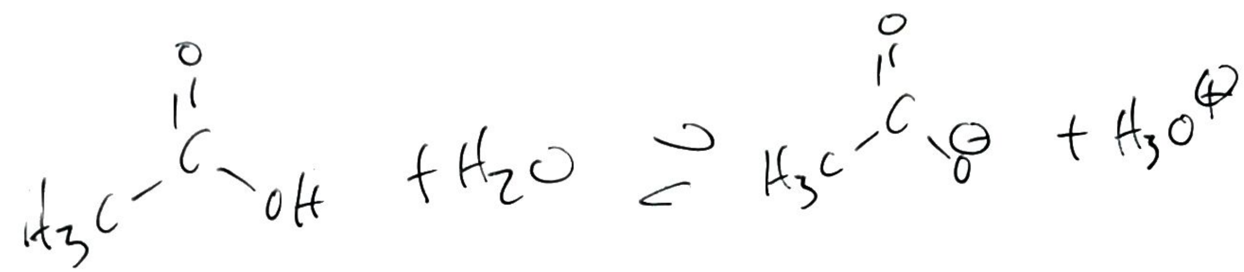
#9 Sept 2019 Weak Acids Bases

The problem set is due Friday 13, Sept. !?



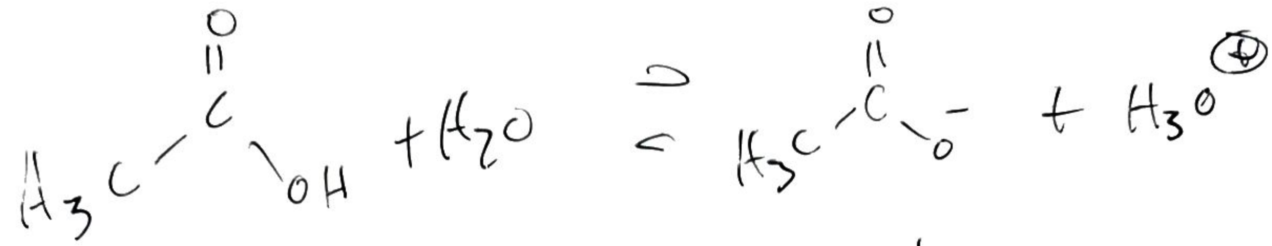
acid dissociation constant K_A

$$H_2O + H_2O \rightleftharpoons {}^{-}OH + H_3O^{\oplus}$$

$$K_w = \frac{[OH^-][H_3O^+]}{[H_2O]} = [OH^-][H_3O^+]$$

$$K_w = 1\times10^{-14} = [{}^{-}OH][H_3O^+]$$

$$\boxed{14 = pOH + pH}$$



Acetic acid Acetate

$$K_A = 1.77 \times 10^{-5} = \frac{[Ac^-][H_3O^+]}{[HAc]}$$

What is the pH of 100 mL of 0.1 M acetic acid ($K_A = 1.77 \times 10^{-5}$)

$$HAc + H_2O \rightleftharpoons Ac^- + H_3O^+$$

$$K_A = \frac{[Ac^-][H_3O^+]}{[HAc]}$$

amount of Ac^- formed = amount of H_3O^+

x x

$$K_A = 1.77 \times 10^{-5} = \frac{[x][x]}{[0.1]}$$

$$x^2 = 1.77 \times 10^{-6}$$

$$x = 1.33 \times 10^{-3} = [Ac^-] = [H_3O^+]$$

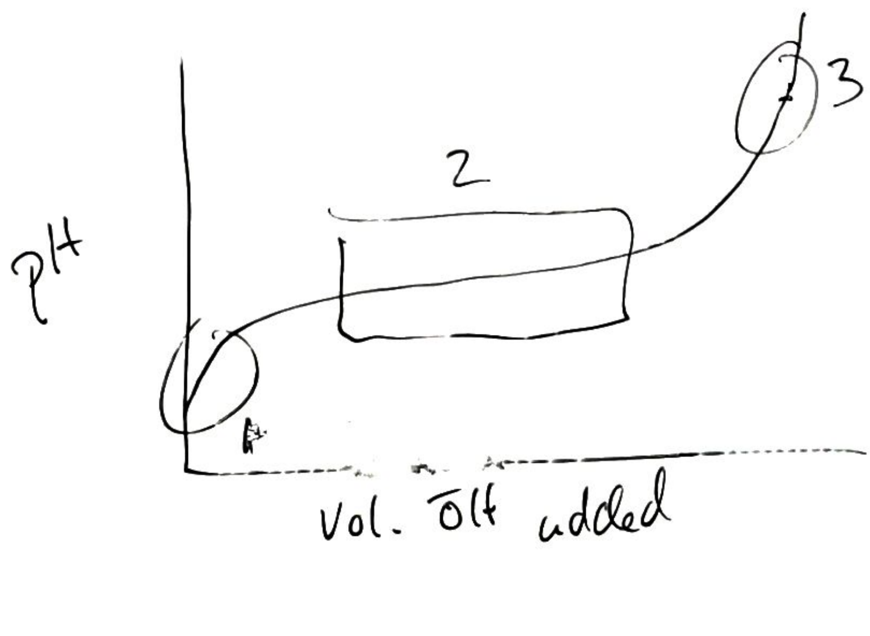
$pH = -\log[H_3O^+]$

$pH = -\log\{1.33\times10^{-3}]$

$\boxed{pH = 2.87}$

To determine the Ka value we have titrate the weak acid with base.

