Chm 1045 Lab
Final Exam
Dr. EAD

Titration of Vinegar

Objective:

To determine the concentration (molarity and weight percent) of acetic acid in a vinegar sample.

Background:

Vinegar is a dilute solution of acetic acid ($HC_2H_3O_2$) in water. The concentration of acetic acid can be determined by titrating the acid with a known concentration of base solution. The end point in the titration can be determined using a phenolphthalein indicator which is colorless in acidic solution and red in basic solution. The endpoint will be the first pale pink that persists with mixing. The reaction is:

$$HC_2H_3O_2 + NaOH \rightarrow H_2O + NaC_2H_3O_2$$

Procedure:

1. Pipet 5 mL of vinegar solution into a 250 mL flask. Add about 20 mL of deionized water (the exact amount is not important).

2. Add 3-5 drops of phenolphthalein indicator.

3. Fill a buret with 0.200 M NaOH. Record the initial volume to the nearest 0.01 mL.

4. Open the buret stopcock fully, add the NaOH to the vinegar solution. Swirl the flask continuously. When the pink color begins to persist, slow down the addition of NaOH. The endpoint is the first palest pink color that permanently persists with swirling. A correct end point occurs when one drop changes the solution from colorless to pink. Record the final volume.

5. Repeat steps 1-4 twice.

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	Trial 1	Trial 2	Trial 3
Data:			
1) Volume of vinegar solution	5.0 mL	___ mL	___ mL
2) Concentration of NaOH solution	0.1 M	___ M	___ M
3) Initial buret reading	0.0 mL	___ mL	___ mL
4) Final buret reading	46.36 mL	___ mL	___ mL
Results:			
5) Volume of NaOH added in mL	46.36 mL	= #4 - #3 mL	___ mL
6) Volume of NaOH added in liters	0.0464 L	= #5 / 1000 L	___ L
7) Moles of NaOH added	0.00464 moles	= #2 x #6	___ moles
8) Moles of acetic acid ($HC_2H_3O_2$) present in vinegar sample	0.00464 moles	= moles NaOH moles	___ moles
9) Molarity of acetic acid in vinegar sample	0.93 M	= #8 / 0.005 L M	___ M
10) Average molarity	N/A M		
11) Mass of acetic acid in vinegar sample	0.279 g	= #8 x 60.06 g	___ g
12) Weight percent acetic acid in vinegar sample. (Assume that 5 mL of the vinegar sample weighs 5 g)	5.6%	#11 / 5g	x 100%
13) Average weight percent	N/A		

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Analysis of a Hydrogen Peroxide Solution

Objective:

To determine the percent H_2O_2 in an unknown solution.

Background:

Hydrogen peroxide is a common household and industrial substance. It is mainly used as an antiseptic and as a bleaching agent. Due to its reactive nature, hydrogen peroxide is usually diluted with water to form a solution. The concentration of an unknown H_2O_2 solution can be determined by a redox reaction involving MnO_4^- in acidic medium. The reaction is:

$$5\,H_2O_2 + 2\,MnO_4^- + 6\,H^+ \rightarrow 2\,Mn^{2+} + 8\,H_2O + 5\,O_2$$

Procedure:

1. Add about 6 g of the unknown H_2O_2 solution to a 125 mL flask followed by 25 mL of water.

2. Add about 15 mL of 2M H_2SO_4 to the flask.

3. Fill a buret with 0.100 M $KMnO_4$ solution.

4. Record the initial buret reading to the nearest 0.01 mL.

5. Slowly add $KMnO_4$ until the pink color begins to persist. Then add dropwise, while swirling, until one drop produces a permanent pink that lasts for 30 seconds. (The color is easier to see if a piece of white paper is placed under the flask). Record the final burette reading.

