

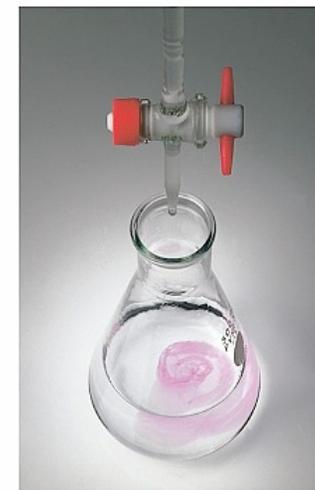
# Exp 07: Acetic Acid Titration



Your job is to measure the concentration of an unknown solution of acetic acid (HOAc). Another name for aqueous solutions of acetic acid is vinegar.

Report the molarity you determine.

You will use double displacement kinetics and your understanding of neutralization reactions to justify your conclusions.

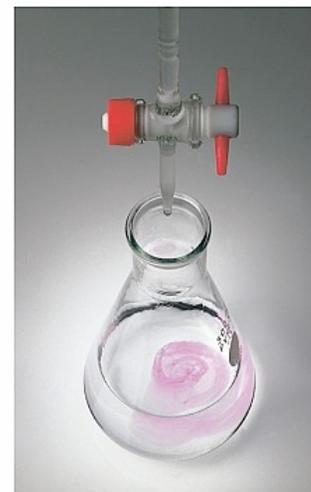


Process: Add exactly enough NaOH of known molarity to the flask to neutralize the AcOH and use a balanced equation to calculate the moles in your original volume of AcOH. Knowing moles and volume, you know molarity.



# Exp 07: Acetic Acid Titration

- ▶ **Titration** is an analytic technique for determining the concentration in one solution by carefully adding a measured quantity of a known solution and observing a clear end point.
- ▶ The unknown is called an **analyte**.
- ▶ The standard solution is called a **titrant** or titrator.
- ▶ The **end point** is the point in the experiment where an indicator suggests the quantities of analyte and titrant are equal.
- ▶ The **equivalence point** is the point where they actually are.
  - ▶ With a good chemical indicator, the two should be close, but your equivalence point is almost always reached before you see the end point.
- ▶ An **indicator** is a chemical added to the mixture that changes color close to the equivalence point.
- ▶ Finding the end point with a chemical indicator requires some skill.



# Exp 07: Acetic Acid Titration

## Procedure:

For this experiment, two students will be assigned to each titration apparatus. Each student will be assigned about 50 mL of vinegar solution of unknown concentration.

## Setup:

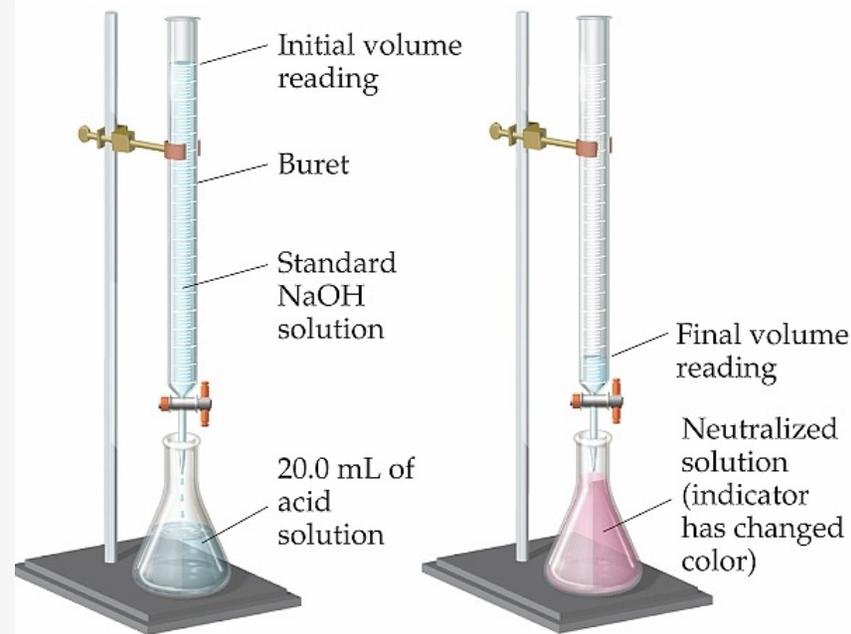
1. Collect about 80 mL each of the standardized sodium hydroxide solution and unknown acetic acid solution (vinegar) in each of your two beakers.
2. Collect the titration apparatus. Thoroughly clean your buret and then rinse it out with about 5 mL of deionized water and then 5 mL of sodium hydroxide solution. Attach the buret clamp to the ring stand and install the buret in the buret clamp.
3. Clean and rinse the pipet with about 5 mL of deionized water and then 5 mL of the acetic acid solution.

