Chem 112L – M_____ Name:_____/Partner_____

GTA:_____ Date: _____

Lab #9 – Concentration

1. Concentration
   - 1.1. Units of concentration
     - 1.1.1. Molarity
     - 1.1.2. Molality
     - 1.1.3. Percent by mass or volume
     - 1.1.4. Parts per million, parts per billion
   - 1.2. Stoichiometry in solutions
     - 1.2.1. How to use units of concentration in calculations.
     - 1.2.2. Conversions between different types of units of concentration.
   - 1.3. Serial dilutions
     - 1.3.1. Using volumetric flasks for serial dilutions
     - 1.3.2. Calculating concentration in serial dilutions
   - 1.4. Titration for measurement of concentration
     - 1.4.1. Beer's Law
     - 1.4.2. Point slope method for determining the concentration of a solution.

2. **Learning Objectives**
   - 2.1. Demonstrate the use of a burette to dispense liquid reagents.
   - 2.2. Determine the end point of a reaction based on an indicator.
   - 2.3. Validate the published concentration of acetic acid in vinegar.
   - 2.4. Practice converting between different units of concentration.
   - 2.5. Practice creating specific dilutions.
   - 2.6. Observation of how Beer's Law is used in chemistry using a calibration curve.
   - 2.7. Practice using the point-slope method for practical applications.

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3. **Equipment**
   - 3.1. Erlenmeyer Flask
   - 3.2. Ring stand set
   - 3.3. Buret
   - 3.4. Magnetic stir bar
   - 3.5. Stir plate
   - 3.6. Cuvette
   - 3.7. Colorimeter

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4. **Chemicals**
   - 4.1. Vinegar (acetic acid)
   - 4.2. NaOH (aq)
   - 4.3. Phenolphthalein
5. **Additional Resources:**
   - 5.1. https://www.thoughtco.com/definition-of-solution-604650
   - 5.2. https://openstax.org/books/chemistry-atoms-first-2e/pages/6-3-molarity
   - 5.3. https://openstax.org/books/chemistry-atoms-first-2e/pages/6-4-other-units-for-solution-concentrations
   - 5.4. https://chem.libretexts.org/Bookshelves/Physical_and_Theoretical_Chemistry_Textbook_Maps/Supplemental_Modules_(Physical_and_Theoretical_Chemistry)/Spectroscopy/Electronic_Spectroscopy/Electronic_Spectroscopy_Basics/The_Beer-Lambert_Law
   - 5.5. https://chemistrytalk.org/beer-lambert-law/
   - 5.6. https://chem.libretexts.org/Bookshelves/Analytical_Chemistry/Molecular_and_Atomic_Spectroscopy_(Wenzel)/1%3A_General_Background_on_Molecular_Spectroscopy/1.2%3A_Beers_Law
   - 5.7. https://youtu.be/pbbHg_9zFbI
   - 5.8. https://byjus.com/chemistry/absorbance-vs-concentration/

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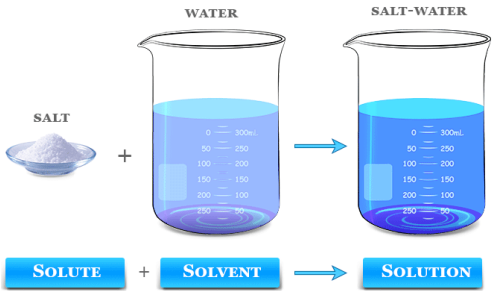
6. Introduction

6.1. Concentration

Concentration is a very important concept in chemistry. It is central to the discussion of solutions chemistry. Solutions are defined as homogeneous mixtures of two or more substances that cannot be mechanically filtered to separate them. Solutions can be in any phase: solid, liquid, or gas. They may start out as two substances in different phases that when put together become a homogeneous mixture all in the same phase, for example salt water. Because the mixture is homogeneous, all samples will be identical.



