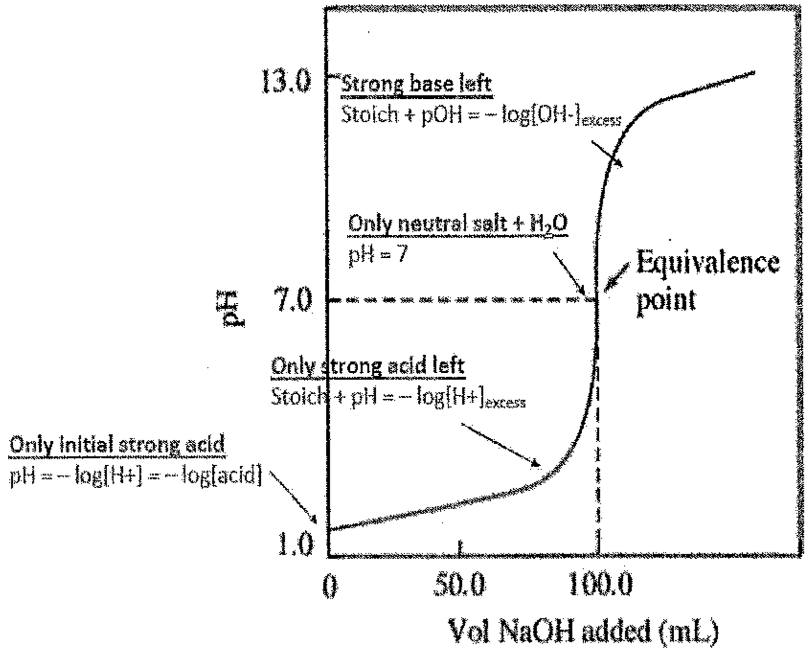
pHun with Titration Calculations

<u>Strong Acid + Strong Base</u> = 3 Situations to Determine pH

1. Initial acid concentration:
   a. only strong acid
   b. pH = − log[H^+] = − log[acid]
2. Equivalence point:
   a. moles of H^+ = moles of OH^-
   b. pH = 7
   c. Use $V_AM_A = V_BM_B$ to calculate volume needed to reach equivalence point.
3. Before or after equivalence point: use stoich (BCA table) to calculate excess moles of H^+ or OH^-, divide by total volume, and calculate the pH based on this value.

Titration Curve of 50.0 mL of 0.200 *M* HCl with 0.100 *M* NaOH



Example: 100.0 mL × 1.00 M = 100. mmol HCl 50.0 mL × 0.500 M = 25.0 mmol NaOH

a. 100.0 mL of 1.00 M HCl is titrated with 0.500 M NaOH. Calculate the pH after 50.0 mL of base has been added.

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

	H^+	OH^-
B	100.	25.0
C	−25.0	−25.0
A	75	∅

$[H^+] = \dfrac{75 \text{ mmol}}{\underbrace{150.0 \text{ mL}}_{100.0 + 50.0}} = 0.500 \text{ M}$

$pH = -\log(0.500) = \boxed{0.301}$

b. Calculate the pH after 200.0 mL of base has been added.

200.0 mL × 0.500 M = 100. mmol NaOH

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

	H^+	OH^-
B	100.	100.
C	−100	−100
A	∅	∅

⇒ @ equivalence pt! (both 100% neutralized) → pH = $\boxed{7.000}$

c. Calculate the pH after 300.0 mL of base has been added.

300.0 mL × 0.500 M = 150. mmol NaOH

$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

H^+	OH^-
100.	150.
−100	−100
∅	50.

