LAB:
Titration of Vinegar/Acid Rain
[25 pts]

Name ______________________________
Lab Partner(s) ______________________
Period ______

Purpose: To determine the molarity and percent acetic acid in ordinary vinegar by titration.
Introduction: If acetic acid is added to NaOH with an exact mole ratio of 1:1, then the acetic acid will be perfectly neutralized as shown below. (Notice that only a salt and water are formed.)

$$\underset{\text{acetic acid}}{\mathbf{HC_2H_3O_2\ (aq)}} + \underset{\text{sodium hydroxide}}{\mathbf{NaOH\ (aq)}} \rightarrow \underset{\text{sodium acetate}}{\mathbf{NaC_2H_3O_2\ (aq)}} + \underset{\text{water}}{\mathbf{HOH\ (\mathit{l})}}$$

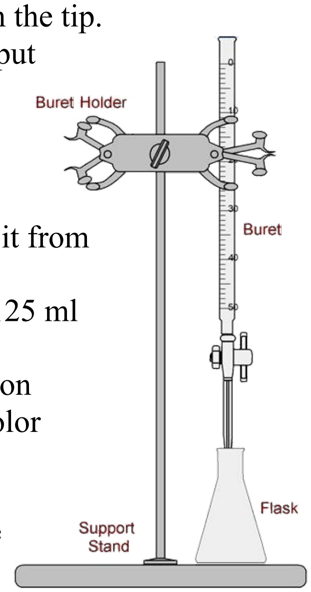
We will use the pH indicator, phenolphthalein, to tell us when the acid has been completely neutralized. Phenolphthalein is **clear when acidic** and it is **pink when basic**, so the solution will stay pink when there is no more acid and a slight excess NaOH is added at the equivalence point. If this is done carefully, the solution will turn *very light* pink with the addition of just one drop of NaOH past the *endpoint*.

Materials

50-mL buret	buret funnel	ring stand
buret clamp	125 mL Erlenmeyer flask	10 mL volumetric pipet
pipet filler	100 mL beaker	50 mL Erlenmeyer flask
50-55 mL standardized NaOH	30 mL vinegar	dropper bottle of phenolphthalein

Procedure:

1) Put about **50-55 ml of the standardized NaOH** solution into a clean, dry 100 ml beaker. (Please conserve the NaOH; we have a limited supply.) Record the molarity of NaOH below.
2) Rinse the buret with ~3-5 mL of the NaOH solution. Discard this solution through the tip.



3) Fill the buret all the way up with the NaOH. ***Slowly*** open the valve to fill the tip (put your beaker of NaOH under it), get any air bubbles out of the tip and refill the buret to about 0 mL. (Make sure you read the buret correctly—it reads downwards—**do NOT subtract from 50!** Read to the tenths place—1 decimal point!) Record initial volume in data table on the top of the next page.
4) Obtain ~30 mL of vinegar in a clean, dry 50-mL Erlenmeyer flask (to distinguish it from the NaOH).
5) Use volumetric pipet to measure **10.0 ml of the acetic acid** solution into a clean 125 ml flask.
6) **Add 2 drops of phenolphthalein** indicator to the vinegar in the flask. (The solution should be clear.) Place a sheet of white paper beneath the flask to better see the color change during the titration.
7) Start adding NaOH to the flask, about 1 mL at a time at first. Swirl the flask after every addition until the pink color disappears. Use shorter length squirts when the pink color persists longer. Try to get it so that one quick squirt suddenly causes the color to stay pink. (Swirl for about 30 seconds to make sure it is permanently pink.) The fainter the pink color, the better.
8) Record final volume level of NaOH in buret at the endpoint. Determine the volume of NaOH added, in mL: $V_{NaOH} = V_{Final} - V_{Initial}$.
9) Pour the products down the sink and rinse your flask with deionized water and do at least two trials. Repeat steps 5-8, starting with a fresh sample of vinegar. Add new indicator too. If you have less NaOH than you needed for the first trial remaining in the buret, refill it so that you do not run out during the titration. Record the current volume level as $V_{Initial}$ for the new trial.
10) You MUST do a third trial if your values for the volume of NaOH are not within 0.4 mL. (I will deduct points if this is not done)
11) When done, drain any leftover NaOH solution into the beaker and pour it back into the original container. Rinse the buret well with distilled water, leave the valve open, and put it back where you got it. Rinse all other glassware and put it away.