Solutions are composed of two parts, the solute(s) and the solvent. The solvent is whichever substance has the larger quantity. A solute is a smaller quantity that is dissolved or dispersed equally throughout the solvent. Water is a common solvent, to the point where it is often called the universal solvent despite that being a misnomer. Electrolytes (acids, bases, and salts) are frequently solutes.

Concentration is the proportion of solute to solvent and can be expressed in many ways. The most common units are molarity, molality, mass percent, volume percent, and parts per quantity (ppm, ppb).[1] Each unit has a different frame of reference, but each is a ratio of solute to either the amount of solvent or the total amount of the solution. We select the unit we are using based on what is most appropriate to the problem we are trying to solve.

[1] There is also normality and molar fraction, but normality is beyond the scope of this course and molar fractions will be discussed with the Gas Laws and Dalton's Law of partial pressure. Normality is the idea of active solute per liter of solution. Not every gram of solute necessarily completely dissolves into the solution and so a unit that takes that into account can be very handy. Mole fraction is the idea that we can have a volume of gas and its composition will be formed by different numbers of moles of each component gas.

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6.1.1. Molarity

Chemists tend to favor molarity (M) as a unit of concentration and many chemical phenomena are mathematically described using this unit. We use molarity when the temperature is constant or otherwise not a factor that could affect the experiment. It is a very easy unit to use because the frame of reference is the volume of the solution. We get the mass of the solute, convert that to moles and then divide the number of moles of solute by the volume of the solution to get the molarity. Please note the spelling and the capitalization.

$$Molarity = M$$

$$M = \frac{moles\ of\ solute}{Volume\ of\ solution\ in\ L}$$

6.1.2. Molality

Molality is favored by chemical engineers and physical chemists because it is unaffected by temperature. Molality is particularly useful when working with colligative properties. Again, we are relating the number of moles of solute, but this time the frame of reference is the mass of the solvent in kilograms. This is very useful when the amount of solute appreciably changes the total volume of the solution. Please note that water has a density of ~1 g/mL at room temperature (25℃) which can make the value for a molar solution very similar to that of a molal solution.

$$molality = m$$

$$m = \frac{moles\ of\ solute}{mass\ of\ solvent\ in\ kg}$$

6.1.3. Percent by mass or volume

Percent concentration is frequently used by the US Food and Drug Administration (FDA), US Department of Agriculture (USDA), and Bureau of Alcohol Tabaco and Firearms (ATF). It is important to note whether the percent concentration is by mass, volume, or mass by volume. In each case, the percentage is relating solute to the total amount of the solution.

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Percent by mass is the mass of solute divided by the mass of the solution times 100%. We use it to talk about the concentration of things like bleach, peroxides, and small molecule catalysts.

Percent by volume is the volume of the solute divided by the volume of the solution times 100%. We use it to talk about the concentration of things that tend to typically be thought of as fluids, vinegar, alcohol, or gas-to-gas solutions.

Percent mass by mass

$$\%\frac{m}{m} = \frac{mass\ of\ solute}{mass\ of\ solution}100\% = \frac{mass\ of\ solute}{mass\ of\ (solute+solvent)}100\%$$

Percent volume by volume

$$\%\frac{v}{v} = \frac{volume\ of\ solute}{volume\ of\ solution}100\% = \frac{volume\ of\ solute}{volume\ of\ (solute+solvent)}100\%$$

Percent mass by volume

$$\%\frac{m}{v} = \frac{mass\ of\ solute}{volume\ of\ solution}100\% = \frac{mass\ of\ solute}{volume\ of\ (solute+solvent)}100\%$$

Percent mass by volume is typically used in health care for things like blood sugar and physiological saline. The percent mass by volume is the mass of the solute divided by the volume of solution times 100%.

