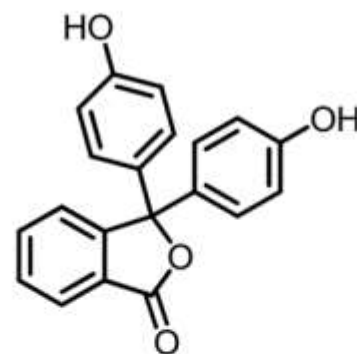
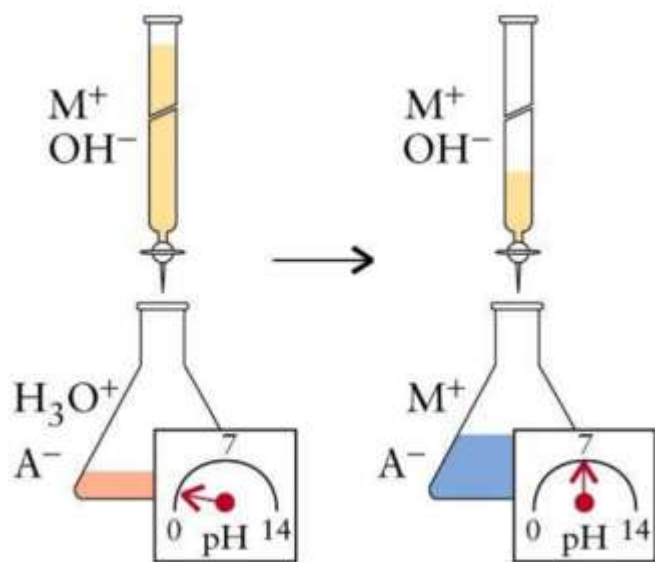


Acid-Base TITRATIONS



Amit Z Chaudhari

FLOW

INTRO

Acid – Base THEORIES

Law of MASS ACTION

IONIZATION PRODUCT of Water

Common ION EFFEC

pH of Buffer

HYDROLYSIS

END POINT DETECTION

INTRO

Defination:

- Titration involving reaction btwn acid and base

[sum15]

