Titration of weak acid with strong base

Remember:

- Before adding the base: It is a weak acid HA (Ka)
- After adding the base: It is a buffer: pH = pK_a + log [base]/[acid]
- At half the volume of the equivalence point: [base]=[acid] or [A-] = [HA]
    - pH = pK_a
    - So, $[H^+] = K_a$
    - Buffer is most effective
- At the equivalence point: It is a weak base ($K_b = 10^{-14}/K_a$
    - pH is governed by the concentration of the buffer base (A-)
    - pH at the equivalence point is greater than 7 (pH > 7).
- After the equivalence point: It is a strong base.
- [ ] = moles / Volume (L)

Note: For the titration of weak base with a strong acid. The pH at the equivalence point is lower than 7 (pH < 7)

Indicator: pH = $pK_a \pm 1$

The pKa of the weak acid to be used in the buffer should be as close as possible to the desired pH.

PH = pKa + log [base]/[acid]. For most effective buffer: [base]/[acid] = 1

Case Study

50 ml (0,05 l) of 0.1M acetic acid solution ($HC_2H_3O_2$, $K_a = 1.8 \times 10^{-5}$) with 0.1 M NaOH.

1) First calculate the volume of the base needed for the equivalence point:

$$M\ acid \times V\ acid = M\ base \times V\ base$$
$$0.1 \times 50 = 0.1 \times V\ base$$
$$V\ base = 50\ ml$$

So, we need 50 ml of NaOH to completely neutralize the 50 mL of acetic acid.

2) At a volume half the volume of the equivalence point, $pH = pK_a$

So, at 25 ml, $pH = pK_a = -\log K_a = -\log 1.8 \times 10^{-5}$

3) Calculate the moles of the acid;

Moles of $HC_2H_3O_2$ = moles of H^+ = Volume x Molarity
= 0.05 x 0.1 = 0.005 moles

A) No base is added:

It is a weak acid

	$HC_2H_3O_2$	=	$C_2H_3O_2^-$	+	H^+
Initial	0.1				
Change	- X		X		X
Equilibrium	0.1 - X		X		X

$K_a = x^2 / 0.1$

$X = 1.3 \times 10^{-3}$ M

pH = **2.87**

B) Add 10 mL of 0.1 M NaOH:

Calculate the moles of NaOH = Molarity x Volume (L)
= 0.1 x 0.01= 0.001mole
Mole of $HC_2H_3O_2$ = 0.005 mole

Total volume is: 50 + 10 = 60 ml = 0.060 Liter

	$HC_2H_3O_2$	+	OH^-	=	$C_2H_3O_2^-$	+	H_2O
Before reaction	0.005 mole		0.001				
After reaction	0.005 - 0.001		0		0.001		
	= 0.004 mole				0.001		
[ ] :	0.004/ 0.06				0.001 / 0.06		

Buffer: pH = pK_a + log [base]/[acid] = -log 1.8 x 10^{-5} + log (0.001/0.06) x (0.06/0.004)

` pH = **4.14**

C) Add 25 mL of NaOH

Half the volume of the equivalence point: pH = pK_a
So, $[H^+] = K_a$

- At half the volume of the equivalence point: [base]=[acid] or [A-] = [HA]
    - pH = pK_a
    - So, $[H^+] = K_a$
    - Buffer is most effective

D) Add 40 mL of NaOH

pH = **5.35**

E) Add 50 mL of 0.1 M NaOH:

Calculate the moles of NaOH = Molarity x Volume (L)
= 0.1 x 0.05 = 0.005 mole

Mole of $HC_2H_3O_2$ = 0.005 mole

Total volume is: 50 + 50 = 100 ml = 0.1 Liter

	$HC_2H_3O_2$	+	OH^-	=	$C_2H_3O_2^-$	+	H_2O
Before reaction	0.005 mole		0.005				
After reaction	0.005 - 0.005 = 0 mole				0.005		
[.] :					0.005/0.1 = 0.05M		

Flip the reaction; $C_2H_3O_2^-$ is a weak base.

	$C_2H_3O_2^-$	+	H_2O	=.	$HC_2H_3O_2$	+	OH^-
Initial	0.05						
Change	- X				X		X
Equilibrium	0.05 -X				X		X

$K_b = 10^{-14}/K_a$

$K_b = 10^{-14}/1.8 \times 10^{-5}$

$K_b = 5.6 \times 10^{-10}$

$K_b = x^2 / 0.05$

$X = 5.3 \times 10^{-6} = [OH^-]$

pH = 14- pOH = **8.72**

- **At the equivalence point: It is a weak base ($K_b = 10^{-14}/K_a$**
    - **pH is governed by the concentration of the buffer base (A-)**
    - **pH at the equivalence point is greater than 7 (pH > 7).**

F) Add 60 mL of 0.1 M NaOH

Calculate the moles of NaOH = Molarity x Volume (L)
= 0.1 x 0.06 = 0.006 mole

Mole of $HC_2H_3O_2$ = 0.005 mole

Total volume is: 50 + 60 = 110 ml = 0.11 Liter

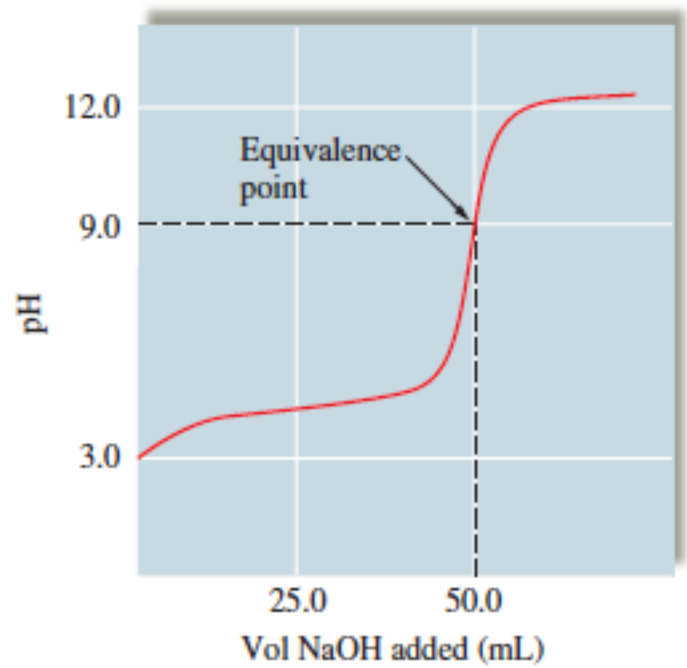
	$HC_2H_3O_2$	+	OH^-	=	$C_2H_3O_2^-$	+	H_2O
Before reaction	0.005 mole		0.006				
After reaction	0.005 - 0.005 = 0 mole		0.006-0.005 = 0.001				
[.] :			0.001/0.11 = 9.1×10^{-3} M				

$[OH^-] = 9.1 \times 10^{-3}$ M

pH = 14 - pOH = **11.96**

G) Add 75 mL of 0.1 M NaOH

pH = 14 - pOH = **12.3**



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At the equivalence point: pH< 7