6.1.4. Parts per million, parts per billion

Very dilute concentrations can be measured in parts per million (ppm) or parts per billion (ppb). The Environmental Protection Agency (EPA) and the World Health Organization (WHO) use ppm and ppb to measure trace amounts of various heavy metals. Trace amounts of toxic substances can build up in the body. Levels of exposure are important in any discussion of toxicity.

$$ppm = \frac{mass\ of\ solute}{mass\ of\ solution}10^6\ ppm = \frac{mass\ of\ solute}{mass\ of\ (solute+solvent)}10^6\ ppm$$

$$ppb = \frac{mass\ of\ solute}{mass\ of\ solution}10^9\ ppb = \frac{mass\ of\ solute}{mass\ of\ (solute+solvent)}10^9\ ppb$$

The parts per quantity is essential the same calculation as percentage mass by mass. The difference is that instead of multiplying the mass of solute divided by the mass of solution by 100%, we use the fact $100 = 10^2$ and instead multiplying that ratio by one million which is 10^6 or one billion which is 10^9.

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6.2. Stoichiometry in solutions

Stoichiometry in solutions is essentially the same as it is in any other situation. The chemical equation gives the proportions for moles of reactant to moles of products. The key is to know the number of moles and convert from moles of one to moles of the other via their molecular ratio in the equation.

6.2.1. How to use units of concentration in calculations.

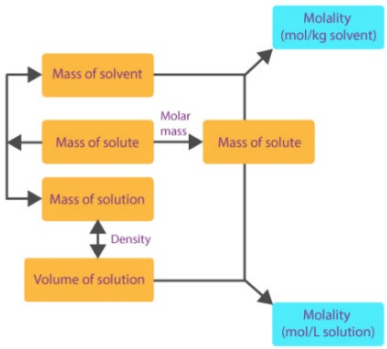
If 355 mL of 0.325 M $Pb(NO_3)_2$ is added to large amount K_3PO_4 solution, how many grams of $Pb_3(PO_4)_2$ precipitate will be formed?
The balanced equation is

$$3\ Pb(NO_3)_2(aq) + 2\ K_3PO_4(aq) \rightarrow Pb_3(PO_4)_2(s) + 6\ KNO_3(aq)$$

$$\frac{355\ \text{mL}\ Pb(NO_3)_2}{1} \times \frac{1\ L}{1000\ \text{mL}} \times \frac{0.325\ \text{mol}\ Pb(NO_3)_2}{1\ L} \times \frac{1\ \text{mol}\ Pb_3(PO_4)_2}{3\ \text{mol}\ Pb(NO_3)_2} \times \frac{811.5\ \text{g}\ Pb_3(PO_4)_2}{1\ \text{mol}\ Pb_3(PO_4)_2} = 31.2\ \text{g}\ Pb_3(PO_4)_2$$

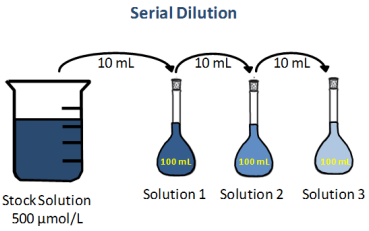
6.2.2. Conversions between different types of units of concentration.

Converting between different units of concentration sometimes needs some extra information. When converting from molarity to molality or vice versa, we need to know the density of the solvent.

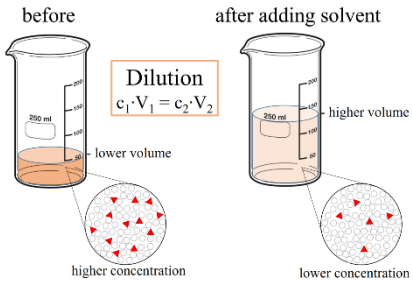


6.3. Serial dilutions

Stock solutions are generally high concentration solutions or solid solutes that are then made into solutions. High concentration solutions are rarely used if not necessary because it can be a serious safety hazard. Instead, we use more dilute solutions because it is safer and because we get the same results with less materials.