Acid

- Strong acid = completely ionized in water
- Weak acid

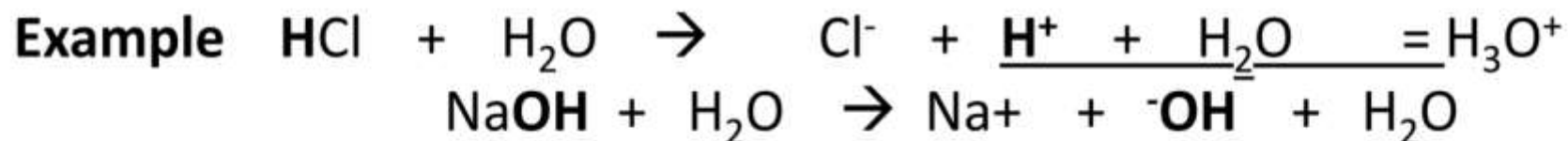
Base

- Strong base =
- Weak base = partially ionized in water e.g. NH_4OH

Acid – Base THEORIES

- Titration involving reaction btwn acid and base

1. Arrhenius concept



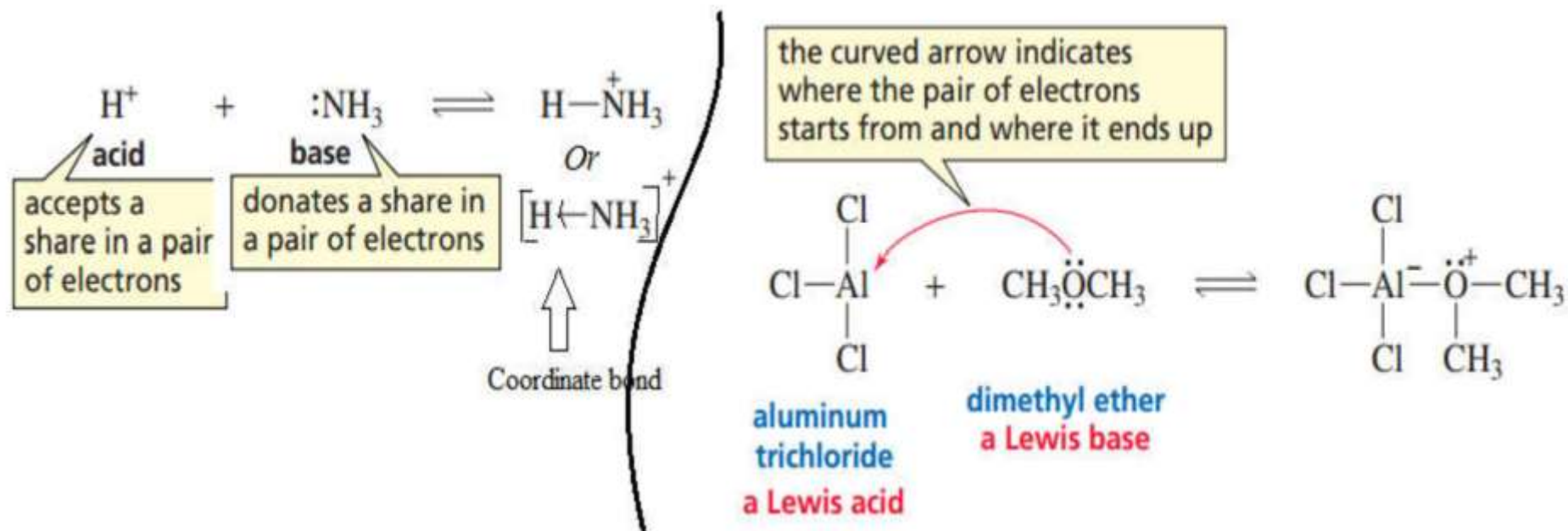
1. Bronsted- Lowry concept



Acid – Base THEORIES

3. Lewis concept

1923



Law of MASS ACTION

[win13, win12]

- In 1864, Guldberg and Waage proposed Law of Mass action statement

‘the rate of a chemical reaction is proportional to the active masses of the reactants involved’

- ‘active mass’ is meant the molar concentration (moles per litre)

Explanation

- Let us consider a general reaction,



- According to the Law of Mass action,

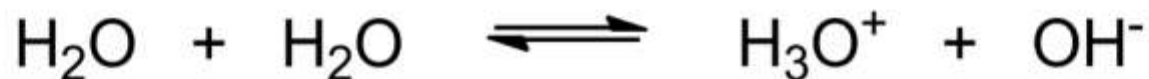
Rate of forward reaction $\propto [A][B] = k_1 [A][B]$

Rate of reverse reaction $\propto [C][D] = k_2 [C][D]$

Where k_1 and k_2 are rate constants for the forward and reverse rxns

IONIZATION PRODUCT of Water

- Amphoteric solvents undergo self-ionization, or autoprotolysis, to form a pair of ionic species.



- Ionization constant of water K_{ionw}

$$K_{\text{ionw}} = \frac{[\text{H}_3\text{O}^+] [\text{OH}^-]}{[\text{H}_2\text{O}]^2}$$

- The $[\text{H}_2\text{O}]$ is constant at 25 °C so combine to it with K_{ionw} gives new constant i.e. self ionization constant K_w

$$K_w = [\text{H}_3\text{O}^+] [\text{OH}^-]$$

IONIZATION PRODUCT of Water

[win13, win12]

- The experimental value for

$$[\text{H}_3\text{O}^+] * [\text{OH}^-] = 1 \times 10^{-14} \dots\dots\dots \text{Eq. 1}$$

- At equilibrium condition,

$$[\text{H}_3\text{O}^+] = [\text{OH}^-]$$

Adding this to Eq. 1

thus
$$[\text{H}_3\text{O}^+] * [\text{H}_3\text{O}^+] = 1 \times 10^{-14}$$
$$[\text{H}_3\text{O}^+] = 1 \times 10^{-7} \dots\dots\dots \text{Eq. 2}$$

Now $\text{pH} = -\text{Log}_{10}[\text{H}^+]$,

$$[\text{H}_3\text{O}^+] = 1 \times 10^{-7} \dots\dots\dots \text{Eq. 2}$$

- $\text{Log}_{10}[\text{H}^+] = 7$

- Conclusion at equilibrium ionization state water pH is 7

Common ION EFFEC

[win12]

Defination:

- The process by which ionization of comp. is suppressed by addition of one of its ion

E.g.



- Removal of proton or acetate will increase the rate of forward rxn.
- Addition of HCl or Sod. acetate decrease the rate of forward rxn.

Reason

- The system try to maintain equilibrium state of [] of all reaction component

pH of Buffer

[win15, win12]

Buffer:

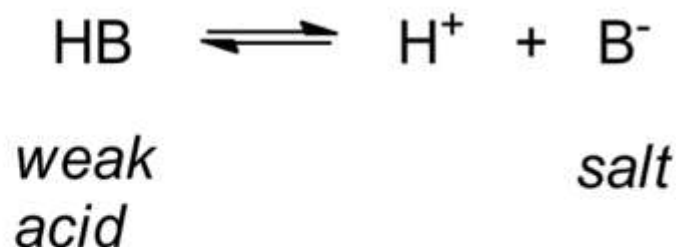
- A soln which resist drastic change in pH upon addition of small amt of acid or alkali.

