

pH after addition of 50.01 ml of
Base.

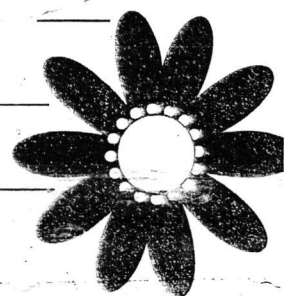
After the addition of 50.01 ml of NaOH, both the excess base and acetate ion are sources of hydroxide ion.

The contribution from the acetate ion is small, however, because the excess of strong base represses the reaction of acetate with water when we have

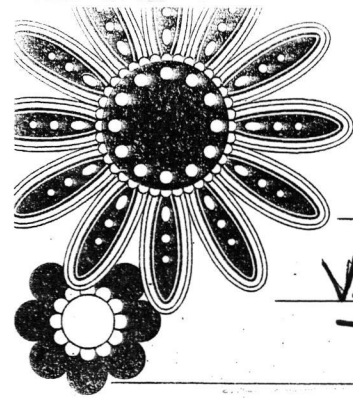
$$[\text{OH}^-] \approx C_{\text{NaOH}} = \frac{50.01 \text{ ml} \times 0.1 \text{ M} - 50 \text{ ml} \times 0.1 \text{ M}}{100.1 \text{ ml}} \\ = 1 \times 10^{-5} \text{ M}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log (1 \times 10^{-5}) = 5$$

$$\text{pH} = \text{pK}_w - \text{pOH} \\ = 14 - 5 = 9$$



3.6 =



V. of NaOH ml

pH
50 ml of 0.1 M HOAc
with 0.1 M NaOH

0.00

2.88

5.00

3.67

25.00

4.76

49.90

7.46

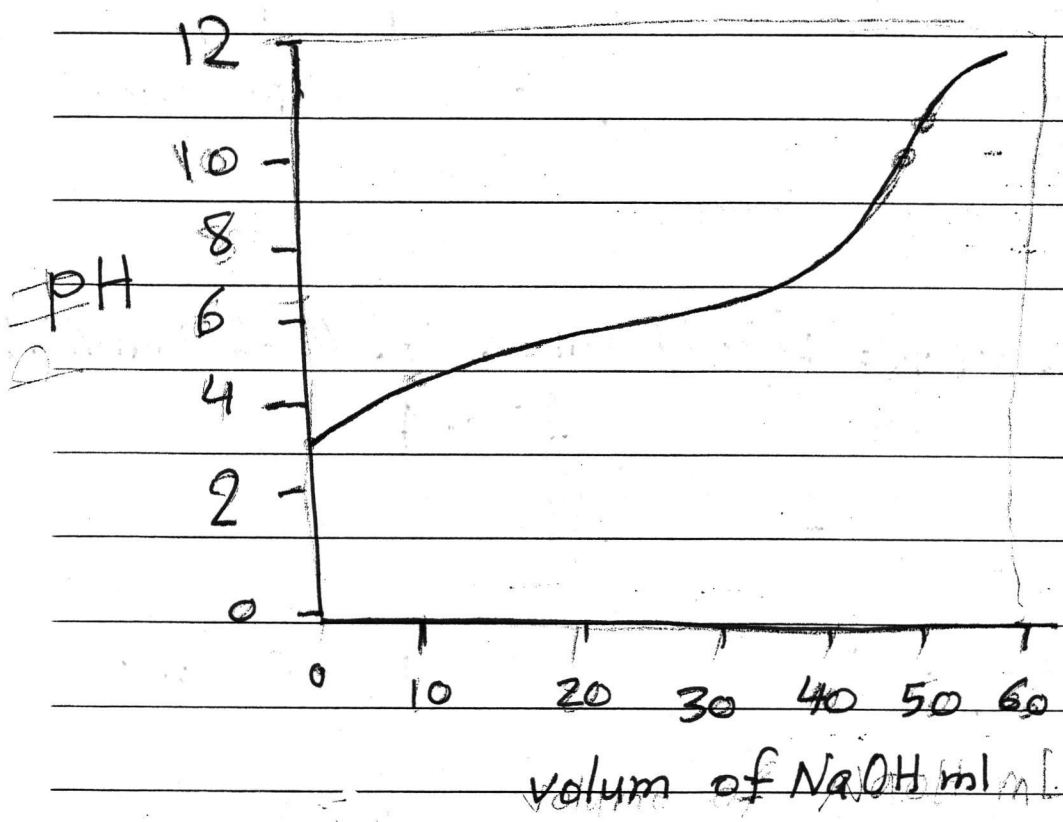
50.00

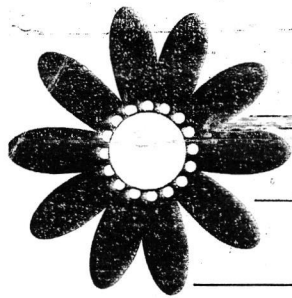
8.73

50.1

10.00

$50 \times 0.05 \neq 1$
95





(53)