To get the more dilute concentration we want, we dilute the solution by taking a small amount of the stock solution and place it into a flask and add solvent to a set volume. This works because we know how much we took



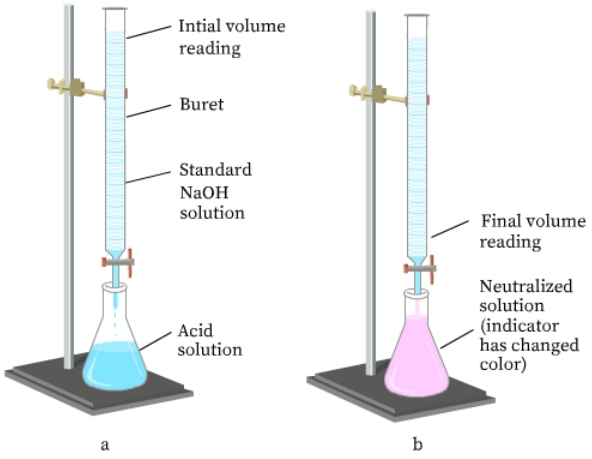
from the stock solution and what the concentration of that solution, which means we know how many moles of solute we have. From there we know how many moles of solute we have, and we are simply adding to the volume of the solution.

6.4. Measurements of concentration

There are multiple methods available for determining a solution's concentration. In fact, there are so many methods that covering all of them could be an entire college course.

6.4.1. Titration

Titration is perhaps the simplest and most well-developed method of determining the concentration of a solution. All titrations work with a set of well defined "knowns" to determine the "unknown" concentration. Titration has a reagent with a known concentration in the buret, called the titrant, which allows for the volume of the titrant to be measured. The analyte, the solution the titrant is being added to, has an unknown concentration, but a known initial volume. Since we know the concentration and the volume of the titrant, we know how many moles of the titrant we have dispensed into the analyte. Molar concentration is in mol/L and we know L, so that just leaves us with mol.



We need some sort of indicator for when we have added enough of the titrant to the analyte. There are color changing chemicals, pH meters, electrical sensors, and more. In this lab, we are doing an acid-base titration with an indicator that changes color.

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When we consider the chemical equation for an acid-base neutralization, we will always see that there is distinct ratio of acid to base. When the analyte is an acid and the titrant a base, that analyte will become less acidic as the basic titrant is added. There is a point at which the acid and the base, the analyte and the titrant, will have equivalency. This is called the endpoint. Past the endpoint, the analyte will wind up with more base than acid. We are looking for the endpoint and that is exactly when the color change indicator changes color. The initial color change is subtle and very faint.

6.4.2. Our Titration Experiment

In this experiment, you will determine the concentration of acetic acid in white vinegar. By law, white vinegar can only contain 5% acetic acid in water. To verify this number, it is easier to determine the Molarity of acetic acid in vinegar by reacting with a solution of NaOH at a known concentration. The reaction of acetic acid and sodium hydroxide is given as Reaction CC.1.

$$HC_2H_3O_2(aq) + NaOH(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(l) \qquad \text{(Reaction CC.1)}$$

This reaction requires the use of an indicator. An indicator is a chemical that marks the end point of any reaction. There are different indicators based on the type of reaction you are performing. For this lab, you will use phenolphthalein as your indicator. When all the sodium hydroxide has reacted with the acetic acid, phenolphthalein will turn the solution from colorless to pink. This happens instantly in the presence of excess sodium hydroxide. You will want to stop the addition of NaOH at the first sign of pink color in the entire solution. At this point, the sodium hydroxide and acetic acid are both the limiting reactants, so you can find the concentration of acetic acid from the number of moles of sodium hydroxide.

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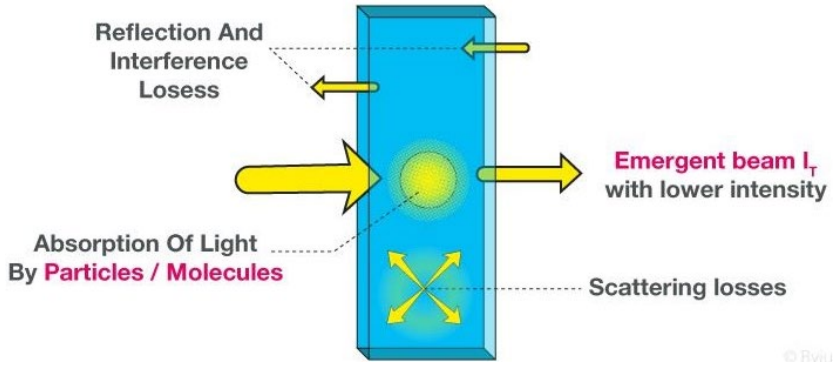
6.4.3. Point-slope method for determining the concentration of a solution.