Buffer composition:

- Weak acid / base and its salt

Finding pH of buffer:

- Suppose an example of **acidic buffer**,



pH of Buffer

- Dissociation constant is,

$$K_a = \frac{[H^+][B^-]}{[HB]}$$

$$\therefore [H^+] = K_a \frac{[HB]}{[B^-]} \quad \text{Henderson eq.n (1908)}$$

- Taking -ve log on both side,

$$\therefore -\log [H^+] = -\log K_a - \log \frac{[HB]}{[B^-]}$$

$$\therefore \text{pH} = \text{pKa} - \log \frac{[HB]}{[B^-]}$$

pH of Buffer

$$\therefore \text{pH} = \text{pK}_a - \log \frac{[\text{HB}]}{[\text{B}^-]}$$

$$\therefore \text{pH} = \text{pK}_a + \log \frac{[\text{B}^-]}{[\text{HB}]}$$

... Hasselbach work (1916)

For **basic buffer**,

$$\text{pOH} = \text{pK}_a + \log \frac{[\text{B}^+]}{[\text{BOH}]}$$

- Generalized the equation,

$$\therefore \text{pH} = \text{pK}_a + \log \frac{[\text{salt}]}{[\text{acid}]}$$

$$\therefore \text{pOH} = \text{pK}_a + \log \frac{[\text{salt}]}{[\text{base}]}$$

- This eq.n is called as *Henderson Hasselbach* eq.n .
- Eq.n gives relation btwn pH and pK_a and ration of acid vs salt

HYDROLYSIS

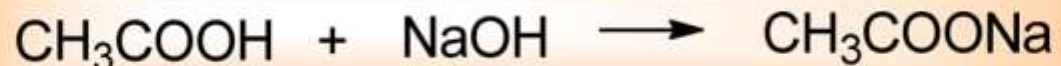
[win13, win12]

⇒ The interaction of sub. with ions of water

[win15]

E.g. of HL of salts:

(1) Salt derived from weak acid & strong base



HL of salt,



- Salt HL gives wa- , which is ionized (reverse rxn) to small degree.
- Thus [sod. acetate] decrease / more salt will get HL than reverse rxn.
- Thus more amt of OH^- are also produced
- Thus the soln will be **alkaline**.

HYDROLYSIS

pH calculation,

[win15]



- the HL constant is,

$$K_h = \frac{[\text{CH}_3\text{COOH}] [\text{OH}^-]}{[\text{CH}_3\text{COO}^-]} \quad \text{..... eq.1}$$

- Now, $K_w = [\text{H}^+] [\text{OH}^-]$ eq. 2

and HAc ionization constant is

$$K_a = \frac{[\text{CH}_3\text{COO}^-] [\text{H}^+]}{[\text{CH}_3\text{COOH}]} \quad \text{..... eq. 3}$$

HYDROLYSIS

- Combining eq. 1, 2 and 3...

$$K_h = \frac{[\text{CH}_3\text{COOH}][\text{OH}^-]}{[\text{CH}_3\text{COO}^-]} \text{ eq.1}$$

$$\frac{K_w}{[\text{H}^+]} = [\text{OH}^-] \text{ fromeq.2}$$

$$\frac{[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]} = \frac{[\text{H}^+]}{K_a} \text{ from eq. 3}$$

$$K_h = \frac{K_w}{[\text{H}^+]} \frac{[\text{H}^+]}{K_a}$$

$$\therefore K_h = \frac{K_w}{K_a} \text{eq. 4}$$

HYDROLYSIS



- The amt. of acetate and hydroxide ions formed are same, thus...

$$K_h = \frac{[\text{OH}^-]^2}{[\text{CH}_3\text{COO}^-]} \quad \text{..... eq. 5}$$

$$K_h = \frac{[\text{CH}_3\text{COOH}] [\text{OH}^-]}{[\text{CH}_3\text{COO}^-]} \quad \text{..... eq.1}$$

- Combining eq. 4 and 5,

$$\frac{[\text{OH}^-]^2}{[\text{CH}_3\text{COO}^-]} = \frac{K_w}{K_a}$$

$$\therefore [\text{OH}^-]^2 = \frac{K_w}{K_a} [\text{CH}_3\text{COO}^-]$$

HYDROLYSIS

$$\therefore [\text{OH}^-] = \frac{K_w^{1/2}}{K_a^{1/2}} [\text{CH}_3\text{COO}^-]^{1/2}$$

$$\therefore [\text{H}^+] = \frac{K_w}{\frac{K_w^{1/2}}{K_a^{1/2}} [\text{CH}_3\text{COO}^-]^{1/2}}$$

$$= \frac{K_w^{1/2} K_a^{1/2}}{[\text{CH}_3\text{COO}^-]^{1/2}}$$

$$\therefore K_w = [\text{H}^+] [\text{OH}^-]$$

- Taking -ve log on both side, gives eq. for pH of salt of wa- & sb-

$$\text{pH} = 1/2 \text{p}K_w + 1/2 \text{p}K_a - 1/2 \text{p}C$$

- Where C represents salt wa- [].