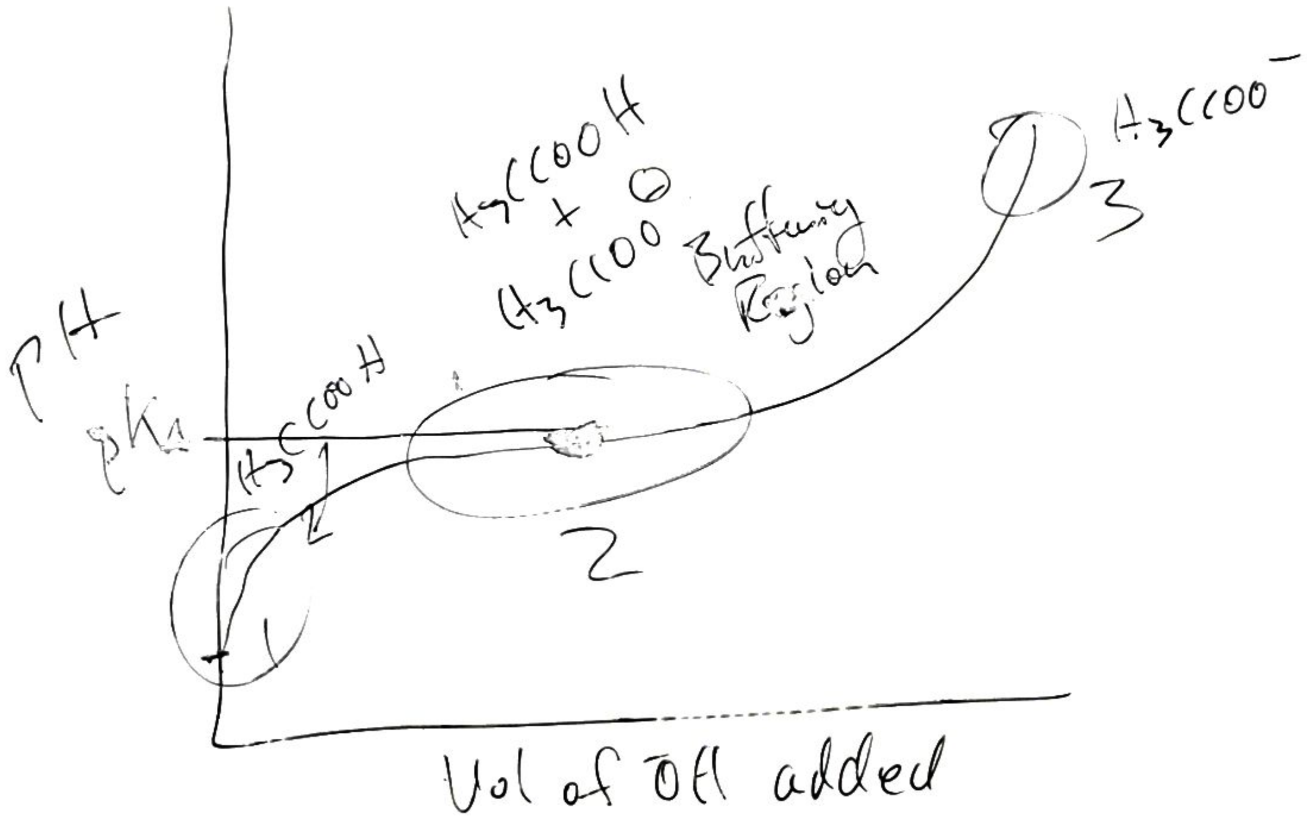
A titration is a chemical procedure where a known volume of acid has a strong base of known conc slowly added to it and the pH is recorded the whole time.



During a titration, two chemical rxns are occurring

(1) Equilibrium rxn. when nothing is being added

(2) Acid/Base reaction occurs when ^-OH reacts with the acid.



$$H_3CCOOH + {}^-OH \rightarrow H_3CCOO^- + H_2O$$

At the pKa,

$$[H_3CCOOH] = [H_3CCOO^-]$$

Things to remember:

i) you are reacting the weak acid with the strong base, so during the titration there are 2 types of chemical rxns occurring

Rxn 1

when nothing is being added
The acid is in equilibrium with H_2O

$$H_3CCOOH + H_2O \rightleftharpoons H_3CCOO^{\ominus} + H_3O^{\oplus}$$

Rxn 2

we add OH^- and it reacts with the acid and creates conjugate base molecules.

$$H_3CCOOH + {}^-OH \longrightarrow H_3CCOO^- + H_2O$$

The # of conjugate base molecules made is equal to the number of

i) # of OH^- ions added (for a 1:1 rxn)

ii) # of acid molecules reacted

The volume has increased after OH^- addition so you recompute the molarities of HA and A^-

Example of a titration problem:

you have 100mL of 0.1M acetic acid ($K_A = 1.77 \times 10^{-5}$) and you want to titrate it with 0.2M NaOH

a) What is the pH of the acid solution after 10mL of 0.2M NaOH have been added?

