1 Techniques and Measurements

Name: ....

Lab. Instructor ....

Section ....

Date ....G.........

Pre-Laboratory Questions

1. A metal sphere weighing 9.48 g is added to 21.27 mL water in a graduated cylinder. The water level rises to 24.78 mL. Calculate the density of the metal.

Volume of metal sphere :- 24.78 - 21.27 = 3.51 mL

Density of metal sphere :- $\frac{9.48}{3.51}$ = 2.70 g/mL

2. An empty beaker weighs 32.4257 g. A 10.00 mL sample of unknown liquid is transferred to the beaker. The total mass of the beaker and liquid sample was 39.4507 g. Calculate the density of the unknown liquid.

mass of the liquid :- 39.4507 - 32.4257 = 7.0250 g

density of the liquid :- $\frac{7.0250}{10.00}$ = 0.702 g/ml

3. A term that is easily confused with density is **specific gravity**. What is meant by specific gravity? What are the units of specific gravity?

Specific gravity is the ratio of the weight of a volume of the substance to the weight of an equal volume of the reference substance (water)

Specific gravity has no units (unitless)

Specific gravity = $\frac{d \text{ substance}}{d\ H_2O}$

1 Techniques and Measurements

Name: ....

Section ....

Lab. Instructor ....

Date ....

Results and Calculation

A. Determination of the Density of Pure Liquid

Pure water	Trial (I)	Trial (II)
Mass of beaker	32.96 ± 0.01 g	32.95 ± 0.01 g
Mass of beaker + water	42.82 ± 0.01 g	42.83 ± 0.01 g
Mass of water	9.86 ± 0.01 g	9.88 ± 0.01 g
Volume of water	10.0 ± 0.1 mL	10.0 ± 0.1 mL
Temperature of water	18.0 ± 0.2 °C	18.0 ± 0.2 °C
Density	0.986 g/mL	0.988 g/mL
Average density	0.987 g/mL	
Handbook density	0.99862 g/mL	
Unknown liquid		
Unknown Nomber	A	
Mass of beaker	32.96 ± 0.01 g	32.95 ± 0.01 g
Mass of beaker + unknown liquid	45.21 ± 0.01 g	45.22 ± 0.01 g
Mass of unknown liquid	12.25 ± 0.01 g	12.27 ± 0.01 g
Volume of unknown liquid	15.9 ± 0.1 mL	16.0 ± 0.1 mL
Density of unknown liquid	0.7704 g/mL	0.7668 g/mL

B. Density of Solutions

m NaCl = 2.50 g
V water = 25 mL

	Trial (I)	Trial (II)
Mass of beaker	32.96 ± 0.01 g	32.95 ± 0.01 g
Volume of solution	10.0 ± 0.1 mL	10.0 ± 0.1 mL
Mass of beaker + solution	43.51 ± 0.01 g	43.51 ± 0.01 g
Mass of solution	10.55 g	10.56 g
Temperature of solution	19.0 ± 0.2 °C	19.0 ± 0.2 °C
Density of solution	1.055 g/mL	1.056 g/mL

C. Density of Solids

	Trial (I)	Trial (II)
Mass of beaker	33.22 ± 0.01 g	33.23 ± 0.01 g
Mass of beaker + solid pieces	42.73 ± 0.01 g	42.73 ± 0.01 g
Mass of solid pieces	9.51 ± 0.01 g	9.50 ± 0.01 g
Initial water level in the graduated cylinder	25.1 ± 0.1 mL	25.3 ± 0.1 mL
Final water level in the graduated cylinder	29.2 ± 0.1 mL	29.1 ± 0.1 mL
Volume of solid pieces	4.1 mL	3.8 mL
Density of the solid substance	2.31 g/mL	2.50 g/mL

Questions

1. What error would be introduced into the density of the metal pellets if you had not shaken the pellets to remove adhering air bubbles? Would the density be too high or too low?

The density would be too low because air bubbles trap air inside them ✓

2. The density of silver is 10.5 g/cm^3 and the density of platinum is 21.45 g/cm^3. If equal masses of silver and platinum were transferred to equal volumes of water in graduated cylinders, which graduated cylinder would have the highest volume reading? Explain.

Silver i because density is inversely proportional with volume ✓

Seen

2 Formula of a Hydrate

Name: ....

Lab. Instructor .... Section ....

Date ....

Pre-Laboratory Questions

0/C

1. What are alums? Give examples other than potassium alum.

Alums are a class of hydrates that consist a double salt with the general Formula ($M'M'''(SO_4)_2 . X\ H_2O$) where M' is a univalent cation and M''' is a trivalent cation.
An example is the potassium alum $KAl(SO_4)_2 . 12H_2O$.

