

## Week 5 Discussion Question Answers

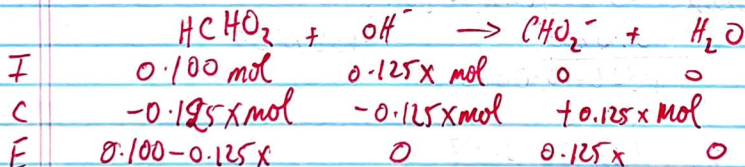
- 1) Imagine you titrate 150 mL of a 0.250 M  $\text{HC}_2\text{H}_3\text{O}_2$  solution with a 0.75 M  $\text{Sr}(\text{OH})_2$  solution. At which of the following points during your titration will a buffer system be present?

$$(.150 \text{ L})(0.250 \text{ M}) = 0.0375 \text{ moles } \text{HC}_2\text{H}_3\text{O}_2$$

Then  $\frac{0.0375 \text{ moles}}{0.75 \text{ M}} = 50.0 \text{ mL } \text{Sr}(\text{OH})_2$  to fully eliminate (reach equivalence point). 25.0 mL to reach  $\frac{1}{2}$  equivalence. Thus, 12.5 is when buffer will form.

$$\boxed{12.5 \text{ mL}}$$

- 2) What volume of 0.125 M  $\text{NaOH}$  must be added to 365 mL of 0.275 M  $\text{HCHO}_2$  to attain a solution with a pH of 3.50? The  $\text{pK}_a$  is 3.74.

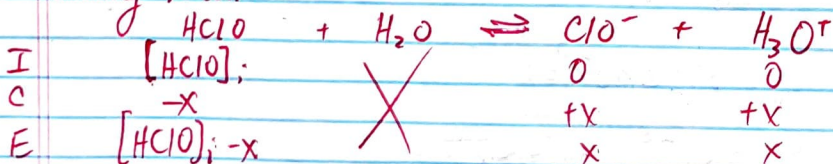


$$\text{So } \text{pH} = \text{pK}_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$3.50 = 3.74 + \log \frac{0.125x}{0.10 - 0.125x} \quad \text{Solving for } \boxed{x = 0.292 \text{ L}}$$

- 3) Consider a 500. mL solution of hypochlorous acid ( $\text{HClO}$ ) with a pH of 3.94. What is the pH of the solution after adding 60.0 mL of 5.00 M  $\text{NaOH}$ ? Assume the solution act ideally.  $\text{pK}_a$  of  $\text{HClO}$  is 7.53.

First determine amount of weak acid ( $\text{H}_3\text{O}^+$ ) in solution before adding  $\text{NaOH}$ .



$$K_a = \frac{[\text{ClO}^-][\text{H}_3\text{O}^+]}{[\text{HClO}]} = \frac{x^2}{([\text{HClO}]_i - x)}$$

$$\text{pH} = -\log_{10} [\text{H}_3\text{O}^+] \Rightarrow [\text{H}_3\text{O}^+] = 1.15 \times 10^{-4} \text{ M} \quad \text{H}_3\text{O}^+ = x$$

$$\text{p}K_a = -\log K_a \quad \text{or}$$

$$K_a = 10^{-\text{p}K_a} = 10^{-7.53} = 2.95 \times 10^{-8}$$

$$\text{Solving for } [\text{HClO}]_i = 0.448 \text{ M}$$

$$(0.448 \text{ M})(0.500 \text{ L}) = \boxed{0.224 \text{ mol HClO}} \text{ initially}$$

Assuming NaOH reacts with HClO fully

$$(0.300 \text{ L}) \left( \frac{5.00 \text{ mol NaOH}}{\text{L}} \right) = 0.300 \text{ mol NaOH}$$

HClO is fully depleted. Hence have 0.076 M excess NaOH.  
Strong base  $\Rightarrow [\text{NaOH}] = [\text{OH}^-]$ .

