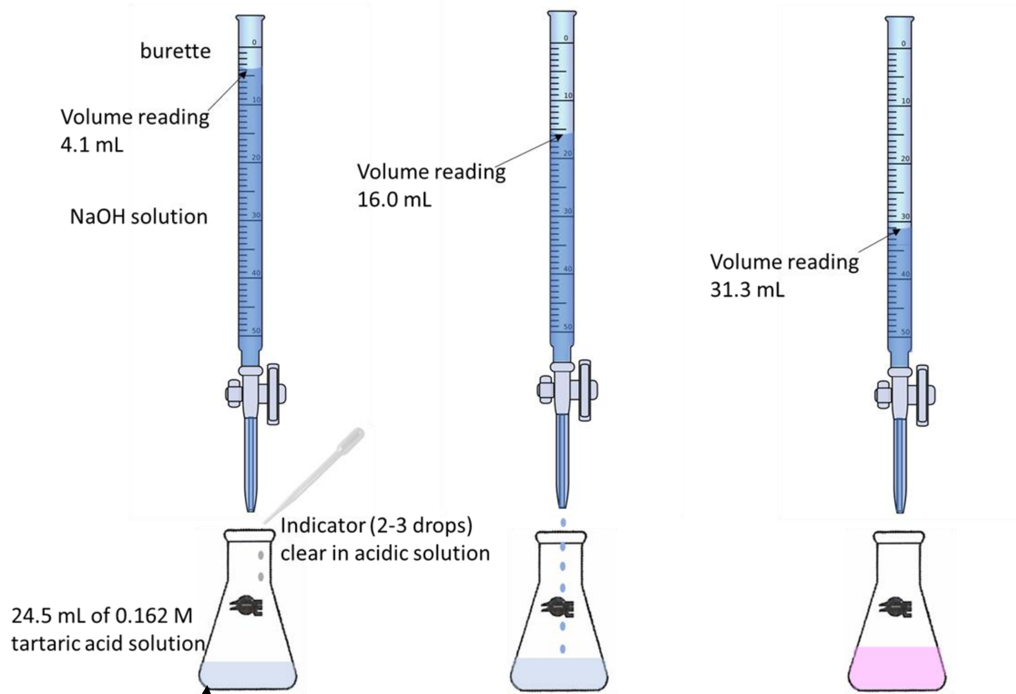
# How Much Acid Is in Lemon/Lime Juice?

**Group member:** Brodric **Group member:** Chris
**Group member:** Kobe **Group member:** Clark

**18 minutes to next solid line**

**Information:** A <u>titration</u> is a technique used to determine the unknown concentration of a reactant in solution. The <u>analyte</u> is the reactant of unknown concentration while the <u>titrant</u> is known. An <u>indicator</u> is used to signal when the reaction is complete (commonly called the <u>endpoint</u>). The endpoint is identified by a significant color change of the indicator. The most common type of titration is between an acid and a base.

<u>Model 1:</u> Determining the concentration of an unknown NaOH solution.




$$H_2C_4H_4O_6(aq) + 2NaOH(aq) \rightarrow Na_2C_4H_4O_6(aq) + 2H_2O(l)$$

$$Equation\ 1:\ 24.5\ mL \times \frac{1\ L}{1000\ mL} \times \boxed{\frac{0.162\ moles\ H_2C_4H_4O_6}{1 L}} = 0.00397\ moles\ H_2C_4H_4O_6$$

$$Equation\ 2:\ 0.00397\ moles H_2C_4H_4O_6 \times \boxed{\frac{2\ mole\ NaOH}{1\ mole\ H_2C_4H_4O_6}} = 0.00794\ moles\ NaOH$$

<u>Key Questions:</u>

1. The titrant for this reaction is tartaric acid ($H_2C_4H_4O_6$). A known volume from a 0.162 *M* solution is added to the Erlenmeyer flask. The calculation to determine the moles in the flask is shown in Equation 1.
    a. What is the tartaric acid volume added to the flask? 24.5 mL

b. Put a square around the conversion factor originating from the molarity.
c. Why was the volume converted to liters?

