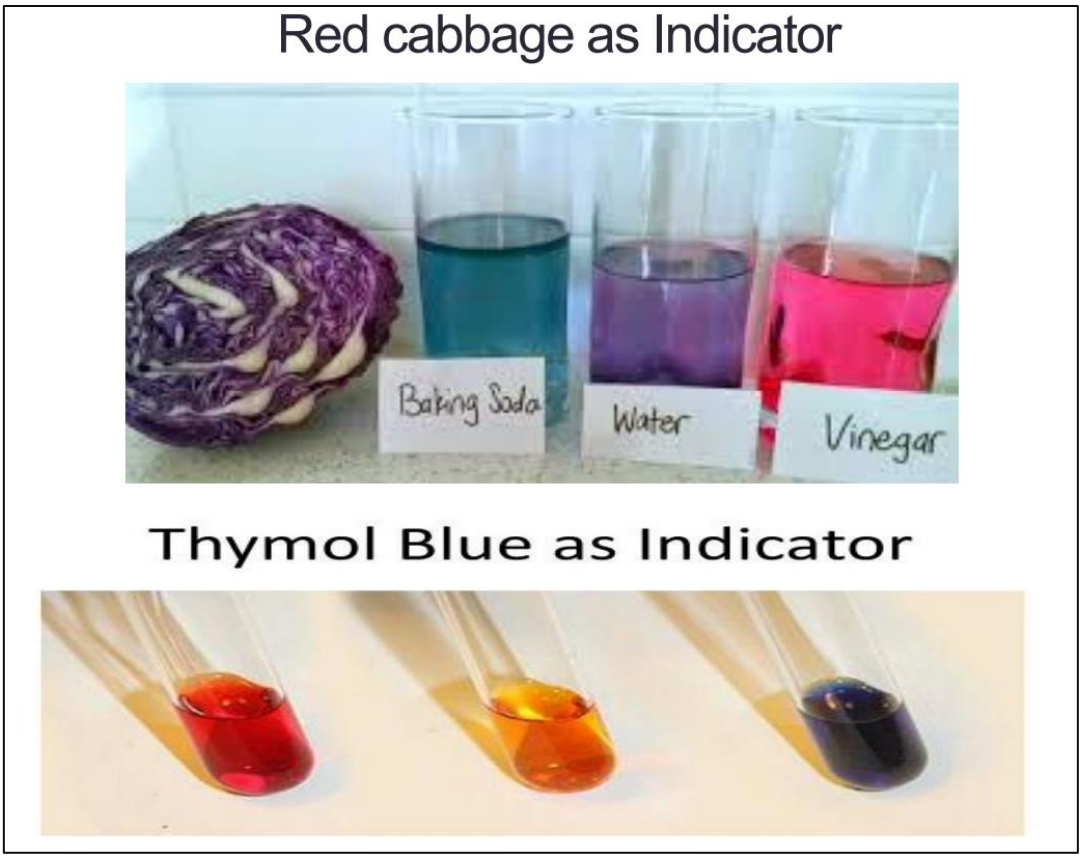
Al-Mustaqbal University College
Department of Medical Instrumentation Techniques Engineering
Class:1st
Subject: Medical Chemistry
Lecturer: Asst.lec. Hadeer Jasem
Lecture: 5

ACID-BASE INDICATORS

An acid-base indicator is either a weak acid or weak base that exhibits a color change as the concentration of hydrogen (H^+) or hydroxide (OH^-) ions changes in an aqueous solution. Acid-base indicators are most often used in a titration to identify the endpoint of an acid-base reaction. They are also used to gauge pH values and for interesting color-change science demonstrations.

Also Known As: pH indicator

Acid-Base Indicator Examples Perhaps the best-known pH indicator is litmus. Thymol Blue, Phenol Red, and Methyl Orange are all common acid-base indicators. Red cabbage can also be used as an acid-base indicator.



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Theories of Acid-base indicators

1. (Ostwald theory) According to this theory:

(a) The color change is due to ionization of the acid-base indicator. The unionized form has different color than the ionized form.

(b) The ionization of the indicator is largely affected in acids and bases as it is either a weak acid or a weak base. In case, the indicator is a weak acid, its ionization is very much low in acids due to common H^+ ions while it is fairly ionized in alkalis. Similarly if the indicator is a weak base, its ionization is large in acids and low in alkalis due to common OH^- ions.

2. (Quinonoid theory) According to this theory:

(a) The acid-base indicators exist in two tautomeric forms having different structures. Two forms are in equilibrium. One form is termed benzenoid form and the other Quinonoid form.