2. What are the hydrates? Give few examples of metal salt hydrates.

Hydrates are solid compounds that have been crystallized from aqueous solutions with weakly bounded water molecules contained in the crystal.
An example is copper sulfate pentahydrate $CuSO_4 5H_2O$

3. Potassium chromic alum has the formula: $KCr(SO_4)_2.xH_2O$. A sample of 1.12 g of this alum was heated in a crucible to get a constant mass. The mass of the anhydrous salt produced ($KCr(SO_4)_2$) was 0.64 g. Calculate the value "x" in the formula of the alum.

mass of H_2O = 1.12g - 0.64g = 0.48g

molar mass of H_2O = 18.002 g/mol

n. of H_2O moles = $\frac{0.48}{18.002}$ = 0.0266 mol

n. of anhydrous salt moles = $\frac{0.64}{283}$ = $2.26 * 10^{-3}$ mol

2 Formula of a Hydrate

Name: ……………………………………………… Section ……………………

Lab. Instructor ……………………………………… Date ……………………

Results and Calculations

A. Potassium Alum:

Mass of empty crucible (m_1)	16.78 ± 0.01	g
Mass of crucible and the alum (m_2)	17.78 ± 0.01	g
Mass of crucible and anhydrous salt (m_3)	17.35 ± 0.01	g
Mass of alum (m_2–m_1)	1.00	g
Mass of anhydrous salt (m_3–m_1)	0.57	g
Mass of water lost upon heating (m_2–m_3)	0.43	g
Number of moles of water lost upon heating	$\frac{0.43}{18.016} = 0.024$	mol
Number of moles of anhydrous salt ($KAl(SO_4)_2$)	$\frac{0.57}{258.22} = 2.21 * 10^{-3}$	mol
Percentage of water of crystallization, by mass $\frac{\text{mass } H_2O}{\text{mass alum}} * 100$	$\frac{0.43}{1.00} * 100 = 43$	%
The value "x" in the formula, (number of moles of water of crystallization / number of moles of anhydrous salt)	$\frac{0.024}{2.21 * 10^{-3}} = 10.8$ $\simeq 11$	

B. Unknown Hydrate:

Unknown number: ____________

Mass of empty crucible (m_4)	16.81 ± 0.01	g
Mass of crucible and the hydrate (m_5)	17.75 ± 0.01	g
Mass of crucible and anhydrous salt (m_6)	17.30 ± 0.01	g
Mass of anhydrous salt (m_6–m_4)	0.49	g
Mass of water lost upon heating (m_5–m_6)	0.45	g
Percentage of water of crystallization, by mass	$\frac{0.45}{0.94}$ x 100	%

= 47.8 %

QUESTIONS

1. What is the effect on the calculated value of "x" if the dehydration of the alum is not complete

The value of "x" will decrease

2. A student heated 1.16 g of hydrated sodium sulfate $Na_2(SO_4)$ in a crucible to get 0.51 g of anhydrous salt. What is the formula of the anhydrous salt? (Show your work)

mass of water = 0.65 g

n moles of water = 0.036 mol

n of anhydrous salt = $\frac{0.51 g}{142.06}$ = $3.59 * 10^{-3}$

$x = \frac{0.036}{3.59 * 10^{-3}} = 10.0$

Formula:- $Na_2SO_4 \cdot 10 H_2O$

3 The Empirical Formula of an Oxide

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. What is the mole and molar mass? What are the molar masses of magnesium and atomic oxygen?

Mole is a unit of measurement for the amount of substance or chemical amount.

Molar mass is a physical property characteristic of a given substance (namely it's mass per amount of substance). MM (Mg) = 24.31 g/mol MM (O_2) = 16 g/mol

2. When 0.192 g of phosphorus is burned, 0.341 g of a white oxide is obtained. (a) Write a balanced chemical equation for the reaction of phosphorus with molecular oxygen based on this empirical formula. (b) Determine the empirical formula of the oxide.

a) $4P + 3O_2 \rightarrow 2P_2O_3$

b) n phosphorus = $\frac{0.192}{30.97}$ = $6.19 * 10^{-3}$ mol

m O_2 = 0.149 g n O_2 = $9.31 * 10^{-3}$ mol

$(P_{6.19*10^{-3}} O_{9.31*10^{-3}})$ % $6.19 * 10^{-3}$

$(P_1O_{1.5}) * 2 \Rightarrow P_2O_3$ empirical Formula

3 The Empirical Formula of an Oxide

Name: ....

Section ....

Lab. Instructor ....

Date ....

Results and Calculations

Mass of empty crucible (after first heating)	15.64	g
Mass of empty crucible (after second heating)	15.63	g
Final mass of empty crucible	15.63	g
Mass of crucible and Mg	15.83	g
Mass of Mg	0.20	g
Moles of Mg (**n1**)	$\frac{0.20}{24.31} = 8.23 * 10^{-3}$	mol
Final mass of crucible and Mg-oxide	15.94	g
Mass of Mg-oxide produced	0.31	g
Mass of oxygen gained	0.11	g
Moles of oxygen atoms (**n2**)	$6.87 * 10^{-3}$	mol
Formula of magnesium oxide ($Mg_{n1}O_{n2}$)	$Mg_{8.23*10^{-3}} O_{6.87*10^{-3}}$	
Empirical formula of magnesium oxide	MgO ✓	
Mass percent of Mg in the oxide (x_1) (experimentally)	64.5	%
Mass percent of Mg in the oxide (x_2) (calculated for MgO) using molar mass	60.3 ✓	%
Percentage error = [|x_2-x_1|/x_2] X 100%	6.9 ≃ 7	%

QUESTIONS

1. If water had not been added to your initial product, what error in the determined percentage of magnesium would have resulted (that is, if part of the product has been magnesium nitride)? Explain.