HYDROLYSIS

(2) Salt derived from weak acid & weak base



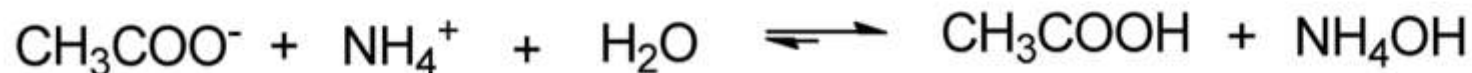
HL of salt,



- Salt HL gives wa- and wb-, which are ionized (reverse rxn) to small degree.
- Thus [amm. acetate] decrease *more rapidly* / more salt will get HL than reverse rxn.
- The soln will be **alkaline** or **acidic** depends on relative strenth of a- or b-.
- If $K_a = K_b$ than soln is neutral
- If $K_a > K_b$ than soln is acidic , , If $K_a < K_b$ than soln is basic

HYDROLYSIS

pH calculation,



- the HL constant is,

$$K_h = \frac{[\text{CH}_3\text{COOH}][\text{NH}_4\text{OH}]}{[\text{CH}_3\text{COO}^-][\text{NH}_4^+]} \quad \text{..... eq. 1}$$

- Now, $K_w = [\text{H}^+][\text{OH}^-] \quad \text{..... eq. 2}$

and HAc ionization constant is

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]} \quad \text{..... eq. 3}$$

and amm. hydroxide ionization constant is

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_4\text{OH}]} \quad \text{..... eq. 4}$$

HYDROLYSIS

- Combining eq. 1, 2, 3 and 4...

$$\therefore K_h = \frac{K_w}{K_a K_b} \quad \text{.....eq. 5}$$



$$K_h = \frac{[\text{CH}_3\text{COOH}][\text{NH}_4\text{OH}]}{[\text{CH}_3\text{COO}^-][\text{NH}_4^+]}$$

- During reaction, The amt. of reactant ions are in same amt and additionally amt of pdct ions are also same thus...

$$K_h = \frac{[\text{CH}_3\text{COOH}]^2}{[\text{CH}_3\text{COO}^-]^2}$$

$$K_h^{1/2} = \frac{[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]} \quad \text{..... eq. 6}$$

HYDROLYSIS

- Combining eq. 5 and 6

$$\frac{K_w^{1/2}}{K_a^{1/2} K_b^{1/2}} = \frac{[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]}$$

$$\therefore \frac{K_w^{1/2}}{K_a^{1/2} K_b^{1/2}} = \frac{[\text{H}^+]}{K_a}$$

$$\therefore \frac{K_a^{1/2} K_w^{1/2}}{K_b^{1/2}} = [\text{H}^+]$$

$$\therefore \frac{[\text{CH}_3\text{COOH}]}{[\text{CH}_3\text{COO}^-]} = \frac{[\text{H}^+]}{K_a} \quad \text{from eq. 3}$$

- Taking -ve log on both side, gives eq. for pH of salt of wa- & sb-

$$\text{pH} = 1/2 \text{pK}_w + 1/2 \text{pK}_a - 1/2 \text{pK}_b$$

HYDROLYSIS

(3) Salt derived from strong acid & strong base



- HL of salt,
- Salts are strong electrolytes so cant have tendency to combine with water ions
- Thus thus equilibrium btwn ^+H and ^-OH will not changed .
- Thus the soln will remains **neutral**.

HYDROLYSIS

(4) Salt derived from strong acid & weak base



HL of salt,



- Salt HL gives wb-, which are ionized (reverse rxn) to small degree.
- Thus [amm. chloride] decrease more rapidly / more salt will get HL than reverse rxn.
- Thus more amt of H^+ are also produced
- Thus the soln will be **acidic**

HYDROLYSIS

pH calculation,



[win15]

- the HL constant is,

$$K_h = \frac{[\text{NH}_4\text{OH}][\text{H}^+]}{[\text{NH}_4^+]} \quad \text{..... eq.1}$$

- Now, $K_w = [\text{H}^+][\text{OH}^-]$ eq. 2

and amm. hydroxide ionization constant is

$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_4\text{OH}]} \quad \text{..... eq. 3}$$