(b) The two forms have different colors. The color change in due to the inter conversation of one tautomeric form into other.

(c) One form mainly exists in acidic medium and the other in alkaline medium.

How an Acid-Base Indicator Works

If the indicator is a weak acid, the acid and its conjugate base are different colors. If the indicator is a weak base, the base, and its conjugate acid display different colors. For a weak acid indicator with the general formula HIn, equilibrium is reached in the solution according to the chemical equation:

$$HIn_{(aq)} + H2O_{(l)} \leftrightarrow In^-_{(aq)} + H3O^+_{(aq)}$$

HIn(aq) is the acid, which is a different color from the base **In^- (aq).** When the pH is low, the concentration of the hydronium ion H3O+ is high and equilibrium is toward the left. At high pH, the concentration of H3O+ is low, so equilibrium tends toward the right side of the equation.

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An example of a weak acid indicator is phenolphthalein, which is colorless as a weak acid but dissociates in water to form a magenta or red-purple anion. In an acidic solution, equilibrium is to the left, so the solution is colorless (too little magenta anion to be visible), but as pH increases, the equilibrium shifts to the right and the magenta color is visible. The equilibrium constant for the reaction may be determined using the equation:

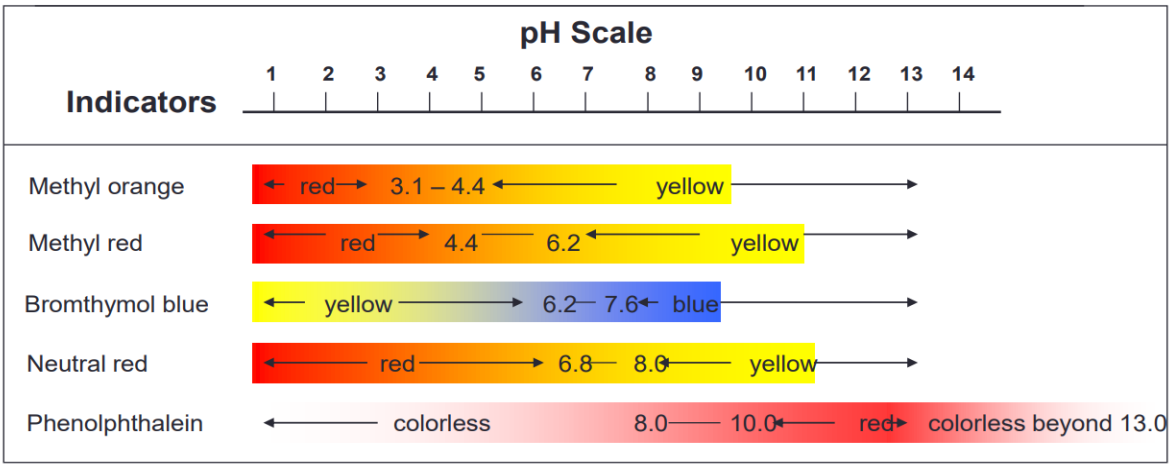
$$KIn = [H3O^{+}][In^{-}] / [HIn]$$

where KIn is the indicator dissociation constant. The color change occurs at the point where the concentration of the acid and anion base is equal:

$$[HIn] = [In^{-}]$$

which is the point where half of the indicator is in acid form and the other half is its conjugate base.

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.

