ACID BASE CONCEPTS: TITRATIONS

Q. 1. **Describe titration curve of a strong acid and strong Base.**

Ans. Let us consider HCl and NaOH reaction.

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

or $$H^+ + OH^- \longrightarrow H_2O$$

- Take 50 ml 0.1 N HCl and add gradually 0.1 NaOH. As reaction proceeds, more and more H^+ from HCl combines with OH^- from NaOH and H^+ will go an decreasing.
- When exactly 50 ml of NaOH has been added, neutralization reaction between both is just complete. Reaction mixture is neutral with pH 7 with formation of NaCl which is salt of strong acid and strong base and it will not under go hydrolysis and pH will not be disturbed. From the previous/ above statement it appears that the indicator which shows color change exactly at pH 7 should be selected. But it is too early to decide the indicator. To know suitable indicator for this titration let us know how pH changes during the titration.

Change in pH during titration of strong acid and strong Base :-

Consider titration of 50 ml 0.1 N HCl with 0.1 N NaOH and calculate its pH at different stages.

(i) pH before addition of any NaOH :-

We have 0.1 N HCl

$\therefore$ $[H^+]$ = 0.1000 g ion/lit = 1/10 g ion/lit OR 10^{-1} g/lit

or log $[H^+]$ = log 10^{-1}

or log $[H^+]$ = -1

or –log $[H^+]$ = -(-1) or

pH = 1

(ii) pH after addition of 10 ml NaOH

Now unreacted HCl remaining will be = 40 ml

and Total volume of solution in Flask = 60 ml

$\therefore$ Normality of HCl is calculated by $N_1V_1 = N_2V_2$

$N_{HCl} \times 60 = 0.1 \times 40$

(Remaining)

$$N_{HCl} = \frac{40 \times 0.1}{60} = 0.0667\ N$$

$\therefore$ 0.0667 g equivalent of HCl is present is 1 lit.

or $[H^+]$= 0.0667 g ion/lit

$\therefore$ - log $[H^+]$ = - (log 0.0667)

or = - (2.82410 = - [- 2.0000 + 0.8241]

pH = - (-1.1759) = 1.1759 = 1.18

$\therefore$ pH = 1.18

$\therefore$ pH rises from 1 to 1.18 when 20% reaction is complete.

(iii) pH after addition of 20 ml NaOH

unreacted HCl remaining will be now 30 ml and Total volume of solution become 70 ml.

$\therefore$ Normality of HCl:- $N_1V_1 = N_2V_2$

$N_{HCl} \times 70 = 0.1 \times 30$

$$N_{HCl} = \frac{0.1 \times 30}{70} = 0.0428\ \text{g ion/lit}$$

pH = - (log 0.428) = 1.37

$\therefore$ pH = 1.37 when reaction is 40% complete.

(iv) pH after addition of 50 ml NaOH

At this stage reaction is just complete i.e. equivalence point is reached.

We have started with 50 ml x 0.100 mmol/ml = 5 mmol HCl

And Now added 50 ml x 0.100 m mol/ml = 5 mmol NaOH

$$H_3O^+ + OH^- \longrightarrow 2H_2O$$

Then equilibrium is $2H_2O \rightleftharpoons H_3O^+ + OH^-$

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

$[H_3O^+] = [OH^-]$

or $[H_3O^+]^2 = 1 \times 10^{-14}$ or $[H_3O^+] = 1 \times 10^{-7}$ or

$\therefore$ pH = 7

(v) pH after addition of 60 ml base:

We started with 50 ml x 0.100 m mol/ml = 5 m mol HCl

Now added 60 ml x 0.100 m mol/ml = 6 m mol NaOH

$$H_3O^+ + OH^- \longrightarrow 2H_2O$$

$$\therefore [OH^-] = \frac{1}{110} = 9.1 \times 10^{-3} M$$

POH = 3 – log 9.1 = 2.04

Or pH = 14.00 – 2.04 = 11.96

The pH values at other points in titration

NaOH added (in ml)	Volume of solution	pH
0.0	50	1
10	60	1.18
20	70	1.37
25	75	1.48
30	80.0	1.60
40	90.0	1.95
49	99.0	3.00
49.90	99.9	4.00
49.95	100.0	4.30
50.05	100.00	4.30
50.05	100.05	9.70
50.10	100.10	10.00
51.00	101.0	11.00
60.00	110.00	11.96
70.00	120.0	12.23

Initially pH rises only gradually as the titrant is added and rises more rapidly as the equivalence point is approached. The pH increases by about 5.40 units for addition of only 0.1 ml of base at equivalence point & pH again increases only slowly as titrant is added after end point.