CEq: $HAc + H_2O \rightleftharpoons Ac^- + H_3O^{+}$

Acid/Base: $HAc + OH^- \rightarrow Ac^- + H_2O$

i) # of moles of OH^- added:

$$0.01L \times \frac{0.2\,moles\,OH^-}{L} = 2 \times 10^{-3}\ moles\ OH^-\ added$$

ii) We know from the acid/base rxn 1 mole OH^- reacts with 1 mole HAc to make 1 mole Ac^-

$$2\times10^{-3}\text{ moles }OH^-\text{ added}\times\frac{1\text{ mole HAc}}{1\text{ mole }OH^-}=\boxed{2\times10^{-3}\text{ moles HAc reacted off}}$$

$$2\times10^{-3}\text{ moles }OH^-\text{ added}\times\frac{1\text{ mole }Ac^-\text{ made}}{1\text{ mole }OH^-\text{ added}}=\boxed{2\times10^{-3}\text{ moles }Ac^-\text{ made}}$$

Done reacting, so what happens?

$$0.1\text{ L HAc}\times\frac{0.1\text{ moles}}{\text{L}}=0.01\text{ moles HAc @ start}$$

$$0.01\text{ moles HAc}-2\times10^{-3}\text{ moles HAc reacted}=\boxed{8\times10^{-3}\text{ moles HAc remaining}}$$

$$\boxed{2\times10^{-3}\text{ moles }Ac^-\text{ made}}$$

(much higher than $\sqrt{1.77\times10^{-5}}$)

iii) Volume changed

New Volume = 100mL + 10mL = 110mL
= 0.11L

New molarity of HAc = $\frac{8\times10^{-3}\text{ moles HAc}}{0.11L} = \boxed{0.072M \text{ HAc}}$

New molarity of Ac^- = $\frac{2\times10^{-3}\text{ moles } Ac^-}{0.11L} = \boxed{0.018M\ Ac^-}$

iv) Equilibrium kicks in

$HAc + H_2O \rightleftharpoons Ac^- + H_3O^+$

$K_a = 1.77\times10^{-5} = \frac{[H_3O^+][Ac^-]}{[HAc]}$

$[H_3O^+] = 7.08\times10^{-3}$ M

$-\log[H_3O^+] = \boxed{pH = 4.15}$

Remember

i) The initial system is @ equilibrium So you can easily determine the starting moles of acid and the pH (if you have the Ka)

ii) Calculate the number of moles of OH^- added, and write the balanced acid/base chemical rxn out

- moles of acid reacted

  $\downarrow$

  mole of acid remaining

- moles of conjugate base produced

iii) Put the system back into equilibrium

USING THE NEW MOLAR VALUES FOR WEAK ACID and CONJUGATE BASE

Example:

Calculate the pH when the following solutions are added to 100ml of 0.1M HClO

The K_A of HClO is 3×10^{-8}

a) 0ml of 0.1M NaOH

b) 75ml of 0.1M NaOH

c) 100ml of 0.1M NaOH

---

a) $\overset{0.1M}{HClO} + H_2O \rightleftharpoons H_3O^{+} + ClO^{-}$

$$K_A = \frac{(\text{Products})}{(\text{Reactant})} = \frac{[H_3O^+][ClO^-]}{[HClO]} = \frac{[x][x]}{[0.1M]}$$

$$\frac{x^2}{0.1} = 3\times10^{-8}$$

$$x^2 = 3\times10^{-9}$$

$$x = 5.477\times10^{-5} = [H_3O^+] = [ClO^-]$$

$\boxed{pH = 4.26}$ $-\log[H_3O^+]$

b) after 75 mL of 0.1 M NaOH have been added

Acid/Base $HClO + OH^- \rightarrow ClO^{\ominus} + H_2O$

moles of OH added $= 0.075\,L \times \frac{0.1\,mol\,OH}{L} = 7.5 \times 10^{-3}$ moles OH^- added

7.5×10^{-3} moles $OH^- \times \frac{1\,mole\,HClO}{1\,mole\,OH^-} = 7.5 \times 10^{-3}$ moles HClO reacted

$\frac{0.1\,mol\,HClO}{\cancel{L}} \times 0.1\,\cancel{L} = 1 \times 10^{-2}$ moles HClO @ start

1×10^{-2} moles HClO $- 7.5 \times 10^{-3}$ moles HClO reacted $=$

$\boxed{2.5 \times 10^{-3} \text{ moles HClO remaining}}$

moles ClO^- got made? $\boxed{7.5 \times 10^{-3} \text{ moles } ClO^-}$