CH 302 – Unit 2 Review 3

ACID AND BASE CALCULATIONS, APPLICATIONS

"THINKING LIKE A CHEMIST"

Acid and Base Question Types (Simplified)

In addition to real world applications of acid/base problems, you should be able to put yourself into the mindset of a chemist doing these experiments in the following scenarios:

- Strong acid, strong base questions
  - Simple relationships converting $[H_3O^+]$ and $[OH^-]$ to pH, pOH
- Weak acid, weak base questions
  - Approximations or quadratic formula (if necessary) with your general formula:
  - $HA \rightleftharpoons H^+ + A^-$ OR $B \rightleftharpoons BH^+ + OH^-$
- Buffer Questions
  - Mixtures of a weak acid and its salt (conjugate base) ; weak base and its salt (conjugate acid)
  - Solve for pH, pOH using Henderson-Hasselbalch equation
- Neutralization reactions (Titration experiments)
  - By adding a titrant to an analyte solution, you can neutralize your original species. Depending where you are in the neutralization experiment, you will make different calculations based on the three scenarios above.

Quantifying Acids and Bases

- The standard units of measurement for acids and bases are pH and pOH

$$\mathrm{pH} = -\log[H^+] \qquad \mathrm{pOH} = -\log[OH^-]$$

- In a strong acid or strong base solution, we can use this relationship:

$$\mathrm{pH} = -\log[C_A] \qquad \mathrm{pOH} = -\log[C_B]^*$$

- pH and pOH can be interconverted using the relationship based on pK_w (14):

$$\mathrm{pH} = 14 - pOH \qquad \mathrm{pOH} = 14 - pH$$

Quantifying Weak Acids and Bases

- For a weak acid, we are dealing with a more interesting equilibrium (additional steps). The approximation formulas are below:

$$[H^+] = \sqrt{C_{HA} \cdot K_a} \qquad [OH^-] = \sqrt{C_B \cdot K_b}$$

- Don't forget that at any time you can convert between different terms:

$$K_w = 1 \cdot 10^{-14} = [H^+][OH^-]$$

$$K_w = K_a K_b$$

$$\frac{K_w}{K_a} = K_b$$

$$\frac{K_w}{K_b} = K_a$$

Another helpful relationship:

- Another way of explaining how much a weak acid or base dissolves is percent ionization. You can memorize the formula:

$$[H^+] = (\%\ \text{ionization})(C_{HA})$$

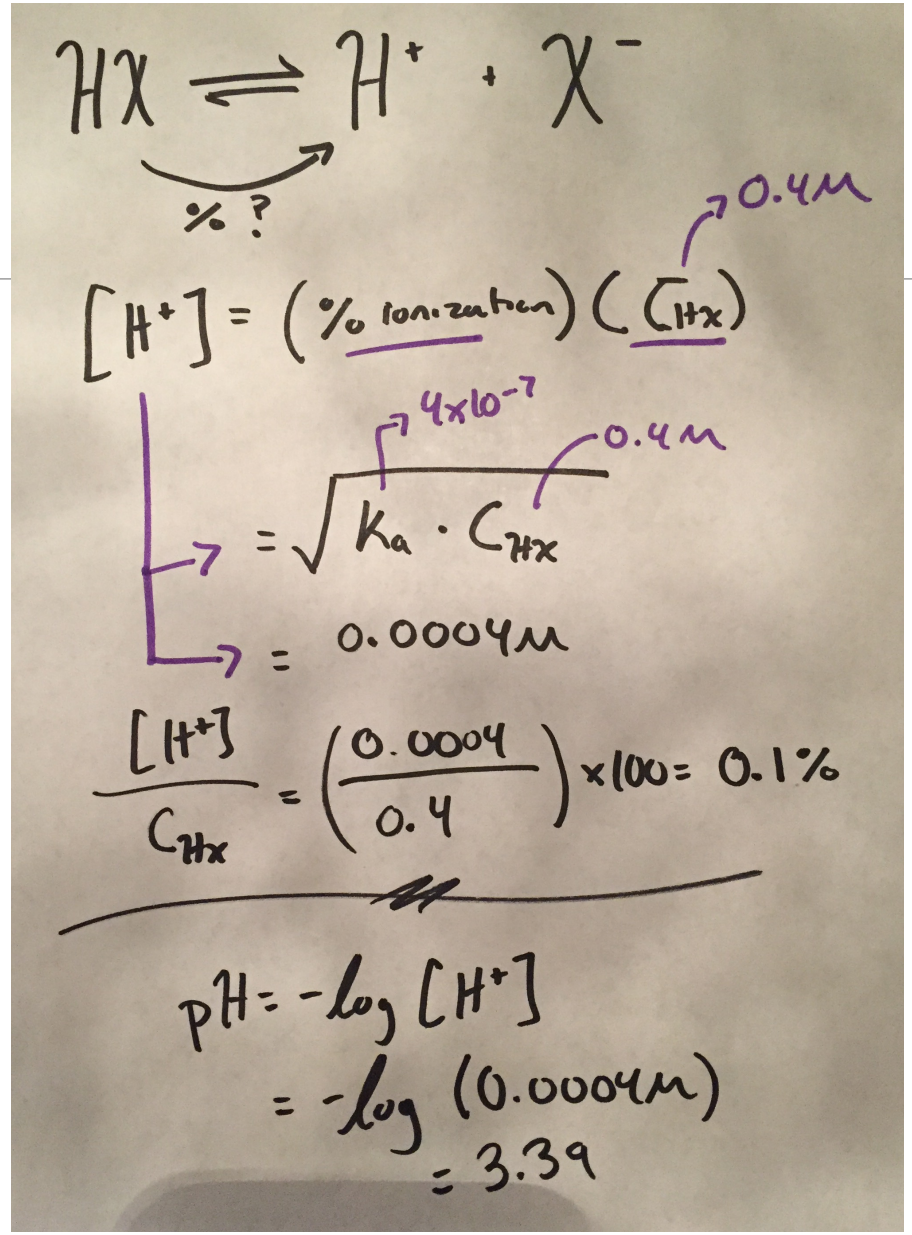
$$[OH^-] = (\%\ \text{ionization})(C_B)$$