The percent would Increase because the mass of Mg would Increase and mass of MgO would decrease.

2. If large amount of magnesium oxide had been lost during the heating of the crucible, would this have made the calculated %Mg in the product too high or too low? Explain.

Too high because the mass of MgO would decrease.

4 Limiting Reactant

Name: ....

Section ....

Lab. Instructor ....

Date ....

Pre-Laboratory Questions

1. $Na_3PO_4.12H_2O$ and $BaCl_2.2H_2O$ are examples of compounds known as hydrates. Calculate the molar masses of Na_3PO_4 and $BaCl_2$ in the hydrated and anhydrous forms.

mm Na_3PO_4 = $3 * 22.99 + 4 * 16.00 + 30.97$ = 163.94 g/mol

mm $Na_3PO_4.12H_2O$ = $216.192 + 163.93$ = 380.18 g/mol

mm $BaCl_2$ = $137.3 + 2 * 35.45$ = 208.2 g/mol

mm $BaCl_2.2H_2O$ = $36.032 + 208.2$ = 244.2 g/mol

2. A mixture of 0.28 g $BaCl_2.2H_2O$ and 0.72 g $Na_3PO_4.12H_2O$ was dissolved in water. Which of the two reactants is the limiting reactant? Calculate the mass of the precipitate formed, $Ba_3(PO_4)_2$.

$3BaCl_2 + 2Na_3PO_4 \rightarrow 6NaCl + Ba_3(PO_4)_2$

$\downarrow$ 0.28 g $\downarrow$ 0.72 g

n $BaCl_2$ = $1.3 * 10^{-3}$ mol   n Na_3PO_4 = $4.4 * 10^{-3}$ mol

→ n $Ba_3(PO_4)_2$ produced From $BaCl_2$:-

$1.3 * 10^{-3}$ mol $BaCl_2$ * $\frac{1 \text{ mol } (Ba_3(PO_4)_2)}{3 \text{ mol } BaCl_2}$ = $4.3 * 10^{-4}$ mol

→ n $Ba_3(PO_4)_2$ produced From Na_3PO_4 :-

$4.4 * 10^{-3}$ mol Na_3PO_4 * $\frac{1 \text{ mol } BaCl_2}{2 \text{ mol } Na_3PO_4}$ = $2.2 * 10^{-3}$

∴ $BaCl_2$ is the ~~limiting~~ reagent

→ m. $Ba_3(PO_4)_2$ = $601.84 * 4.3 * 10^{-4}$ = 0.22 g

4 Limiting Reactant

Name: ....

Lab. Instructor .... Section ....

Date ....

Results and Calculations

A. Precipitation of $Ba_3(PO_4)_2$:

Unknown Number: 9

Mass of salt mixture (m_1)	0.70	g
Mass of filter paper (m_2)	0.58	g
Mass of filter paper and $Ba_3(PO_4)_2$ (m_3)	0.87	g

B. Determination of the Limiting Reactant:

Limiting reactant in salt mixture is ....$BaCl_2.2H_2O$

Excess reactant in salt mixture is .$Na_3PO_4.12H_2O$

Mass of $Ba_3(PO_4)_2$ precipitated (m_3-m_2)	0.29	g
Number of moles of $Ba_3(PO_4)_2$ precipitated (n_1)	$4.85 * 10^{-4}$	mol

(a) If the limiting reactant is $BaCl_2$. $2H_2O$:

Number of moles of $BaCl_2.2H_2O$ reacted (n_2) $n_1 \times 3$	$1.44 * 10^{-3}$	mol
Number of moles of Na_3PO_4 .$2H_2O$ reacted (n_3) $n_1 \times 2$ 2	$9.62 * 10^{-4}$	mol
Mass of $BaCl_2.2H_2O$ reacted (m_4) $n_2 \times Mw$	0.35	g
Mass of Na_3PO_4 reacted (m_5) $n_3 \times Mw$	~~0.16~~ 0.36	g
Mass of excess Na_3PO_4 [m_1-(m_4+m_5)]	~~0.19~~ zero	g
Mass percentage of $BaCl_2.2H_2O$	49	%

m_4 ×100%

0.5

QUESTIONS

1. Calculate the mass of $Ba_3(PO_4)_2$ produced from the reaction of 0.78 g $Na_3PO_4.12H_2O$ with excess $BaCl_2.2H_2O$. What is the purpose of heating the mixture in step 3 for 20 minutes? In order For the reaction (coagulation) to occur

$mm\ Na_3PO_4.12H_2O = 380.2\ g/mol$

$n\ Na_3PO_4.12H_2O = \frac{0.78}{380.2} = 2.1 * 10^{-3}\ mol$

$2.1 * 10^{-3}\ mol * \frac{1\ mol}{2\ mol} = 1.02 * 10^{-3}\ mol\ Ba_3(PO_4)_2$

$m\ Ba_3(PO_4)_2 = 601.84 * 1.02 * 10^{-3} = 0.61\ g$

0.62

2. What is the purpose of washing the precipitate with hot water in step 4? How would the reported percentage of the excess reactant be affected if the precipitate was not washed in this step?