The concentration is in moles/Liter so you should either convert the concentration to moles/mL or the way it was done here with volume from mL to L

2. The indicator used to signal the end of a reaction is colorless in solutions that are acidic and pink in solutions that are basic. The acid-base reaction will be complete when the indicator just barely changes from being colorless to pink. Based on the appearance of the flask on the left, is the solution acidic or basic? __acidic__ How do you know this?

   it's colorless, looks like water

3. Some of the NaOH solution in the burette is added to the flask (middle of the figure).
   a. Determine the volume of the NaOH solution that was added. __4.1mL__

   b. After adding this amount of NaOH, the contents in the flask are still acidic. How do you know this?

   it's still colorless

Since the indicator has not changed color, the reaction is still incomplete (endpoint has not been reached).

4. For the flask on the right, just enough NaOH has been added to change the color of the indicator.
   a. Is the reaction complete? __Yes__ How do you know this?

   It changed color from colorless to pink meaning it's complete

   b. Determine the total volume of the NaOH solution that was added. __31.3 mL__

5. The reaction between the tartaric acid and sodium hydroxide is shown in the model. Note: Tartaric acid is *diprotic* since it donates *two protons* for the acid-base reaction.
   a. By examining the formula for tartaric acid, explain why 2 NaOH's are needed for every 1 $H_2C_2H_4O_6$.

   To complete the entire reaction since there needs to be twice as many moles of NaOH as there needs to be H2C2H4O6

   b. Equation 2 shows the calculation to determine the moles of NaOH neutralized by the tartaric acid solution.

i. What is the mole ratio of NaOH to $H_2C_2H_4O_6$ in the balanced chemical equation? 2:1

ii. Put a square around this mole ratio conversion factor in Equation 2.

c. The molarity of the NaOH solution can now be determined by dividing the moles NaOH neutralized by the volume (in liters) dispensed by the burette. Perform this calculation. Show your work.

$$\frac{0.00794mol}{0.0313L} = 0.254M$$

You have just performed a titration. The concentration of the NaOH solution was determined by using a known amount of tartaric acid.

6. The burette was filled by a NaOH solution prepared in a plastic bottle. The solution in the bottle was prepared by using 4 mL of a concentrated 8 *M* NaOH solution followed by filling the bottle to the shoulder. The total volume of the bottle is 100 mL. The equation for calculating the diluted solution concentration is shown below.

$M_1V_1 = M_2V_2$

| |
|---|
| $M_1$ = 8 M |
| $V_1$ = 4 mL |
| $M_2$ = ? |
| $V_2$ = 100 mL |

a. What is the diluted solution molarity, $M_2$? 0.3 M

b. Why does it only have one significant figure?

M1 and V2 only have one significant figure and since its multiplication/division the answer should only have one as well

7. Compare the molarity determined in the previous question with that determined by performing a titration. What is the obvious advantage in performing a titration?

A titration is more exact

Exercises:

8. A burette is filled with a NaOH solution having a concentration of 0.150 M (the initial volume reading on the syringe is 2.1 mL). The NaOH solution is added to a solution that contains tartaric acid until the reaction is complete. The final reading on the burette is 39.5 mL. Calculate the moles of tartaric acid in the solution. Show your work.

$$\frac{0.150mol}{L} \cdot (0.0395 - 0.0021)\,L = 0.00561mol, NaOH$$

$$0.00561mol, NaOH \cdot \frac{1mol, tartaricacid}{2mol, NaOH} = 0.00281mol, tartaricacid$$

Entire class discussion:

Method development:

Items needed for each member of the group: 125 mL plastic drop bottle, 6 *M* sodium hydroxide (1 container per group), 10 mL graduated cylinder, tartaric acid (1 container per group), spatula, 50 mL volumetric flask, small funnel, burette attached to stand, adjustable pipette (one per group, return when not in use), pipette tip (located on first bench), 50 mL beaker, 50 mL Erlenmeyer (pear-shaped) flask, phenolphthalein indicator (1 container per group).