HYDROLYSIS

- Combining eq. 1, 2, and 3...

$$\therefore K_h = \frac{K_w}{K_b} \quad \text{.....eq. 4}$$



$$K_h = \frac{[\text{NH}_4\text{OH}][\text{H}^+]}{[\text{NH}_4^+]}$$

- The amt. of amm. Hydroxide and hydroxide ion formed are same thus from eq.1...

$$K_h = \frac{[\text{H}^+]^2}{[\text{NH}_4^+]} \quad \text{..... eq.5}$$

HYDROLYSIS

- Combining eq. 4 and 5

$$\therefore \frac{[\text{H}^+]^2}{[\text{NH}_4^+]} = \frac{K_w}{K_b}$$

$$\therefore [\text{H}^+] = \frac{K_w^{1/2}}{K_b^{1/2}} [\text{NH}_4^+]^{1/2}$$

- Taking -ve log on both side, gives eq. for pH of salt of wa- & sb-

$$\therefore \text{pH} = 1/2 \text{pK}_w + 1/2 \text{pC} - 1/2 \text{pK}_b$$

- Where C represents salt of wb- [] .

END POINT DETECTION

- Total two methods
 - (1) to use acid-base indicator (visual end point detection)
 - (2) Instrumental method

[win15]

1. VISUAL END POINT DETECTION

1.1 THEORY OF ACID-BASE INDICATOR

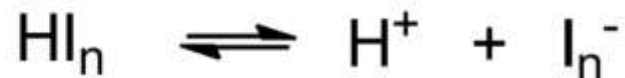
[win14, win13, win13]

1.1.1 Ostwald theory:

- Acc. to this theory,
'The undissociated acid [HI_n] / [I_nOH] base indicator has a colour diff. than its ion'

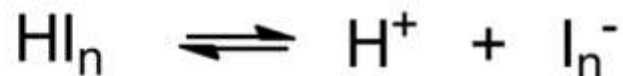
E.g.

1.1.1.1. Acid indicator equivalence,



END POINT DETECTION

- Acid indicator equivalence,



Conditions,

- Burtte : alkali
- Co. flask : acid soln
- Indicator : HI_n
- Here acid indicator added to acidic soln.
- Due to common ion effect, ionization of HI_n is suppressed by H^+ of acid soln. Thus before titration solution appears as the colour of HI_n
- As the titration proceeds, alkali neutralize H^+ in soln. , and thus rate of ionization of acid indicator increase.
- At the end pt of titration, that time soln gives the color of ionized indicator.

END POINT DETECTION

- Applying law of mass action,

$$K_{I_n a} = \frac{[H^+] [I_n^-]}{[HI_n]}$$

$$\therefore [H^+] = K_{I_n a} \frac{[HI_n]}{[I_n^-]}$$

- Taking -ve log on both side, \therefore

$$-\log [H^+] = -\log K_{I_n a} - \log \frac{[HI_n]}{[I_n^-]}$$

$$\therefore \text{pH} = \text{p}K_{I_n a} - \log \frac{[HI_n]}{[I_n^-]}$$

\therefore

$$\text{pH} = \text{p}K_{I_n a} + \log \frac{[I_n^-]}{[HI_n]}$$

END POINT DETECTION

$$\text{pH} = \text{pK}_{\text{In}}^{\text{a}} + \log \frac{[\text{I}_\text{n}^-]}{[\text{HI}_\text{n}]}$$

Experiment study shows that,

- If $\frac{[\text{I}_\text{n}^-]}{[\text{HI}_\text{n}]} = \frac{1}{10}$ ratio, then color of unionized indicator will appear. Here pH value becomes,

$$\boxed{\text{pH} = \text{pK}_{\text{In}}^{\text{a}} - 1} \quad \left[\because \log \frac{[\text{I}_\text{n}^-]}{[\text{HI}_\text{n}]} = \log \frac{1}{10} = -1 \right]$$

- If $\frac{[\text{I}_\text{n}^-]}{[\text{HI}_\text{n}]} = 10$ ratio, then ionized indicator color will appear.

Here pH value becomes,

$$\boxed{\text{pH} = \text{pK}_{\text{In}}^{\text{a}} + 1} \quad \left[\because \log \frac{[\text{I}_\text{n}^-]}{[\text{HI}_\text{n}]} = \log 10 = 1 \right]$$

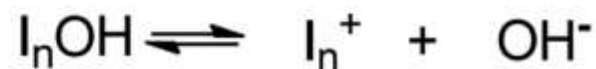
END POINT DETECTION

Conclusion,

- pH range for colour change is observed is $\text{pH} = \text{pK}_{\text{In}} \pm 1$
- This is the **transition interval** of the indicator.
- For particular titration, the indicator should be choose such that the pH equivalence pt. falls within the transition interval of indicator.