To collect all the precipitate The percentage of the ~~limiting~~ reagent would increase causing the excess reagent to decrease (the limiting reactant will be a mixture of both LR and ER)

5 Determination of Acetic Acid in Vinegar

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. Why was the standard NaOH solution not prepared by calculating the amount of solid NaOH needed for 100 mL of solution, weighing it accurately, and making it up to exactly 100 mL of total volume?

Because NaOH is a secondary standard solution which means that its concentration changes with time and it may react with vapor water and absorb H_2O.

2. Why not simply rinse the buret with distilled water rather than the solution to be used in it?

This assures that the solution to be used in the buret will not be dilluted with distilled water adhering to the buret coat

3. Why does the volume of water added to potassium hydrogen phthalate not have to be measured carefully?

Because the basis of this reaction is that if one reactant reacts completely with another reactant with no excess remaining., the number of moles of each reactant is the same

5 Determination of Acetic Acid in Vinegar

Name: ………………………………… Section ……………………

Lab. Instructor …………………… Date ……………………

Results and Calculations

A. Standardization of NaOH Solution

	Trial (I)	Trial (II)
Mass of flask	122.08 g	122.08 g
Mass of Flask + KHP	122.28 g	122.29 g
Mass of KHP	0.20 g	0.21 g
Molar mass of KHP	204.22 g/mol	204.22 g/mol
Moles of KHP	$9.8 * 10^{-4}$ mol	$1.0 * 10^{-3}$ mol
Initial buret reading	20.00 mL	18.10 mL
Final buret reading	27.00 mL	24.10 mL
Volume of NaOH	7.00 mL	6.00 mL
Moles of NaOH	$9.8 * 10^{-4}$ mol	$1.0 * 10^{-3}$ mol
Molarity of NaOH	0.14 M	0.17 M
Average Molarity of NaOH	0.16 M	

$M = \frac{mol}{V}$

B. Mass percent of acetic acid in vinegar. (15)

	Trial (I)		Trial (II)	
Volume of vinegar	10.00	mL	10.00	m
Initial buret reading	22.00	mL	20.80	m
Final buret reading	36.60	mL	35.00	m
Volume of NaOH used	14.60	mL	14.20	m
Average molarity of NaOH, From Part A:			0.16	
Moles of NaOH used $n = cV$	2.3	mol	2.3	m
Moles of CH_3COOH in vinegar Reacted with the NaOH	2.3	mol	2.3	m
Molarity of CH_3COOH in vinegar	~~0.023~~ 0.23	M	~~0.023~~ 0.23	
Average molarity of acetic acid in vinegar			~~0.023~~ 0.23	
Molar Mass of acetic acid, CH_3COOH $m = M * n$			60.052	g/m
Mass of CH_3COOH per liter of vinegar (NaOH) Volume * d			14 ~~$138 * 10^{-3}$~~	
Mass percent acetic acid in vinegar (assume that vinegar has a density of 1.00 g/mL)			0.14 1.4	

$d = \frac{m}{V}$

$m = 1.00 * 10.00$

$= 10.00$ g

QUESTIONS

$$\frac{|\text{real} - \square|}{\text{reel}} \times 100\%$$

1. Calculate the percent error would have been in a titration that used 32.75 mL of a solution if a bubble with a volume of 0.25 mL had been swept out of the tip during the titration.

$$\frac{0.25}{\cancel{32.75}\ 32.50} * 100\% = 0.76\%$$

2. The label on the vinegar bottle used in this experiment claims that the vinegar contains 3 % acetic acid by weight. Use your results and a density of 1.0 g/mL to investigate this claim. $M = 60.052$ g/mol

$$3\% = \frac{\text{Conc} * 60.052}{d * 1000} * 100\%$$

concentration = 0.499 ≈ 0.5 M

6 The Neutralizing Capacity of Antacid Tablets

Name: ....

Lab. Instructor ....

Section ....

Date ....

Pre-Laboratory Questions

1. Write a balanced equation for the reaction of HCl and $NaHCO_3$.

$NaHCO_3 + HCl \rightarrow NaCl + \underbrace{CO_2 + H_2O}_{H_2CO_3}$

2. How many moles of HCl are needed to react with 0.47 g of $NaHCO_3$?

$n\ NaHCO_3 = n\ HCl$

$mm\ NaHCO_3 = 84.008\ g/mol$

$n = 5.6 \times 10^{-3}\ mol = n\ HCl$

3. Why we do not use strong bases as active ingredients of antacid tablet?

Because strong bases go to completion and in this case we don't want the stomach to become neutralized or to become basic. instead we need it to get back to it's normal acidity

6 The Neutralizing Capacity of Antacid Tablets

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

Name of antacid tablet: ....

Ingredients:..