Data: [4 pts] **Molarity of standardized NaOH (from board): ______ M**

	Trial 1	**Trial 2**	**Trial 3 (if necessary)**
Volume of Vinegar (mL)			
Initial volume of NaOH (mL)			
Final volume of NaOH (mL)			
Volume of NaOH added (mL)			

Average Volume of NaOH added (average your ***closest two*** trials): ________ mL

Calculations: Show all work! Keep THREE Sig Figs throughout your calculations!!! UNITS on ***every*** number!

1) Follow these steps to determine the **molarity of the acetic acid in vinegar**.
   a. [1 pt] Calculate the **moles of NaOH** added to the acetic acid solution from the average volume and the molarity.

   b. [1 pt] How many **moles of $HC_2H_3O_2$** must have been in the vinegar you used? How do you know this? *[Explain using stoichiometry or by referring to the balanced chemical equation.]*

   c. [1 pt] Calculate the **molarity of $HC_2H_3O_2$** in vinegar. *[Remember to convert volume to L!]*

2) Follow these steps to determine the **percent by mass of acetic acid in vinegar.**
   a. [1 pt] Find **mass of acetic acid** in your vinegar sample. *(moles of acetic acid was calculated above)*

   b. [1 pt] Find the **mass of your total vinegar** sample. *Hint: You used 10.0 mL of vinegar. Since vinegar is mostly water, we can assume that the density of vinegar is essentially that of pure water (1 g/mL)*

c. [1 pt] Calculate the **percent** by mass of acetic acid in vinegar. *[mass of acetic acid/ mass of vinegar × 100]*

d. [1 pt] The % acetic acid reported on vinegar bottle is 5%. What was your %error?

Post Lab Questions:

3) [1 pt] Why is it necessary to use an indicator when doing a titration? To do the most accurate titration, at what pH should the indicator change color (i.e. where should the *end point* be)?

4) A bottle of H_2SO_4 is found in a lab cabinet, but the bottle is not labeled with any molarity. Thus, you decide to do a titration with a standardized solution of NaOH. It is found that it takes 45.8 mL of the 2.65 M NaOH solution to titrate 30.0 mL of the H_2SO_4 solution to a sharp endpoint. The ***unbalanced*** equation is:

___ H_2SO_4 (aq)	+	___ NaOH (aq)	→	___ Na_2SO_4 (aq)	+	___ HOH (*l*)
30.0 mL of ?? M		*45.8 mL of 2.65 M*				

a) [1 pt] How many moles of NaOH were used to neutralize the H_2SO_4? (Remember to convert volume to liters.)

b) [1 pt] How many moles of H_2SO_4 were titrated? Show your calculation. *[Hint: stoichiometry! Balance your equation FIRST! It's not a 1:1 mole ratio!!]*

c) [1 pt] What is the molarity of the H_2SO_4 solution? (Remember to convert volume to liters.)

Read the attached article on Acid Rain and answer the following questions.

Air pollutants which cause acid rain:

1) [1 pt] What gas in the air makes rain **naturally** slightly acidic? ____________ Write the equation below:

2) [2 pts] What three gases released into the air cause acid rain? ____________________________________ Write the 3 equations:

Harmful effects of Acid Rain

3) [1 pt] Why does acid rain disintegrate buildings and statues made of limestone, marble and concrete? Write equation and explain.

4) [1 pt] Why is it harmful when acid rain falls on lakes?

5) [1 pt] What can be done to help an overly acidic lake?

Why SO_2 (g) and SO_3 (g) are released into the air

6) [1 pt] What is the major cause of SO_2 (g) and SO_3 (g) pollution? Write the chemical equations.

7) [1 pt] What can be done to reduce the amount of SO_2 (g) released into the air?

Why NO_2 (g) is released into the air.

8) [1 pt] What is the major source of NO_2 (g) pollution? Write equations that show its formation.