$$[CN^-] = [NaCN] - [OH^-] \approx [NaCN] = 0.05 M$$
$$[OH^-] = [HCN]$$

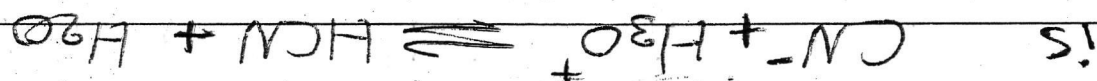
$$K_b = \frac{[OH^-][HCN]}{[CN^-]} = \frac{K_w}{K_a} = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-10}} = 1.61 \times 10^{-5}$$
$$CN^- + H_2O \rightleftharpoons HCN + OH^-$$

(a) 0.00 ml of Reagent

ml of acid

of (a) 0.00, (b) 10.00, (c) 25.00, and (d) 26.00

Calculate the pH after the addition




titrated with 0.1 M HCl. The reaction

A 50 ml aliquot of 0.05 M NaCN is

EX:-

The calculations needed to draw the titration curve for a weak base are analogous to those for a weak acid.

Titration Curve for weak Base with strong Acid



Substitution into the dissociation constant expression gives,

$$[\text{OH}^-] = \sqrt{K_b C_{\text{NaCN}}} = \sqrt{1.61 \times 10^{-5} \times 0.05}$$
$$= 8.97 \times 10^{-4}$$

$$\text{pOH} = -\log [\text{OH}^-] = -\log (8.97 \times 10^{-4}) = 3.05$$

$$\text{pH} = \text{p}K_w - \text{pOH} = 14 - 3.05 = 10.95$$

(b) 10 ml of Reagent

Addition of acid produces a buffer with a composition given by:

$$C_{\text{NaCN}} = \frac{\text{original no. mmol NaCN} - \text{no. mmol HCl added}}{\text{total volume soln.}}$$

$$C_{\text{NaCN}} = \frac{50 \text{ ml} \times 0.05 \text{ M} - 10 \text{ ml} \times 0.1 \text{ M}}{60 \text{ ml}}$$

$$= 0.025 \text{ M}$$

$$C_{\text{CN}^-} = \frac{10 \text{ ml} \times 0.1 \text{ M}}{60 \text{ ml}} = 0.017 \text{ M}$$

These values are then substituted into equation of the pH of basic buffer.

$$pH = pK_a + \log \frac{[salt]}{[acid]}$$

$$pH = -\log (6.2 \times 10^{-10}) + \log \frac{[0.025]}{[0.017]}$$
$$pH = 9.04$$

(C) 25.00 ml of Reagent

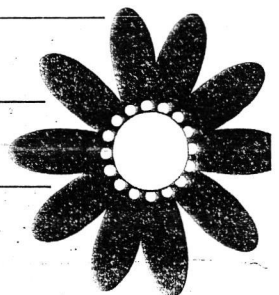
This volume corresponds to the equivalence point, where the principal solute species is the weak acid HCN.

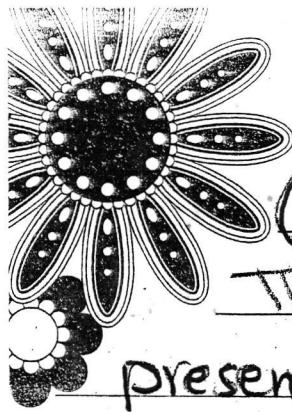
Thus:

$$C_{HCN} = \frac{25 \text{ ml} \times 0.1 \text{ M}}{75 \text{ ml}} = 0.033 \text{ M}$$

$$[H_3O^+] = \sqrt{K_a C_{HCN}} = \sqrt{6.2 \times 10^{-10} \times 0.033}$$
$$= 4.45 \times 10^{-6} \text{ M}$$

$$pH = -\log [H_3O^+] = -\log (4.45 \times 10^{-6})$$
$$= 5.34$$





(d) 26.00 ml of Reagent

The excess of strong acid now present represses the dissociation of the HCN, to the point where its contribution to the pH is negligible.

Thus,

$$[H_3O^+] = C_{HCl} = \frac{\text{no. mmol HCl added} - \text{original no. mmol NaCN}}{\text{total volume Soln.}}$$

$$= \frac{26.00 \times 0.1 - 50 \text{ ml} \times 0.05}{76} = 1.13 \times 10^{-3} \text{ M}$$

$$\text{pH} = -\log [H_3O^+] = -\log (1.13 \times 10^{-3}) = 2.90$$

<u>Volume of HCl</u>	<u>pH</u>
0	10.25
10	9.4
15	8.3
20	7.3
25	5.34
26	2.90

Q // Try to drawing the curve?