Name: ________________________________ Section: ____________________

Data and Calculations for Experiment 1

Be sure to record all numerical values with the correct number of significan figures and note the units are already recorded for each in this first experiment!

A. Temperature
1. Water at room temperature ______________ ^{o}C

2. Boiling point ______________ ^{o}C

3. Ice water
   Unstirred ______________ ^{o}C

   Stirred ______________ ^{o}C

4. Ice water with salt added ______________ ^{o}C

B. Mass
1. 100 mL beaker ________________ g

2. 250 mL Erlenmeyer flask ________________ g

3. Weighing boat ________________ g

4. Mass of weighing boat + sodium chloride ________________ g

   Mass of sodium chloride (show calculation setup) ________________ g

C. Length
1. Length of ⟵⟶ ______________ cm

2. Height of 250 mL beaker ______________ cm

3. Length of test tube ______________ cm

D. Volume

1. 200 mL mark (from Erlenmeyer flask) water
   transferred to graduated cylinder ______________ mL

2. Height of 5.0 mL of water in test tube ______________ cm

3. Height of 10.0 mL of water in test tube ______________ cm

Name: ____________________ Section: ____________________

Data and Calculations for Experiment 2

Data Sheet for Density of an Element sample

Name of Element: ____________________

Sample #	Object Mass (g)	Initial mL H_2O	mL H_2O w/ Object	Volume object (mL)	Density (g / mL)	Cumulative Sample #s	*Using Microsoft Excel®, graph the following:* Cumulative volume (mL) (x-axis)	Cumulative object mass (g) (y-axis)
1	____	____	____	____	____	1	____	____
2	____	____	____	____	____	1 + 2	____	____
3	____	____	____	____	____	1 + 2 + 3	____	____
4	____	____	____	____	____	1 + 2 + 3 + 4	____	____
5	____	____	____	____	____	1 + 2 + 3 + 4 + 5	____	____

Average Density from Table = ________

Average Density from Graph (slope of line) = ________

Be sure to show your properly formatted graph on the laptop or your device to your instructor to receive credit for this part of the experiment OR they may instruct you to take a photo of your graph to submit with this report to Canvas.

Name: ______________________________ Section: ____________________

Questions

1. Which would work better in this experiment as an unknown solid whose density is to be determined, wood chips or small quartz rocks? Explain your choice.

2. Why is it best to use a smaller graduated cylinder as opposed to a larger graduated cylinder for this experiment?

3. How well does the average density from the table and density from the slope of the graph compare? Which value is closer to the accepted density of your metal in the table below (from the *CRC Handbook of Chemistry and Physics*)? Calculate the percent error between your better value and the accepted value.

Element	Accepted Density (g/mL)
Aluminum, Al	2.70
Copper, Cu	8.96
Lead, Pb	11.34
Zinc, Zn	7.14

4. What is the density of a 9.343 gram piece of metal that causes the level of water in a graduated cylinder to rise from 5.1 to 8.1 mL when the metal is submerged in the water? Consider significant figures when doing the calculation.

Name: ___________________________________ Section: ______________________

Data and Calculations for Experiment 3

Use your graphs to answer the following questions. Note: The accepted freezing point of acetic acid, according to the *CRC Handbook of Chemistry and Physics*, is: 16.6 °C

1. a. What is the experimental freezing point of acetic acid in Trial 1? ________________

   b. What is the experimental freezing point of acetic acid in Trial 2? ________________

   c. Calculate the percent error in the Trial 1 measurement of the freezing point:

   SHOW CALCULATION:

   ____

2. What is supercooling? Did you observe supercooling in your experiment? Explain.

3. What is the difference between melting and freezing a substance?

Name: ___________________________________ Section: ______________________

Data for Experiment 4

Record your observations for each combination below. If a reaction occurs, write balanced MOLECULAR, IONIC, and NET-IONIC equations. If no reaction occurs, write NR. Make sure to include the physical states of all the products.

