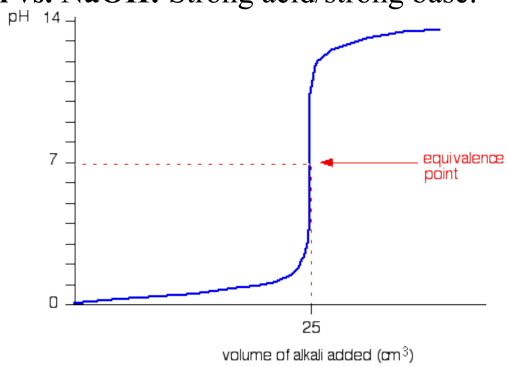
Ch. 16.4-16.5: Acid-Base Titrations and Indicators

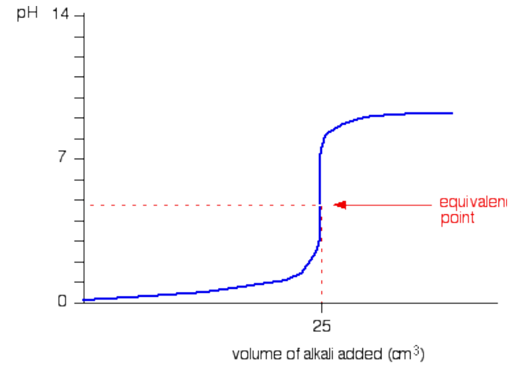
Homework 16-5: Problems pg. 722 #22, 24, 29, 35; Additional Problem: How many mL of 0.0850 M NaOH are required to titrate 25.0 mL of 0.128 M $CH_2ClCOOH$ (chloroacetic acid, $K_a = 1.38\times10^{-3}$) to the equivalence point? What is the pH of the solution at the equivalence point?

22 Sketch titration curves for the following acid-base titrations:

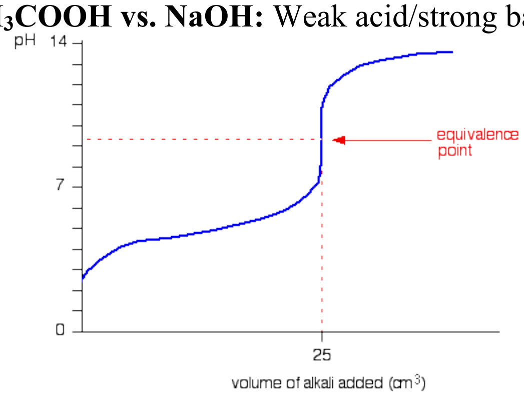
(a) HCl vs. NaOH: Strong acid/strong base:



(b) HCl vs. CH_3NH_2: Strong acid/weak base:



(c) CH_3COOH vs. NaOH: Weak acid/strong base:



24 What is the molar mass of 5.00 g of diprotic acid (H_2A) dissolved to make 250. mL solution and titrated with 11.1 mL of 1.00 M NaOH?

want to calculate — given

$$\text{molar mass of } H_2A = \frac{\text{g } H_2A}{\text{mol } H_2A}$$

need to find

$$2KOH(aq) + H_2A(aq) \longrightarrow K_2A(aq) + 2H_2O(l)$$

$$11.1\text{ mL KOH} \times \frac{1.00\text{ mol KOH}}{1000\text{ mL}} \times \frac{1\text{ mol } H_2A}{2\text{ mol KOH}} = 5.55\times10^{-3}\text{ mol } H_2A$$

$$\mathcal{M}\,(\mathbf{H_2A}) = \frac{0.500\text{ g } H_2A}{5.55\times10^{-3}\text{ mol } H_2A} = \mathbf{90.1\ g/mol}$$