The point-slope method is a form of colorimetry. Using a colorimeter, we see how well the light at a specific wavelength transmits through a sample. Clear as mud? Just as thick mud makes it hard to shine a light through, and extremely dilute it is easy to see a light through, this is our transmittance. This is what the colorimeter is measuring, but it is doing so with very dilute samples where the transmittance is very high.

Since certain molecules absorb light at specific wavelengths, we can determine the concentration of a solution based on how much light is absorbed. The means we use to find the concentration of our solute is the Beer-Lambert law. We can plot a graph showing the absorption with respect to the concentration for known concentrations. We call this plot of known concentration samples the calibration curve. We can then take samples with unknown concentrations and see where they fall along that line which tells us the concentration of those samples.

$$A = \varepsilon b C$$

A = absorbance
ε = molar absorptivity
b = length of light path
c = concentration



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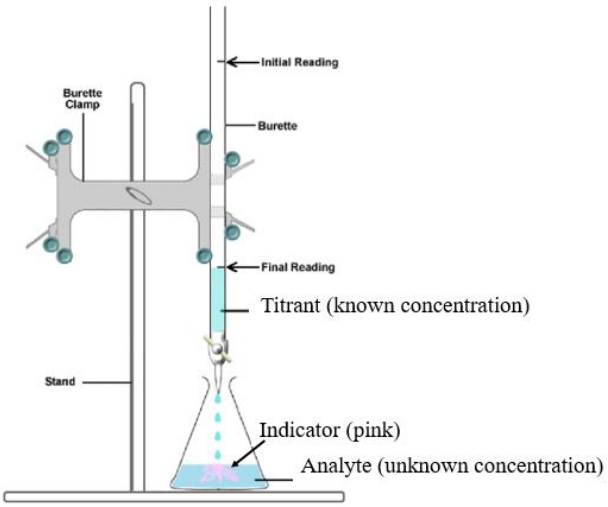
7. Procedure

7.1. Finding Concentration from titrating NaOH into commercial vinegar

7.1.1. Set up your buret and stand as shown.

- Raise the buret sufficiently for the stir plate and Erlenmeyer flask to fit below the buret.



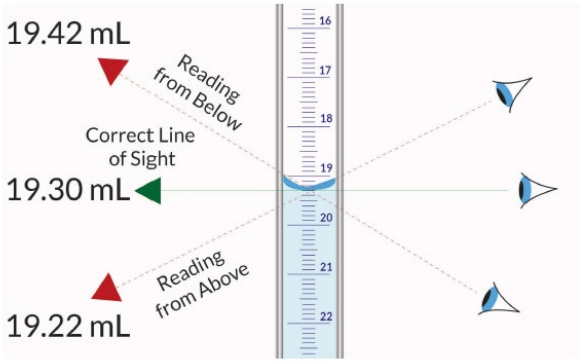
7.1.2. Record the concentration of the NaOH solution.

7.1.3. Wash the buret with NaOH by running a small volume of the NaOH solution along the sides and down the stopcock to a waste container.

7.1.4. Fill the buret with NaOH and remove the air pocket below the stopcock by running a small volume of solution all the way through to a waste container.

7.1.5. Record the initial volume of NaOH in the buret.

- This should be near 50mL worth of volume in the buret.
- Be mindful of the line of sight when reading the buret.



7.1.6. Record the mass of an empty 125 mL Erlenmeyer flask.

7.1.7. Place 5 mL of white vinegar in your 125 mL Erlenmeyer flask.

7.1.8. Record the mass of the vinegar and the Erlenmeyer flask.

7.1.9. Place a magnetic stir bar in the flask.

7.1.10. Add enough DI water to cover the stir bar completely.

7.1.11. Add 3 drops of phenolphthalein to the flask.

7.1.12. Place the flask on a stir plate under the buret.

7.1.13. Place the tip of the buret just inside the mouth of the flask.

- This is to prevent spillage.