9. The estimated digit for the burette is in the hundredths place. Make a volume reading from the burette in the provided picture. Everybody in your group should have a similar measurement. Whenever you make a measurement with the provided burette it will be to the hundredths place.
   ___0.28___ mL
10. What will be the molarity of a diluted solution if 5 mL of a 6 *M* NaOH solution is added to a 125 mL container followed by filling it with water? Show your work and pay attention to significant figures. This diluted solution will be prepared later as indicated.

$$\frac{\left(\frac{6mol}{L} \cdot 0.005L\right)}{0.125L} = 0.24M$$

11. A titration will be performed to determine the molarity of the diluted sodium hydroxide solution with more precision and accuracy. Briefly describe a method to perform this titration using a tartaric acid solution of known concentration. The information in Model 1 will also assist you in describing this method.

Add an amount on NaOH solution to the tartaric acid until the reaction is complete and then that amount of NaOH solution add and the amount of moles of tartaric acid can be used to calculate the molarity of the NaOH solution

Individual data collection (each person in the group will acquire their own data):

12. The concentration determined previously for the diluted NaOH solution is a rough estimate. You will perform a titration to determine its concentration with more precision and accuracy. The following steps will assist you.

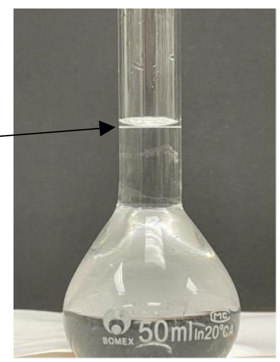
a. Pour approximately 5 mL of 6 *M* sodium hydroxide into the 125 mL plastic container. Fill the container to the shoulder with distilled water, screw the cap on snuggly, and shake to mix. **CAUTION: 6 *M* sodium hydroxide is extremely corrosive. Be sure to rinse your hands after usage.** The approximate molarity of the solution was calculated above.

b. Measure out between 0.66 and 0.68 grams of tartaric acid using a weigh boat and record the mass below. Using a small plastic funnel transfer the tartaric acid to a 50.0 mL volumetric flask. Rinse the funnel with distilled water to verify that all the tartaric acid is transferred to the flask. Convert this mass to moles.

| | |
|---|---|
| **Mass of tartaric acid** | 0.670 g |
| **Moles of tartaric acid** | 0.00456 mol |

$$\frac{0.670}{150.087} = 0.00456$$

c. Fill the flask about half full with distilled water and swirl until the solid is fully dissolved (*do not* put the cap on and invert the flask at this point). Now, carefully adjust the volume until the bottom of the meniscus is on the calibration ring (this is the 50.0 mL mark). Stopper the flask, and while holding the stopper, invert the flask completely and shake the contents. Repeat the inversion and shake several times until the solution is thoroughly mixed.



The volume in the volumetric flask is 50.0 mL. How many significant figures are in this volume measurement? ___3___ This is important!!!

d. Calculate the molarity of the tartaric acid solution. Show your work below.

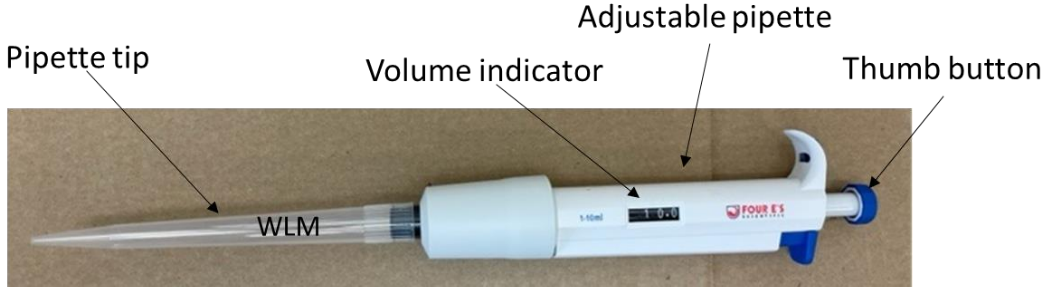
$$\frac{0.00456 mol}{0.050 L} = 0.0912 M$$

| | |
|---|---|
| **Tartaric acid solution molarity** | 0.0912 M |

e. Open the valve, empty the 10 mL burette, and rinse with several small portions of your diluted NaOH solution. Allow the rinse to drain completely through the stopcock into a waste beaker (150 mL). Close the valve on the burette until the liquid just barely stops flowing through the tip. Note: DO NOT OVERTIGHTEN

