CHE116

Experiments and Worksheets
Part 1: E & W 1 - 7

Fall 2024
Spring 2025
Summer 2025

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Experiment 1: Density of Metals

This experiment was adapted from Colby College, CHE141,
Experiment 1: Scientific Measurements and Introduction to Excel

Purpose: To become familiar with the lab room and lab equipment by determining the density of metals and practicing pipetting.

Background: The determination of density for any sample is a good introduction of basic laboratory techniques. For the determination of the density of various metals, one would have to use a laboratory balance and graduated cylinder. The proper use of each of these is expected of all STEM students.

This lab involves the density determination of copper, nickel, and brass. The percent composition of the brass will be calculated. Each of these samples are supplied by RCBC. Please return all samples, clean and dry.

The mass of each sample will be determined with the lab balance. The volume of each sample will be determined by the volume by difference approach. Please make sure to look over the video in Blackboard regarding the proper use of a laboratory balance, and the video regarding the volume by difference approach.

Chemicals:

Copper metal
Zinc metal
Brass alloy
Deionized water

Equipment:

10 mL graduated cylinder graduated cylinder
Pipet, 10 mL, & pipettor
Lab balance
Beakers (50 mL or 100 mL)

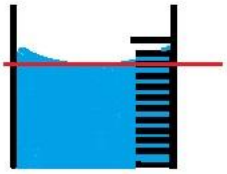
Procedure:

Part A Density of Metals

1. Obtain your samples.
   a) You will need enough of each metal to give a large enough volume change. So, what is large enough? To get a good number of significant figures, without having to use an excessive amount of sample, make sure the volume change is over 1 mL. For copper, zinc, and brass, that would be approximately 10 grams of each.
   b) You will need to transport your samples in some type of container. A clean, dry, beaker for each metal will work well.

2. Determine the mass of your first sample. (Work with one sample at a time.)
   a) Use a clean, dry 50 mL beaker.
   b) Tare the beaker on a lab balance so the mass reads 0.000 g.
   c) Remove the beaker from the balance, and add your metal sample.
   d) Place the beaker with the metal sample on the balance and record the mass of the metal sample. Record the mass directly in your lab worksheet, not on a piece of scrap paper. Mention in your worksheet that the balance was tared.
   e) Return to your lab bench so other students can use the balance.

3. Determine the volume of your first sample.
   a) Use a 10 mL graduated cylinder.
   b) Add approximately 6 mL of DI water to the cylinder. Record the volume of water with the correct number of significant figures.



Read at the bottom of the meniscus, at eye level.

   c) Add the metal sample to the graduated cylinder, gently. Use all of your sample that gave the recorded mass.
   d) Record the volume of water and metal now in the graduated cylinder. Just like in step 3b, use the correct number of significant figures in your volume reading.

4. Pour the water and metal from the graduated cylinder, then repeat steps 2 and 3 for your other metal samples.
5. Please return your metal samples to the designated area, clean and dry. Be careful not to let the metal go down the sink drain.
6. Consult page 6 for calculation instructions.

Part B Practice Pipetting

1. Obtain approximately 20 mL of deionized water in a small beaker.
2. Use the 10 mL pipet and pipettor to pipet the water into a 10 mL graduated cylinder.
3. Practice pipetting 5.0 mL of water into the graduated cylinder. Do this 3 times; reuse your water for each trial. Each person in the group should do the pipetting three times.
Read the volume of water in the graduated cylinder for each trial.
4. Organize your data in a table in your data section. You do not need your lab partners data. Calculate the average volume of water that was in the graduated cylinder.

Part C Practice Reading Volumes with a Graduated Cylinder

1. Locate the prefilled graduated cylinders in the lab.
2. Read the volume of liquid in each cylinder, use the correct significant figures and units. Write each volume in your notebook.
3. Compare the volume you read to the volume on the given answer key. Record the given answer in your notebook.

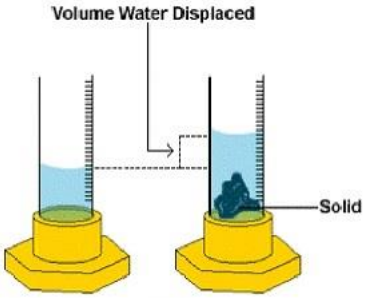
(continued on the next page)

Calculations for Part A:

1. Calculate the volume and density for each sample.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad \boxed{d = \frac{m}{v}}$$

   a) Report each density with the correct number of significant figures, and organized in a table.
   b) Show your calculations below the table.



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2. Calculate the percent composition of the brass as % Cu and % Zn.
   a) Copper and zinc are the only elements in brass. So, the % Cu and % Zn will add to 100%.
   b) Use the density you calculated for your sample of brass.
   c) Follow the the math equations below, so you are solving for only one unknown at a time.

$$\text{density}_{\text{Brass}} = \frac{p\,d_{Cu} + q\,d_{Zn}}{100}$$

p = the % Cu by mass and q = the % Zn by mass

You need to solve for p and q.