## Trial Process (repeat for 3-4 trials):

1. Fill or refill the buret with the standardized sodium hydroxide solution to within 3 mL of its capacity. Record the initial buret reading to two decimal places ( $\pm 0.01$  mL).
2. Clean and rinse the 125 mL Erlenmeyer flask with deionized water. Dry the outside of the Erlenmeyer flask and weigh it.
3. Weigh and record the flask weight to within  $\pm 0.01$  g. It is not necessary for the inside of the flask to be dry. You must weigh the flask again with every trial as droplets of deionized water will change the weight.
4. Pipet roughly 10.0 mL of the acetic acid solution into the Erlenmeyer flask from your beaker of acetic acid solution. Record the volume of acetic acid to within 0.01 mL.
5. Re-weigh the flask with acetic acid solution. Record the combined weight of the flask and acetic acid solution. Subtract the weight of the empty flask from this number to find the weight of the acetic acid solution.
6. Add 1 drop of phenolphthalein indicator to the acetic acid solution.
7. Add roughly 20 mL of deionized water to the acetic acid solution.
8. Place a sheet of white paper underneath the Erlenmeyer flask (so you can more easily spot the persistent pink tinge of solution that will indicate the titration is complete).
9. Titrate the vinegar solution to the endpoint using the standardized sodium hydroxide solution. Record the buret reading at the endpoint, to two decimal places ( $\pm 0.01$  mL).

After the first titration, dump the contents of the beaker down the sink, rinse the beaker well with tap water, shake it dry, and carefully dry off the outside of the flask with a paper towel. The flask does not have to be absolutely dry, but it should be dry on the outside. You will need to weigh it before each trial, because it will contain a slightly different amount of water each time. Weigh the flask again, and continue with the next trial. You should conduct at least three trials on the vinegar solution. Your objective is to generate at least three results that do not differ by more than 1.5%. Use the formula below to calculate the % difference between runs:

$$\% \text{ difference} = [ \text{Molarity}_{\text{high}} - \text{Molarity}_{\text{low}} \times 100 ] / \text{Molarity}_{\text{average}}$$



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## Interpreting Data (Calculations & Reasoning):

You will need to complete the following calculations to complete your data table. In your calculations section show your work for each calculation below for one example trial.

C1: Calculate the moles of NaOH used in each trial. To do this multiply the volume of sodium hydroxide solution you added with your buret by the molarity of that solution. Be careful to convert your volume from mL to L before multiplying.

C2: Calculate the moles of AcOH in the flask, for each trial. Using the balanced equation above, determine the mole ratio between NaOH and AcOH. Use the mole ratio to determine the moles of AcOH neutralized by the sodium hydroxide you added.

