

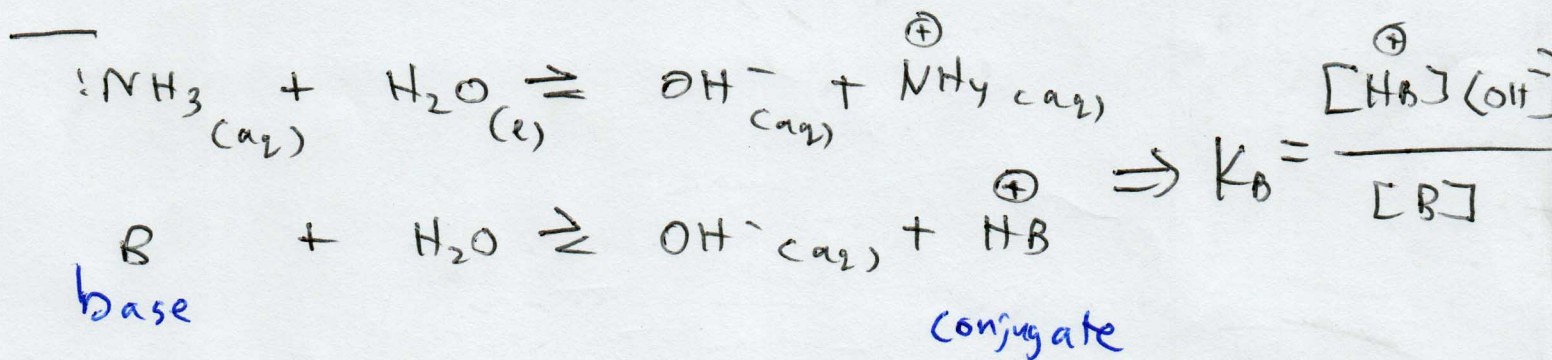
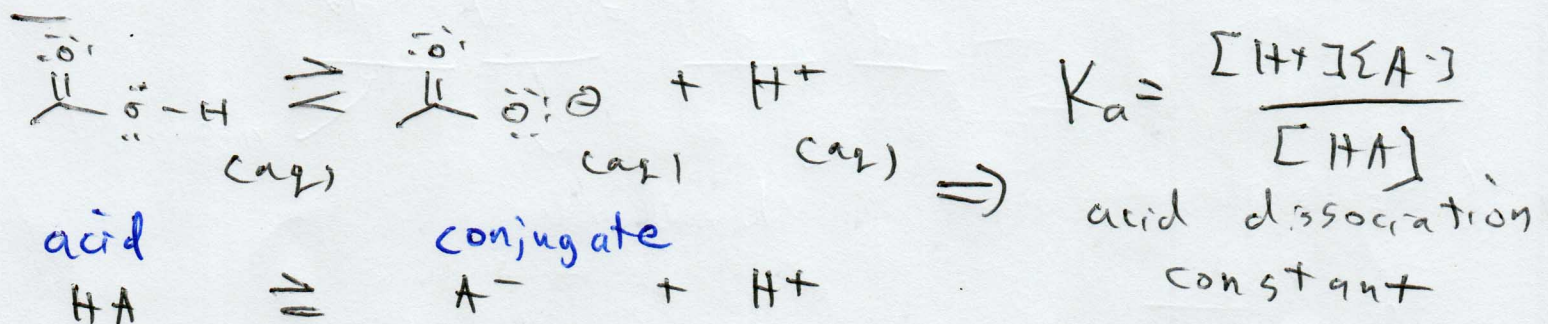
The auto-ionization of water is an unfavorable process that does occur to some degree at any temperature since there will always be a fraction of molecules that have the energy to dissociate.

$$K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14} \text{ at only one temperature}$$

In pure water at this temperature (), $[\text{H}^+] = [\text{OH}^-]$

$$\Rightarrow [\text{H}^+] = 1.0 \times 10^{-7} \text{ M} \quad * \text{ pH} = 7 \text{ means neutral}$$

$$\text{pH} \equiv -\log_{10} [\text{H}^+] = 7.00 \text{ at only one temperature.}$$



If $A = B$ (an acid/base conjugate pair)

$$K_a \cdot K_b = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \cdot \frac{[\text{OH}^-][\text{HA}]}{[\text{A}]} = [\text{H}^+][\text{OH}^-] = K_w$$