Rearrange to solve for one unknown at a time.

$$p + q = 100\ \%$$

$$q = (100 - p)$$

rewrite as

$$\text{density}_{\text{Brass}} = \frac{p\,d_{Cu} + (100 - p)\,d_{Zn}}{100}$$ Solve for p

$$q = (100 - p)$$ Solve for q

Name: _____

RCBC CHE116 Notebook Worksheet for Experiment 1: Density of Metals

Section 1

Experiment work was performed on _____.

Lab Partner: _____.

Section 2 Raw Data

Part A Density of Metals

Copper Sample

Mass of metal in tared beaker: _____

Volume of water in the 10 mL graduated cylinder: _____

Volume of water and metal in the graduated cylinder: _____

Zinc Sample

Mass of metal in tared beaker: _____

Volume of water in the 10 mL graduated cylinder: _____

Volume of water and metal in the graduated cylinder: _____

Brass Sample

Mass of metal in tared beaker: _____

Volume of water in the 10 mL graduated cylinder: _____

Volume of water and metal in the graduated cylinder: _____

Part B Pipet Practice

Volume reading for the 5.0 mL of water pipetted into the 10 mL graduated cylinder:

Trial 1: _____ mL	Trial 2: _____ mL	Trial 3: _____ mL

Part C Reading the Volume in a Graduated Cylinder

Size of Graduated Cylinder	Your Volume Reading (include units)	Answer Key Volume (mL)

Section 3: Experimental Procedure

Part A: A 50 mL beaker was tared on the lab balance. The _____ metal was put into the beaker and the mass was recorded. The volume of this metal sample was then determined by mass by difference. The volume of only water was measured in a _____ mL graduated cylinder. The _____ metal sample was then put into the graduated cylinder, and the volume of _____ and _____ was measured. These steps were repeated for the other metals: _____ and _____.

Part B: A 10 mL pipet and pipettor were used to dispense _____ mL of deionized water into a _____ mL graduated cylinder. The volume of the water in the graduated cylinder, as read from the graduated cylinder graduation marks, was recorded. The pipetting and volume reading were repeated _____ more times for a total of _____ readings.

Part C: Graduated cylinders were prefilled with water by the lab instructor. The volume of water in each cylinder was recorded. The student-read volume was compared to the answer-key volume.

Section 4: Data Tables

Part A (show units)

Sample	Mass in tared beaker	Volume of only water	Volume of water & metal	Volume of metal
Copper				
Zinc				
Brass				

Part B (show units)

Trial Number	Volume of water pipetted into cylinder	Volume of water read from cylinder graduation marks
1		
2		
3		

Part C (show units)

Cylinder Size	Student-read volume of water	Answer Key volume of water

Section 5: Calculations & Results

Part A: Show the calculation for the volume of the metal and for the density of the metal.

Volume of Copper:
Density of Copper
Volume of Zinc:
Density of Zinc:
Volume of Brass:
Density of Brass:

% Composition of Brass, reported as % Cu and % Zn: (show calculations and final answers in this box)

Part B: Calculate the average volume of H_2O in the graduated cylinder.

Section 6: Conclusion

Part A: How did your experimental density for copper and for zinc compare to the literature density for copper and zinc (list the literature densities)?

Part B: How did your graduated cylinder volumes compare to the pipetted volumes?

Part C: How did your volumes compare to the answer key volumes?

Experiment 2: Laboratory Measurements

This experiment was written with information from the following sources:
Santa Monica College, 1: Measurements in the Laboratory (Experiment)
CC-BY Torres & González-Urbina, Volumetric Experiments

Purpose: People working in a lab need to know how to use the instruments and glassware, how to organize their data, perform their calculations, and properly report their results. This week's experiment provides experience with all of these aspects of lab work.

Background: There are several different parts in this Laboratory Measurements experiment.

(A) General Information: This area discusses significant figures, accuracy and precision, volume by difference, and mass by difference.

(B) Determination of Density: Unknown Liquid: This involves the measurement of both the mass and volume of the unknown liquid sample. The volume is determined by use of a graduated cylinder, and the mass is determined by use of the lab balance and mass by difference. Density is calculated using the mass and volume data.

(C) Determination of Solubility: Unknown Liquid: This experiment has the unknown liquid mixed into three different liquid solvents. An observation is made as to whether or not the unknown liquid dissolves in each solvent.

(D) Determination of Solubility: Unknown Solid: This experiment has a few crystals of solid mixed into three different liquid solvents. An observation is made as to whether or not the solid dissolves in each solvent.