6. Repeat steps 1-5.

1) **Data:**	**Trial 1**	**Trial 2**
1) Mass H_2O_2 solution	10.500 g	______ g
2) Concentration $KMnO_4$ solution	0.100 M	______ M
3) Initial buret reading	0.0 mL	______ mL
4) Final buret reading	31.5 mL	______ mL

Results:	**Trial 1**	**Trial 2**
5) Volume of $KMnO_4$ solution added	31.5 mL = #4 − #3	______ mL
6) Moles of $KMnO_4$ added = (#2 × #5 / 1000 mL) moles →	0.00315 moles	
7) Moles of H_2O_2 in unknown = #6 × $\frac{5 \text{ mol } H_2O_2}{2 \text{ mol } MnO_4^-}$ moles →	0.007875 moles	
8) Mass of H_2O_2 in unknown = #7 × 34.02 g →	0.268 g	
9) Percent H_2O_2 in unknown = #8 / #1 × 100% →	2.55%	
Average		______

Determination of Specific Heat

Objective:

To determine the specific heat of an unknown metal sample.

Background:

Heat and temperature are different. Heat is the amount of energy in a body, measured in calories or joules, whereas temperature is the intensity of hotness, measured in degrees Fahrenheit or Celsius.

The *specific heat* of a substance is the amount of heat required to change the temperature of 1 g of the substance by 1°C and has units of joules (or calories) /g°C. For example, the value for water is 4.184 joules per gram per °C and is written 4.184 J/g°C.

Procedure:

1. Carefully weigh approximately 45 g of water into a styrofoam cup. Place it inside another cup for insulation purposes.

2. Weigh an unknown metal.

3. Add water to a beaker and bring it to a boil. Gently place the metal sample in the boiling water for 5 minutes; the metal must be completely submerged.

4. Measure the temperature of the boiling water (and metal).

5. Measure the temperature of the water in the styrofoam cup. Leave the thermometer in the styrofoam cup.

6. Using tongs, quickly transfer the metal sample to the water in the styrofoam cup.

7. Use the thermometer to stir the water containing the metal sample. Record the highest temperature.

8. Repeat steps 1-7

9. Calculate the specific heat of the metal.

Useful Information:

Since the styrofoam cup is thermally insulated, very little heat energy is lost to the surroundings. The heat gained by the cold water in the styrofoam cup is equal to the heat lost by the hot metal sample. In equation form:

A. Heat lost by metal = Heat gained by cold water
B. Heat lost by metal = (specific heat metal) (mass metal) (change T metal)
C. Heat gained by water = (specific heat water) (mass water) (change T water)

Set equation B equal to equation C and solve for the specific heat metal. Use 4.18 J/g.°C for the specific heat of water.

Data:

1. Mass cold water	45 g	g
2. Mass unknown metal	60 g	g
3. Initial temperature of hot metal	100 °C	°C
4. Initial temperature of cold water in cup	22.6 °C	°C
5. Final temperature of water-metal mixture	32.8 °C	°C

Results:

6. Change in temperature of metal	= (#5 − #3)	−67.2 °C	°C
7. Change in temperature of cold water	= #5 − #4	10.2 °C	°C
8. Heat gained by water	= #1×(specific heat H_2O)×7	1918.6 J	J
9. Heat lost by metal	= − #8	−1918.6 J	J
10. Specific heat of metal	$= \frac{\#9}{(\#2)\times(\#6)}$	0.476 J/g°C	J/g°C

Average specific heat ______ J/g°C

Calorimetry and Hess's Law

Objective:

To measure reaction enthalpies (ΔH) and to verify Hess's Law.

Background:

An exothermic reaction releases energy into the surroundings. Three exothermic reactions will be studied in this experiment. A calorimeter will be used to determine the heat energy released (enthalpies). The reactions whose enthalpies (ΔH) will be measured are:

A. $HCl_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

B. $NaOH_{(s)} \rightarrow NaOH_{(aq)}$

C. $HCl_{(aq)} + NaOH_{(s)} \rightarrow H_2O_{(l)} + NaCl_{(aq)}$

Reactions A and B can be added to obtain reaction C. According to Hess's Law,

$$\Delta H_A + \Delta H_B = \Delta H_C$$

Procedure:

Reaction A:

1. Add 50.0 mL of 1.0M HCl to a styrofoam calorimeter.

2. Add 50.0 mL of 1.0M NaOH to second styrofoam calorimeter.

3. Measure the temperature of the cup with the HCl solution. (since both solutions are stored in the same room, it is reasonable to assume both are at the same temperature). Keep the thermometer in the HCl solution.