1. $NaCl(aq)$ and $KNO_3(aq)$

   Observations:

   Molecular:

   Ionic:

   Net-Ionic:

2. $NaCl(aq)$ and $AgNO_3(aq)$

   Observations:

   Molecular:

   Ionic:

   Net-Ionic:

3. $Na_2CO_3(aq)$ and $HCl(aq)$

   Observations:

   Molecular:

   Product of *secondary* reaction:

4. $NaOH(aq)$ and $HCl(aq)$

   Observations:

   Molecular:

   Ionic:

   Net-Ionic:

Name: ___________________________________ Section: ______________________

5. $BaCl_2$(aq) and H_2SO_4(aq)

Observations:

Molecular:

Ionic:

Net-Ionic:

6. NH_4OH(aq) and H_2SO_4(aq)

Observations:

Molecular:

Ionic:

Net-Ionic:

7. $CuSO_4$(aq) and $Zn(NO_3)_2$(aq)

Observations:

Molecular:

Ionic:

Net-Ionic:

8. Na_2CO_3(aq) and $CaCl_2$(aq)

Observations:

Molecular:

Ionic:

Net-Ionic:

Name: ______________________________ Section: ____________________

9. $CuSO_4$(aq) and NH_4Cl(aq)

<u>Observations</u>:

<u>Molecular</u>:

<u>Ionic</u>:

<u>Net-Ionic</u>:

10. NaOH(aq) and HNO_3(aq)

<u>Observations</u>:

<u>Molecular</u>:

<u>Ionic</u>:

<u>Net-Ionic</u>:

11. $FeCl_3$(aq) and NH_4OH(aq)

<u>Observations</u>:

<u>Molecular</u>:

<u>Ionic</u>:

<u>Net-Ionic</u>:

Name: ________________________________ Section: ____________________

Data and Calculations for Experiment 5

Be sure to record all masses to 0.001g and include the units on all numerical values!

Mass of $CuSO_4$/sand mixture ______________

Mass of empty evaporating dish ______________

AFTER separation and drying both solids:

Mass of evaporating dish and dry $CuSO_4$ ______________

Mass of dry $CuSO_4$ ______________

Mass of dry sand ______________

CALCULATIONS:

Total mass of recovered sand and $CuSO_4$ ______________

Calculated total percent recovery ______________

Show Calculation

Percent by mass of $CuSO_4$:

Show Calculation ______________

Percent by mass of sand:

Show Calculation ______________

Name: ________________________________ Section: ____________________

Questions

1. Many students do NOT recover 100% of the original mixture. Describe at least TWO possible problems that could cause LESS than 100% recovery of the mixture.

2. A student obtained the following data:

Mass of beaker	25.87 g
Mass of beaker with mixture sample	28.12 g
Mass of evaporating dish	146.36 g
Mass of evaporating dish with dried salt	147.10 g
Mass of beaker with dried sand	???

However, this student spills her sand sample out of the beaker before weighing it. If the student believes in the Law of Conservation of Mass, what should have been the weight of the beaker with the dried sand in it? Show all your work.

3. A student receives a sample of a mixture with three components: (1) solid iodine that is first removed from the mixture by evaporation, (2) solid salt that is dissolved to separate it from the third component, and (3) solid sand. The salt and sand are dried and weighed, but the iodine escapes as a gas and is not recovered. The student starts with 4.25 g of the mixture and recovers 1.16 g of salt and 2.40 g of sand. What is the mass percent of each component in the original mixture? Show all your work.

Name: ________________________________ Section: ____________________

Data and Calculations for Experiment 6

Be sure to record all numerical values to the correct significant figures and include units!

Identification of metal __________

Mass of DRY metal __________

Mass empty calorimeter __________

Mass calorimeter with water __________

Mass water __________

Initial temperature of water in calorimeter __________

Initial temperature of hot metal
(before adding it to calorimeter) __________

Final temperature of water + metal
in calorimeter __________

ΔT_{water} __________

ΔT_{metal} __________

1. Calculate the experimental specific heat of the metal. Show Calculation and units:

2. Find the accepted value for the specific heat of your metal in the table below (from the *CRC Handbook of Chemistry and Physics*). Calculate the % error of your experimental specific heat.

Metal	Accepted s_{metal} (J/g °C)
Al	0.903
Cu	0.385
Zn	0.387

Name: ________________________________ Section: ____________________

Data and Calculations for Experiment 7

A. Quantitative Determination of Percent Composition

1. When solid $KClO_3$ is heated above 400 °C, it decomposes to solid potassium chloride and elemental oxygen gas. Write the balanced equation for the decomposition of $KClO_3$ solid.