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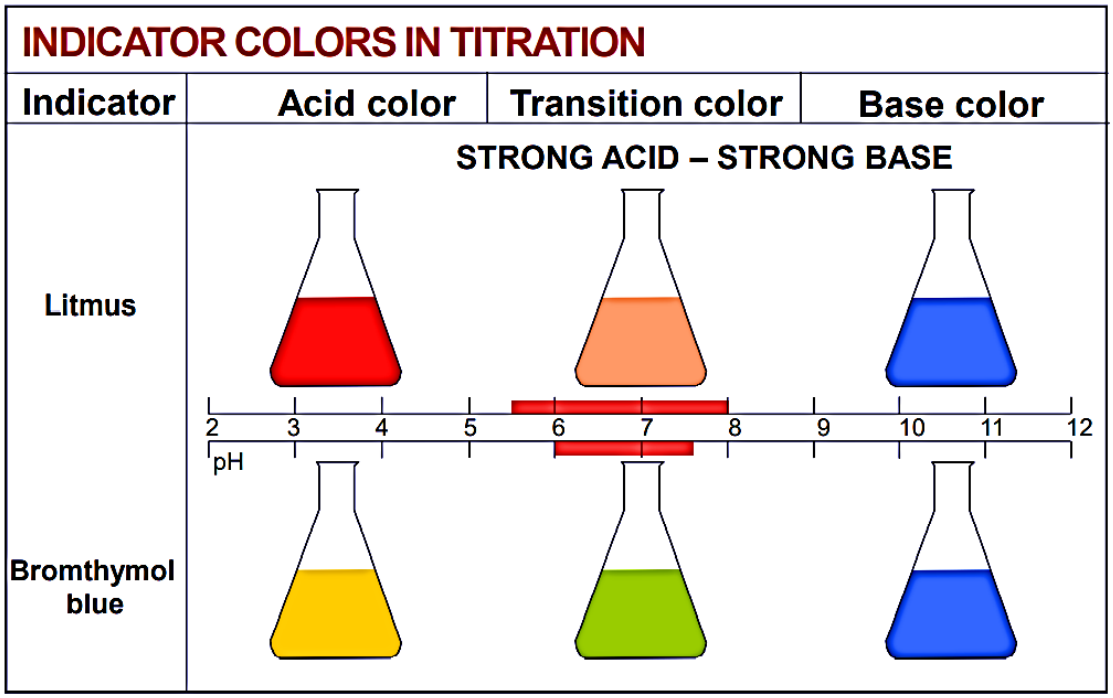

