Chem 160: Titrations

1. You start off with 50mL of a .3 M acetic acid solution

Ka of Acetic Acid=1.8×10^{-5}

A) What is the pH of the solution before you add any of the NaOH?

B) Suppose you have 50 mL of a .3 M solution of acetic acid (CH_3COOH) and you titrate it with 20 mL of .2 M solution of sodium hydroxide (NaOH). Calculate the pH of the solution?

C) Calculate the pH when 50 mL of .3 M CH_3COOH is titrated with 75 mL of .2 M NaOH

D) Calculate the pH when 50 mL of .3M CH_3COOH is titrated with 100 mL of .2 M of NaOH

2. What is the pH of a solution when you titrate 25 mL of a .3M NaOH solution into a 50 mL .4 M solution of HCl?
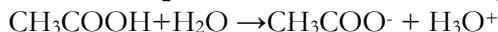
3. How much NaOH of a .3 M solution would you need to titrate to reach equivalence of a 15 mL .5 M HCl solution.

Solutions

1. You start off with 50mL of a .3 M acetic acid solution

Ka of Acetic Acid=1.8x10^{-5}

A) **What is the pH of the solution before you add any of the NaOH?**
$CH_3COOH + H_2O \rightarrow CH_3COO^- + H_3O^+$

	CH_3COOH	H_2O	CH_3COO^-	H_3O^+
I	.3 M	-	0	0
C	-X	-	+x	+x
E	.3 - X	-	x	X

$\frac{X^2}{.3-X}$ = 1.8x10^{-5} , we check for the rule of 1000 to determine if we can omit the -X.

To check rule of 1000: $\frac{Initial\ Concentration}{Ka}$

$\frac{.3}{1.8x10^{-5}}$ = 166666.67

New equation: $\frac{X^2}{.3}$ = 1.8x10^{-5} , X=.00232, plug into ice table and solve at equilibrium

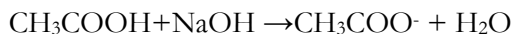$[CH_3COO^-]=[H_3O^+]=.00232$ M
$[CH_3COOH]=.298$ M
-log $[H_3O^+]$=pH
-log[.00232]=2.63=pH

B) **Suppose you have 50 mL of a .3 M solution of acetic acid (CH_3COOH) and you titrate it with 20 mL of .2 M solution of sodium hydroxide (NaOH). Calculate the pH of the solution?**

$CH_3COOH + NaOH \rightarrow CH_3COO^- + H_2O$
Interaction of the string base with the weak acid species forces the reaction. Because of this we focus on the moles of the substances rather than the molarity. Acid with a base forms a neutralization reaction which is why we get water as a product now.

Start off by converting everything to moles:
.3 mol/L x .05L = .015 mol CH_3COOH
.2 mol/L x .02 L = .004 mol NaOH

Once the number of moles has been solved for the reaction is treated like a limited reactants problem from chem 150. You directly subtract the least number of moles from both sides, the molar ratio is 1:1. The amount subtracted from the reactants is the amount of product made.

CH_3COOH	NaOH	CH_3COO^-	H_2O
.015 mol	.004 mol	0	-
- .004 mol	- .004 mol	+ .004 mol	-
.011 mol	0 mol	.004 mol	-

After the calculation, it's observed that the amount that both acetic acid and its conjugate base are existing in the same solution after titrating. The use of the Henderson Hasselbach equation is needed since the solution is now a buffer. The number of moles at equilibrium are used, no need to convert to molarity since the 2 moles share the same overall volume.

Solve for the pka: $-\log(Ka)=pka$

HH equation: $pH=pka+\log(-A/HA)$

$pH=4.74+\log(.004/.011)$

$pH=4.30$

C) Calculate the pH when 50 mL of .3 M CH_3COOH is titrated with 75 mL of .2 M NaOH.

