Lab 10.2
Detailed Explanation and Calculations

I performed the experiments twice, using different bottles of ammonia.
I found that one was a weaker base. The first clue was in the amount of vinegar needed to neutralize the base and push the solution to indicating the acid in it.

Use the information below to write up your data and pull your conclusion from.

Objective: Using a know concentration of an acid (acetic acid in vinegar) to determine through experimentation and calculations the concentration of a base (ammonia).

Molarity is concentration; so all calculations are aimed at being able to use the formula for Molarity to determine the concentration of the base.

This is where knowing your formulas or having your reference cards out to look over help you to know what to do first. We have to get the acid and the base into moles in order to use the Molarity formula, which is:

Concentration (M) = $\frac{\text{\# moles}}{\text{\# liters}}$

	acetic acid		ammonia				
Reaction:	$C_2H_4O_2$	+	NH_3	$\rightarrow$	NH_4^+	+	$C_2H_3O_2^-$
	acid		base		ion		ion

	1st Trial	**2nd Trial**
Mass of Graduated Cylinder:	11.2 g cylinder	11.2 g cylinder
Mass of 50.0 mL Vinegar + Cylinder:	60.9 g vinegar&cylinder	61.10 g vinegar&cylinder

Determine Mass of Vinegar: **(**mass of vinegar + cylinder**)** **-** **(**mass of cylinder**)**

	1st Trial	2nd Trial
	60.9 g	61.10 g
-	11.2 g	11.2 g
	49.7 g vinegar	49.8 g vinegar

(Step #13) Density = Mass / Volume
Determine the Density of Vinegar:

1st Trial	2nd Trial
$\frac{49.7 \text{ g}}{50.0 \text{ mL}}$ = 0.994 g/mL vinegar	$\frac{49.8 \text{ g}}{50.0 \text{ mL}}$ = 0.996 g/mL vinegar

Your data should now list the two volumes of vinegar used in the titration. The first time was so you would have a general idea of how much vinegar was needed to neutralize the base and indicate that acid was now present. It is your rough titration. You then repeated the procedure and stopped adding vinegar quickly when you were within 5 mL of the rough titration; however, if the original amount is small, then you stop within 2-3 mL, as is the case of the 1st trial detailed here. The key to both trials is stirring as vinegar was added to properly see when the indicator in the solution showed the changing from a base to an acid solution.

	1st Trial	**2nd Trial**
Rough Titration – volume of vinegar added:	5 mL vinegar	13 mL vinegar
Precise Titration – volume of vinegar added:	4.5 mL vinegar	12.2 mL vinegar

For remaining calculations the precise titration is the measurement to be used.

		1st Trial		2nd Trial
Convert mL of vinegar added to grams:		4.5 mL		12.2 mL
(volume of vinegar) x (density)	x	0.994 g/mL	x	0.996 g/mL
This converts mL to g, since the mL cancel.		4.473 g		12.151 g
Apply significant figure -		4.5 g vinegar		12.2 g vinegar

The density is basically 1 g/mL, so the end result after applying significant figure rules is that the grams of vinegar equals the mL of vinegar. They are two different units and represent two different things even though they end up being the same number, so you have to follow each step and not take shortcuts that could give you incorrect results in the next experiment you do.

Determine mass (g) of acid added:
Remember, vinegar is a solution of an acid in water. We need to know the amount of acid that is in the vinegar we used, so we go to the concentration (% of acid), which is 5 %.
Read after your lab on pg 344, the text states to use 0.0500 for the 5%. (5% = 5.00 / 100.0 - The precision of the concentration is represented by using 0.0500 and NOT 0.05. Remember the last two zeros are significant and tell us the precision of this measurement.)

	acetic acid		ammonia				
Reaction:	$C_2H_4O_2$	+	NH_3	→	NH_4^+	+	$C_2H_3O_2^-$
	acid		base		ion		ion

Mass of acid ($C_2H_4O_2$) used = (Mass of vinegar) x (concentration of vinegar)

	1st Trial		2nd Trial
	4.5 g		12.2 g
x	0.0500	x	0.0500
	0.20 g $C_2H_4O_2$		0.610 g $C_2H_4O_2$

Determine mass (g) of one molecule of $C_2H_4O_2$, so you have a way to convert to moles:
As learned previously, using the formula and the periodic chart, add up each amu that applies for the elements making up the formula. The result is 60.0 amu, which equals 60.0 g $C_2H_4O_2$.