$$[\text{OH}^-] = \frac{0.076 \text{ mol}}{0.500 \text{ L} + 0.300 \text{ L}} = 0.136 \text{ M OH}^-$$

$$\text{pOH} = -\log_{10} [\text{OH}^-] = 0.87$$

$$\text{pH} = 14 - \text{pOH} = \boxed{13.13}$$

- 4) Acetic acid ( $\text{CH}_3\text{COOH}$ ) is a weak monoprotic acid with  $\text{p}K_a = 4.75$ .  
Its MW is 60.06 g/mol and the MW of  $\text{NaCH}_3\text{COO}$  is 82.03 g/mol.
- a) What is the pH of a solution of 0.685 g  $\text{NaCH}_3\text{COO}$  in 50.0 mL of water.

$$\frac{0.685 \text{ g}}{0.0500 \text{ L}} \left( \frac{1 \text{ mol}}{82.03 \text{ g}} \right) = 0.167 \text{ M NaCH}_3\text{COO}$$

$$K_b = \frac{[\text{CH}_3\text{COO}^-][\text{OH}^-]}{[\text{CH}_3\text{COOH}]}$$

$$K_a = \frac{K_w}{K_b} \Rightarrow K_b = \frac{K_w}{K_a} = \frac{10^{-14}}{10^{-4.75}}$$

$$K_b = 5.62 \times 10^{-10} \Rightarrow$$

$$5.62 \times 10^{-10} = \frac{x^2}{0.167 - x}$$



Assuming  $x$  small enough  $\Rightarrow$

$$x = 9.69 \times 10^{-6} M$$

Checking:

$$\frac{(9.69 \times 10^{-6} M)^2}{(0.167 - 9.69 \times 10^{-6} M)} = 5.62 \times 10^{-10}$$

assumption is good

$$pOH = -\log(5.62 \times 10^{-10}) = 9.25$$

$$pH = 14 - pOH = \boxed{4.75}$$

b) What would the pH be if you ~~add~~ added 0.485g  $CH_3COOH$  to the mixture in part a?

$$\left( \frac{0.485g \text{ } (CH_3COOH)}{0.0500L} \right) \left( \frac{1 \text{ mol}}{60.05g} \right) = 0.162 M \text{ } CH_3COOH$$

$$K_b = \frac{[CH_3COO^-][OH^-]}{[CH_3COOH]}$$

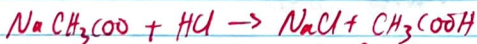
$$5.62 \times 10^{-10} = \frac{(0.162 + x)x}{(0.167 - x)}$$

Assuming  $x$  small  $\Rightarrow x = 5.81 \times 10^{-10} M$  (can confirm assumption works)

$$pOH = -\log(5.81 \times 10^{-10} M) = 9.23$$

$$pH = 14 - 9.23 = \boxed{4.77}$$

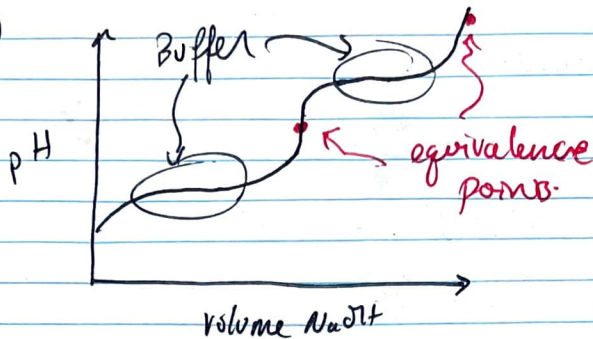
c) How many mL of 0.750 M HCl would you need to add to the solution from part a to get the same pH you got for part b?



Adding HCl makes  $CH_3COOH$ . Reacts in 1:1 ratio.  $\therefore$  moles of HCl = moles  $CH_3COOH$  from before (0.00808 mol)

$$\frac{0.00808 \text{ mol HCl}}{0.750 M HCl} \times \frac{1000 \text{ mL}}{1 L} = \boxed{10.8 \text{ mL HCl}}$$

5) a)



- b) How are buffers made over the course of a titration like this one (how much base do you need to add)? How do you know how many buffer regions you will have?

During a titration, buffers are made when you have approximately equal moles of weak acid & its conjugate base. For a diprotic acid, there are 2 buffer regions since 2 sets of conjugate base/acid pairs.

- c) Explain what the equivalence point is. How do you know how many equivalence points you are going to have.  
 Equivalence when equal weak acid and strong base.  
 For diprotic acids  $\rightarrow$  2 equivalence points.

- d) Sketch titration of  $A^{2-}$  with HCl.