This reaction is special in the case that the number of moles of both base and acid will be in equilibrium. This will allow for an additional dissociation. The additional dissociation occurs with the CH_3COO^- (conjugate base) and react with water. The only other species in the solution. Again, because we start out with a weak acid with strong base you can need to convert everything all species to moles.

CH_3COOH	NaOH	CH_3COO^-	H_2O
.015 mol	.015 mol	0	-
- .015 mol	- .015 mol	+ .015 mol	-
0 mol	0 mol	.015 mol	-

Since all you have is a weak base and water, treat the problem as a weak base reaction. Calculate kb and the concentration of the CH_3COO^-. Note, you need to use the total volume of the solutions for the new concentration.

$.015/.(075L+.05L)= .12$ M CH_3COO^-
$Kb= (1.0x10^{-14})/Ka$
$Kb=5.56x10^{-10}$

	CH_3COO^-	H_2O	CH_3COOH	NaOH

I	.12 M	-	0	0
C	-X	-	+x	+x
E	.12 - X	-	x	X

$$\frac{x^2}{.12-X} = 5.56 \times 10^{-10}$$

Can ignore -X, it passes through rule of 1000. Use the kb instead of the ka this time.

$$\frac{x^2}{.12} = 5.56 \times 10^{-10}, \quad X = 8.17 \times 10^{-6}$$

The strong species is the NaOH, therefore the main influencer of the reaction will be pH.
$[OH] = 8.17 \times 10^{-6}$
$-\log(8.17 \times 10^{-6}) = pOH = 5.09$
$14 - 5.09 = 8.91 \, pH$

D) Calculate the pH when 50 mL of .3M CH_3COOH is titrated with 100 mL of .2 M of NaOH.

After going over equivalence the acid is completely neutralized, and you'll have excess NaOH floating in solution. The excess NaOH will determine the pH of the solution since it's the strongest species in solution. The total volume is the accumulation of both substances added. Limiting reagent will now be acetic acid since NaOH will be in excess once the number of moles is calculated.

CH_3COOH	NaOH	CH_3COO^-	H_2O
.015 mol	.02 mol	0	-
- .015 mol	- .015 mol	+ .015 mol	-
0 mol	0.005 mol	.015 mol	-

.005 mol/ (.1 L+.05 L)= .0333 M OH-
$-\log(.0333) = 1.477$
$14 - 1.477 = 12.52 \, pH$

2. What is the pH of a solution when you titrate 25 mL of a .3M NaOH solution into a 50 mL .4 M solution of HCl?

When working with a strong acid strong base titration, you don't focus on partial dissociation. The reactions go to completion based on the number of reagents available. In a sense this will be another limiting reagent problem. We focus on what strong species is left over.
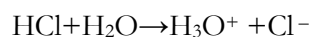
Calculate out the moles of each substance:

.025 L * .3mol/L = .0075 mol NaOH

.05 L * .4 mol/L =.02 mol HCl

HCl	NaOH	H_2O	NaCl
.02 mol	.0075 mol	-	-
- .0075 mol	- .0075 mol	-	-
.0125 mol	0 mol	-	-

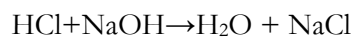The stronger species fully dissociates in water and will determine the pH.

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

.0125 mol / .075 L(total solution) = .16667 M [H_3O^+]

-log (.1667)=.78 pH

3. How much NaOH of a .3 M solution would you need to titrate to reach equivalence of a 15 mL .5 M HCl solution.

To reach the equivalence point you need the number of moles to be the same. The strategy is to calculate the moles of what you find and do a molar conversion to match that amount. The molecular formula is needed for this process.

$HCl + NaOH \rightarrow H_2O + NaCl$

.015L x .5 M = .0075 mol HCl

.0075 mol HCl x $\frac{1\ mol\ NaOH}{1\ mol\ HCl}$ = .0075 mol NaOH

Use the molarity of NaOH to calculate the volume amount.

.3M= .0075 moles NaOH / L, solve for L

.025L = L, 25mL