	Trial (I)		Trial (II)	
Mass of antacid sample	0.20	g	0.21	g
Volume of 0.150 M HCl solution	25.00	mL	25.00	mL
Moles of HCl (used to dissolve antacid)	$3.75 * 10^{-3}$	mol	$3.75 * 10^{-3}$	mol
Initial buret reading	0.00	mL	0.00	mL
Final buret reading	16.20	mL	16.40	mL
Volume of NaOH added	16.20	mL	16.40	mL
Moles of NaOH (used to titrate the excess acid)	$2.43 * 10^{-3}$	mol	$2.46 * 10^{-3}$	mol
Moles of excess HCl	$2.43 * 10^{-3}$	mol	$2.46 * 10^{-3}$	mol
Moles of HCl(needed to neutralize the antacid tablet)	$1.32 * 10^{-3}$	mol	$1.29 * 10^{-3}$	mol
Neutralizing capacity of antacid	$6.60 * 10^{-3}$	mol HCl/g	$6.14 * 10^{-3}$	mol HCl/g
Average (mol HCl/g antacid)	$6.37 * 10^{-3}$		mol HCl/g antacid	

$\frac{10}{10}$

QUESTIONS

1. Assume the concentration of stomach acid is 0.10M and the neutralizing capacity of the acid is 0.013 mol/g, calculate the volume of stomach acid needed to neutralize 1.0 g antacid tablet.

n moles = $0.013 * 1.0 = 0.013$ mol

$V = \frac{n \text{ mole}}{M} = 0.13$ L

2. A 0.333 g-sample of antacid was dissolved in 40.00 mL of 0.135 M HCl solution, then back-titrated to the end-point with 9.28 mL of a 0.0203 M NaOH solution.

a) Calculate number of moles of acid in the original 40.00 mL of HCl?

n moles = $40 * 10^{-3} * 0.135$

$= 5.46 * 10^{-3}$ mol

b) How many moles of base were used in the back titration of excess HCl?

n moles = $0.0203 * 9.28 * 10^{-3}$

$= 1.88 * 10^{-4}$ mol

c) How many moles of excess HCl I?

$1.88 * 10^{-4}$ mol

d) How many moles of HCl reacted with the antacid sample?

$n_A - n_B$

$= 5.21 * 10^{-3}$ mol

Date ....

Pre-Laboratory Questions

1. Explain why cooked fruits and vegetables have lower vitamin C content than fresh fruits and vegetables.

Cooking leaches water solluble vitamin C from the vegetables and high temperatures accelerate it's degradation by air oxidation

2. What are the oxidizing agents in this experiment?

KI , I_3^- IO_3^-

3. What will be the color change of the starch indicator that indicates the end of the titration in this experiment?

blue → Colorless

4. Vitamin C is an acid (ascorbic acid) and a reducing agent. Which property is utilized for this analysis in this experiment?

reducing agent

5. If 39.50 ml of a vegetable juice contains 45% of the recommended daily allowance of vitamin C (equal to 60 mg). How many milliliters of the vegetable juice will provide 100% of the recommended daily allowance?

39.50 mL → 0.45

? → 1

V = 87.70 mL

88 mL

7 Vitamin C Analysis

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

	Trial (I)	Trial (II)
Mass of sample	0.10 g	g
Volume of 0.010 M KIO_3 added	50.0 mL	mL
Moles of IO_3^- added	$5 * 10^{-4}$ mol	mol
Moles of I_3^- generated, total n * 3	$1.5 * 10^{-3}$ mol	mol
Initial buret reading	11.00 mL	mL
Final buret reading	24.00 mL	mL
Volume of 0.10 M $Na_2S_2O_3$ added	16.0 mL	mL
Molar concentration of $Na_2S_2O_3$	0.1 mol/L	mol/L
Moles of $S_2O_3^{2-}$ added	$1.6 * 10^{-3}$ mol	mol
Moles of I_3^- reduced by $S_2O_3^{2-}$ $(\frac{n}{2})$	$8 * 10^{-4}$ mol	mol
Moles of I_3^- reduced by $C_6H_8O_6$	$4 * 10^{-4}$ mol	mol
Moles of $C_6H_8O_6$ in the sample	$4 * 10^{-4}$ mol	mol
Mass of $C_6H_8O_6$ in the sample	0.12 g	g
Mass percent of $C_6H_8O_6$ in the sample	123 %	%
Average percent of $C_6H_8O_6$ in the sample	123 %	

QUESTIONS

1. A 25.0 mL volume of 0.010 M KIO_3, containing an excess of KI, is added to a 0.346 g sample of a lemon solution containing vitamin C. The red brown solution, caused by the presence of excess I_3^-, is titrated to a colorless starch end point with 7.70 mL of 0.100 M $Na_2S_2O_3$.

a. Calculate the moles of I_3^- that are generated from KIO_3.

$n\ KIO_3 = 2.5 * 10^{-4}$ mol

$n\ I_3^- = (2.5 * 10^{-4}) * 3 = 7.5 * 10^{-4}$ mol

b. How many moles of I_3^- reacted with the 0.100 M $Na_2S_2O_3$ in the titration?

$n\ S_2O_3 = 7.7 * 10^{-4}$

$n\ I_3^- = \left(\frac{7.7 * 10^{-4}}{2}\right) = 3.8 * 10^{-4}$ mol

c. How many moles of I_3^- had reacted with the vitamin C in the lemon sample?