THE VALVE WHEN CLOSING! Fill the burette with your diluted NaOH solution. Open the stopcock and allow the burette tip to fill with solution (no air bubbles in the tip). Record the initial burette reading in the table below. The reading needs to be between 0.00 mL and 0.50 mL. Add more diluted NaOH solution if it is not. Note: Each student should have their own burette.

f. Transfer 10.0 mL of your tartaric acid solution into the Erlenmeyer flask using the adjustable pipette following the instructions below.
   i. Pour 12-15 mL of the tartaric acid solution into a *clean* 50 mL beaker.
   ii. Label a pipette tip with your initials and secure it snuggly to an adjustable pipette. Make sure the reading on the adjustable pipette is 10.0 mL.

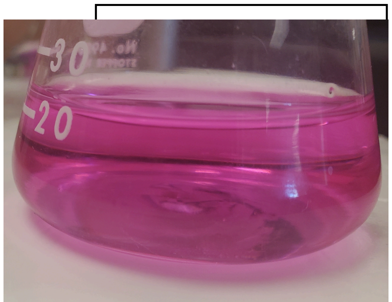



   iii. Place the pipette tip into the solution and depress the thumb button to the first stop (see picture). Allow the thumb button to return back slowly to the original position *without* introducing any air in the tip.
   iv. Transfer the entire solution in the tip to the 50 mL Erlenmeyer flask by slowly depressing the thumb button to the first stop, holding for a second, and then depressing to the second stop. The second stop is slightly beyond the first stop.
   v. Remove and keep the tip for another trial.

g. Add 1-2 drops of the phenolphthalein indicator to the tartaric acid solution in the Erlenmeyer flask. What is the color of the indicator in the solution (record this in the table below)?

13. You will now add the diluted NaOH solution from the burette to the tartaric acid solution in the Erlenmeyer flask until the acid-base reaction is complete (i.e., the endpoint is acquired). Remember, the endpoint is acquired when the solution just barely turns pink (and stays pink when mixed).
    a. You can safely add 5.0 mL of the diluted NaOH solution from the burette without reaching the endpoint. Make sure to swirl the solution after each addition.
    b. After adding 5.0 mL of NaOH solution from the burette, you will carefully add a the NaOH solution dropwise (drop by drop) and swirl the contents in the flask after each addition.

c. Keep adding small amounts of NaOH solution and swirling until the endpoint is acquired. The solution should barely turn pink and stay that color. Record the final volume on the burette. This should not take more than 9.0 mL.
d. What is the color of the solution at the endpoint (record in the table below)? Take a clear picture of this solution *at the endpoint*.




e. Determine the volume of NaOH solution added to the flask and place it in the table.
f. Discard the contents of the Erlenmeyer flask in the sink, rinse the flask with purified water, fill with another 10.0 mL of your tartaric acid solution using an automatic pipettor, refill the burette with your dilute NaOH solution, and perform another titration (Trail 2).

| | **Trial 1** | **Trail 2** |
|---|---|---|
| **Initial burette reading for NaOH solution** | 0.00 mL | 0.00 mL |
| **Volume of tartaric acid solution** | 10.0 mL | 10. 0 mL |
| **Color of indicator in tartaric acid solution before endpoint.** | colorless | colorless |
| **Final burette reading for NaOH solution** | 9.73 mL | 9.81 mL |
| **Color of indicator in solution at endpoint** | Pink | Pink |
| **Volume of NaOH solution added to flask** | 9.73 mL | 9.81 mL |

14. Calculate the moles of tartaric acid in the 10.0 mL added to the flask. This will be the same for each trial. Show your work and use dimensional analysis. Place these values in the table below.

$$\cancel{0.00456\,\frac{mol}{L}} \cdot 0.01L = 0.0000456mol$$

15. Using the balanced chemical equation below, calculate the moles of NaOH that were neutralized by the tartaric acid for each trial. Show your work and use dimensional analysis. Place the answers in the table.
$H_2C_4H_4O_6(aq) + 2NaOH(aq) \rightarrow Na_2C_4H_4O_6(aq) + 2H_2O(l)$

$$0.0000456 mol, tartaric acid \cdot \frac{2 mol, NaOH}{1 mol, tartaric acid} = 0.0000912 mol, NaOH$$

16. Now, determine the molarity of the diluted NaOH solution in the plastic container for each trial. Show your work. Place the answer in the table below. Your value should be fairly close to the approximate value calculated previously. If it's not, you probably did something wrong.