29 What is the pH at the equivalence point of a solution of 0.20 M methylamine (CH_3NH_2) titrated with 0.20 M HCl?

$HCl(aq) + CH_3NH_2(aq) \rightleftharpoons CH_3NH_3^+(aq) + Cl^-(aq)$

Since the concentrations of acid and base are equal, equal volumes of each solution will need to be added to reach the equivalence point. Therefore, the solution volume is doubled at the equivalence point, and the concentration of the conjugate acid from the salt, $CH_3NH_3^+$, is:

$$\frac{0.20\ M}{2} = 0.10\ \boldsymbol{M}$$

	$CH_3NH_3^+(aq)$ + $H_2O(l)$ $\rightleftharpoons$	$H_3O^+(aq)$ +	$CH_3NH_2(aq)$
Initial (M):	0.10	0	0
Change (M):	$-x$	$+x$	$+x$
Equilibrium (M):	$0.10 - x$	x	x

$$\boldsymbol{K}_a = \frac{[H_3O^+][CH_3NH_2]}{[CH_3NH_3^+]},\ 2.3\times10^{-11} = \frac{x^2}{0.10 - x} \approx \frac{x^2}{0.10}$$

$x = [H_3O^+] = 1.5 \times 10^{-6}\ M$, **pH** = **5.82**

35 From Table 16.1, what indicator(s) would you use for the following titrations:

(a) HCOOH vs. NaOH: HCOOH is a weak acid and NaOH is a strong base. Suitable indicators are cresol red and phenolphthalein.

(b) HCl vs KOH: HCl is a strong acid and KOH is a strong base. Suitable indicators are all those listed with the exceptions of thymol blue, bromophenol blue, and methyl orange.

(c) HNO_3 vs. CH_3NH_2: HNO_3 is a strong acid and CH_3NH_2 is a weak base. Suitable indicators are bromophenol blue, methyl orange, methyl red, and chlorophenol blue.

Additional Problem

How many mL of 0.0850 M NaOH are required to titrate 25.0 mL of 0.128 M $CH_2ClCOOH$ (chloroacetic acid, $K_a = 1.38\times10^{-3}$) to the equivalence point? What is the pH of the solution at the equivalence point?

$$\text{mol NaOH} = \text{mol } CH_2ClCOOH = \frac{0.128 \text{ mol } CH_2ClCOOH}{1000 \text{ mL}} \times 25.0 \text{ mL} = 0.00320 \text{ mol NaOH}$$

$$V_{NaOH} = \frac{\text{mol NaOH}}{\text{M NaOH}} = 0.00320 \text{ mol NaOH} \times \frac{1000 \text{ mL}}{0.0850 \text{ mol NaOH}} = \boxed{37.6 \text{ mL NaOH}}$$

To determine pH, first determine the concentration of CH_2ClCOO^- at the equivalence point, use ICE to re-equilibrate the system:

$$[CH_2ClCOO^-] = \frac{0.00320 \text{ mol}}{(25.0 + 37.6) \text{ mL}} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 0.0511 \text{ M}$$

	CH_2ClCOO^- (aq) + H_2O (l) $\rightleftharpoons$	$CH_2ClCOOH$ (aq) +	OH^- (aq)
I	0.0511 M	0	0
C	$-x$	$+x$	$+x$
E	$0.0511 - x$ M	x M	x M

$K_b = K_w/K_a = 1.0\times10^{-14}/1.38\times10^{-3} = 7.24\times10^{-12}$

$$K_b = \frac{[\mathrm{CH_2ClCOOH}][\mathrm{OH^-}]}{[\mathrm{CH_2ClCOO^-}]} = \frac{x^2}{0.0511 - x} \approx \frac{x^2}{0.0511} = 7.24 \times 10^{-12};$$

$$x = [\mathrm{OH^-}] = \sqrt{\left(7.24 \times 10^{-12}\right)\left(0.0511\right)} = 6.08 \times 10^{-7}\ \mathrm{M}$$

pOH = 6.22; $\boxed{\textbf{pH = 14.00 – 6.22 = 7.78}}$