4. Pour the NaOH solution into the cup containing the HCl and stir using the thermometer. Watch as the temperature increases and record the highest reading.

Reaction B:

1. Add 50.0 mL of water to a styrofoam cup.

2. Measure the temperature of the water. Leave the thermometer in the water.

3. Quickly weigh about 2.0 g solid NaOH on a watch glass. Immediately add the solid NaOH to the water and stir vigorously. Record the highest temperature reading. (Note: solid NaOH absorbes moisture from the air and must be added to the water cup as quickly as possible).

Reaction C:

1. Add 50.0 mL of 1.0M HCl to a styrofoam cup.

2. Measure the temperature of the HCl solution. Keep the thermometer in the HCl solution.

3. Quickly weigh about 2.00 g solid NaOH and add to the HCl solution with stirring.

4. Record the highest temperature reading.

Useful information:

The heat energy released by reactions A, B and C can be calculated using the equation below.

Heat Energy = (specific heat) (mass) (change in temperature)

Since the acid and base solutions are dilute, one can use the specific heat of pure water (4.18 J/g°C) and one can estimate that 50.0 mL of solution equals 50.0 g of solution when calculating the heat released.

Data:

Reaction A:

1. Volume of 1.0M HCl	50	mL
2. Volume of 1.0M NaOH	50	mL
3. Initial temperature of HCl solution	25.4	°C
4. Final temperature of mixture	31.8	°C

Reaction B:

5. Volume of water	50	mL
6. Initial temperature of water	24.5	°C
7. Mass solid NaOH	2.02	g
8. Final temperature of mixture	33.6	°C

Reaction C:

9. Volume of 1.0M HCl	100	mL
10. Initial temperature of HCl solution	23.9	°C
11. Mass solid NaOH	2.04	g
12. Final temperature of mixture	35.2	°C

Results: $\boxed{mol = (L)(M)}$

$\boxed{1 = \frac{mol}{L}}$

Reaction A:

Total mass	50g HCl, 50g NaOH	100 g
Change in temperature	31.8°C − 25.4°C	6.4 °C
Heat released	$-q = ms\Delta t = (100g)(4.18 J/g°C)(6.4°C)$	−2675.20 J
Moles HCl added	(1)(0.050L)	0.05 moles
Moles NaOH added	(1)(0.050L)	0.05 moles
ΔH_A (Heat released when one mole HCl and one mole NaOH reacts)	$-q/mol = \frac{-2675.2}{0.05 \text{ mols}}$	−53504 J

Reaction B:

$\boxed{1W = \frac{g}{mol}}$

Total mass	50g H_2O + 2.02g NaOH	52.02 g
Change in temperature	33.6°C − 24.5°C	9.1 °C
Heat released	$-q = ms\Delta t = (52.02g)(4.18 J/g°C)(9.1°C)$	−1978.74 J
Moles NaOH that dissolved	2.02g NaOH × (1mol / 40g NaOH)	0.0505 moles
ΔH_B (Heat released when one mole NaOH dissolves in water)	$-q/mol = \frac{-1978.74}{0.0505}$	−39183 J

Reaction C:

mol HCl = (0.100L)(1)