An acid's strength (K_a) is inversely proportional to its conjugate base's strength.

strong acid —

extensive spontaneous dissociation

→ K_a large → pK_a small / $pK_a \equiv -\log_{10} K_a$

1 M 10^0 HCl → $pH \approx 0$
 0.1 M 10^{-1} HCl → $pH \approx 1$
 0.01 M 10^{-2} HCl → $pH = 2$

when solutions are highly concentrated, interactions between solute particles can cause deviations from predicted solution behavior → activity

Weak acids — acids that undergo only partial spontaneous dissociation

Given a 1.0 M of acetic acid (and ignoring activity), ~~can~~ predict the pH of the sol'n,

$$K_a = 1.76 \times 10^{-5}$$

$$pK_a = 4.76$$

	HA	H ⁺	A ⁻	$K_a = \frac{[H^+][A^-]}{[HA]}$
Initial	1.00	10^{-7}	0	
Change	-x	+x	+x	$1.76 \times 10^{-5} = \frac{(x)(x)}{1-x}$
End	1-x	x	x	1-x

If $x \ll 1$, the x can be ignored, →

$$\Rightarrow x^2 = 1.76 \times 10^{-5} \Rightarrow x = 4.19 \times 10^{-3} = [H^+]$$

$$\Rightarrow pH = 2.38$$

0.500 M acetic acid + 0.500 M sodium acetate

pH = ?

	HA	H ⁺	A ⁻
I	0.500	"0"	0.500
C	-x	+x	+x
E	.5-x	x	.5+x

$$K_a = 1.76 \times 10^{-5} = \frac{(x)(0.5-x)}{0.5-x} \quad \text{If } x \ll 0.5, \quad 0.5-x \approx 0.5$$

$$K_a = 1.76 \times 10^{-5} = x = [H^+] \Rightarrow pK_a = pH$$

* This simplification only is valid if the concentrations of an acid and its conjugate do not shift substantially once the solution is prepared.

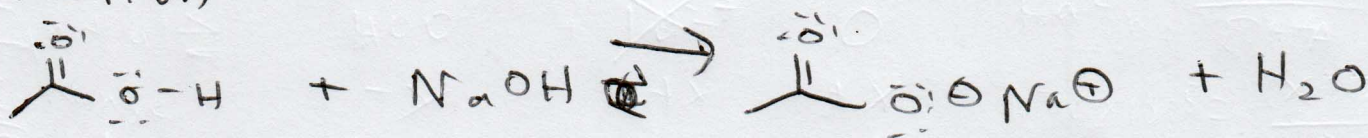
$$K_a = \frac{[H^+][A^-]}{[HA]} \Rightarrow -\log_{10}(K_a) = -\log_{10}\left(\frac{[H^+][A^-]}{[HA]}\right)$$

$$pK_a = -\log_{10} [H^+] + \left[-\log_{10} \left(\frac{[A^-]}{[HA]} \right) \right]$$

$$pK_a = pH - \log_{10} \frac{[A^-]}{[HA]}$$

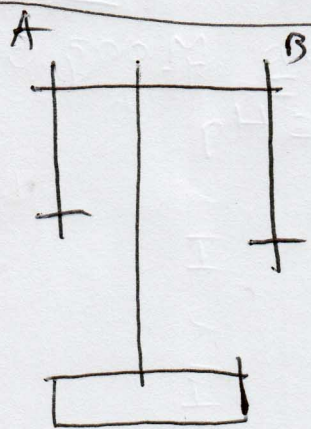
$$pH = pK_a + \log_{10} \frac{[A^-]}{[HA]} \quad \text{Henderson-Hasselbalch eq}$$

Titration



@ Equivalence point: $n_{\text{acid}} = n_{\text{base}}$

$$\underbrace{M_{\text{acid}} V_{\text{acid}}}_{\text{unknown}} = \underbrace{M_{\text{base}} V_{\text{base}}}_{\text{calculated or standardized}} \quad \text{measured}$$