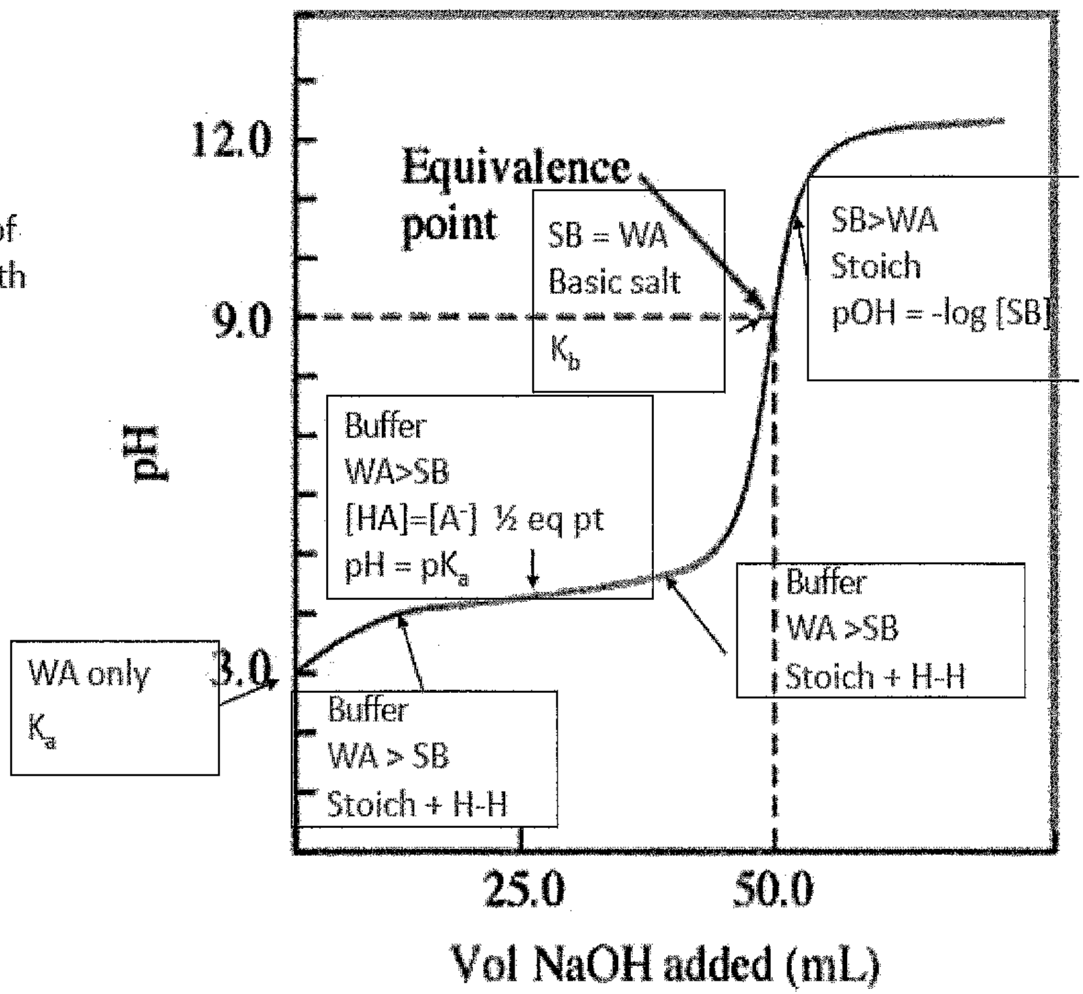
$[OH^-] = \dfrac{50.0 \text{ mmol}}{400.0 \text{ mL}} = 0.125 \text{ M}$