Total mass	100g HCl + 2.04g NaOH	102.04 g
Change in temperature	35.2°C − 23.9°C	11.3 °C
Heat released	$-q = ms\Delta t = (102.04g)(4.18 J/g°C)(11.3°C)$	−4819.76 J
Moles HCl added		0.100 moles
Moles NaOH added	$2.04g\,\cancel{NaOH} \times \frac{1 mol}{40g\,\cancel{NaOH}}$	0.051 moles
ΔH_C (Heat released when one mole NaOH dissolves and reacts with one mole HCl)	$\frac{-4819.76J}{0.051}$	−94506 J

Verification of Hess's Law:

ΔH_A + ΔH_B = ΔH_C

−53504 J + −39183 J = −92687 J

−94,506 J

Atomic Weight of a Metal from Hydrogen Gas Production

Objective:

To produce hydrogen gas by reaction of a metal with acid.
To use the ideal gas equation to determine the atomic weight of the metal.

Background:

Many metals react with hydrochloric acid to produce hydrogen gas and metal ions.

For example:

$$2\,Al_{(s)} + 6\,H^{+}_{(aq)} \rightarrow 3\,H_{2(g)} + 2\,Al^{3+}_{(aq)}$$

$$Mg_{(s)} + 2\,H^{+}_{(aq)} \rightarrow H_{2(g)} + Mg^{2+}_{(aq)}$$

The moles of hydrogen gas produced can be determined using the Ideal Gas equation, PV = nRT. From the moles of hydrogen and the stoichiometry (reacting ratios) of the reaction, it will be possible to calculate how many moles of metal reacted. Finally the atomic weight of the metal will be determined.

Procedure:

1. A buret will be used to hold the hydrogen gas produced in the reaction. First, however, the uncalibrated volume at the bottom of your buret must be estimated. To do this add water to a 10 mL graduated cylinder and carefully read the volume to the nearest 0.1 mL. Pour water from the cylinder into the buret until the "uncalibrated volume" is full with water. Read the new volume of water in the graduated cylinder. The difference represent the "uncalibrated" volume of the buret.

2. Weigh the metal. For Al use 0.030 - 0.032 g; for Mg use 0.038 - 0.042 g.

3. Wrap the metal sample with copper wire into a "cage". Leave about 5 cm of unwrapped Cu wire to serve as a handle.

4. Add concentrated HCl to the empty buret. For Al use about 20 mL HCl; for Mg use about 3 mL HCl. These amounts are more than enough to completely react the metal.

5. Using a wash bottle, slowly add water down the side of the buret. Do not disturb the bottom acid layer. Fill the buret completely to the top.

6. Fill a 600 mL beaker about half full with water.

7. Do this next step quickly and carefully. Insert the metal sample cage into the top of the buret, leaving the Cu wire handle on the outside. Push a one-hole rubber stopper into the buret top to hold the Cu wire in place. Try to leave no air bubbles. Place your thumb over the hole in the rubber stopper, turn the buret upside down, place the "top" of the buret below the water level of the 600 mL beaker, and clamp the upside-down buret in place. The denser, concentrated HCl will slowly migrate downward to the metal where the reaction will take place producing hydrogen gas. When the reaction is complete the buret should be nearly full of hydrogen gas and water vapor.

8. Measure the temperature of the water in the beaker. We will assume that the water and gases are at the same temperature.

9. Adjust the water level in the beaker or move the buret up or down (keep buret end submerged) so that the water levels in the beaker and the buret are even with each other. If not level, measure the difference in mm, and convert mm of water to mm of mercury by dividing by 13.6. Subtract from barometric pressure to give hydrogen gas pressure. At this point the gas pressure in the buret is equal to the outside (barometric) pressure. Record the volume of the gas in the buret when the water levels are even. **Note:** The buret may need to be transferred to a larger breaker in order to adjust the water levels.