$n\ I_3^- = 7.5 * 10^{-4} - 3.8 * 10^{-4}$

$= 3.7 * 10^{-4}$

d. Calculate the moles and grams of vitamin C in the sample.

n moles $= 3.7 * 10^{-4}$ mol

mass $= n * mm = 0.065$ g

e. Calculate the percent (by mass) of vitamin C in the lemon sample.

% mass = 18.8 %

2. If the blue color does not appear when the starch solution is added during the titration, should you continue titrating or discard the sample? Explain.

Discard because this means that the end-point was passed

blue

8 Bleach Analysis

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. Write balanced equation(s) for the reactions involved in standardization of sodium thiosulfate.

$6S_2O_3^{-2} + Cr_2O_7^{-2} + 14H^+ \rightarrow 3S_4O_6^{-2} + 2Cr^{+3} + 7H_2O$

2. In today's chemical analysis of bleach solution, what substance is oxidized by the hypochlorite ion?

Iodide ion (I^-)

3. What are the formulas of sodium thiosulfate and hypochlorite ion?

$Na_2S_2O_3$, ClO^-

4. Does sodium thiosulfate serve as an oxidizing agent or as a reducing agent? What does sodium thiosulfate reduce?

It acts as a reducing agent. It reduces $I_3^- \rightarrow I^-$

8 Bleach Analysis

Name: ....

Section ....

Lab. Instructor ....

Date ....

Results and Calculations

A. Standardization of thiosulfate solution

	Trial (I)		Trial (II)	
Volume of 0.010 M $K_2Cr_2O_7$ solution	25.0	mL	25.0	mL
Initial buret reading	0.00	mL	15.0	mL
Final buret reading	15.0	mL	30.2	mL
Volume of $Na_2S_2O_3$ solution	15.0	mL	15.2	mL
Number of moles of $K_2Cr_2O_7$	$2.5 * 10^{-4}$	mol	$2.5 * 10^{-4}$	mol
Number of moles of $Na_2S_2O_3$ $n * 6$	$1.5 * 10^{-3}$	mol	$1.5 * 10^{-3}$	mol
Molarity of $Na_2S_2O_3$	0.1	M	0.1	M
Average molarity of $Na_2S_2O_3$	0.10			M

B. Analysis of bleach solution

	Trial (I)	Trial (II)
Dilution Factor	20	20
Volume of bleach solution	25.0 mL	25.0 mL
Initial buret reading	21.0 mL	31.2 mL
Final buret reading	31.2 mL	41.1 mL
Volume of $Na_2S_2O_3$ solution	10.2 mL	10.1 mL
Number of moles of $Na_2S_2O_3$ oxidized	$1.02 * 10^{-3}$ mol	$1.01 * 10^{-3}$ mol
Number of moles of ClO^- reduced $(\frac{n}{2})$	$5.10 * 10^{-4}$ mol	$5.05 * 10^{-4}$ mol
Molarity of diluted bleach solution	0.0204 M	0.0202 M
Molarity of original bleach solution	0.408 M	0.404 M
Average molarity of original bleach solution	0.406 M	
Mass% of NaClO(assume density of bleach solution = 1.00g/mL)	3.02 %	

- $M\ dil = \frac{n_{ClO^-}}{V}$

- $M\ original = M\ dil * dilution\ factor$

- $mass\ \% = \frac{M\ org * Mw}{1000 * d} * 100\%$

QUESTIONS

1. A 10.0 mL bleach sample is diluted to 100 mL in a volumetric flask. A 25.0 mL of this solution is analyzed according to the procedure in this experiment. If 12.5 mL of 0.30 M $Na_2S_2O_3$ is needed to reach the end point, calculate the mass percent of NaClO in the original sample? (Assume the density of bleach solution is 1.084 g/mL).

$n\ Na_2S_2O_3 = 3.75 * 10^{-3}$ mol

$n\ ClO = \frac{(3.75 * 10^{-3})}{2} = 1.88 * 10^{-3}$

$M = 0.018$ mol/L

$C_1V_1 = C_2V_2$

$C_1 = 0.76$ mol/L

% mass = 5.22 %

2. A bleach solution that is 6.58% NaClO (density= 1.10 g/mL) was diluted to 0.056 M concentration. Calculate dilution factor.

$\%\ 6.58 = \frac{M * 74.44}{1.10 * 1000} * 100\ \%$

$M_i = 0.97$

$D.F = \frac{M_i}{M_F} = 17.32$

9 Molar Mass of a Volatile Liquid

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. Dumas method assumes the gas behaves ideally, in general, when gases behave ideally?

   At high temperatures and low pressure

2. The vapor from an unknown volatile liquid occupies a 279 mL Erlenmeyer flask at 98.5 °C and 745 torr. The mass of vapor is 0.841 g.

   1 atm → 760 torr
   ?? → 745 torr

   a) What is the molar mass of the volatile liquid?

   $PV = \frac{m}{M} RT$

   $M = \frac{mRT}{PV} = \frac{(0.841)(0.0821)(98.5+273)}{(745/760)(0.279)}$

   $= 93.8$ g/mol

   b) What is the density of vapor at STP?