9) [1 pt] What part of your car significantly reduces the amount of NO_2 (g) released into the air? _________ _____________ Write equations involved. What catalysts are most commonly used? _________________

A Short Reading on Acid Rain

Acid Rain has become a widespread problem in the industrialized world. A variety of problems have been attributed to acid rain. Acid rain causes concrete buildings and marble statues to crumble; fish no longer populate some major lakes; and in some areas crops grow more slowly and forests begin to die. Some questions we want to be able to answer are the following: What air pollutants cause acid rain? How are these pollutants produced? What can be done to decrease air pollutants that cause acid rain? Why does acid rain have such a harmful effect on concrete and living things?

Air pollutants which cause acid rain

Good, clean rainwater is naturally slightly acidic (pH ~ 5.6) because carbon dioxide and other naturally occurring acidic substances are dissolved in the rainwater. The reaction with carbon dioxide is shown here:

$$\mathbf{CO_2\ (g) + H_2O\ (\mathit{l}) \rightarrow \underset{\textit{carbonic acid}}{H_2CO_3\ (aq)} \rightleftharpoons H^+\ (aq) + HCO_3^-\ (aq)}$$

weakly acidic solution ($K_a = 4.5\times10^{-7}$)

Only rain with a pH below 5.6 is considered **"acid rain."** Acid rain is most often caused by three common air pollutants, *SO_2 (g), SO_3 (g) and NO_2 (g).* All three of these gases form strong or relatively strong acids when they dissolve into rainwater.

$$\mathbf{SO_2\ (g) + H_2O\ (\mathit{l}) \rightarrow H_2SO_3\ (aq) \rightleftharpoons H^+\ (aq) + HSO_3^-\ (aq)}$$

sulfurous acid ($K_a = 1.54\times10^{-2}$, *among the strongest of weak acids*)

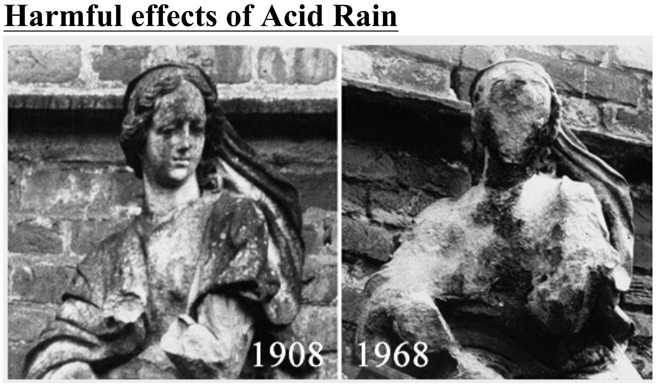
$$\mathbf{SO_3\ (g) + H_2O\ (\mathit{l}) \rightarrow H_2SO_4\ (aq) \rightarrow H^+\ (aq) + HSO_4^-\ (aq)}$$

sulfuric acid (strong acid)

$$\mathbf{2\ NO_2\ (g) + H_2O\ (\mathit{l}) \rightarrow HNO_3\ (aq) + HNO_2\ (aq) \rightarrow 2\ H^+\ (aq) + NO_3^-\ (aq) + NO_2^-(aq)}$$

nitric acid (strong) + nitrous acid ($K_a = 4.0\times10^{-4}$)

Harmful effects of Acid Rain



Acid rain is particularly harmful to buildings and statues because the acid reacts with insoluble $CaCO_3$ in limestone, marble and concrete to form $CaSO_4$—a more soluble solid. For example:

$$\mathbf{CaCO_3\ (s) + H_2SO_4\ (aq) \rightarrow CaSO_4\ (aq) + H_2O\ (\mathit{l}) + CO_2\ (g)}$$

Thus, the $CaCO_3$ (s) will be eroded away as it is turned into $CaSO_4$ and dissolves into the water.