Note that these are fractions, 0-1, so you need to multiply by 100 to get the “percent” value

Warm-Up Questions

What is the percent ionization for a weak acid HX that is 0.40M? $K_a = 4.0 \times 10^{-7}$

What is the pH of this weak acid solution?



Buffers: Conceptual, Calculation

- **The purpose of a buffer is to resist changes in pH.**
- Here's the idea:
  - If you add one mole of NaOH to pure water, you are adding one mole of OH^-. This results in a pretty big change in pH
  - If you add one mole of NaOH to a solution of acetic acid and sodium acetate, you are just creating one more mole of sodium acetate. **This *barely* increases the pH.**
- **By definition, a buffer is made of:**
  - **A weak acid and its salt (conjugate base)**
  - **A weak base and its salt (conjugate acid)**
- By definition, a buffer is effective when:
  - The pH is within about +/- 1 of the pK_a or pK_b
  - This favorable ratio of the weak acid : salt (or weak base : salt) is known as the "Buffer Zone

$$pH = pK_a + \log\left(\frac{A^-}{HA}\right)$$

$$pOH = pK_b + \log\left(\frac{BH^+}{B}\right)$$

If you can identify that you have a buffer based on this definition, you know to use the Henderson-Hasselbalch equation (above)

Choosing a Buffer

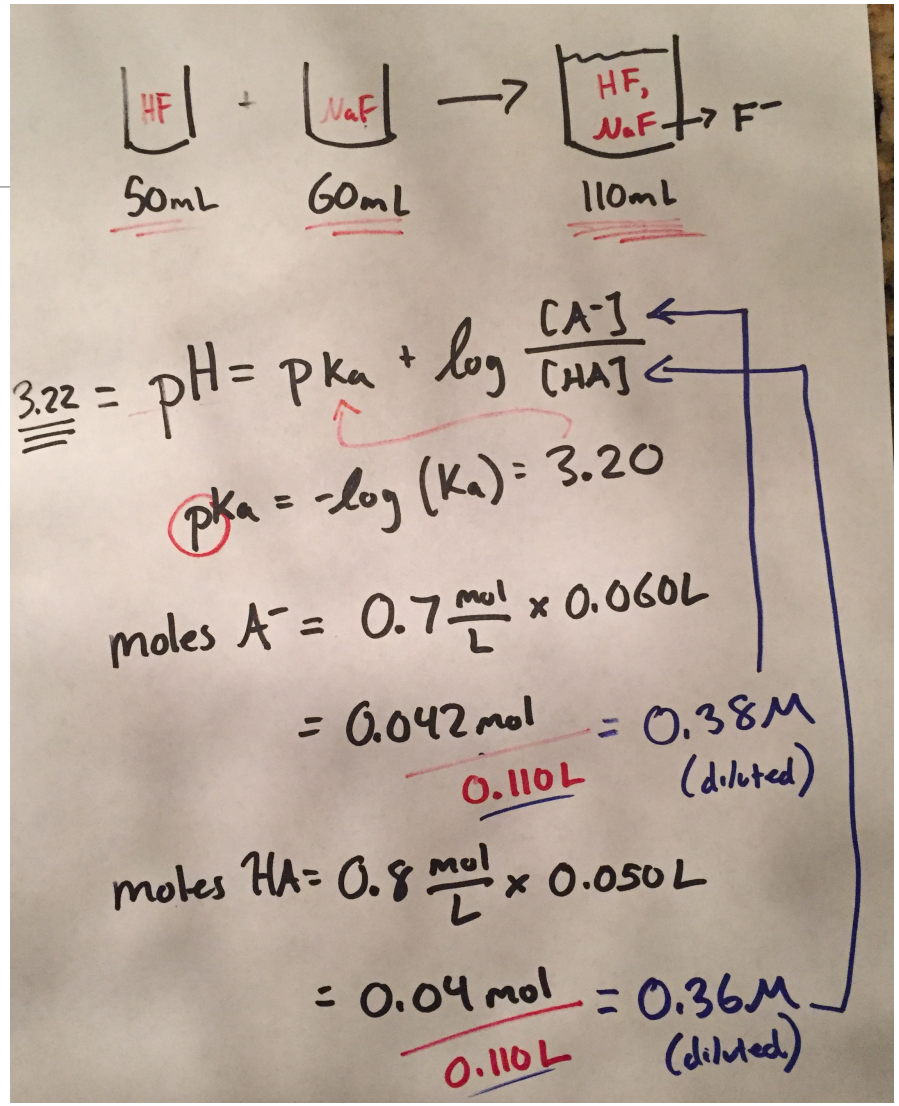
- **A buffer only functions in the "Buffer Zone," which is +/- 1 pH point of the pK_a (for a weak acid buffer) or pK_b (for a weak base buffer). Let's stick with a weak acid buffer.**
- Buffers are used in the real world (reaction chemistry, physiology, pharmacology, etc.) to maintain a stable pH environment. You choose a buffer with a pK_a closest to the pH environment you want to hold constant.

Example Question: The human bloodstream is held constant by a buffer system at pH = 7.35. Which of the following buffer systems is likely found in the bloodstream?

a. **Carbonic Acid, pKa = 6.37**
b. Acetic Acid, pKa = 4.75
c. Hydrofluoric Acid, 3.14
d. Ammonium, 9.21

Question

What is the pH of a solution consisting of 50mL 0.8M HF and 60mL 0.7M NaF (Don't try this at home)? K_a for HF is 6.3×10^{-4}.



Neutralization Reactions: Salts

- **The product of a neutralization reaction is a salt. In acid/base chemistry, your salt can be neutral, acidic, or basic depending on the reaction:**

1. Strong Acid, strong base: results in a neutral salt

$$HCl(aq) + NaOH(aq) \rightleftharpoons NaCl(aq) + H_2O(l)$$

2. Strong Acid to weak base: results in an acidic salt

$$HCl(aq) + NH_3(aq) \rightleftharpoons NH_4^+(aq) + Cl^-(aq)$$

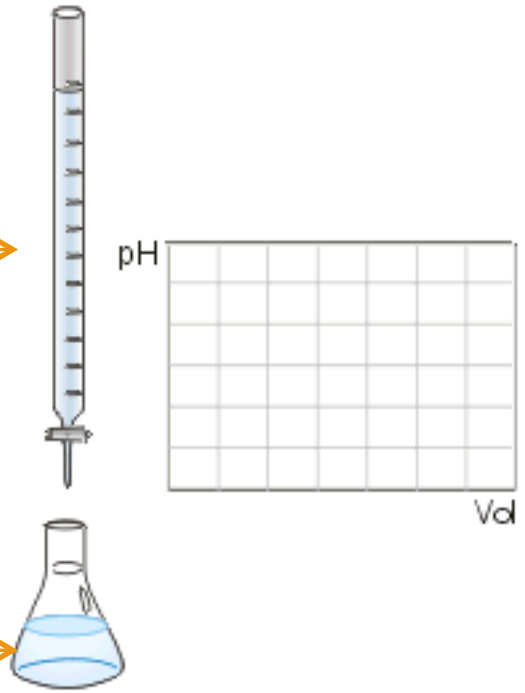
3. Strong base to weak acid: results in a basic salt

$$NaOH(aq) + CH_3COOH(aq) \rightleftharpoons CH_3COO^-(aq) + Na^+(aq) + H_2O(l)$$

Titration Experiments: NaOH added to HCl

This is the titrant. For this experiment, the titrant is NaOH (strong base). It is added drop-wise to the analyte. We know the concentration of the titrant.