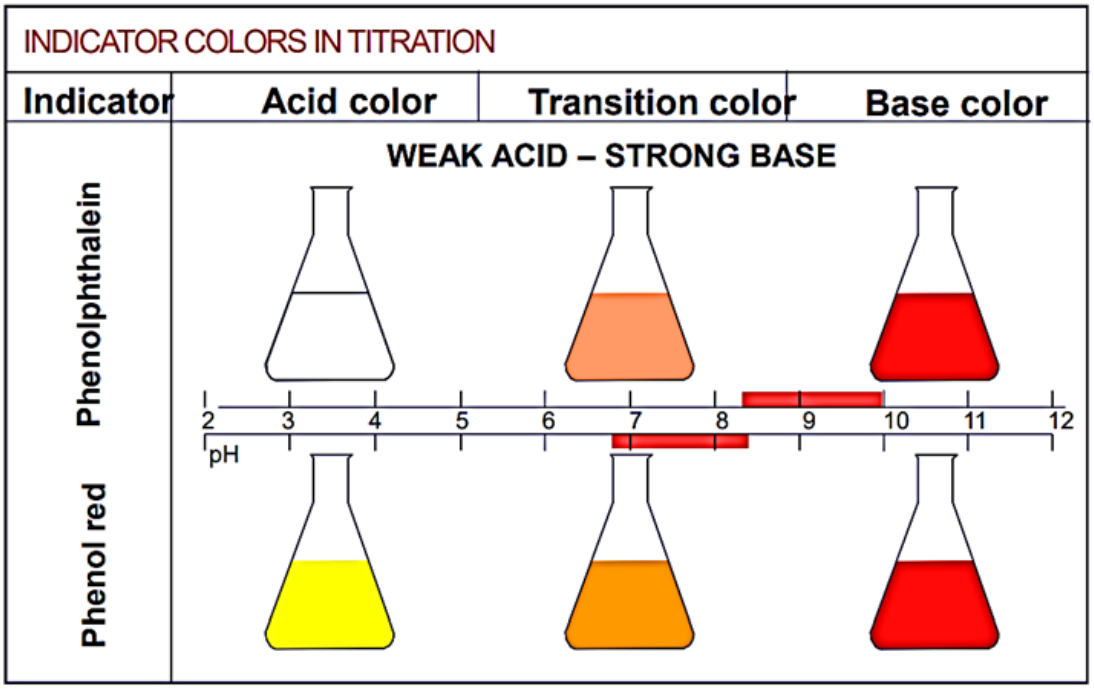

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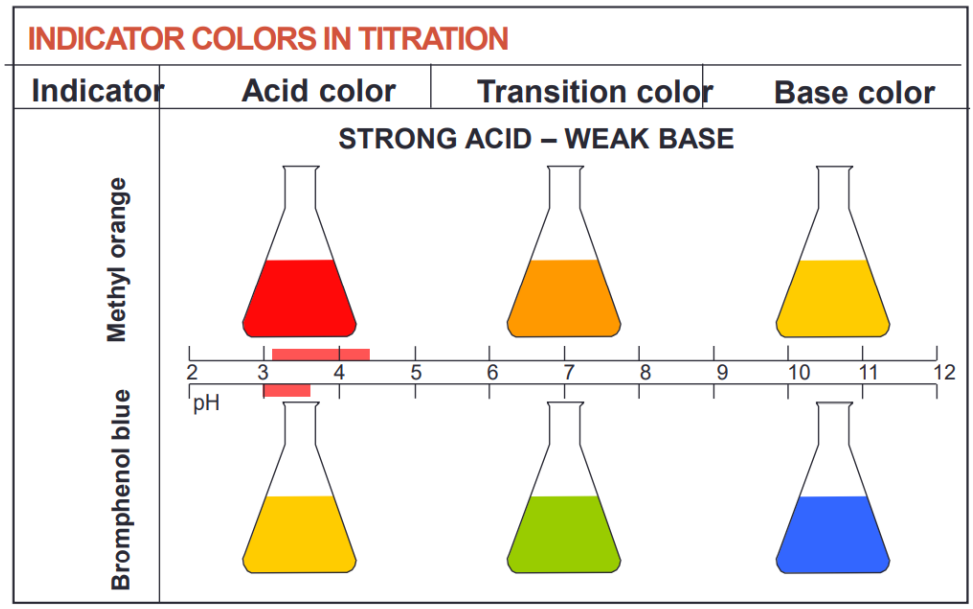

Indicator	Acid Color	Base Color	pH Range	pK_{In}
thymol blue (first change)	red	yellow	1.2 - 2.8	1.5
methyl orange	red	yellow	3.2 - 4.4	3.7
bromocresol green	yellow	blue	3.8 - 5.4	4.7
methyl red	yellow	red	4.8 - 6.0	5.1
bromothymol blue	yellow	blue	6.0 - 7.6	7.0
phenol red	yellow	red	6.8- 8.4	7.9
thymol blue (second change)	yellow	blue	8.0 - 9.6	8.9
phenolphthalein	colorless	magenta	8.2 -10.0	9.4

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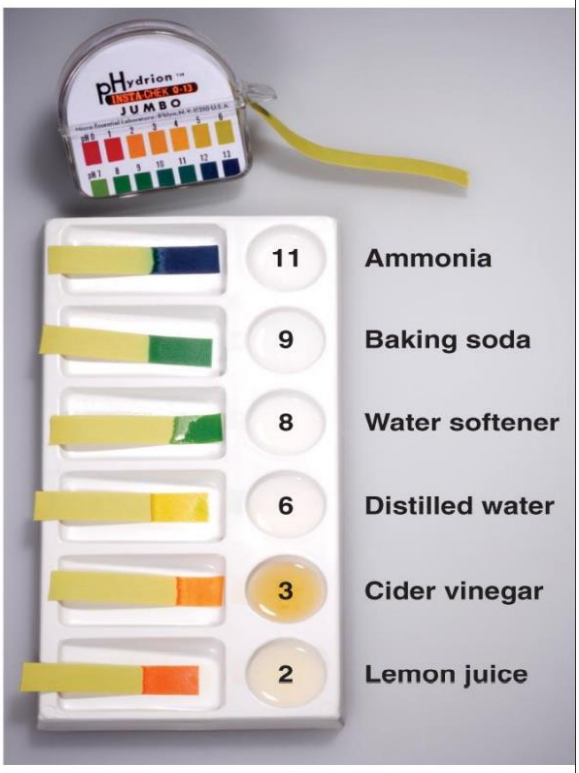
pH indicator

A pH indicator is a halochromic chemical compound added in small amounts to a solution so the pH (acidity or basicity) of the solution can be determined visually. Hence, a pH indicator is a chemical detector for hydronium ions ($H3O^+$) or hydrogen ions (H^+) in the Arrhenius model. Normally, the indicator causes the color of the solution to change depending on the pH. Indicators can also show change in other physical properties; for example, olfactory indicators show change in their odor. The pH value of a neutral solution is 7.0 at 25°C (standard laboratory conditions).

Solutions with a pH value below 7.0 are considered acidic and solutions with pH value above 7.0 are basic (alkaline). As most naturally occurring organic compounds are weak protolytes, carboxylic acids and amines, pH indicators find many applications in biology and analytical chemistry.



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Applications

pH indicators are frequently employed in titrations in analytical chemistry and biology to determine the extent of a chemical reaction. Because of the subjective choice (determination) of color. For applications requiring precise measurement of pH, a pH meter is frequently used. Sometimes, a blend of different indicators is used to achieve several smooth color changes over a wide range of pH values. These commercial indicators (e.g., universal indicator and Hydrion papers) are used when only rough knowledge of pH is necessary.