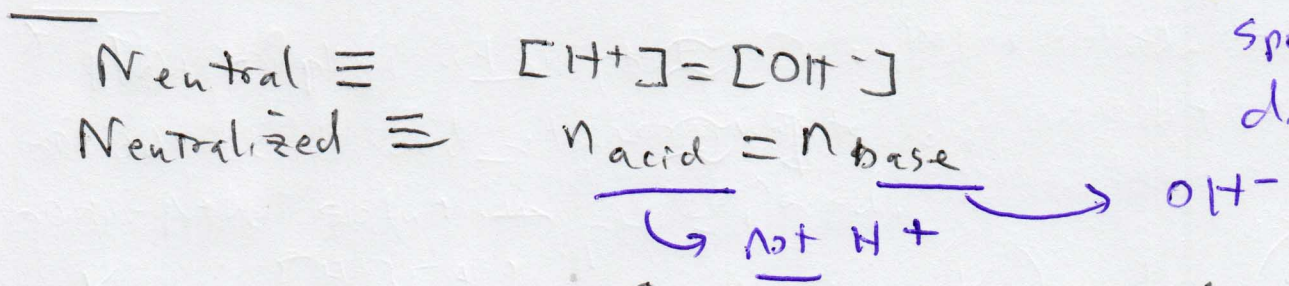
acid; H_2O ; indicator

An indicator is chosen so that the pK_a of the indicator matches the anticipated pH of the titration sol'n at equivalence,

Equivalence point - $n_{\text{acid}} = n_{\text{base}}$

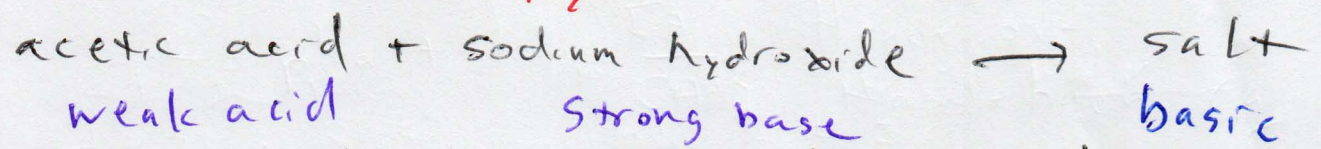
Endpoint - The visual indication that a titration is complete (color change)

* by spontaneous dissociation

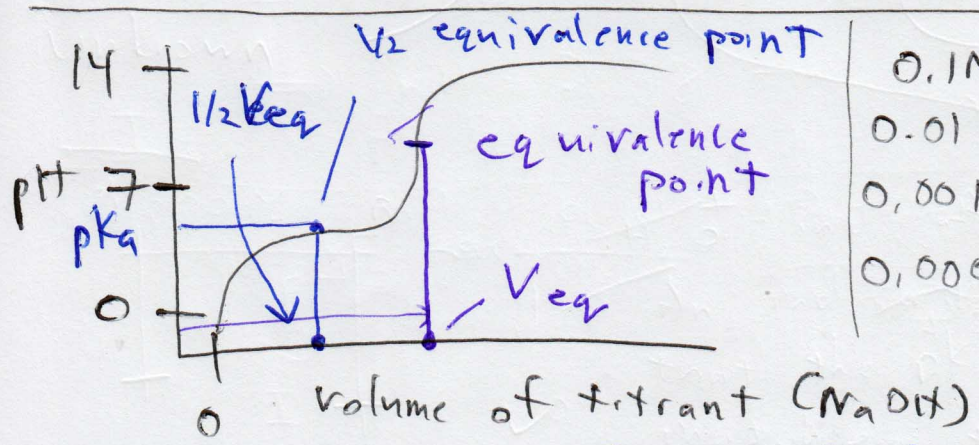


If a weak acid (or base) is used in titration, the # of moles of H^+ (or OH^-) that are effectively generated will be less than the # of moles of acid (or base) itself,

\therefore Neutral \neq Neutralized



Once acetic acid is quantitatively neutralized by sodium hydroxide, a salt (sodium acetate) is produced that tries to re-establish equilibrium with the parent acid, acetic acid. To do this, acetate reacts with ~~hydroxide~~ water, producing hydroxide, producing a basic solution,



0.1 M HCl	\rightarrow pH = 1	90%
0.01 M HCl	\rightarrow pH = 2	$\Delta 90\%$
0.001 M HCl	\rightarrow pH = 3	99%
0.0001 M HCl	\rightarrow pH = 4	$\Delta 99\%$ 99.9%

@ 1/2 equivalence: $[HA] \approx [A^-]$ \therefore pH = pKa