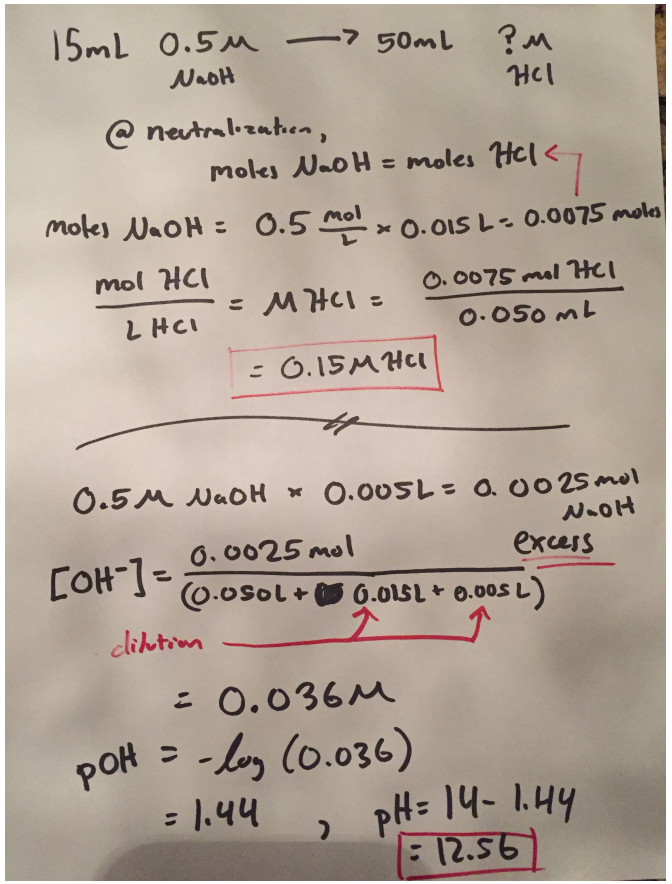
This is the analyte. For this experiment, the analyte is HCl (strong acid). We don't always know the concentration of the analyte. There is an indicator in the analyte that changes color at a specific pH.

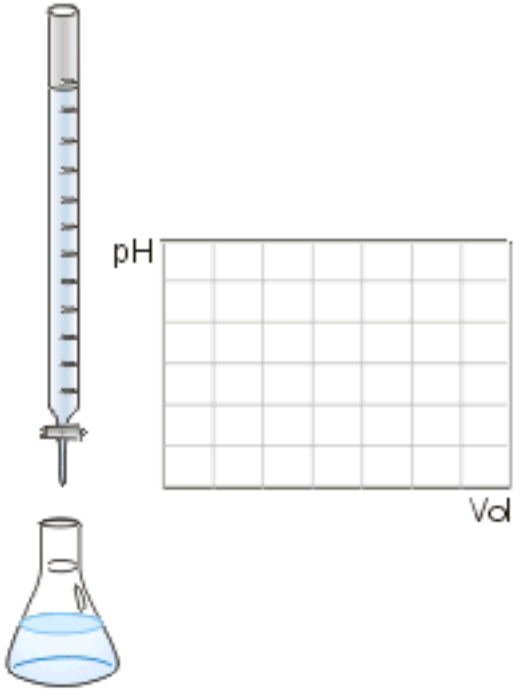


Question: Suppose it took 15mL of 0.5M NaOH to completely neutralize 50mL of HCl analyte. What was your original HCl concentration? What is the pH of your sample if you then add 5mL MORE of NaOH?

The objective of a titration experiment is to fully neutralize the analyte solution. This will form a neutral, basic, or acidic salt, depending on the experiment.