$pOH = -\log(0.125) = 0.903 \Rightarrow pH = 14 - 0.903 = \boxed{13.097}$

<u>Weak Acid + Strong Base</u> = 5 Situations to determine pH

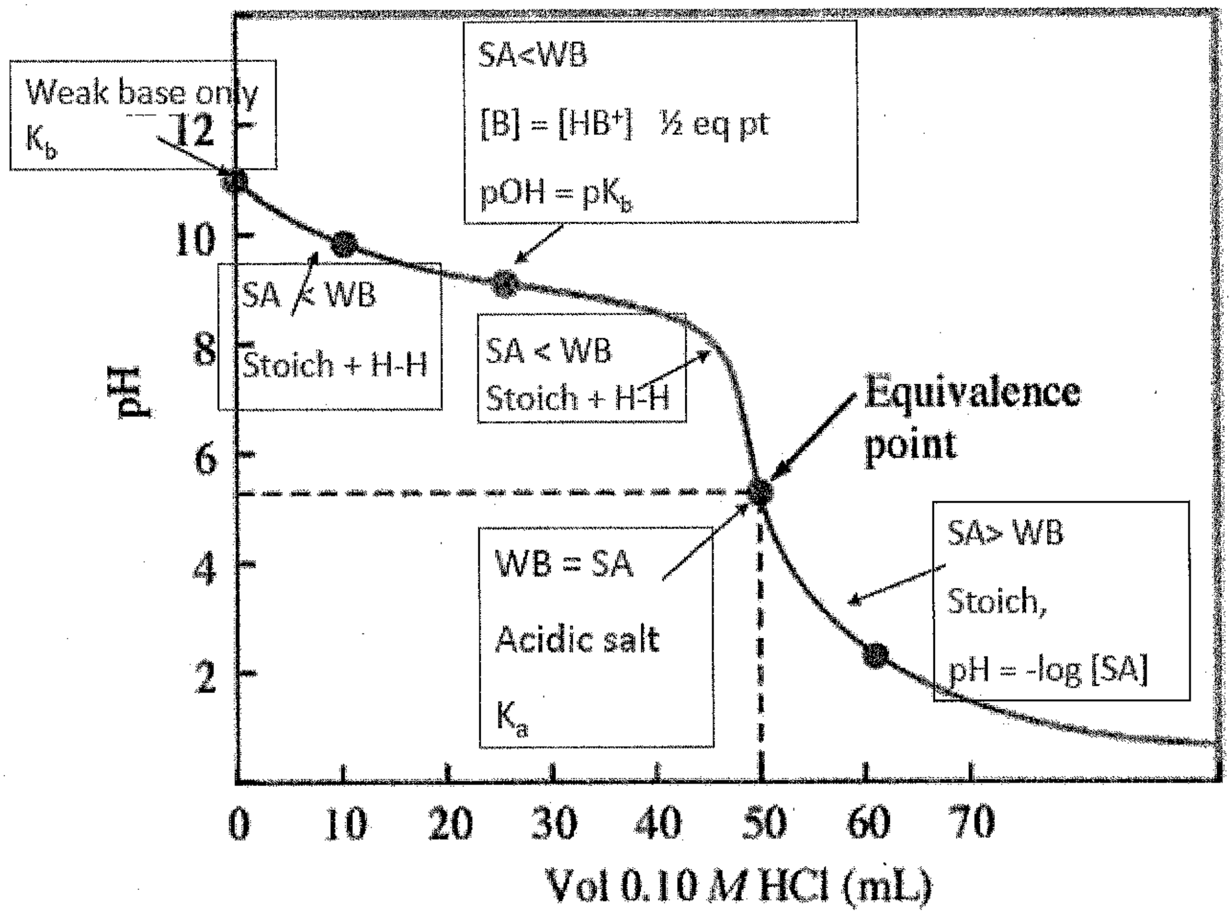
1. <u>The pH before the titration begins:</u> only weak acid.
2. <u>During titration but before equivalence point</u>: this is a buffer problem!!!
   - Once the titration begins, weak acid has reacted with the strong base to produce salt (the conjugate base of the weak acid).
3. <u>½ Equivalence point</u> (<u>midpoint</u>). A perfect buffer!
   - **[HA] = [A⁻], thus [H_3O^+] = K_a** and pH = pKa
4. <u>Equivalence point</u>**.** All acid and base neutralized: only species left are Salt and water.
   - pH is based on basic properties of the salt (conjugate base of weak acid.) ☺
5. <u>Beyond the equivalence point</u>: all about the base !
   - Calculate the amount of excess strong base added beyond the equivalence point.

The pH Curve for the Titration of 50.0 mL of 0.100 *M* $HC_2H_3O_2$ with 0.100 *M* NaOH



Weak Base + Strong Acid: Just like Weak Acid/ Strong Base, just flip flop the process!