2. What is the remaining residue in the crucible after heating?

3. What substance is lost during the heating?

Be sure to record all numerical values to the correct significant figures and include units!

4. Mass of crucible and cover ____________
5. Mass of crucible, cover and sample ____________
6. Mass of crucible, cover and sample after heating ____________
7. Mass of original sample ____________
8. Mass of the residue ____________
9. Mass lost upon heating ____________

10. Experimental percentage of KCl in the $KClO_3$ sample.

11. Experimental percentage of oxygen in the $KClO_3$ sample.

Name: ___________________________ Section: __________________

12. Using the atomic masses from the periodic table, solve for the molar mass of $KClO_3$.

____

13. Theoretical percentage of KCl in the $KClO_3$ sample ______________

14. Theoretical percentage of oxygen in the $KClO_3$ sample ______________

15. Percent error in oxygen determination

B. Qualitative Examination of the Residue

1. Record what you observed when $AgNO_3$ solution was added to the following:
   i. KCl
   ii. $KClO_3$
   iii. Residue
2. What does the evidence lead you to believe about the residue?
3. Does the evidence from the $AgNO_3$ test prove conclusively (without a doubt) that the residue is KCl? Explain.

Name: ________________________________ Section: ____________________

Experiment 8 – Precipitation of Strontium Sulfate

In this experiment, you will study a precipitation reaction between sodium sulfate and strontium chloride. You will collect, dry, and weigh the precipitate and compare this experimental yield to the theoretical yield.

Procedure

Weigh a clean, dry, 100-mL beaker. Add about 0.25 g (0.350 g max!) of solid sodium sulfate to the beaker and weigh it again. Completely dissolve the sodium sulfate in about 20 mL of D.I. water. Add 5.0 mL of 0.50 M strontium chloride solution.

Putting the beaker in an ice-cold water bath. Your precipitate should settle to the bottom, leaving a relatively clear solution above it. Obtain a piece of filter paper and weigh it on the analytical balance.

Set up a vacuum filtration apparatus with a Büchner funnel and your weighed filter paper (your instructor will show you how). Using a stirring rod to guide the stream of liquid, pour the contents of the beaker into the Büchner funnel. Use your wash bottle (filled with D.I. water) to rinse any solid out of the beaker and into the filter. Make sure no precipitate remains in the beaker or on the stirring rod. Fill the beaker with 15 mL of D.I. water, swirl it around, and then pour it into the filter. Repeat the washing process, and then draw air through the funnel for a few minutes to help dry the crystals.

Turn off the vacuum, carefully remove the filter paper containing your precipitate with a spatula, and place it over a watch glass. Fill a 100-mL beaker half-way with water, place the watch glass with filter paper over the beaker, and heat to boil for twenty minutes to dry the precipitate (alternatively, you can place the watch glass with filter paper in a drying oven at 130 °C for twenty minutes). Allow to cool, then determine the mass of your precipitate.

Data and Calculations for Experiment 8

Be sure to record all numerical values to the correct significant figures and include units!

1. Weight of empty beaker ____________________

2. Weight of beaker and sodium sulfate ____________________

3. Weight of sodium sulfate ____________________

   Show Calculation

Name: ______________________________ Section: ____________________

4. Moles of sodium sulfate:

Show Calculation ____________________

5. Moles of strontium chloride

$$\text{moles } SrCl_2 = 5.0 \text{ mL } SrCl_2 \left(\frac{10^{-3} \text{ L } SrCl_2 \text{ solution}}{1 \text{ mL } SrCl_2 \text{ solution}} \right) \left(\frac{0.50 \text{ mol } SrCl_2}{1 \text{ L } SrCl_2 \text{ solution}} \right) =$$