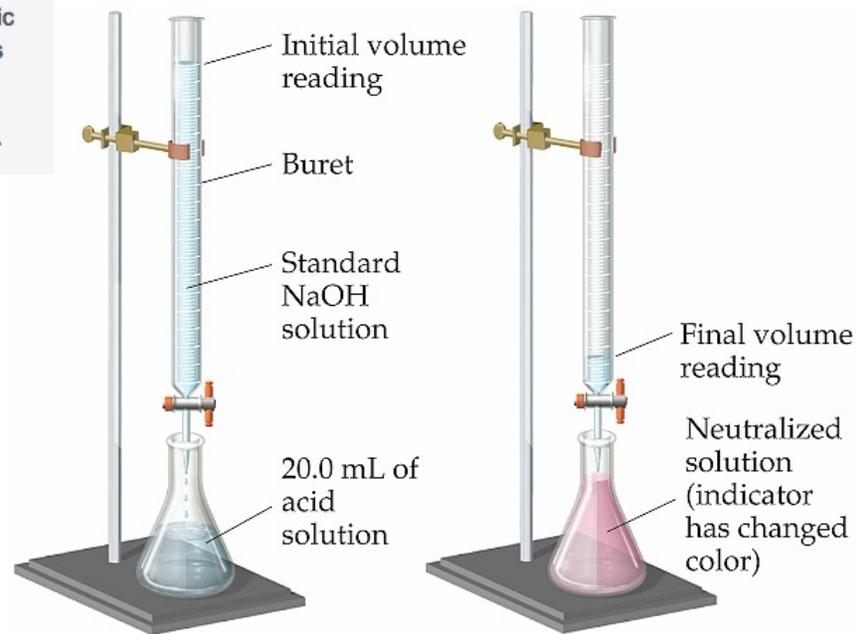
C3: Calculate the molarity of the AcOH, for each trial. You determined the moles of AcOH in the acetic acid sample you added to your Erlenmeyer flask. You carefully measured the volume. Divide the number of moles by the volume (in Liters) to determine the molarity.

C4: Calculate the average molarity and the percent difference for your trials.

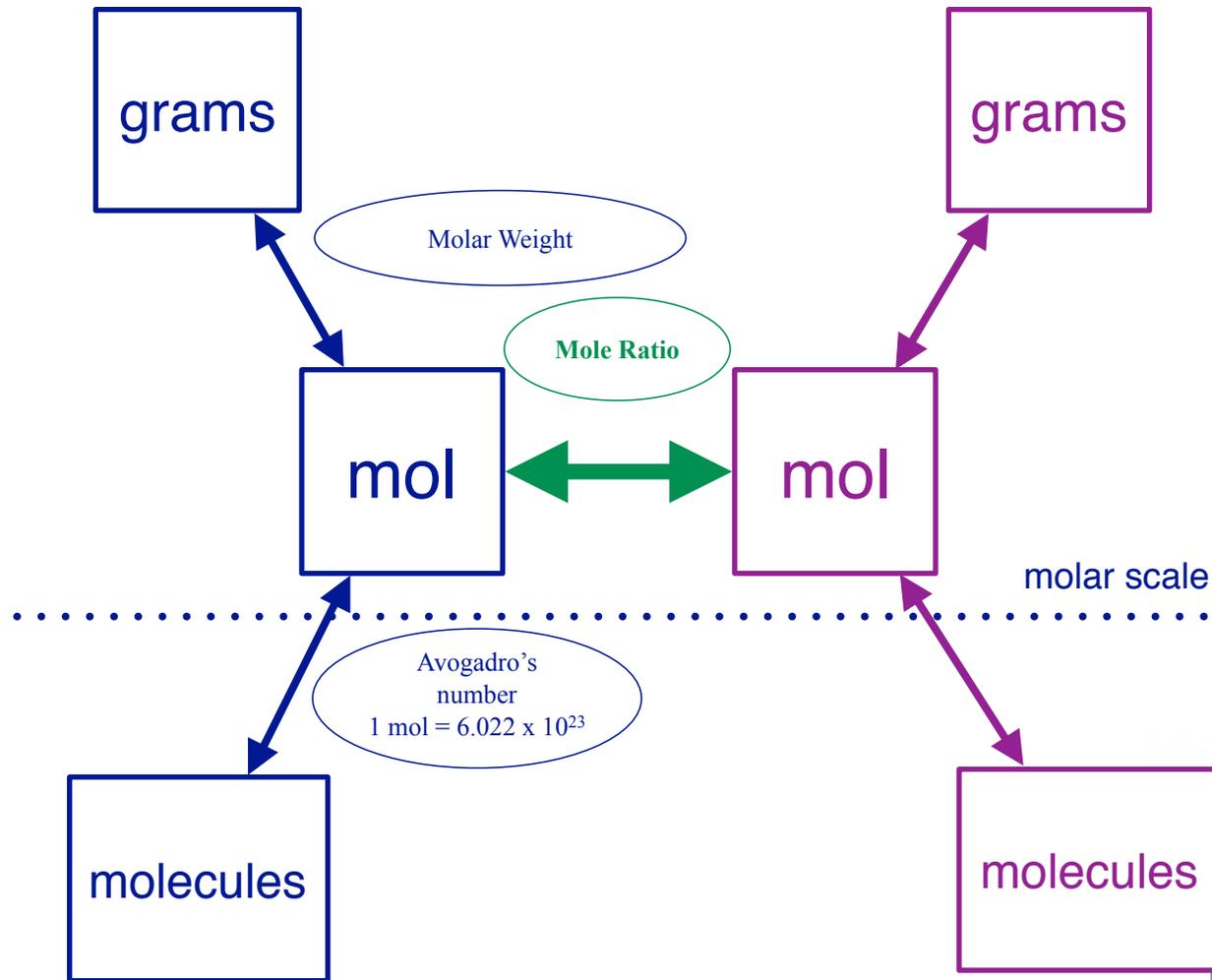
C5: Calculate the mass of acetic acid in the sample of acetic acid solution (vinegar), for each trial. Use the moles of acetic acid calculated in C2 and the molar mass of acetic acid to calculate the mass of acetic acid in your sample.