   T = 273 k
   P = 1 atm

   $PV = nRT$

   $PV = \frac{m}{M} RT$ $d = \frac{PM}{RT} = 4.19$ g/L

   $MP = \frac{m}{V} RT$

9 Molar Mass of a Volatile Liquid

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

	Trial (I)	Trial (II)
Boiling point of water	97.5 ºC	ºC
Atmospheric pressure	682 mmHg	mmHg
Mass of empty flask	78.53 g	g
Mass of flask and condensed vapor	79.18 g	g
Mass of condensed vapor	0.65 g	g
Volume of flask	205 mL	mL
Boiling point of water	370.5 K	K
Atmospheric pressure	0.897 atm	atm
Volume of flask	0.195 L	L
Gas constant (R)	0.0821 L.atm /mol.K	
Molar mass of unknown	113 g/mol	g/mol
Average molar mass	g/mol	

10/10

QUESTIONS

1. If the outside of the flask is not dried after vaporizing the liquid, will the calculated molar mass be too high or too low? Explain.

Too high ; because the mass of the Flask would increase due to the water on the walls of the Flask

2. Consider the following experimental data: Mass of condensed vapor = 0.395 g at 96 °C and 755 mm Hg occupies 137 mL. What is the molar mass of the liquid?

$$M = \frac{mRT}{PV}$$

$$= \frac{0.395(0.0821)(96+273)}{(0.134)(755/760)}$$

$$= 87.9 \text{ g/mol} \simeq 88 \text{ g/mol}$$

10 Thermochemistry and Hess's Law

Name: .... Section ....

Lab. Instructor .... Date ....

Pre-Laboratory Questions

1. What is the definition of the *joule* in terms of the basic SI units?

It is the amount of energy it takes to move an object that weighs one newton up to a height (distanc of one meter. (Kg.m²/s²)

2. In a calorimeter calibration experiment, a sampe of 51.682 g of water at 55.2 °C is added to a calorimeter containing 50.220 g of water at 23.5 °C. After stirring and waiting for the system to equilibrate, the final temperature reached is 36.6 °C. Calculate the calorimeter constant.

$S_{water} = 4.184$

$q_{warm\ water} = q_{cold\ water} + q_{cal}$

$ms\Delta t_{ww} - ms\Delta t_{cw} = C\Delta t_{cal}$

$(51.682)(4.184)(36.6 - 55.2) - (50.220)(4.184)(36.6 - 23.5) = C\ (36$

23

$4022 - 2752 = 13.1\ C$

$C_{cal} = 96.9\ J/C°$

3. Give chemical equations for the reactions that will occur during this Experiment

$NaOH + HCl \rightarrow NaCl + H_2O$

$NH_3 + HCl \rightarrow NH_4Cl$

$NaOH + NH_4Cl \rightarrow NaCl + H_2O + NH_3$

10 Thermochemistry and Hess's Law

Name: ……………………………………… Section ……………………

Lab. Instructor …………………………… Date ……………………

Results and Calculations

A. Determination of a Calorimeter Constant.

	Trial (I)	Trial (II)
Mass (or volume) of cold water	75 g	g
Temperature of cold water	25.0 °C	°C
Mass (or volume) of hot water	75 g	g
Temperature of hot water	70 °C	°C
Final temperature reached	46.0 °C	°C
Increase of cold water temperature (ΔT)	21.0 °C	°C
Decrease of hot water temperature (ΔT)	24.0 °C	°C
Heat lost by hot water	7531.2 J	J
Heat gained by cold water	6590 J	J
Heat gained by calorimeter water	941.2 J	J
Calorimeter constant (C)	44.8 J/°C	J/°C
Average value of calorimeter constant		J/°C

$-q_{warm} = q_{cold} + q_{cal}$

$7531.2 - 6590 = C\Delta t$

B. Heat of Acid/Base Reactions

1. HCl/NaOH

	Trial (I)	Trial (II)
Volume of 1.0 M NaOH used	50 mL	mL
Initial temperature of NaOH	22.0 °C	°C
Volume of 1.0 M HCl used	50 mL	mL
Initial temperature of HCl	22.0 °C	°C
Final temperature reached	27.5 °C	°C
Total volume of mixture	100 mL	mL
Total mass of mixture	100 g	g
Temperature change, Δt	5.5 °C	°C
Average value of the Calorimeter constant (C)	44.8	J/°C
Heat gained by the solution* msΔt	2239 J	J
Heat gained by the calorimeter CΔt	246 J	J
Heat of the reaction	-2485 J	J
Moles of NaOH reacted CV	0.05 mol	mol
Moles of HCl reacted CV	0.05 mol	mol
Moles of water produced	0.05 mol	mol
ΔH $\frac{2485}{0.05}$	49.7 kJ/mol H_2O	kJ/mol H_2O
Average value of ΔH	49.7	kJ/mol H_2O
Literature value of ΔH		- 55.9 kJ/mol H_2O