7.1.14. While stirring, add the NaOH dropwise at a rate of about 2-3 drops per second.

7.1.15. When the solution just starts to turn pink, slow the drop rate to about 1 drop per second.

- This is very faint, and the pink tinge may disappear.
- Consider using a white piece of paper behind the Erlenmeyer flask to make the pink color more visible.

7.1.16. Stop the buret when the solution turns faintly pink and stays pink.

- This is the equivalence point or endpoint of the titration.
- If the solution returns to clear after 5 seconds, add one more drop of NaOH.

7.1.17. When the solution stays pink for more than 10 seconds, record the final volume of NaOH added to the solution.

7.1.18. Discard the vinegar solution and clean the flask with water.

- The discarded material may be rinsed down the sink.

7.1.19. Repeat Steps 7.1.4-18 twice.

7.1.20. After you have completed all three trials, calculate the mass percent of acetic acid in your vinegar using the mass of vinegar that you recorded in Step 8.

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7.2. **Finding concentration from Beer's Law and a calibration curve.**

7.2.1. Pay attention to your GTA's discussion of colorimetry.

7.2.2. Record the identity of your sample on the datasheet.

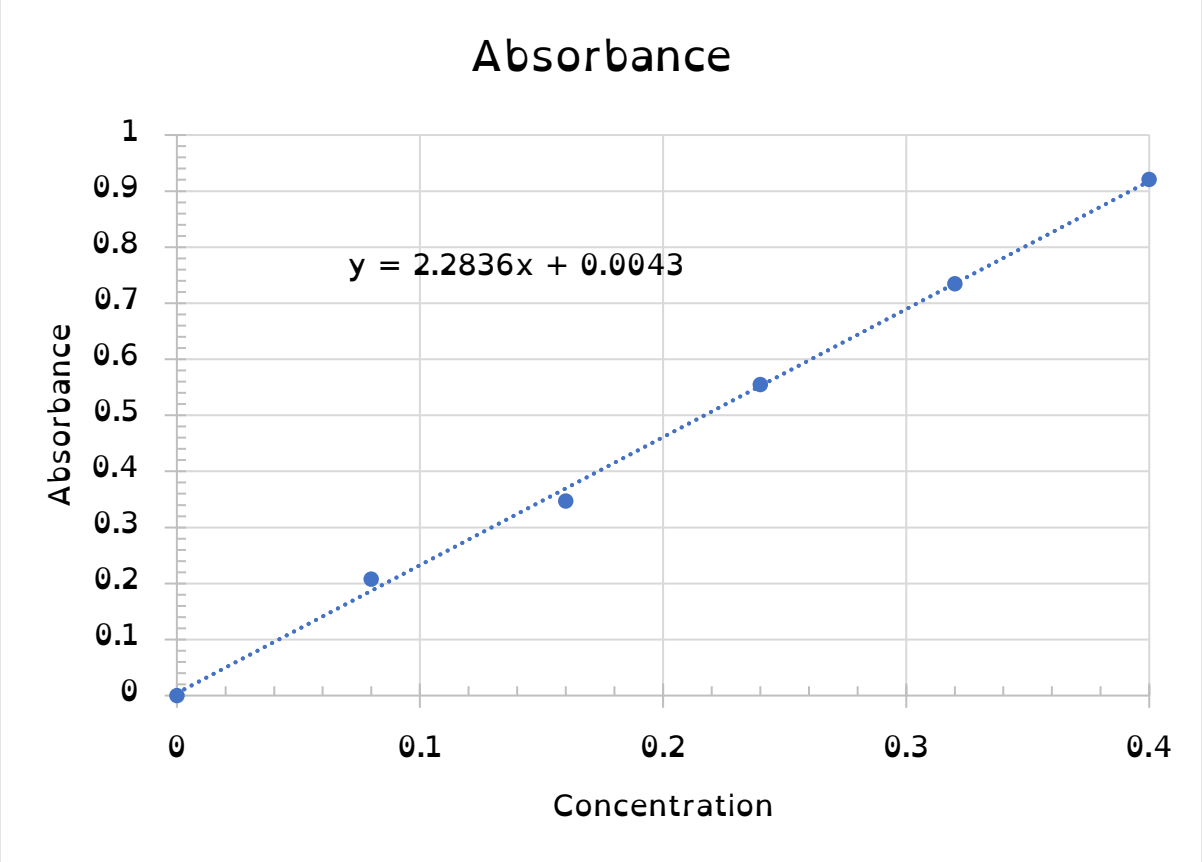
7.2.3. Follow directions to find the absorbance of your sample at the colorimetry station.