Titration Experiments: NaOH added to HCl





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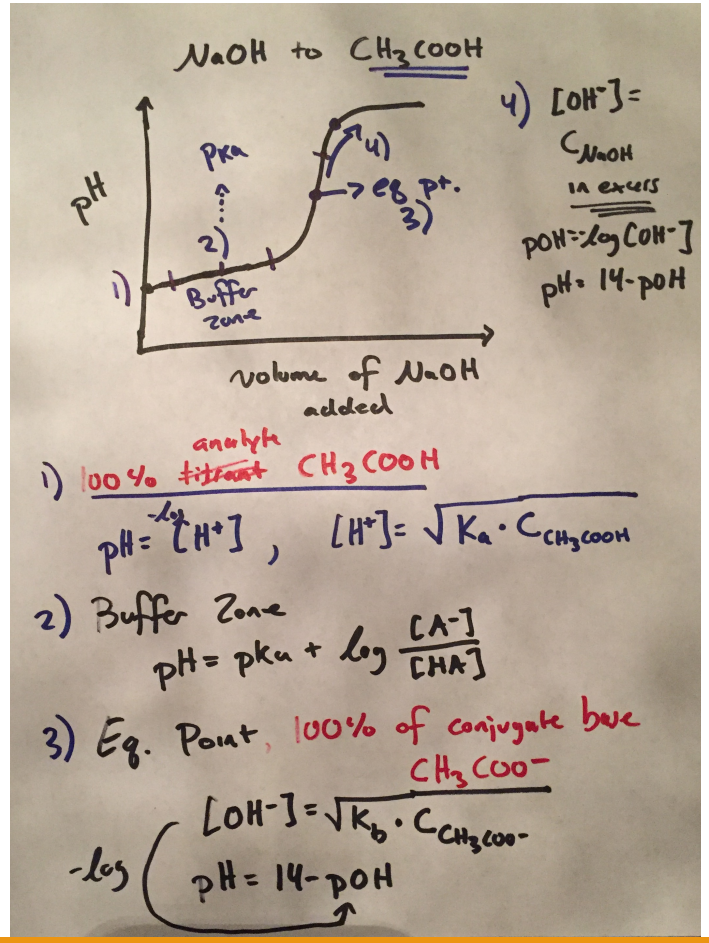
Titrations of Buffered Solutions



A) Before you have started:

Calculate pH using the approximation formula or using $K_a = [H^+][A^-] / [HA]$.

B) Within the buffer zone:

$pH = pKa + \log([A^-]/[HA])$

C) At the equivalence point:

You only have the salt of your weak acid in solution. Use $K_w = K_a K_b$, then solve for $[OH^-]$, pOH, then convert to pH

D) After the equivalence point:

Find the pH using the excess base in solution

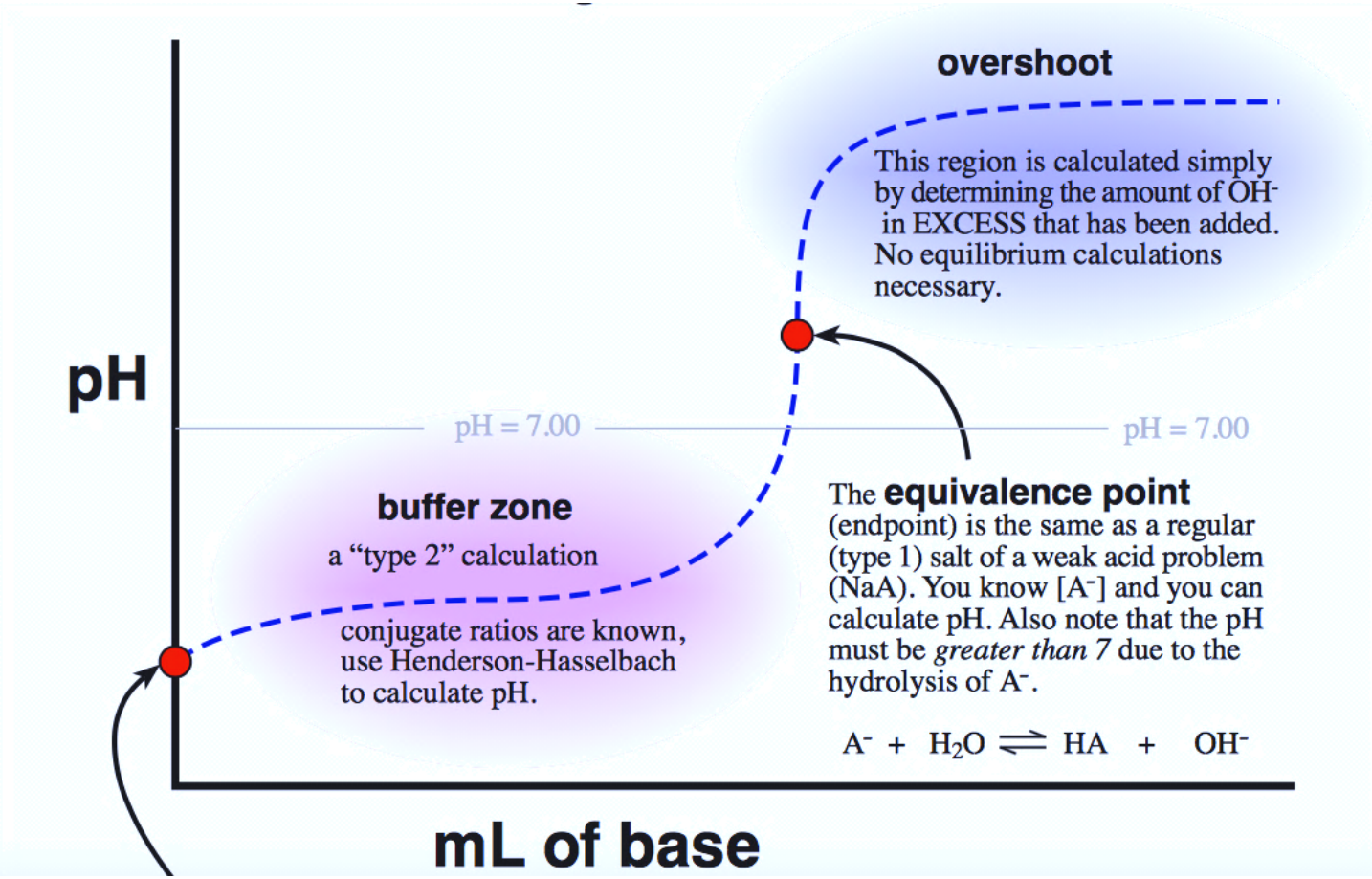
pOH = -log ([excess base])