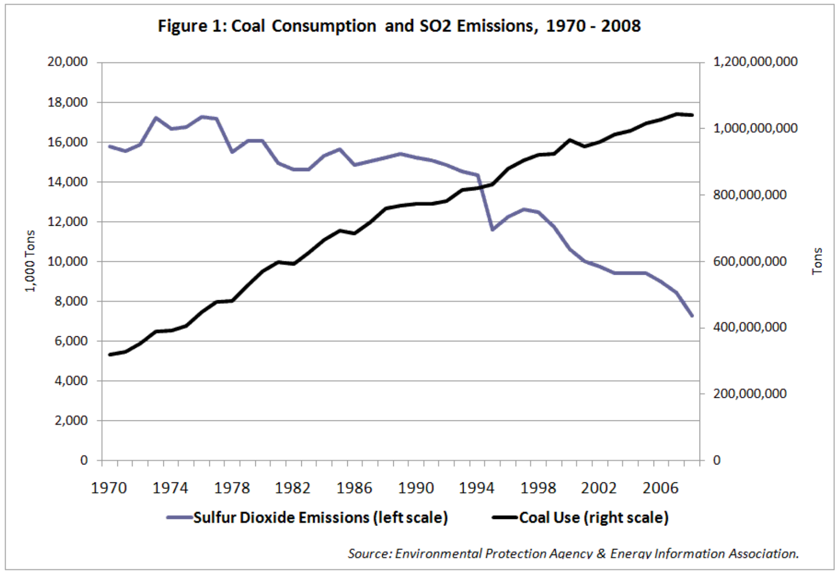
Acid Rain is also particularly harmful to fish and other aquatic life living in lakes. In the northeast, the pH of many lakes has been lowered to a pH of 4-4.5—the acidity of orange juice. Sometimes acid rain can lower the pH of a lake as low as pH of 3—the pH of vinegar. At these acidic levels, fish eggs die and many species of fish cannot live. To counteract the acid rain, sometimes lime (CaO) is added to lakes. CaO (like $CaCO_3$) can neutralize acids.

Why SO_2 (g) and SO_3 (g) are released into the air

SO_2 (g) and SO_3 (g) is released into the air mainly due to the burning of coal and oil (in power stations and in homes). This is because coal and oil contain varying quantities of sulfur. Thus, when the coal or oil is burned in oxygen, SO_2 and SO_3 gases are produced as shown:

$$\textbf{S (in coal or oil)} + O_2\ (g) \rightarrow SO_2\ (g)$$
$$\textbf{S (in coal or oil)} + \tfrac{3}{2}\ O_2\ (g) \rightarrow SO_3\ (g)$$

The amount of SO_2 and SO_3 gases released into the air can be greatly reduced by removing most of the sulfur from coal before burning it. Also, some coals and oils naturally contain less sulfur. In recent years, the United States has been successful in decreasing SO_2 pollution. Since 1970 there has been a dramatic decrease in the amount of SO_2 (g) released into the air even though coal burning has increased.



Why NO_2 (g) is released into the air.

Today NO_2 (g) is most commonly released into the air by the large amount of cars and trucks on the roads. NO_2 (g) is released into the air due to two reactions:

(1) At the high temperatures and pressures present inside an engine, nitrogen gas and oxygen gas react to form NO (g).

$$N_2\ (g) + O_2\ (g) \xrightarrow{\text{heat and pressure}} 2\ NO\ (g)$$

(2) In the atmosphere, NO from the car's exhaust then is oxidized to form visible orange-brown NO_2 (g):

$$2\ NO\ (g) + O_2\ (g) \rightarrow 2\ NO_2\ (g)$$

Catalytic converters have significantly reduced the amount of NO_2 (g) released into the air due to auto exhaust. Almost 7 million metric tons less of NO_2 (g) were emitted in 1991 than in 1980 Basically, the catalytic converters reduce the formation of NO_2 (g) by decreasing the amount of NO (g) in auto exhaust. NO (g) is removed from auto exhaust by these two reactions:

$$2\ NO\ (g) + 2\ CO\ (g) \rightarrow N_2\ (g) + 2\ CO_2\ (g)$$
$$2\ NO\ (g) + 2\ H_2\ (g) \rightarrow N_2\ (g) + 2\ H_2O\ (g)$$

In order to get the rates of these reactions fast enough, catalysts must be used. The most common catalysts for catalytic converters are platinum, palladium and rhodium—all relatively expensive, rare metals.