* Assume density of solution = 1.00 g/ml

2. HCl/NH_3

	Trial (I)	Trial (II)
Volume of 1.0 M NH_3 used	50.0 mL	mL
Initial temperature of NH_3	22.0 °C	°C
Volume of 1.0 M HCl used	50.0 mL	mL
Initial temperature of HCl	22.0 °C	°C
Final temperature reached	27.0 °C	°C
Total volume of mixture	100 mL	mL
Total mass of mixture	100 g	g
Temperature change, Δt	5.0 °C	°C
Average value of the Calorimeter constant (C)	44.8 J/°C	
Heat gained by the solution* ms Δt	2035 J	J
Heat gained by the calorimeter	224 J	J
Heat of the reaction	-2259 J	J
Moles of NH_3 reacted	0.05 mol	mol
Moles of HCl reacted	0.05 mol	mol
Moles of NH_4Cl produced	0.05 mol	mol
ΔH	45.2 kJ/mol	kJ/mol
Average value of ΔH	45.2 kJ/mol	
ΔH for the reaction: $NaOH + NH_4Cl \rightarrow NaCl + NH_3 + H_2O$	4.5 kJ	

*** Assume density of solution = 1.00 g/mL and specific heat of solution = 4.07 J/g. °C.**

Questions

1. What effect on the calculated calorimeter constant would be observed if the calorimeter cup were made of conducting material (such as metal) rather than plastic foam?

It would increase

2. Why is water typically used as the heat-absorbing liquid in calorimeter?

Because water has a very high specific heat capacity than other liquids

3. Given that:

$$H_2O_{2(\ell)} \rightarrow H_2O_{(\ell)} + \tfrac{1}{2} O_{2(g)} \quad \Delta H = -96.0 \text{ kJ/mol}$$

$$\left(2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(\ell)} \quad \Delta H = -571.6 \text{ kJ/mol} \right) \div 2$$

Use Hess's law to calculate ΔH for the reaction:

$$H_{2(g)} + O_{2(g)} \rightarrow H_2O_{2(g)}$$

$H_2 + \tfrac{1}{2}O_2 \rightarrow \cancel{H_2O}$ $\Delta H = -285.8$ KJ/mol

$\cancel{H_2O} + \tfrac{1}{2}O_2 \rightarrow H_2O_2$ $\Delta H = 96.0$ KJ/mol

$H_2 + O_2 \rightarrow H_2O_2$ $\Delta H = -189.8$ KJ/mol

11 Molar Mass from Freezing Point Depression

Name: ....

Section ....

Lab. Instructor ....

Date ....

Pre-Laboratory Questions

1. Define what we mean by the colligative property.

Colligative property :- are properties of solution that depend on the number of molecules (of solute) in a given volume of solvent and not on the properties - identity (size/mass) of the molecules

2. A 0.36-g sample of an unknown substance was dissolved in 30.0 mL of cyclohexane. The density of cyclohexane is 0.779 g/mL. The freezing-point depression was 2.50 °C. Calculate the molar mass of the unknown substance. [k_f (cyclohexane) = 20.0 °C/m]

①

$\Delta T_f = K_f . m$

$2.5 = 20 m$

$m = 0.125$ mol/kg

③

$m = \frac{n}{mass}$

$n = 0.125 * 23.37 * 10^{-3}$

$= 2.921 * 10^{-3}$ mol

②

$d = \frac{m}{V}$

$m = 0.779 * 30$

$= 23.37g$

$= 23.37 * 10^{-3}$ kg

④

$n = \frac{m}{M}$

$M = \frac{0.36}{2.921 * 10^{-3}}$

$= 123.2$

1.2×10^{2} g/mol

11 Molar Mass from Freezing Point Depression

Name: .... Section ....

Lab. Instructor .... Date ....

Results and Calculations

A. Freezing point of cyclohexane

	Trial (I)	Trial (II)
Mass of empty small tube	29.84 g	29.84 g
Freezing point of cyclohexane	7.0 ± 0.2 °C	7.0 ± 0.2 °C
Average value of freezing point	7.0 °C	

$$M_\omega = \frac{K_f * m_{sollute}^{\to g}}{\text{mass of solvent (kg)} * \Delta T_f}$$

B. Freezing point of solution

26

	Data
Mass of empty small tube	29.84 g
Mass of solute	0.21 g
Mass of empty small tube + cyclohexane	37.49 g
Mass of cyclohexane	7.65 g
First freezing point	4.0 °C
$(\Delta T_f)_1$ (7.0 - 4.0)	3.0 °C
Molar mass of solute	183 g/mol
Mass of second sample of solute	0.15 g
Total mass of solute	0.36 g
Second freezing point	3.0 °C
$(\Delta T_f)_2$	4.0 °C
Molar mass of solute	137 g/mol
Average molar mass of solute	160 g/mol

$\frac{9.5}{10}$

QUESTIONS

1. What would be the effect of each of the following on the calculated molar mass of the solute?

   a) Some cyclohexane evaporated while the freezing point of pure cyclohexane was being measured.

   ....no effect..

   b) Some cyclohexane evaporated after the solute was added

   ....would in~~d~~ecrease..

   c) A foreign solute was already present in the cyclohexane.

   ....no effect..

   d) The thermometer is not calibrated correctly. It gives a temperature that is 1.5 °C too low at all temperature.

   ....no effect..