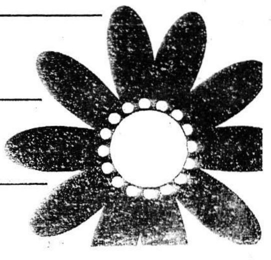
Lec (4)

principles of Neutralization Titrations.

Neutralization titrations are widely employed to determine the amounts of acids and bases Like all titrations, neutralization titrations depend on achemical reaction between the analyte and a standard reagent.

The standard reagents used in acid / base titrations are always strong acids or strong bases, most commonly HCl, $HClO_4$, H_2SO_4, $NaOH$, KOH



Titration Curve.

Titration of strong acids and strong Bases.

In Solutions of astrong acid that are more concentrated than about 1×10^{-6} M, we can assume that the equilibrium concentration of H_3O^+ is equal to the analytical concentration of the acid. The same is true for $[OH^-]$ in Solutions of strong bases.

$$[H_3O^+] = C_{HCl} + [OH^-] \approx C_{HCl}$$

↑ يهمل لأنه معدوم تقريباً

$$[OH^-] = C_{NaOH} + [H_3O^+] \approx C_{NaOH}$$

↗ يهمل

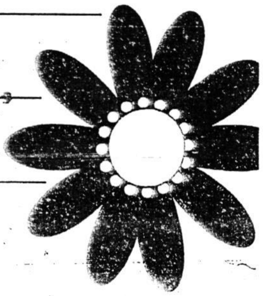
$$HCl \longrightarrow H^+ + Cl^-$$

Titrating a strong acid with a strong Base.

Three types of Calculations must be done to construct the hypothetical curve for titrating a Solution of a strong acid with a strong base

① pre equivalence
② equivalence
③ post equivalence

In the pre equivalence stage, we compute the concentration of the acid from it's starting concentration and the amount of base added. At the equivalence point, the hydronium and hydroxide ions are present in equal concentration, and the hydronium ion concentration is derived directly from the ion-product constant for water.



In the post equivalence stage. the analytical concentration of the excess base is computed, and the hydroxide ion concentration is assumed to be equal to or a multiple of the analytical concentration.

Ex : [illegible]

Generate the hypothetical titration curve for the titration of 50 ml of 0.05 M HCl with 0.1 M NaOH.

Initial point

Before any base is added, the Solution is 0.05 M in H_3O^+

$pH = -\log [H_3O^+] = -\log 0.05 = 1.3$

After addition of 10 ml of the reagent ⟵ (NaOH)

The hydronium ion concentration is decreased as a result of both reaction with the base and dilution. So the analytical calculations of the acid concentration which it residue in the solution carried out by:

$$C_{HCl} = \frac{\text{no. mmol HCl remaining after addition of NaOH}}{\text{total Volume Solution}}$$

$$= \frac{\text{original no. mmol HCl} - \text{no. mmol NaOH added}}{\text{total Volume Solution}}$$

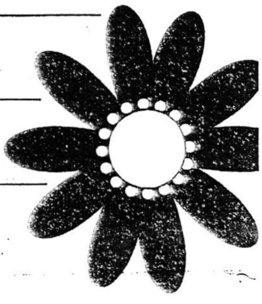
$$= \frac{(50\ \text{ml} \times 0.05\ \text{ml}) - (10\ \text{ml} \times 0.1\ \text{M})}{50\ \text{ml} + 10\ \text{ml}}$$

$$= \frac{2.5\ \text{mmol} - 1\ \text{mmol}}{60\ \text{ml}} = 2.5 * 10^{-2}\ \text{M}$$

$[H_3O^+] = 2.5 * 10^{-2}$ M

and $pH = -\log [H_3O^+] = -\log (2.5 \times 10^{-2})$

$= 1.60$



After addition of 25 ml of Reagent :-

The equivalence point

At the equivalence point, neither HCl nor NaoH is in excess, and so the concentration of hydronium and hydroxide ions must be equal. Substituting this equality into the ion product constant for water yields

$$[H_3O^+] = \sqrt{K_w} = \sqrt{1 \times 10^{-14}} = 1 \times 10^{-7}$$

$$pH = -\log [H^+] = -\log 10^{-7} = 7$$

After addition of 25.1 ml of Reagent. The Solution now contain excess of NaoH, and we can write

$$C_{NaOH} = \frac{\text{no. mmol NaoH added} - \text{original no. mmol HCl}}{\text{total volum soln}}$$

$$= \frac{25.1 \times 0.1 - 50 \times 0,05}{75.10} = 1.33 \times 10^{-4}\ M$$