END POINT DETECTION

1.1.1.2. Base indicator equivalence,



- Applying law of mass action,

$$K_{I_n b} = \frac{[I_n^+] [\text{OH}^-]}{[I_n\text{OH}]}$$

$$\therefore [\text{OH}^-] = K_{I_n b} \frac{[I_n\text{OH}]}{[I_n^+]}$$

END POINT DETECTION

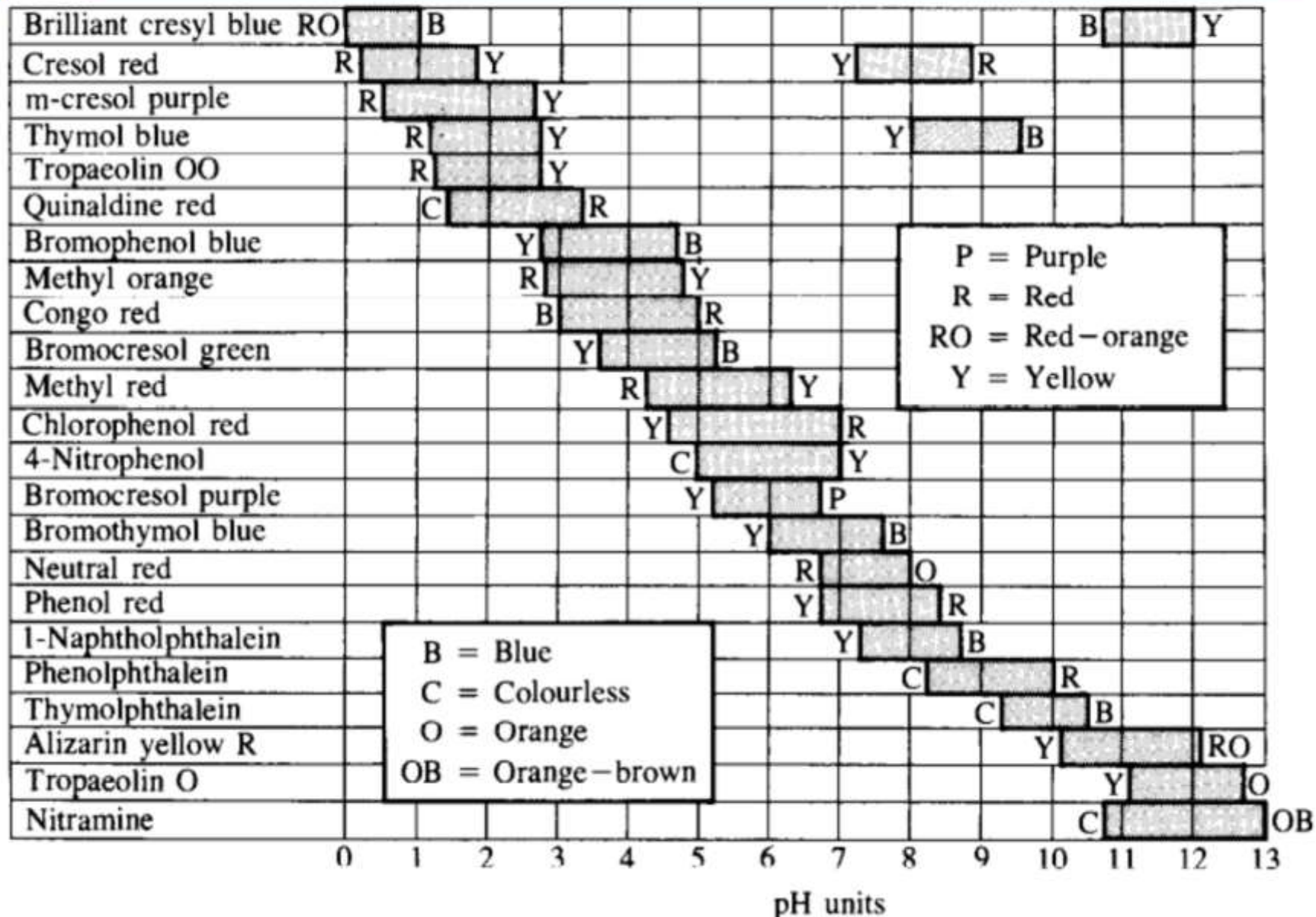
$$[H^+] = \frac{K_w}{K_{I_n b}} \frac{[I_n^+]}{[I_n OH]} \quad \left[\because [H^+] = \frac{K_w}{[OH^-]} \right]$$