Trial 1:

$$\frac{0.0000912 mol}{0.00973 L} = 0.00937 M$$

Trial 2:

$$\frac{0.0000912 mol}{0.00981 L} = 0.00930 M$$

17. Determine an average molarity for your diluted NaOH solution and place it in the table.

| | **Trial 1** | **Trail 2** |
|---|---|---|
| **Moles of tartaric acid for titration** | 0.0000456 mol | 0.0000456 mol |
| **Moles of NaOH neutralized** | 0.0000912 mol | 0.0000912 mol |
| **Molarity of NaOH solution in container** | 0.00937 M | 0.00930 M |
| **Average molarity of NaOH solution** | | 0.00933 M |

Compilation of group data and group analysis:

18. In the table below, input the molarity for the NaOH solution for each group member (including your own). Calculate an average value.

| | **Molarity** |
|---|---|
| **Group member** | 0.00933 M |
| **Group member** | 0.00935 M |
| **Group member** | 0.0987 M |
| **Group member** | 0.0949 M |
| **AVERAGE** | 0.0531 M |

19. The reported molarities of the group members are not expected to be exactly the same (even though they should be somewhat similar). Why is this?

We didn't dilute the NaOH the same measured amount, and our tartaric acid was different amounts of grams

20. If there is a molarity *significantly* different from the others, suggest a reason for this.

2 of us were really close and the other 2 were really close, and then both groups were off by a factor of 10 so it's likely it was a mistake in calculations

*Ask the teacher or lab assistant to check your work for the calculations above before proceeding.*

**Class data: In the class data sheet (link in Canvas), input the average molarity value.**

Method development:

Items needed for *each* member of the group: lemon juice (one container per group) and lime juice (one container per group).

21. Examine the containers of lemon juice and lime juice. What do you think is the first ingredient for each juice? water
22. The second ingredient is citric acid (this is what primarily produces the sour taste). It has the formula $H_3C_6H_5O_7$. In a reaction with a strong base such as NaOH, it will donate *three* protons (triprotic). These protons ($H^+$s) are indicated first in the formula. Write the balanced chemical equation for the acid-base reaction between $H_3C_6H_5O_7$ and NaOH.

$$H_3C_6H_5O_7 + 3NaOH \rightarrow Na_3C_6H_5O_7 + 3H_2O$$

23. Two members of your group will perform a titration to determine the amount of citric acid in lemon juice and two in lime juice. Make these assignments and record them below. For groups of only three, only one person will be assigned lime juice.

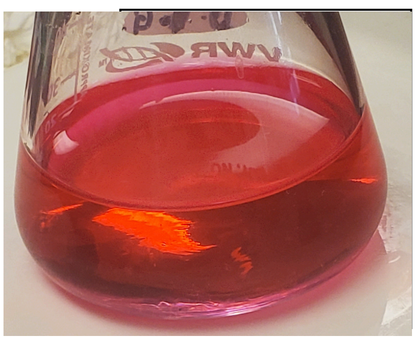
| Group member | Juice |
|---|---|
| Chris | Lemon juice |
| Clark | Lemon juice |
| Kobe | Lime juice |
| Brodric | Lime juice |

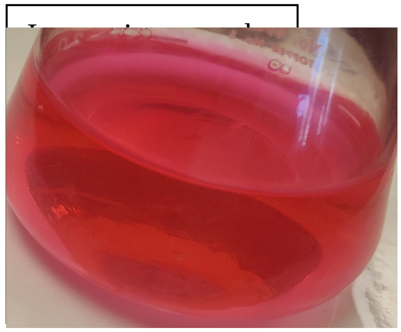
24. A titration will be performed with the NaOH solution to determine the amount of citric acid in lemon or lime juice. Describe a method to perform this titration using the provided items.

