Chemical Equilibrium

Name

Date of Work

TOTAL: 25 points

CHM 116 POST-LAB

Chemical Equilibrium: Calculating K_a of a Weak Acid

1. Insert ONE picture of yourself in full PPE here (**include the turmeric indicator**). ***Remember to (1) show your full body so that we see you are wearing shoes; (2) wear your safety glasses, buttoned lab coat and gloves; (3) cover lower legs including ankles (socks are required, even in Arizona); (4) tie back long hair in a ponytail or a bun; (5) remove jewelry.*

Insert Picture Here

2. Insert <u>three</u> pictures of your titration (either trial): **before any titrant was added** (i.e. after the addition of the turmeric indicator but before you add NaOH), **at the endpoint** (i.e. right when/just after the turmeric indicator changes color), **at the end of the titration** (i.e. when the pH stops changing). *Make certain that your pictures show the beaker and the pH meter, with the pH reading and color of solution clearly visible.*
*Include a figure legend below your picture (include the **color** and the **pH** for each) You may create one figure legend for all three pictures, or an individual legend for each picture.*

BEFORE TITRANT	**AT ENDPOINT**

AFTER TITRATION

3. Include the **data from the Trial 1 titration** here *(Table 1 from your procedure).* You may need to add or remove rows as needed to fit your data. **Indicate your endpoint by highlighting the endpoint pH yellow.**

Table 1. Trial 1 Equilibrium Constant Data

Drops NaOH Added (each interval)	mL NaOH Added (each interval)	mL NaOH Added (total)	pH	Color Observation
0	**0**	0	3.66	yellow
5	0.06*5=0.3	0.3	3.70	yellow
5	0.06*5=0.3	0.6	4.13	yellow
5	0.06*5=0.3	0.9	4.38	yellow
5	0.06*5=0.3	1.2	4.62	yellow
5	0.06*5=0.3	1.5	4.93	yellow
5	0.06*5=0.3	1.8	5.12	yellow
5	0.06*5=0.3	2.1	5.34	yellow
5	0.06*5=0.3	2.4	6.05	yellow
5	0.06*5=0.3	2.7	6.80	yellow
5	0.06*5=0.3	3	10.33	orange
5	0.06*5=0.3	3.3	10.59	orange
5	0.06*5=0.3	3.6	10.77	orange
5	0.06*5=0.3	3.9	10.90	orange
5	0.06*5=0.3	4.2	11.13	orange
5	0.06*5=0.3	4.5	11.23	orange
5	0.06*5=0.3	4.8	11.30	orange
5	0.06*5=0.3	5.1	11.41	orange
5	0.06*5=0.3	5.4	11.51	orange

4. In Question 3, you identified the pH of your endpoint. What is the value of the pH of your endpoint? Use your experimental observations to **explain why, in 1 – 2 sentences,** you chose this pH as your endpoint.

pH of endpoint: ___10.33___. I chose this as the endpoint because after adding five drops of NaOH, the color permanently changed from yellow to orange. The lab guide indicates that the change in color from yellow to dark pink, orange or red indicates the endpoint. This indicates a change from acidic to basic pH.

5. Include the **data from the Trial 2 titration** here *(Table 2 from your procedure).* You may need to add or remove rows as needed to fit your data. **Indicate your endpoint by highlighting the endpoint pH yellow.**

Table 2. Equilibrium Constant Data for **Trial 2**

Drops NaOH Added (each interval)	mL NaOH Added (each interval)	mL NaOH Added (total)	pH	Color Observation
0	**0**	0	3.66	yellow
5	0.06*5=0.3	0.3	3.91	yellow
5	0.06*5=0.3	0.6	4.12	yellow
5	0.06*5=0.3	0.9	4.34	yellow
5	0.06*5=0.3	1.2	4.52	yellow
5	0.06*5=0.3	1.5	4.65	yellow
5	0.06*5=0.3	1.8	4.73	yellow
5	0.06*5=0.3	2.1	4.95	yellow
5	0.06*5=0.3	2.4	5.16	yellow
5	0.06*5=0.3	2.7	5.26	yellow
5	0.06*5=0.3	3	5.48	yellow
5	0.06*5=0.3	3.3	5.77	yellow
5	0.06*5=0.3	3.6	6.17	yellow
5	0.06*5=0.3	3.9	7.18	yellow
5	0.06*5=0.3	4.2	10.24	orange
5	0.06*5=0.3	4.5	10.68	orange
5	0.06*5=0.3	4.8	10.91	orange
5	0.06*5=0.3	5.1	11.28	orange
5	0.06*5=0.3	5.4	11.35	orange

6. In Question 5, you identified the pH of your endpoint. What is the value of the pH of your endpoint? Use your experimental observations to **explain why, in 1 – 2 sentences,** you chose this pH as your endpoint.

[redacted]

7. **Which Trial** do you think resulted in **more accurate results** (a more accurate endpoint)? Use your experimental observation to explain your choice, **in 2 – 3 sentences.**

Trial # ___2___
I chose trial 2 because it was conducted while knowing the range of the endpoint and was, therefore, more careful. Trial I was a trial run and, therefore, has some errors.