Solve the Equation Shown ____________________

6. Write a balanced MOLECULAR equation for the reaction:

7. Write a balanced NET-IONIC equation for the reaction:

8. Weight of empty filter paper ____________________

9. Weight of filter paper and dried precipitate ____________________

10. Weight of dried precipitate (Actual Yield) ____________________

11. Determine the limiting reactant and excess reactant for your reaction.

Limiting Reactant: ____________________ Excess Reactant: ____________________

12. Calculate the Theoretical Yield (in grams) of strontium sulfate.

Show Calculation ____________________

Name: ________________________________ Section: ____________________

13. Determine the percentage yield of your reaction.

Show Calculation ____________________

14. Briefly describe how you could have improved your percentage yield in this experiment.

Name: ______________________________ Section: ____________________

Data and Calculations for Experiment 9

A. Qualitative Determination of the Released Liquid

1. Record observations regarding the solid before, during, and after heating the copper(II) sulfate pentahydrate.

2. Compare and record observations after adding liquid to the anhydrous cobalt(II) chloride test strips.

3. Compare and record observations after adding liquid to the residue on the watch glass.

4. What conclusions can you draw from the above observations?

5. Write the balanced chemical equation for the decomposition of copper(II) sulfate pentahydrate, include phases.

B. Quantitative Determination of Mass Lost in a Hydrate

Sample number: ____________________

1. Mass of crucible and cover ____________________
2. Mass of crucible, cover and sample ____________________
3. Mass of crucible, cover and sample after heating ____________________
4. Mass of sample after final heating ____________________
5. Mass of original sample ____________________
6. Total mass lost by sample ____________________

7. Percentage of water in sample ____________________

Name: ______________________________ Section: ____________________

8. Ask your instructor for the name of the anhydrous salt of your residue and solve for the formula and name of your original unknown hydrate.

Formula: ____________________ Name: ______________________________

9. Is it possible that the decrease in mass from heating is something other than water?

Yes or *No* Explain and include an example.

Name: ________________________________ Section: ____________________

Experiment 10 – Atomic Spectra

The purpose of this experiment is to show that different elements give off unique colors of light when atoms of the elements are excited by heating. By identifying the unique colors the element can be identified.

Part I

There will be three gas discharge tubes set up in the lab. Observe the color of light given off by each discharge tube. Record the colors in the data table below. After recording the color of the light, observe the light through the diffraction grating (look off to one side). Draw a picture of the spectral lines. Identify the color of each line.

1. Element __________________________ Color of light _______________________

   Spectral diagram:

   Violet Red

2. Element __________________________ Color of light _______________________

   Spectral diagram:

   Violet Red

3. Element __________________________ Color of light _______________________

   Spectral diagram:

   Violet Red

Name: ________________________________ Section: ____________________

Part II-A:

There will be seven containers with wooden splints soaking in salt solutions. Each solution will be labeled with its chemical name.

Procedure

Light the bunsen burner and adjust its flame until you see a blue inner cone. Use crucible tongs to remove a splint from the soaking solution and place it in the flame. Observe the color of the flame and record the color in the table below. If two solutions give colors which seem similar, repeat the experiment until you can notice the differences in color well enough that you can describe the differences. Always use the same burner with a given chemical. If you mix up burners, you can contaminate the colors.

Chemical	Color of Flame
lithium chloride	
calcium chloride	
potassium chloride	
copper(II) chloride	
strontium chloride	
sodium chloride	
barium chloride	

Part II-B.

There will be seven containers having unknown chemicals. Repeat the procedure that you used for known chemicals and identify the unknown chemicals. Record results below.

Unknown #	Color of Flame	Chemical
1		
2		
3		
4		
5		
6		
7		

Name: ________________________________ Section: ____________________

Formula	Number of valence electrons	Lewis dot structure	VSEPR structure	Molecular Geometry	Bond angle on central atom(s)	Polar? Nonpolar?
I_2						
H_2S						
PBr_3						
CS_2						

Name: ________________________________ Section: ____________________

Formula	Number of valence electrons	Lewis dot structure	VSEPR structure	Molecular Geometry	Bond angle on central atom(s)	Polar? Nonpolar?
$CHCl_3$						
PO_3^{-3}						
PO_4^{-3}						
CH_2O						

Name: ______________________________ Section: ____________________

Formula	Number of valence electrons	Lewis dot structure	VSEPR structure	Molecular Geometry	Bond angle on central atom(s)	Polar? Nonpolar?
SO_3						
SO_3^{-2}						
SO_4^{-2}						
NO_2^{-}						

Name: ___________________________________ Section: ______________________

Formula	Number of valence electrons	Lewis dot structure	VSEPR structure	Molecular Geometry	Bond angle on central atom(s)	Polar? Nonpolar?
CH_2Cl_2						
CO_2						
PH_3						
O_2						

Name: ______________________________ Section: ____________________

Data and Calculations for Experiment 12

Pressure of the air in the room: ________________________________

Temperature of the air in the room: ________________________________

Actual Volume (mL)	**1 / Volume (mL^{-1})**	**Pressure * Vol. = *k* (mmHg·mL)**	**Plunger Position**	**Pressure (mmHg)**
			5	
			6	
			7	
			8	
			9	
			10	
			11	
			12	
			13	
			14	
			15	
			16	
			17	
			18	
			19	
			20	

These will be calculated in Excel®. Submit your spreadsheet and your graphs.