- Taking -ve log on both side,

$$\therefore -\log [H^+] = \frac{-\log K_w}{-\log K_{I_n b}} - \log \frac{[I_n^+]}{[I_n OH]}$$

$$\therefore pH = pK_w - pK_{I_n b} + \log \frac{[I_n^+]}{[I_n OH]}$$

END POINT DETECTION



END POINT DETECTION

1.1.2 Resonance theory:

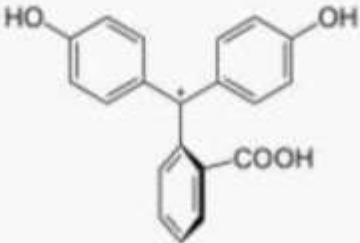
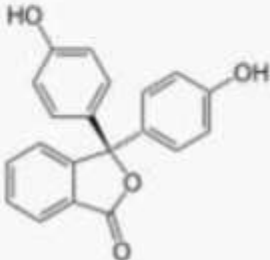
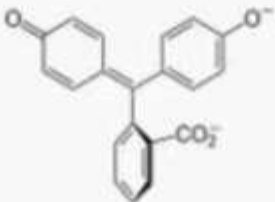
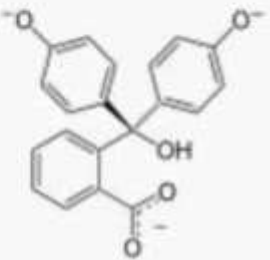
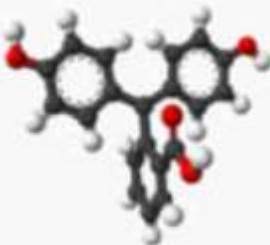
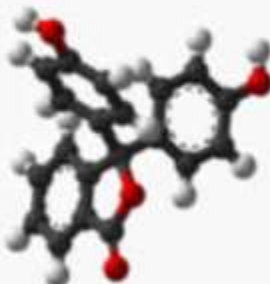
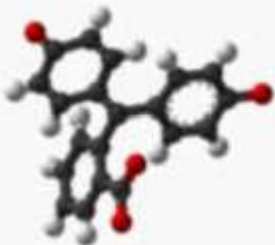
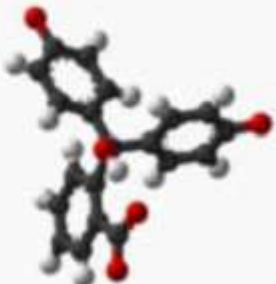
- Acc. to this theory,
'the diff. in color of the indicator in acid and base medium is apparently due to diff. in its organic structure of two forms '
- Color of comp. (indicator) is due to absorption of visible light, the absorption and emission of light depends on electronic structure of compound
- Change in electronic structure will change color of compound (indicator)

E.g.

1.1.2.1. Phenolphthalein,

Phenolphthalein is colorless below pH 8.3 and above pH 12. [win13] [sum15]

END POINT DETECTION

| Species | H_3In^+ | H_2In | In^{2-} | $In(OH)^{3-}$ |
|------------|---|--|---|---|
| Structure |  |  |  |  |
| Model |  |  |  |  |
| pH | <0 | 0–8.2 | 8.2–12.0 | >13.0 |
| Conditions | strongly acidic | acidic or near-neutral | basic | strongly basic |
| Color | orange | colorless | pink to fuchsia | colorless |
| forms of | Cation | Benzenoid | Quinonoid | Open Benzenoid |

END POINT DETECTION

1.2 MIXED INDICATOR

- Use of two indicator
- To obtain sharp color change
- Both indicator selected in such a way that their K_{in} are close.

| Indicator | pH | Color change |
|-----------------------------------|-----|-----------------------|
| Neutral red + Methylene blue | 7 | Violate-blue to green |
| Bromocresol green + Methyl orange | 4.3 | Orange to blue-green |

END POINT DETECTION

1.3 UNIVERSAL INDICATOR

or

MULTIPLE RANGE INDICATOR

- Mixture of various indicators

⇒ Is a mixture of suitable indicators, color change may be obtained over a considerable portion of the pH range.

E.g.

- Mixture: 0.1g Phth, 0.2 Me. red, 0.3 me. yellow, 0.4g bromothy. blue, 0.5g thy. blue in 500ml absolute alcohol and addition of NaOH soln to obtained color up to yellow.
- Color change:

| pH 2 | 4 | 6 | 8 | 10 |
|------|--------|--------|-------|------|
| Red | Orange | Yellow | Green | Blue |

END POINT DETECTION

2. INSTRUMENTAL METHOD

- Two methods

2.1 Conductometric titration

- Measure the change in the conductance during titration

2.2. pH metric titration

- Measure pH during titration
- Calomel e-de is used as reference e-de and glass or H e-de is used as indicator electrode.
- Plot of ml of titrant vs pH gives sigmoidal shape curve where end point can be detected.