C6: Calculate the mass percent of acetic acid in the acetic acid sample, for each trial. Divide the mass of acetic acid you calculated in C5 by the total mass of the acetic acid solution sample. The total mass of the solution is just the mass of the flask with vinegar less the mass of the empty flask.

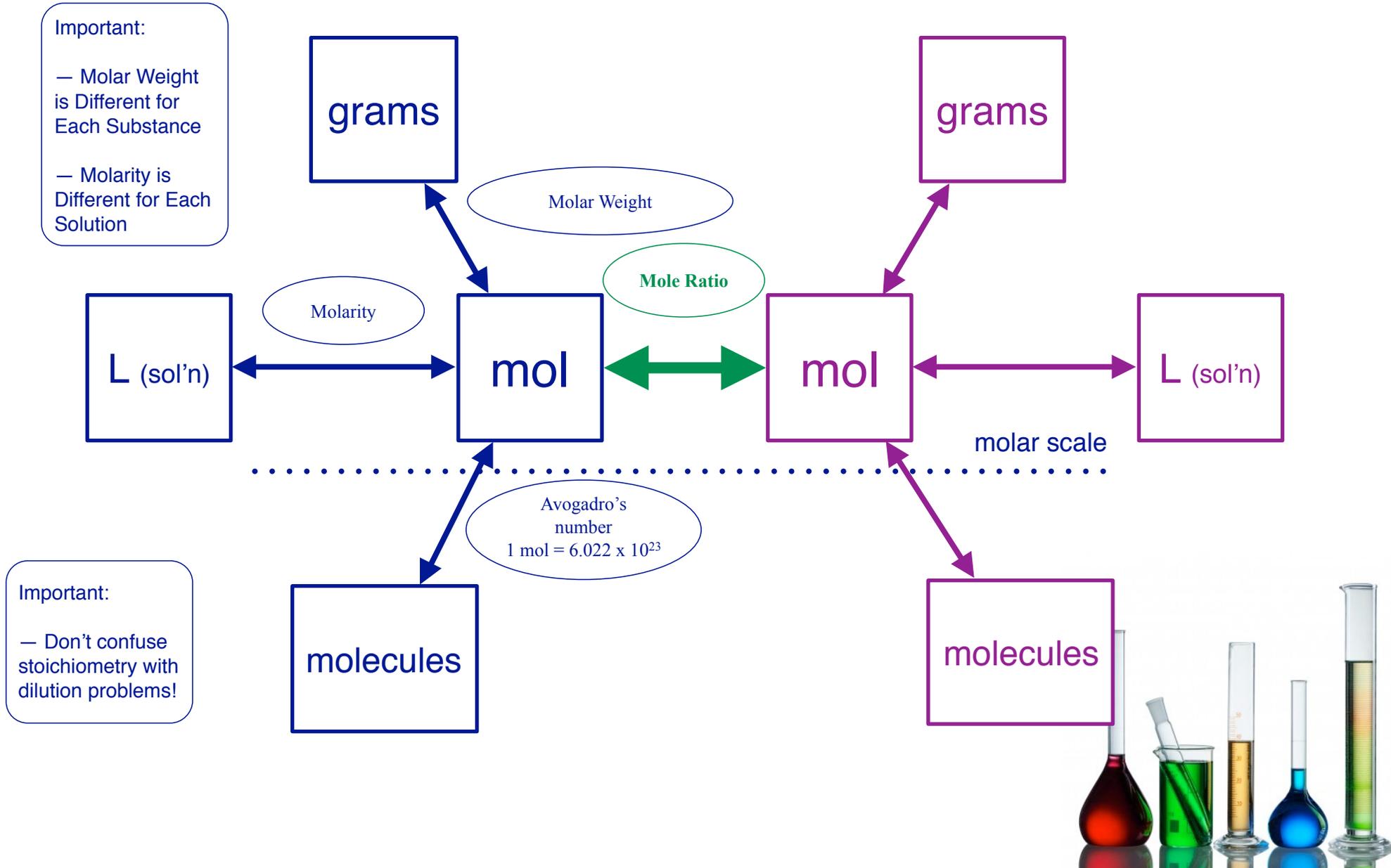
C7: Calculate the average of the mass percents and also the percent different of mass percents for your trials.