Get the mass of the juice and the volume of how much NaOH solution is used to complete the reaction then we can find the moles of citric acid using that and then using that and the juice's mass you can find the percent of it in there

Individual data collection (each person in the group will acquire their own data):

25. The below information will assist you in performing the titration to determine the amount of citric acid in lemon juice or lime juice.
    a. Shake the container of lemon juice or lime juice and then open it. Place between 2.2 and 2.4 grams of lemon juice or lime juice in a *clean* Erlenmeyer flask (record the mass of the drops in the table below).
    b. Dilute the juice with approximately 10 mL of water.
    c. Add 1-2 drops of indicator and record the color of the solution in the table. The color of the solution did not change at this point. Why not?
    d. Fill the burette with the dilute NaOH solution and record the initial burette reading in the table (reading needs to be between 0.0 and 0.5 mL).
26. You will now add the NaOH solution from the burette to the juice until the acid-base reaction is complete (the endpoint is acquired).
    a. You can safely add around 4 mL of NaOH solution without reaching the endpoint.
    b. After adding 4 mL of NaOH solution, you will carefully add a small amount of NaOH solution (drop by drop) from the burette and swirl the contents in the flask after each addition. If the reaction is incomplete, the color will be the same.
    c. Keep adding a small amount of NaOH and swirling until the endpoint is acquired (record the final reading in the table below). There will be an obvious color change.
    d. What is the color of the solution at the endpoint (record in the table below)? Take a clear picture of this solution *at the endpoint*.



    e. Determine the volume of NaOH solution added to the flask and place it in the table.
    f. Discard the contents of the flask in the sink, rinse the flask with purified water, add another 2.3 to 2.5 grams of juice, refill the burette with dilute NaOH solution, and perform another titration (Trail 2).

| | **Trial 1** | **Trail 2** |
|---|---|---|
| **Mass of lemon juice or lime juice** | 2.388 g | 2.322 g |
| **Color of juice solution** | cloudy white | cloudy white |
| **Initial burette reading for NaOH solution** | 0.00 mL | 0.00 mL |

| | | |
|---|---|---|
| **Final burette reading for NaOH solution** | 9.91 mL | 9.82 mL |
| **Color of solution at endpont** | Pink | Pink |
| **Volume of NaOH solution added to flask** | 9.91 mL | 9.82 mL |

27. Using your individual average molarity of the diluted NaOH, determine the moles of NaOH that was added from the burette for this reaction. Show your work and use dimensional analysis. Place these values in the table.

Trial 1:

$$0.00933\,\frac{mol}{L}\cdot 0.00991L = 0.0000925mol, NaOH$$

Trial 2:

$$0.00933\,\frac{mol}{L}\cdot 0.00982L = 0.0000916mol, NaOH$$

28. Using the chemical equation below, calculate the *grams* of citric acid that was neutralized by the NaOH. Show your work and use dimensional analysis. Place these values in the table.

$H_3C_6H_5O_7(aq) + 3NaOH(aq) \rightarrow 3H_2O(l) + Na_3C_6H_5O_7(aq)$

Trial 1:

$$0.0000925mol, NaOH\cdot\frac{1mol, H_3C_6H_5O_7}{3mol, NaOH} = 0.0000308mol, H_3C_6H_5O_7\cdot 192.124\,\frac{g}{mol} = 0.00592g$$

Trial 2:

$$0.0000916mol, NaOH\cdot\frac{1mol, H_3C_6H_5O_7}{3mol, NaOH} = 0.0000305mol, H_3C_6H_5O_7\cdot 192.124\,\frac{g}{mol} = 0.00586g$$

29. From the mass of the juice you added to the flask, determine the percent citric acid in the juice. Show your work. Place these values in the table.

Trial 1:

$$\frac{0.00592}{2.388}\cdot 100\% = 0.248\%$$

Trial 2:

$$\frac{0.00586}{2.322}\cdot 100\% = 0.252\%$$

30. Determine an average percent citric acid in the juice and place it in the table.

$$\frac{0.248+0.252}{2} = 0.250\%$$

| | **Trial 1** | **Trail 2** |
|---|---|---|
| **Moles of NaOH neutralized** | 0.0000925 mol | 0.0000916 mol |
| **Grams of citric acid neutralized** | 0.00592 g | 0.00586 g |
| **Percent citric acid in juice** | 0.248% | 0.252% |
| **Average percent citric acid in juice** | | 0.250% |