Average *k* = __________

Be sure to show your 2 properly formatted graphs on the laptop or your device to your instructor to receive credit for this part of the experiment OR they may instruct you to take a photo of your graphs to submit with this report to Canvas.

Name: ________________________________ Section: ____________________

Questions

1) On your linear graph, do any points deviate from the straight line?

2) Write down the equation of the trendline ($y = mx + b$) from your linear graph. How does the slope (m) compare to the average $P*V=k$ value from the table of data?

3) Using the equation of your trendline, solve for the pressure at a volume of 2.0 mL. *Hint:* $x = 1/V$ in your equation!

4) Why must the temperature be constant during this experiment? Use observations from your experiment and the graphs to support your answer!

5) If you repeated this experiment at a higher temperature, how would the P vs. V curve obtained differ from the curve on your 1st graph?

6) You have a 1.00 L sample of Argon gas at 700.0 mmHg. You decrease the pressure to 500.0 mmHg. What is the new volume?

7) Describe (quantitatively) what you would do to the volume of a container of gas if you wanted to double the pressure inside.

Name: ______________________________ Section: ____________________

Data and Calculations for Experiment 13

A. Relative Solubility of a Solute in Two Solvents

1. a) Which liquid is denser, decane or water? ______________

   b) How did you decide which layer was water?

2. What is the color of iodine in water? ______________

   What is the color of iodine in decane? ______________

3. Which solvent dissolves more iodine? How did you decide this?

B. Miscibility of Liquids

1. Which liquids were miscible with each other?

2. Which liquids were immiscible with each other?

C. Ionic Reactions in Solution

1. Write the formulas for the following substances. Include states of matter (e.g. $_{(aq)}$ or $_{(s)}$) based on the results of your experiment:

   barium sulfate ______________________

   barium chloride ______________________

   sodium sulfate ______________________

   sodium chloride ______________________

2. Write the equation that shows the reaction of barium chloride and sodium sulfate. Use state indicators (e.g. $_{(aq)}$ or $_{(s)}$) for all compounds.

3. Which compound is the white precipitate? How do you know this?

Name: ________________________________ Section: ____________________

Data and Calculations for Experiment 14

Be sure to record all numerical values to the correct significant figures and include units!

Mass of flask and KHP	
Mass of empty flask	
Mass of KHP	
Initial buret reading	
Final buret reading	
Volume of base used	

CALCULATIONS – Show your work!

1. Moles of acid (KHP, Molar mass = 204.2)

2. Moles of base used to neutralize acid

3. Molarity of base (NaOH)

4. *Class Average* Molarity of Base:

Name: ___________________________________ Section: ______________________

Observations for Experiment 15

Place an "X" on the label that properly describes each compound below:

	Non-Electrolyte	Strong Electrolyte	Weak Electrolyte
1. Tap water			
2. Deionized water			
3. Sugar solution			
4. NaCl solution			
5a. Pure (glacial) acetic acid			
5b. Diluted acetic acid			
5c. Twice diluted acetic acid			
6a. 1 M acetic acid			
6b. 1 M HCl			
6c. 1 M NH_4OH			
6d. 1 M NaOH			
7a. $NaNO_3$			
7b. NaBr			
7c. $Ni(NO_3)_2$			
7d. $CuSO_4$			
7e. NH_4Cl			

1. In 5a. to 5c., what happens to the glacial acetic acid as it is diluted? How does this explain the changes in light intensity?

2. In the final demonstrate, write a balanced chemical reaction (with states) for the mixing of aqueous solutions of barium hydroxide and sulfuric acid.

3. What type of reaction is this? Hint: More than 1 term may apply

4. Explain why the light becomes dimmer as two strong electrolytes are mixed with each other.

5. Why does the light come back on after more of the electrolyte is added?