Determine the moles of $C_2H_4O_2$ used:

Rule-	grams of a molecule based on its formula	=	1 mole of molecule
	60.0 g $C_2H_4O_2$	=	1 mole $C_2H_4O_2$

This will allow you to convert the grams of $C_2H_4O_2$ used to moles used in the reaction.

1st Trial

$$\frac{0.20\ \cancel{g\ C_2H_4O_2}}{1} \times \frac{1\ \text{mole}\ C_2H_4O_2}{60.0\ \cancel{g\ C^2H^4O^2}} = 0.0033\ \text{moles}\ C_2H_4O_2$$

2nd Trial

$$\frac{0.610\ \cancel{C_2H_4O_2}}{1} \times \frac{1\ \text{mole}\ C_2H_4O_2}{60.0\ \cancel{g\ C^2H^4O^2}} = 0.010\ \text{moles}\ C_2H_4O_2$$

Determine the moles of NH_3 based on the reaction (1 mole of $C_2H_4O_2$ reacts with 1 mole of NH_3):

1st Trial

$$\frac{0.0033\ \text{mo}\cancel{\text{les}\ C_2H_4}O_2}{1} \times \frac{1\ \text{mole}\ NH_3}{1\ \cancel{\text{mole}\ C_2H_4}O_2} = \underline{0.0033\ \text{moles}\ NH_3}$$

2nd Trial

$$\frac{0.010\ \text{mo}\cancel{\text{les}\ C_2H_4}O_2}{1} \times \frac{1\ \text{mole}\ NH_3}{1\ \cancel{\text{mole}\ C_2H_4O_2}} = \underline{0.010\ \text{moles}\ NH_3}$$

Determine the concentration of ammonia:

To determine the concentration, we need to remember we used 10.0 mL of ammonia NH_3 when we diluted it with the 90.0 mL of water. This is what we poured the indicator into, and then titrated the vinegar into.

$$\text{Concentration (M)} = \frac{\text{\# moles (based on what was used in reaction that we just determined)}}{\text{\# liters (based on volume used, which was 10.0 mL of ammonia)}}$$

The formula in is liters and the volume we used was mL, so you must convert the 10.0 mL.

$$10.0\ \text{mL}\ NH_3 \times \frac{0.001\ \text{L}}{1\ \text{mL}} = 0.010\ \text{L}\ NH_3$$

Note: The conversion formula is exact, so it is not considered when determining significant figures.

Plug your knowns into the concentration formula:

1st Trial

$$\text{Concentration (M)} = \frac{\text{\# moles } NH_3}{\text{\# liters } NH_3} = \frac{0.0033\ \text{moles}}{0.010\ \text{L}} = 0.30\ \text{M}$$

2nd Trial

$$\text{Concentration (M)} = \frac{\text{\# moles } NH_3}{\text{\# liters } NH_3} = \frac{0.010\ \text{moles}}{0.010\ \text{L}} = 1.0\ \text{M}$$

You can see that the strength of the ammonia from the first trial (0.3 M) is a lot less potent (strong) than the ammonia from the second trial (1.0 M).

In our MicroChem chemicals, each bottle is labeled with its Molarity. It tells us the strength of that chemical, which in turn will tell us how potentially dangerous it is, such as if it is a strong acid or base. It can also indicate if you have the potential for a strong reaction that could be dangerous or be so weak that a reaction will not be noticeable.

Think about it. When I tell you a base has a 0.30 M and to compare it to one that has a 1.0 M that it is easy to quickly assess their strengths. It is also easy to see that we can affect the concentration by increasing the number of moles, which will increase the Molarity or by increasing the liters, which will decrease the Molarity.

Titration is used extensively in chemistry and it is important you understand the concept of how it is done and how to apply the stoichiometry to the results to determine Molarity.

PRINT THIS OUT AND FILE IT IN YOUR LAB SECTION OF YOUR NOTEBOOK. WHEN YOU GET LAB 10.2 BACK AFTER IT IS GRADED, FILE IT WITH THIS SUPPLEMENT.