<u>Compilation of group data and group analysis:</u>

31. Record the percent citric acid in lemon juice or lime juice for each group member. Calculate an average for your group.

| | **Percent citric acid in lemon juice** | | **Percent citric acid in lime juice** |
|---|---|---|---|
| **Group member** | 1.8% | **Group member** | 0.250% |
| **Group member** | 1.7% | **Group member** | 0.251% |
| **AVERAGE PERCENT CITRIC ACID** | 1.75% | | 0.2505% |

32. Are these values determined by your group members for percent citric acid precise? No

a. Explain how you made this determination.

2 are precise and the other 2 are precise but altogether its far from precise because they differ by about 1.5%

b. If the percent citric acid values aren't precise, suggest a reason.

It's likely a mistake in calculations or could be doing the experiment differently or with different amounts

**Class data: In the class data sheet (link in Canvas), input the average value for percent citric acid.**

Make sure all items are clean and return them to the drawer or bench. Finish cleaning any glassware used by rinsing with distilled water. Before you leave, ask the teacher or lab assistant to examine your data and laboratory space. Note: Your data must be entered into the spreadsheet before you leave.

<u>Application of principles (to be completed individually)</u>:

33. Lemon juice and lime juice also contain malic acid, $H_2C_4H_4O_5$ (albeit in much smaller amounts). It will donate two protons (diprotic) in a reaction with a strong base such as NaOH. Write the balanced chemical equation for the acid between malic acid and NaOH.

$$H_2C_4H_4O_5 + 2NaOH \rightarrow Na_2C_4H_4O_5 + 2H_2O$$

34. Lactic acid, $HC_3H_5O_3$ (monoprotic acid), is produced in muscle and red blood cells. Determine the grams of lactic acid which will be neutralized by 0.00329 moles of NaOH. You must show your work by writing below or inserting a picture of your work. Use dimensional analysis.

$$0.00329 mol, NaOH \cdot \frac{1 mol, HC_3H_5O_3}{1 mol, NaOH} = 0.00329 mol, HC_3H_5O_3 \cdot 90.08 \frac{g}{mol} = 0.296 g$$

35. Vinegar contains a small percentage of acetic acid, $HC_2H_3O_2$. A titration was performed to determine the amount of acetic acid, $HC_2H_3O_2$, in vinegar. In a flask, 2.85 grams of vinegar was titrated with 0.086 M sodium hydroxide. The initial volume of the sodium hydroxide solution in a burette was 53 mL. At the end point, the final volume in the burette was 25 mL. Determine the percent acetic acid in the vinegar. You must show your work by writing below or inserting a picture of your work. Note: It is needful to write the balanced chemical equation for the acid-base reaction between NaOH and $HC_2H_3O_2$.

$$0.086 \frac{mol}{L} \cdot (0.053 - 0.025)\, L = 0.0024 mol, NaOH \cdot \frac{1 mol, HC_2H_3O_2}{1 mol, NaOH} = 0.0024 mol, HC_2H_3O_2$$

$$0.0024 mol \cdot 60.052 \frac{g}{mol} = 0.14 g, HC_2H_3O_2 \qquad \frac{0.14}{2.85} \cdot 100\% = \underline{4.9\%}$$

36. You are now going to compare your individual value with the class average value in the *Class Data Sheet*. Make sure the *Class Data Sheet* is complete before doing this. You will have to wait until the end of the lab period. Fill in the following table below.

| | **Your individual value** | **Average value of class** |
|---|---|---|
| **Percent Citric Acid** | 0.250% | 3.50% |

The agreement between a measured value and an accepted value can be analyzed by performing a percent error calculation (equation given below). For this calculation, the measured value is your individual value, and the accepted value is the class average (in Google Sheet).

$$Percent\ Error = \frac{(measured\ value - accepted\ value)}{accepted\ value} \times 100.0\%$$

a. Calculate a percent error for your individual value.

| | **Percent error** |
|---|---|
| **Percent Citric acid** | |

$$\frac{0.250 - 3.50}{3.50} \cdot 100\% = -92.9\%$$

b. If your percent error is greater than 10% (or less than -10%), suggest at least one reason for your individual value being significantly different than the accepted value.