8. Assuming the density of your initial acid solution is 1.0 g/mL, what is the **concentration (in molarity)** of the diluted acetic acid, **[CH_3COOH]**, solution? **Show your work and include units.**
Hint: An initial 4.5% w/w solution (w/w means 4.5 g of acid per 100 g of solution) of CH_3COOH (MW = 60.05 g/mol) was diluted by placing 0.5 mL into 80 mL of water to produce a final diluted solution with a total volume of 80.5 mL. You will also need to add in the volume of the indicator (from Part 4, Step 6) to your total volume, so it will be slightly higher than 80.5 mL. Use this information to first solve for the % w/w of the diluted acetic acid solution (this is a dilution problem: $M_{conc}V_{conc} = M_{dil}V_{dil}$) and then converting % w/w to molarity. (OR you can first convert the initial % w/w to molarity and then solve for the molarity of the diluted acid using the dilution equation.)

First, find total volume of dilute acetic acid:

V_{dil}=0.5 mL acid + 80 mL water + 0.75 mL

V_{dil} =81.25mL

Now calculate molarity of dilute acetic acid using hints above:
4.5g are in 100 mL solution since density is 1g/mL
Y g are in 0.5mL

$$y = \frac{4.5g \times 0.5mL}{100mL} = 0.0225g$$

Number of moles;

$$Moles = \frac{Mass}{Molar\ Mass} = \frac{0.0225g}{60.05g/mol} = 0.000375mol$$

$$Initial\ Molarity = \frac{0.000375 \times 1000}{0.5} = 0.75M$$

$$V1C1 = V2C2;\ C2 = \frac{V1C1}{V2} = \frac{0.5mL \times 0.75M}{81.25mL} = 0.00462M$$

9. Write the **balanced chemical equation** for the **equilibrium** that is established when acetic acid is dissolved in water. **Include the chemical formulas for all reactants and products, any stoichiometric coefficients and all physical states.** *Recall that when a substance is dissolved in water, the physical state in a chemical reaction is aqueous (aq).*

CH_3COOH (aq) ⇌ CH_3COO^- (aq) + H^+ (aq)

10. Based on the reaction in Question 9, write the **expression** for the equilibrium constant (K_a) of acetic acid.

$$Ka = \frac{[CH3COO^-] \times [H^+]}{[\mathrm{CH3COOH}]}$$

11. Using your **first pH measurement** from the **Trial you chose in Question 7**, what is the **initial [H$^+$] (in M)** present in your acetic acid solution before titrating? *Identify your first pH measurement in the blank provided below.* **Show your work and include units.**

Hint: Use the mathematical definition of pH to solve for [H$^+$].

[redacted]

12. Using your answers from Questions 8, 10, and 11, ***as well as an ICE table***, determine the **K_a** of acetic acid. **Show your work and include units.**

Hint: See Sample Exercise 16.10—Calculating K_a from Measured pH *in your textbook/ebook.*

[redacted]

$$Ka = \frac{[redacted]}{0.0044} = 1.088 \times 10^{-5}$$

13. Based on your answer in Question 12, what is the **pK_a** of acetic acid? **Show your work.**

$pKa = -LogKa = -Log(1.088 \times 10^{-5})$ =4.96

14. Addition of NaOH to the acetic acid solution neutralizes the acid. Write the **balanced chemical equation** for this ***neutralization reaction*** (*this is the reaction you performed in the experiment*). **Include the chemical formulas for all reactants and products, any stoichiometric coefficients and all physical states.**

CH_3COOH (aq) + NaOH (aq) → CH_3COONa (aq) + H_2O (l)

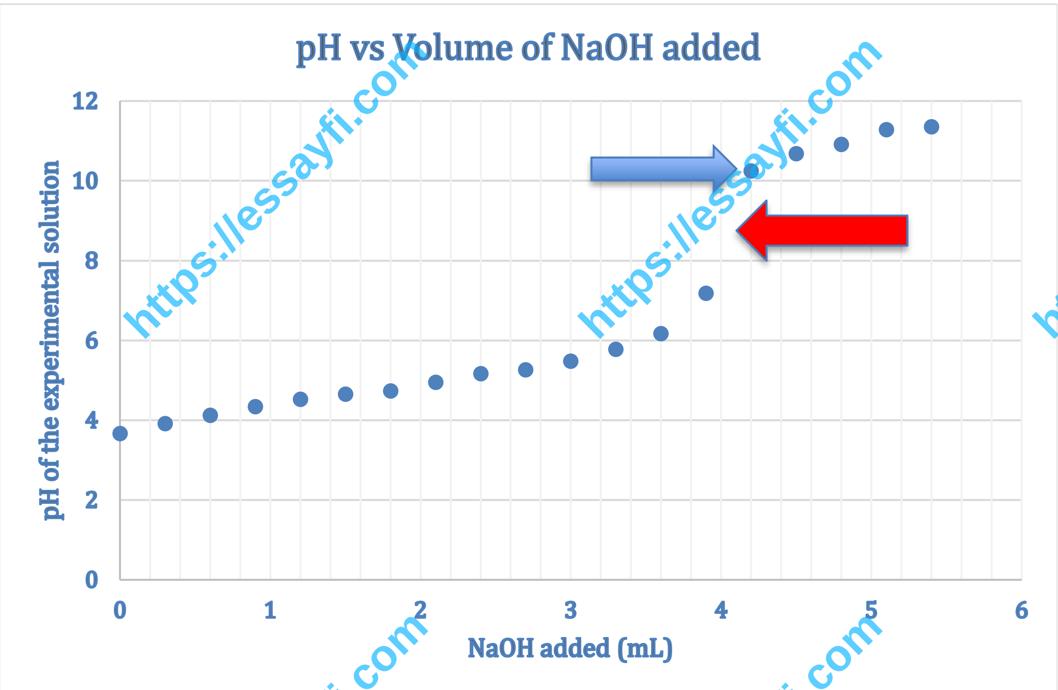
15. ***For the Trial you chose in Question 7*, *use graphing software to create an x-y scatter plot showing the pH of the experimental solution as a function of the volume of NaOH (the titrant) added (do NOT connect your points; no trend line needed).*** On your graph, the **x-axis should be the volume of NaOH added (mL)**, and the **y-axis should be the pH of the experimental solution** at that time. *****The endpoint and equivalence point are not necessarily the same point.***
Draw a BLUE arrow (use 'Insert Shape' feature to do this once you have copied your graph from Excel into your Post-Lab) on your graph indicating the data point that represents your endpoint .
Draw a second RED arrow on your graph indicating the point of the curve that represents the equivalence point.

