Titration

Titration is the process of adding a solution of known concentration, the titrant, to an unknown solution until the two solutions are just reacted. The endpoint of the reaction (the stoichiometric or equivalence point) is generally indicated by a color change in the solution. There are three types of acid/base titrations:

Strong acids and strong base titrations are the easiest, because strong acids and bases dissociate completely there is no equilibrium calculations that need to be done, only stoichiometry.

Weak acids and strong bases, these types of reactions are very similar to the buffer problems that we were doing earlier. We first do a stoichiometry calculation and then a weak acid equilibrium problem.

Strong acids and weak bases are very similar to the weak acid/strong base problem. Stoichiometry and then equilibrium.

Halfway Point

First stoichiometry-

$$\mathrm{OH^- + HC_2H_3O_2 \iff H_2O + C_2H_3O_2^-}$$

0.025	0.050		0
-0.025	-0.025		+0.025
0	0.025		0.025

Then equilibrium-

$$\mathrm{HC_2H_3O_{2(aq)} \iff H^+_{(aq)} + C_2H_3O^-_{2(aq)}}$$

0.17	0	0.17
-x	+x	+x
0.17-x	x	0.17+x

$$pH = -log(1.8e-5) + log(\frac{0.17}{0.17}) \qquad \mathrm{pH = 4.74}$$

This is the half way point of the titration. The concentration of acid and conjugate base are equal and the pK_a=pH.

Equivalence Point

First stoichiometry- $OH^- + HC_2H_3O_2 \Longleftrightarrow H_2O + C_2H_3O_2^-$

0.050	0.050		0
-0.050	-0.050		+0.050
0	0		0.050

Then equilibrium- $C_2H_3O_2^- + H_2O_{(l)} \Longleftrightarrow HC_2H_3O_2 + OH^-$

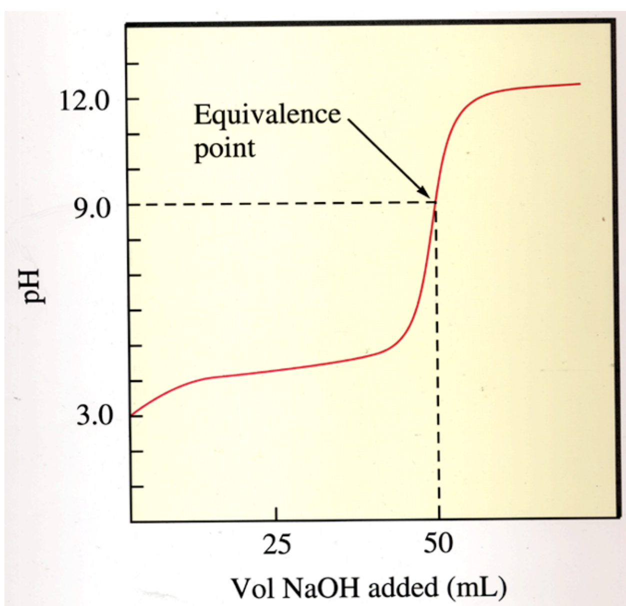
0.25		0	0
-x		+x	+x
0.25-x		x	x

$$K_b = \frac{K_w}{K_a} = \frac{1.0 \cdot 10^{-14}}{1.8 \cdot 10^{-5}} = 5.6 \cdot 10^{-10}$$

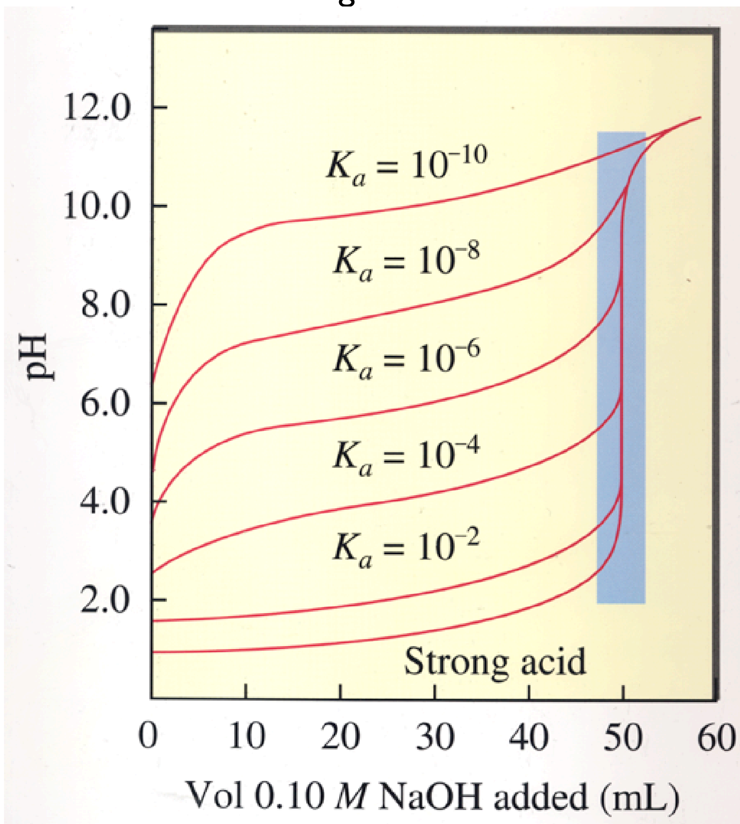
$$K_b = 5.6e - 10 = \frac{x \cdot x}{[0.25 - x]} \approx \frac{x^2}{0.25}$$

$x = 1.2 \cdot 10^{-5} = [OH^-]$ $\quad pOH = 4.9 \quad pH = 9.1$

Titration curves



Strong vs. Weak



Indicators