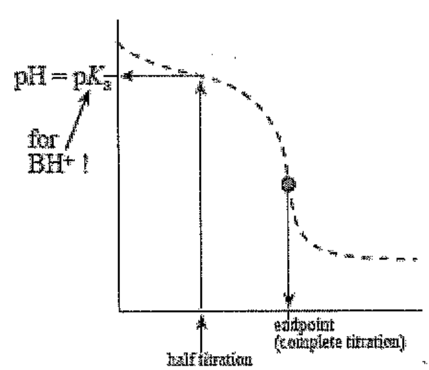
The pH Curve for the Titration of 100.0 mL of 0.050 *M* NH_3 with 0.10 *M* HCl



The half-way point is important!

After you have determined the equivalence point (endpoint) of the titration, go to half that value. The pH at the half-titration point is equal to the pK_a of the weak acid, BH^+. To get the pK_b of the base (B) you MUST subtract the pK_a from 14. The reason for this is that the pOH is actually what equals the pK_b.

$$pK_b = 14 - pK_a$$



$M_aV_a = M_bV_b$

Examples: $(0.200)(100.0) = (0.100)V_b \Rightarrow V_b = 200.$ mL needed to reach eq. pt

1. Consider the titration of 100.0 mL of 0.200M acetic acid ($K_a = 1.8 \times 10^{-5}$) by 0.100 M KOH. Calculate the pH of the resulting solution after the following volumes of KOH have been added.

a. 0.0 mL Only $HC_2H_3O_2 \Rightarrow K_a$ problem

$$\underset{\uparrow\ 1.8E-5}{K_a} = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = \underbrace{\frac{x^2}{0.200-x}}_{\text{x negligible b/c } K_a \ll 1} \approx \frac{x^2}{0.200}$$

$x = \sqrt{(0.200)(1.8E-5)}$

$= 0.0019\ M = [H_3O^+]$

$pH = -\log(0.0019) = \boxed{2.72}$

b. 100.0 mL $\frac{1}{2}$ eq. pt!!

$pH = pK_a = -\log(1.8E-5) = \boxed{4.74}$

c. 150.0 mL b4 eq. pt = buffer! ★ calculate [HA], [A⁻] first

$HC_2H_3O_2 = 100.0\ mL \times 0.200\ M = 20.0$ mmol

$OH^- = 150.0\ mL \times 0.100\ M = 15.0$ mmol

	$HC_2H_3O_{2(aq)}$ +	$OH^-_{(aq)}$ →	$C_2H_3O_{2(aq)}^-$ +	$H_2O_{(l)}$
B	20.0	15.0	∅	
C	-15.0	-15.0	+15.0	
A	5.0	∅	15.0	

$[HC_2H_3O_2] = 5.0\ mmol / 250.0\ mL = 0.0200M$

$[C_2H_3O_2^-] = 15.0\ mmol / 250.0\ mL = 0.0600M$

$pH = pK_a + \log\frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]}$

$= 4.74 + \log\left(\frac{0.0600}{0.0200}\right) = \boxed{5.22}$

d. 200.0 mL

↳ @ eq. pt! only basic salt (K_b problem)

$K_b = \frac{K_w}{K_a} = \frac{1E-14}{1.8E-5} = 5.6E-10$

$[C_2H_3O_2^-] = \frac{(100.0\ mL)(0.200\ M)}{300.0\ mL} = 0.0667\ M$

$$K_b = \underbrace{\frac{x^2}{0.0667-x}}_{\text{x negl., } K_b \ll 1} \approx \frac{x^2}{0.0667}$$

$\Rightarrow x = \sqrt{(0.0667)(5.6E-10)}$

$= 6.1E-6\ M = [OH^-]$

$pOH = -\log(6.1E-6) = 5.21$

$pH = 14 - 5.21 = \boxed{8.79}$

e. 250.0 mL

↳ past eq. pt = only strong base! $OH^- = 250.0\ mL \times 0.100\ M = 25.0$ mmol

	$HC_2H_3O_{2(aq)}$ +	$OH^-_{(aq)}$ →	$C_2H_3O_{2(aq)}^-$ +	H_2O
B	20.0	25.0	∅	
C	-20.0	-20.0	20.0	
A	∅	5.0	20.0	

negligible (weak base) contribution

$[OH^-] = 5.0\ mmol / 350.0\ mL = 0.0143\ M$

$pOH = -\log(0.0143) = 1.85$

$pH = 14 - 1.85 = \boxed{12.16}$