pH = 14- pOH

Points to know:

- Equivalence point
- End Point
- Half-equivalence point
- Buffer Zone

Titrations of Buffered Solutions



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Points to know:
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Titration Questions

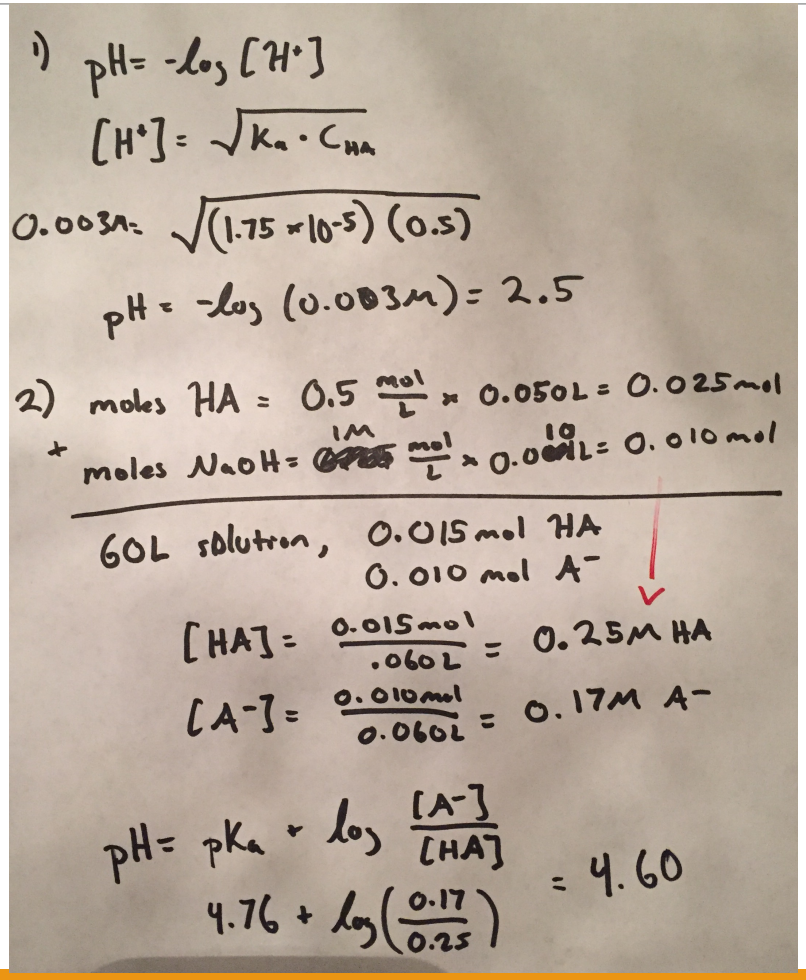
You titrate 50mL of 0.5M acetic acid with 1M NaOH.