CALCULATIONS – Neutralization curve

- Q.3** (a) 50 ml of 0.1M acetic acid is titrated with 0.1 M sodium hydroxide. Calculate the pH of solution when following volume of NaOH is added, 1) 0 ml, 2) 10 ml, 3) 25 ml, 4) 50 ml. K_a of acetic acid = 1.8×10^{-5} **06**

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]}$$

[win12]

Before titration , added 0.1N NaOH = 0 ml

- Here ,

$$[\text{CH}_3\text{COO}^-] = [\text{H}^+]$$

$$K_a = \frac{[\text{H}^+]^2}{[\text{CH}_3\text{COOH}]}$$

$$\therefore 1.82 \times 10^{-5} = \frac{[\text{H}^+]^2}{0.1}$$

CALCULATIONS – Neutralization curve

$$\therefore [\text{H}^+]^2 = 1.82 \times 10^{-6}$$

$$\therefore [\text{H}^+] = 0.001349074$$

$$\therefore \text{pH} = 2.87$$

added 0.1N NaOH = 10 ml

- All NaOH neutralize acetic acid to form acetate ions
- Thus produced acetate,

$$[\text{CH}_3\text{COO}^-] = \frac{10 \text{ ml} \times 0.1 \text{ M}}{60 \text{ ml}} = 0.017$$

- Now acetic acid remained,

$$[\text{CH}_3\text{COOH}] = \frac{40 \text{ ml} \times 0.1 \text{ M}}{60 \text{ ml}} = 0.067$$

CALCULATIONS – Neutralization curve

added 0.1N NaOH = 10 ml

- All NaOH neutralize acetic acid to form acetate ions
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$$[\text{CH}_3\text{COO}^-] = \frac{10 \text{ ml} \times 0.1 \text{ M}}{60 \text{ ml}} = 0.017$$

- Now acetic acid remained,

$$[\text{CH}_3\text{COOH}] = \frac{40 \text{ ml} \times 0.1 \text{ M}}{60 \text{ ml}} = 0.067$$

Now ,

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]} \quad 1.82 \times 10^{-5} = \frac{0.017 [\text{H}^+]}{0.067}$$

$$[\text{H}^+] = 0.00007173$$

$$\text{pH} = 4.14$$

CALCULATIONS – Neutralization curve

added 0.1N NaOH = 25 ml

- All NaOH neutralize acetic acid to form acetate ions
- Thus produced acetate,

$$[\text{CH}_3\text{COO}^-] = \frac{25 \text{ ml} \times 0.1 \text{ M}}{75 \text{ ml}} = 0.033$$

- Now acetic acid remained,

$$[\text{CH}_3\text{COOH}] = \frac{25 \text{ ml} \times 0.1 \text{ M}}{75 \text{ ml}} = 0.033$$

$$\text{Now, } K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]} \quad 1.82 \times 10^{-5} = \frac{0.033 [\text{H}^+]}{0.033}$$

$$[\text{H}^+] = 0.0000182$$

$$\text{pH} = 4.73$$

CALCULATIONS – Neutralization curve

added 0.1N NaOH = 50 ml

- All NaOH neutralize acetic acid to form acetate ions
- Thus produced acetate,

$$[\text{CH}_3\text{COO}^-] = \frac{50 \text{ ml} \times 0.1 \text{ M}}{100 \text{ ml}} = 0.05$$

- Now acetic acid remained,

$$[\text{CH}_3\text{COOH}] = \frac{0 \text{ ml} \times 0.1 \text{ M}}{75 \text{ ml}} = 0$$

$$\text{Now, } K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}^+]}{[\text{CH}_3\text{COOH}]} \quad 1.82 \times 10^{-5} = \frac{0.05 [\text{H}^+]}{0}$$

$$[\text{H}^+] = 0$$

$$\text{pH} = \text{!!!!}$$

CALCULATIONS – Neutralization curve

Now ,

- Applying salt of weak acid and strong base pH eq.

$$\text{pH} = 1/2 \text{pK}_w + 1/2 \text{pK}_a - 1/2 \text{pC}$$

$$= 1/2 \ 14 + 1/2 [-\log (1.8 * 10^{-5})] - 1/2 [-\log (0.05)]$$

$$= 7 + 2.37 - 0.65$$

$$= 8.72$$

End of topic

Reference:

Ph'cal Analysis-I – Dr. Dimal Shah

Quantitative Chemical Ana.- VOGEL's

Ph'cal Analysis – Dr. A.V. Kasture