Likely to be an error in my calculations or I might have added too much NaOH solution to the reaction as well

| 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 1 **H** 1.008 | | | | | | | | | | | | | | | | | 2 **He** 4.0026 |
| 3 **Li** 6.94 | 4 **Be** 9.0122 | | | | | | | | | | | 5 **B** 10.81 | 6 **C** 12.011 | 7 **N** 14.007 | 8 **O** 15.999 | 9 **F** 18.998 | 10 **Ne** 20.180 |
| 11 **Na** 22.990 | 12 **Mg** 24.305 | | | | | | | | | | | 13 **Al** 26.982 | 14 **Si** 28.085 | 15 **P** 30.974 | 16 **S** 32.06 | 17 **Cl** 35.45 | 18 **Ar** 39.948 |
| 19 **K** 39.098 | 20 **Ca** 40.078 | 21 **Sc** 44.956 | 22 **Ti** 47.867 | 23 **V** 50.942 | 24 **Cr** 51.996 | 25 **Mn** 54.938 | 26 **Fe** 55.845 | 27 **Co** 58.933 | 28 **Ni** 58.693 | 29 **Cu** 63.546 | 30 **Zn** 65.38 | 31 **Ga** 69.723 | 32 **Ge** 72.630 | 33 **As** 74.922 | 34 **Se** 78.97 | 35 **Br** 79.904 | 36 **Kr** 83.798 |
| 37 **Rb** 85.468 | 38 **Sr** 87.62 | 39 **Y** 88.906 | 40 **Zr** 91.224 | 41 **Nb** 92.906 | 42 **Mo** 95.95 | 43 **Tc** (98) | 44 **Ru** 101.07 | 45 **Rh** 102.91 | 46 **Pd** 106.42 | 47 **Ag** 107.87 | 48 **Cd** 112.41 | 49 **In** 114.82 | 50 **Sn** 118.71 | 51 **Sb** 121.76 | 52 **Te** 127.60 | 53 **I** 126.90 | 54 **Xe** 131.29 |
| 55 **Cs** 132.91 | 56 **Ba** 137.33 | 57-71 * | 72 **Hf** 178.49 | 73 **Ta** 180.95 | 74 **W** 183.84 | 75 **Re** 186.21 | 76 **Os** 190.23 | 77 **Ir** 192.22 | 78 **Pt** 195.08 | 79 **Au** 196.97 | 80 **Hg** 200.59 | 81 **Tl** 204.38 | 82 **Pb** 207.2 | 83 **Bi** 208.98 | 84 **Po** (209) | 85 **At** (210) | 86 **Rn** (222) |
| 87 **Fr** (223) | 88 **Ra** (226) | 89-103 # | 104 **Rf** (265) | 105 **Db** (268) | 106 **Sg** (271) | 107 **Bh** (270) | 108 **Hs** (277) | 109 **Mt** (276) | 110 **Ds** (281) | 111 **Rg** (280) | 112 **Cn** (285) | 113 **Nh** (286) | 114 **Fl** (289) | 115 **Mc** (289) | 116 **Lv** (293) | 117 **Ts** (294) | 118 **Og** (294) |

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|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| * Lanthanide series | 57 **La** 138.91 | 58 **Ce** 140.12 | 59 **Pr** 140.91 | 60 **Nd** 144.24 | 61 **Pm** (145) | 62 **Sm** 150.36 | 63 **Eu** 151.96 | 64 **Gd** 157.25 | 65 **Tb** 158.93 | 66 **Dy** 162.50 | 67 **Ho** 164.93 | 68 **Er** 167.26 | 69 **Tm** 168.93 | 70 **Yb** 173.05 | 71 **Lu** 174.97 |
| # Actinide series | 89 **Ac** (227) | 90 **Th** 232.04 | 91 **Pa** 231.04 | 92 **U** 238.03 | 93 **Np** (237) | 94 **Pu** (244) | 95 **Am** (243) | 96 **Cm** (247) | 97 **Bk** (247) | 98 **Cf** (251) | 99 **Es** (252) | 100 **Fm** (257) | 101 **Md** (258) | 102 **No** (259) | 103 **Lr** (262) |