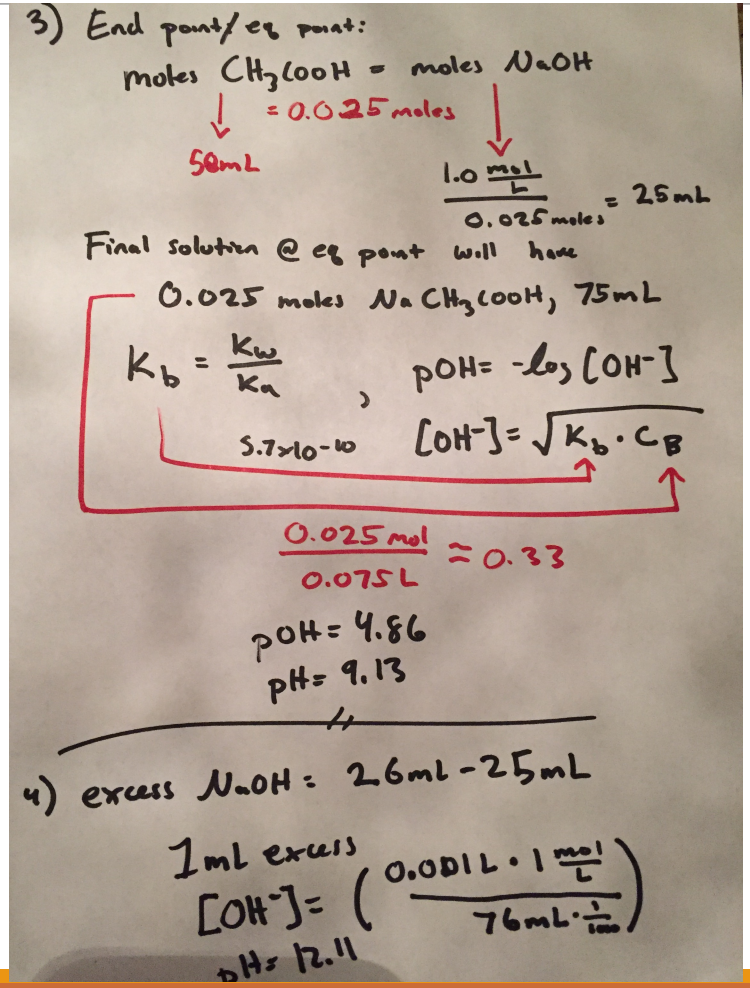
K_a for acetic acid is 1.75×10^{-5}.

What is the pH when:

1. You start the experiment.
2. You have added 10mL NaOH
3. You reach the end point
4. You have added 26mL NaOH

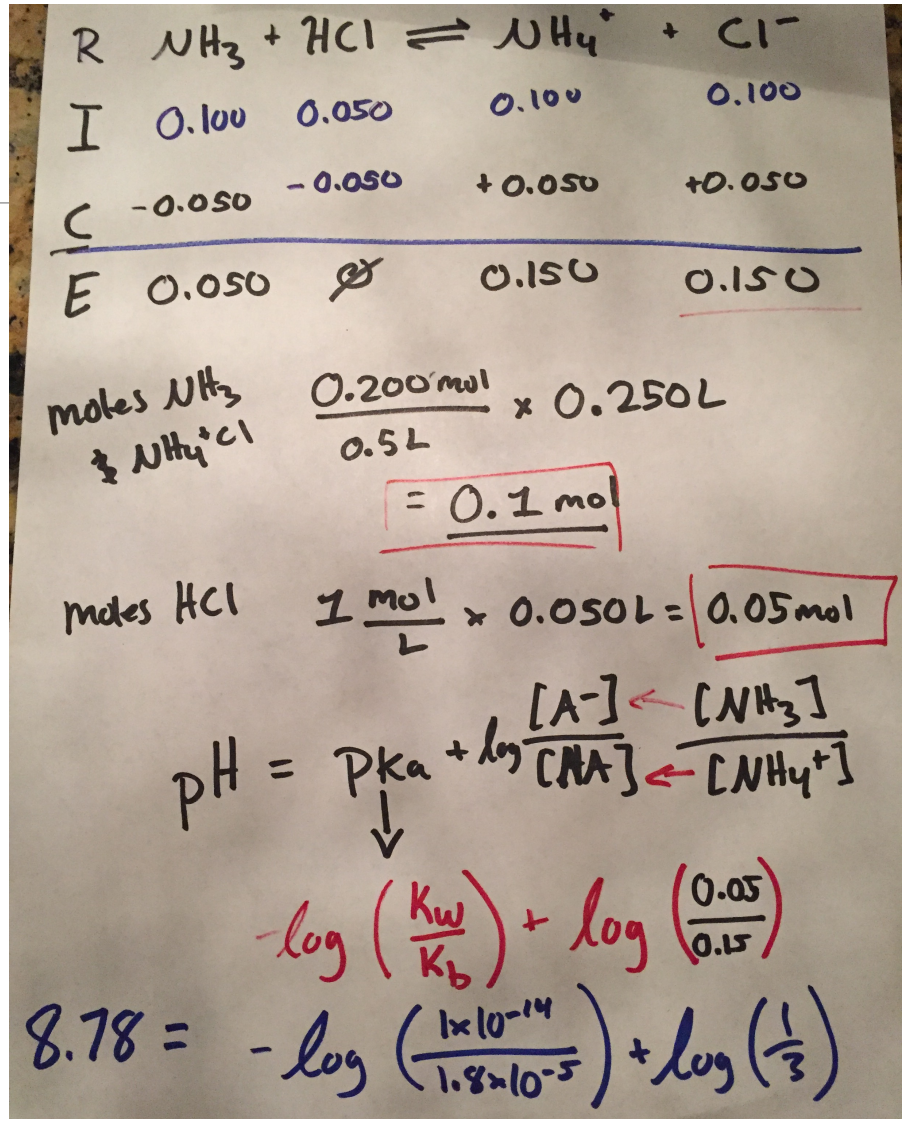
Titration Questions

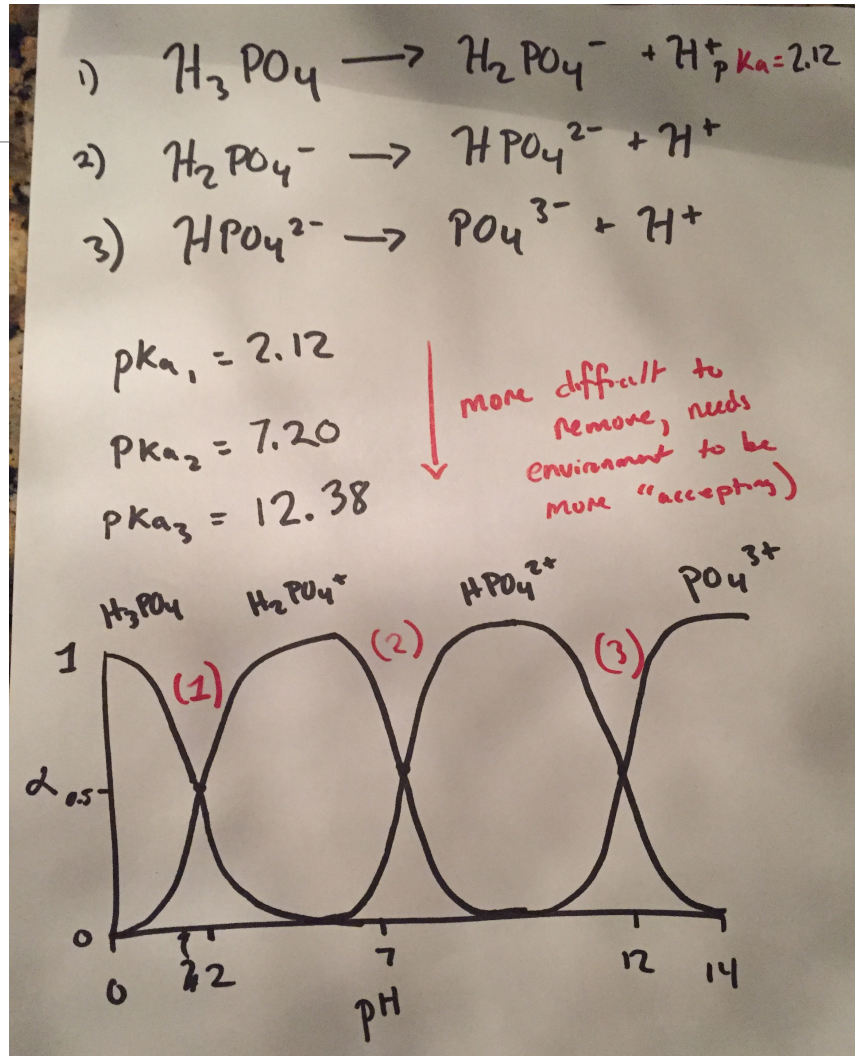




Exam Question

A buffer was prepared by mixing 0.200 mole of ammonia ($K_b = 1.8 \times 10^{-5}$) and 0.200 mole of ammonium chloride to form an aqueous solution with a total volume of 500 mL. To 250 mL of this solution was added 50.0 mL of 1.00 M HCl. What is the pH of this solution?



Fraction of Species

- **For weak acids, especially for organic compounds and polyprotic acids, you can determine the "protonation-state" of your molecule at any given pH.**
- A few simple rules go a long way with answering these types of questions:
  1. Think about your molecule. Your pK_a represents the pH where you have equal protonated and deprotonated states.
  2. Think about your environment. The pH compared to the pK_a will tell you whether your molecule is protonated or deprotonated.
  3. If pH < pK_a ; you have too many hydrogen ions in solution for it to be favorable for a hydrogen ion on your molecule to "pop off"
     - Think about it: when the environment is more acidic, it is more "proton-donating,"
  4. If pH > pK_a ; the environment around your molecule is thirsty for hydrogen ions, so it is favorable for a hydrogen ion on your molecule to pop off
     - Think about it: when the environment is more basic, it is more "proton-accepting."