Make sure to:

- **Label the axes including the units in parentheses.**
- **Include a figure legend (text below the graph) explaining the data in your graph that includes the arrow pointing to the end point and equivalence point.** In general, scientific graphs show the independent variable on the x-axis and the dependent variable on the y-axis.
- **Remove any default title, legend, and gridlines.**

You will not receive credit if you draw your graph by hand. Excel or Google Sheets are good choices that you can learn how to use quickly if you don't already have a favorite graphing program.

See other general tips for making graphs in the How to Make a Graph in Excel document located in the Introductory Materials for this lab.



Scatter plot for the pH of the experimental solution vs the volume of NaOH added. The Blue arrow shows the endpoint, while the red one indicates the point of equivalence.

16. In Question 15, you indicated the endpoint and equivalence point with arrows. Answer the questions below.

a. What **volume of NaOH (in mL)** did you add to reach the *endpoint* of the titration?

[redacted]

b. What **volume of NaOH (in mL)** did you add to reach the *equivalence point* of the titration?

4.1mL

c. Are these two points the same? **Explain, in 2 – 3 sentences,** using your experimental data as well as their definitions.

[redacted]

17. In this type of titration, half-way to the equivalence point (or, when half of the titrant has been added) is when **pH = pK$_a$**. Use your graph from Question 15 and the answer to Question 16a to find the half-equivalence point of the titration. Based on your data, what is the **pK$_a$** of acetic acid?

At endpoint, the volume of NaOH consumed was 4.2mL. Half this value, 2.1mL, indicates the half-equivalence point. At 2.1mL, the graph shows the pH was 4.95. Thus, pKa=4.95

18. Putting it all together! Look up the *actual value* (from a reputable source) for the pK_a of acetic acid at 25°C (be sure to include the source). Compare your answers to Questions 13 and 17 to the value in the literature using percent error; then discuss TWO potential *sources* of error for this experiment. **Show your work for each percent error and include units.**

Hint: "Human error" is NOT an acceptable source of error as it is too vague. Be specific as to the errors you, as the human, may have made during the experiment that resulted in a pK_a that differed from the literature value.

$$Percent\ Error = \frac{|Experimental\ Value - Actual\ Value|}{Actual\ Value} \times 100\%$$

Actual pK_a of acetic acid: __4.76__

Source: Libretexts (2023)

Percent Error for Experimental pK_a from Q13:

$$Percent\ Error = \frac{|Experimental\ Value - Actual\ Value|}{Actual\ Value} \times 100\%$$
$$= \frac{|4.96 - 4.76|}{4.76} \times 100 = 4.2\%$$

Percent Error for Experimental pK_a from Q17:

$$Percent\ Error = \frac{|Experimental\ Value - Actual\ Value|}{Actual\ Value} \times 100\%$$
$$= \frac{|4.95 - 4.76|}{4.76} \times 100 = 3.99\%$$

Two Sources of Error:

1. **Inaccurate measurements. There is a potential that the instruments used for measurement including the pH meter, pipette and measuring cylinder, were inaccurate. An error when measuring the volume of a drop is reflected in the subsequent calculations and in graphing.**
2. **Poor determination of the endpoint. The precision when determining the endpoint determines the calculations involving the endpoint and half equivalence point hence pKa. By adding five drops, I may have overshot the endpoint which could explain my slightly higher pKa.**

Reference

Libretexts. (2023, March 19). *E1: acid dissociation constants at 25°C*. Chemistry LibreTexts. https://chem.libretexts.org/Ancillary_Materials/Reference/Reference_Tables/Equilibrium_Constants/E1%3A_Acid_Dissociation_Constants_at_25C