weak base
The colour of an indicator is a function of the pH of the solution. The dissociation of an acidic indicator is given in simplified form as:

$$\underset{\substack{\text{Acid}_1 \\ \text{(Acid color)}}}{HIn} + \underset{\text{Base}_2}{H_2O} \rightleftharpoons \underset{\text{Acid}_2}{H_3O^+} + \underset{\substack{\text{Base}_1 \\ \text{(Alkaline color)}}}{In^-}$$

— HIn is the un-ionized form of the indicator, which gives the acid color, and In- is the ionized form, which produces the basic color.

— If an acid is added to a solution of the indicator, the hydrogen ion concentration term on the right-hand side of equation is increased, and the ionization is repressed by the common ion effect. The indicator is then **predominantly** in the form of **HIn, the acid color.**

— If base is added, $[H_3O^+]$ is reduced by reaction of the acid with the base, reaction proceeds to the right, yielding **more ionized indicator (In^-), and the base color is predominate.**

— Several indicators can be combined to yield so-called universal indicators just as buffers can be mixed to cover a wide pH range**(1-12)**

— Example of universal indicator is a mixture of **methyl yellow**, **methyl red, bromothymol blue ,thymol blue & phenolphthalein** which covers PH range 1-11.

Characteristics of colorimetric method

① less accurate

② less convenient but less expensive than electrometric method

③ difficult to apply for the un-buffered pharmaceutical preparation (change the pH -indicator itself is acids or base)

④ error may be introduced by the presence of salts & proteins

- **Various buffer systems have been suggested for different pharmaceutical solutions:**
- A. Sorensen phosphate
- B. Acetate buffer
- **Experimental work**

Part l: prepare

- 0.2 M HAC, (solution A)
- 0.2 M NaAC (Solution B)
- 0.1 M NaOH.
- **Part l1:**
- 1- prepares acetate buffer solution according to the following table then measure the pH in each case

Solution (A)ml + solution (B)ml $\xrightarrow{\text{D.W to 100ml}}$ pH

46.3	3.7
30.5	19.5
25.5	24.5
20	30
14.8	35.2
4.8	45.2

Part III

— measuring the pH, using pH meter: Put the electrode of the pH meter in the buffer solution & read the pH.

— Take a certain volume of acetate buffer solution; add 0.0004 M sodium hydroxide portions (0.1 ml of 0.1 M) to it. Then, measure the pH and calculate the buffer capacity.