7.2.4. Record the absorbance.

7.2.5. Using the calibration curve provided in lab, find the concentration of your sample.

- The one below may not be accurate for your lab but is provided as an example.

7.2.6. Record the concentration for your sample.



Lab #9 – Concentration

8. Data Sheet

8.1. Finding Concentration from titrating NaOH into commercial vinegar

8.1.1. Record concentration of NaOH(aq)

8.1.2. Record titration data with units and significant figures.

	Trial 1	Trial 2	Trial 3 (optional)
Volume of vinegar			
Mass of flask			
Mass of flask + vinegar			
Mass of vinegar solution			
Initial buret volume			
Final buret volume			
Solution color			
NaOH volume			

Show your calculations with significant figures and units below.

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8.2. Finding concentration from Beer's Law and a calibration curve.

8.2.1. Record Sample Identification. _____

8.2.2. Record Absorbance. _____

8.2.3. Find and record sample concentration based on the calibration curve. _____

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9. Post-Lab Questions

9.1. What is the molarity of your vinegar (acetic acid)?

Show your work.

$$\frac{\text{ml NaOH}}{} \left|\frac{\text{1 L}}{\text{1000ml}}\right| \frac{\text{mol NaOH}}{\text{L NaOH}} \left|\frac{\text{1 mol acetic acid}}{\text{1 mol NaOH}}\right| \frac{}{\text{mL acetic acid}} \left|\frac{\text{1000mL}}{\text{1 L}}\right| = \text{M acetic acid}$$

9.2. What is the molality of your vinegar (acetic acid)? Assume the density of the vinegar is 1.02 g/mL.

Show your work.

$$\frac{\text{mol acetic acid}}{\text{L}} \left|\frac{\text{1 L}}{\text{1000ml}}\right| \frac{\text{mL acetic acid}}{} \left|\frac{\text{1.02 g acetic acid}}{\text{1 mL acetic acid}}\right| = \text{g acetic acid}$$

Mass of solution – mass of acetic acid = _____g H_2O

$$\frac{\text{mol acetic acid}}{\text{g water}} \left|\frac{\text{1000 g}}{\text{1 kg}}\right| = \text{m acetic acid}$$

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9.3. Find the percent mass by mass of the vinegar from your lab experiment.

9.3.1. Convert moles of acetic acid to grams using the molar mass (60.1 g/ mol).

9.3.2. Determine the total mass in grams from the vinegar volume in your sample in mL using the density 1.02 g/mL for vinegar.

9.3.3. Calculate the mass percent of acetic acid in vinegar from the mass in grams of the CH_3OOH (acetic acid) and the mass of the vinegar.

9.3.4. Does your value for the percent by mass concentration of acetic acid in vinegar match the 5% stated on the label? If not, why not?

9.4. Using the calibration curve, does your sample concentration seem reasonable?

9.5. Thought experiment: Considering that the colorimeter is only measuring the absorption of light at one specific wavelength, would it be possible to measure the concentration of a solute in a sample that has multiple chemicals in it? If so, what would you need to be careful of to make sure your measurement is good? If not, why would the other materials besides the solute of interest be a problem?

10. Summary and Conclusions

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GTA:_____ Date: _____

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