# Stoichiometry

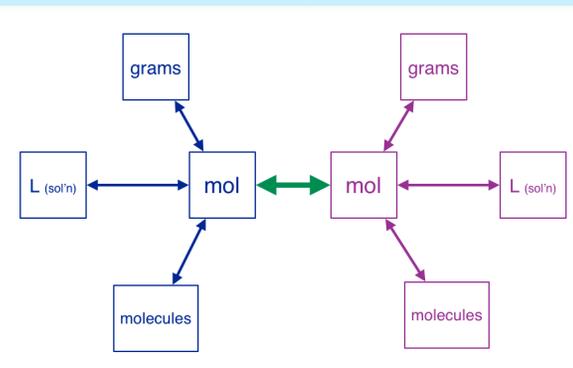


# Stoichiometry

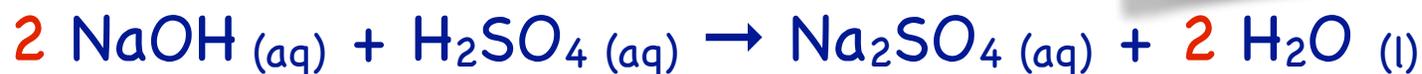


## Problem:

A 20.0 mL sample of  $\text{NaOH}_{(aq)}$  is titrated to an end point with 45.7 mL of 0.500 M  $\text{H}_2\text{SO}_4_{(aq)}$ , what is concentration of the  $\text{NaOH}$  solution?



Solution



$$\underbrace{\text{mL} \xrightarrow{\textcircled{1}} \text{L} \xrightarrow{\textcircled{2}} \text{mol}}_{\text{H}_2\text{SO}_4} \xrightarrow{\textcircled{3}} \text{mol NaOH} \quad ; \quad \frac{\text{mol}}{\text{vol}} = \text{molarity}$$

①  $1000 \text{ mL} = 1 \text{ L}$

②  $0.500 \text{ mol} = 1 \text{ L}$

③  $2 \text{NaOH} = 1 \text{H}_2\text{SO}_4$

Part A

$$45.7 \text{ mL} \cdot \frac{1 \text{ L}}{1000 \text{ mL}} \cdot \frac{0.500 \text{ mol}}{1 \text{ L}} \cdot \frac{2 \text{ NaOH}}{1 \text{ H}_2\text{SO}_4} = 4.57 \times 10^{-2} \text{ mol}$$

Part B

$$20.0 \text{ mL} = 0.0200 \text{ L}$$

$$\frac{4.57 \times 10^{-2} \text{ mol}}{0.0200 \text{ L}} = \boxed{2.29 \text{ M}}$$

# Questions?