10. Record the barometric pressure for the day.

11. Look up the vapor pressure of water at the appropriate temperature.

Useful Information:

$PV = nRT$

$T_K = T^{o}_{C} + 273$

$R = 0.082 \dfrac{\text{atm} \circ \text{L}}{\text{mole} \circ \text{K}}$

$1.0 \text{ atm} = 760 \text{ torr}$

Data:

1. Volume of water in graduated cylinder before addition to buret. 9.0 mL

2. Volume of water in graduated cylinder after addition to buret 5.5 mL

3. Mass of metal sample 0.040 g

4. Temperature of water (and thus temperature of gases). 24.5 °C

5. Buret reading of gases 6.5 mL

6. Difference in water levels 0.0 mm

7. Barometric pressure 760.0 torr

Results:

8. "Uncalibrated volume" of buret = #1 − #2 3.5 mL

9. Total volume of hydrogen gas = (50.0 mL − #5) + #8 47.0 mL = 0.0470 L

10. Temperature of hydrogen gas #4 + 273 297.5 K

11. Mercury equivalent of water column 0.0 torr

12. Total pressure of gas sample 760.0 torr
(P_{H_2O} = 23.0 torr)

13. Partial Pressure of hydrogen gas = #12 − P_{H_2O} 737.0 torr = 0.970 atm

14. Moles of hydrogen gas produced $= \dfrac{\#13 \times \#9}{0.0821\,\frac{L\cdot atm}{mol\cdot K} \times \#10}$ 0.00187 mole

15. Moles of metal that reacted = moles H_2 = #14 0.00187 moles

16. Calculated atomic weight of metal $= \dfrac{\#3}{\#15}$ 21.4 g/mole

Determination of Molar Mass

Objective:

To measure the temperature, volume, pressure and mass of an unknown volatile liquid, and use the ideal gas law to determine the molar mass of the unknown.

Procedure:

1. Set up a boiling water bath using a 600 mL beaker about ¾ filled with water. Use a Bunsen burner as a heat source. Record the temperature of the boiling water.

2. Obtain the mass of a dry 125 mL flask with a piece of aluminum foil to be used as a cap.

3. Add about 2 mL of the unknown volatile liquid to the flask. The exact amount is not important.

4. Place the foil tightly over the top of the flask and secure with a rubber band. Make the smallest possible hole in the foil using a small pin.

5. Clamp the flask at its very top and place the flask into the boiling water. Be careful! Cover as much of the flask as possible with the boiling water.

6. Carefully watch the unknown liquid in the flask as it vaporizes. After the unknown liquid is not longer visible, keep the flask submerged in the boiling water for about 2 minutes.

7. Remove the flask from the boiling water, after cooling, thoroughly dry the outside with paper towels and weigh the flask containing the unknown and foil.

8. Record the barometric pressure. This was the pressure of the unknown gas in the boiling water.

9. Fill the flask as full as possible with water. Pour the water into a large graduated cylinder and measure the volume. The volume of water is equal to the volume of the flask and thus the volume of the gas.

10. Repeat using a second dry flask and piece of foil.

Useful Information:

$PV = nRT$ $\quad\quad R = 0.082 \frac{atm \circ L}{mole \circ K}$

$T_K = T_{\circ C} + 273$

Data:

	Trial 1	Trail 2
1. Temperature of boiling water = temperature of gas	100.3 °C	____ °C
2. Mass of flask plus foil	124.57 g	____ g
3. Mass of flask plus foil plus unknown gas	125.80 g	____ g
4. Barometric pressure	769 torr	____ torr
5. Volume of flask = volume of gas	273.5 mL	____ mL

Results:

	Trial 1	Trail 2
6. Temperature of unknown gas = #1+273	373.3 K	____ K
7. Mass of unknown gas = #3 − #2	1.23 g	____ g
8. Pressure of unknown gas = #4 ÷ 760 torr	1.012 atm	____ atm
9. Volume of unknown gas = 5 ÷ 1000 mL	0.274 L	____ L
10. Moles of unknown gas = $\frac{\#8 \times \#9}{(0.0821 \frac{L \cdot atm}{mole \cdot K}) \times \#6}$	0.00905 mole	____ mole
11. Molar mass of unknown = #7 ÷ #10	135.9 g/mole	____ g/mole
Average	____ g/mole	