(E) Determination of Density: Pennies: This involves the measurement of both the mass and volume of a sample of 10 pennies (either pre-1982 or post-1982 pennies). The volume is determined by volume by difference, and the mass is determined by use of the lab balance and mass by difference. Density is calculated using the mass and volume data. The % composition of the pennies is then calculated.

(A) General Information:

Significant Figures: The rules for identifying significant figures in written values and working with them in calculations are given in chapter 1 of your CHE115 textbook. The rules are summarized below:

1. All non-zero digits are significant.
2. Zeros at the beginning of a decimal number (these are called leading zeros) are not significant: 0.0123 has 3 SF.
3. Terminal zeros at the right of the decimal point (these are called trailing zeros) are significant: 0.120 has 3 SF
4. Terminal zeros in a whole number (a number without a decimal point shown) are not significant: 1,230 has 3 SF.
5. For a number in scientific notation, all numbers before the power of ten are significant: 1.50×10^{-5} has 3 SF.

The number of significant figures in a measured property is determined by the tool used for the measurement. When the data is used in subsequent calculations the uncertainty must be carried on with the calculated results. The following two rules hold when determining the correct number of significant figures to report for the answer to a calculation:

1. When adding or subtracting measured quantities, give the same number of decimal places in the answer as there are in the measurement with the least number of decimal places.

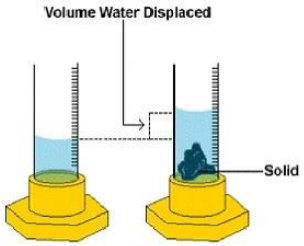
2. When multiplying or dividing measured quantities, give as many significant figures in the answer as there are in the measurement with the least number of significant figures.

Accuracy & Precision: The definitions for accuracy and precision are given in chapter 1 of your CHE115 textbook. The definitions are summarized below.

Accuracy is how close the result is to the true value. Precision refers to the repeatability of the results when more than one trial is done.

Precision is sometimes used to describe equipment. Equipment that is made to give more significant figures is considered more precise than equipment that is made to give less significant figures. This definition of precision is used when referring to equipment, not for results. When more than one trial is made for measurement of or calculation of the same thing, precision describes how close the trial results are to each other.

Volume by Difference: The volume of (water + sample) minus the volume of only water gives the volume of the sample.



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Mass by Difference: The mass of (beaker + sample) minus the mass of the empty beaker gives the mass of the sample.
This is a good technique to use when you have to walk away from the lab balance during your experiment.

Chemicals:
Unknown Solid (most likely flammable)
Unknown Liquid (most likely flammable)
Pennies; pre-1982, Pennies; post-1982)
Deionized water
Ethanol
Cyclohexane

Equipment:
Laboratory balance
10 mL graduated cylinder for the unknown liquid
50 mL or 100 mL graduated cylinder for the pennies
50 mL beaker for the pennies (mass by difference)
Test tubes
Transfer pipets (plastic, disposable)
Parafilm

Procedure:

(B) Determination of Density; Unknown Liquid

Perform this experiment three times; Trial 1, Trial 2, and Trial 3.
Do one trial at a time.

1. Determine the mass and volume of your first trial.
   a) Record the mass of your empty, dry, 10 mL graduated cylinder. Return to your lab bench so other people can use the balance.
   b) Add 10 mL of unknown liquid into the cylinder, so that the bottom of the meniscus is on the 10.0 mL line. Record the volume of liquid with the correct number of significant figures from the calibration marks on the cylinder.
   c) Return to the same balance used in step 1a. Record the mass of the graduated cylinder and unknown liquid.
2. Repeat step 1(a, b, c) for your next two trials.

(C) Determination of Solubility: Unknown Liquid

Three different solvents are used; water, ethanol, and cyclohexane.
1. Obtain 3 small test tubes.
2. Put approximately 2 mL of a solvent in a test tube (a little less than half-full). Each solvent should be in its own test tube.
3. Put about 6 drops of the unknown liquid in each test tube. Put Parafilm over the opening of the tube. Hold the Parafilm in place and shake the tube side-to-side. Try to keep the Parafilm dry. *(Be careful not to dissolve the Parafilm.)*
4. Observe whether or not the liquid dissolves (look for separate layers). If the unknown liquid dissolves (no layers, clear), it is soluble in that solvent. If the unknown liquid does not dissolve (two distinct layers), it is insoluble in that solvent. If there isn't two distinct layers but the solvent is now cloudy, the unknown liquid is slightly soluble in that solvent. Record your observations for each solvent.

(D) Determination of Solubility: Unknown Solid

Three different solvents are used; water, ethanol, and cyclohexane.
1. Obtain 3 small test tubes.
2. Put approximately 2 mL of a solvent in a test tube (a little less than half-full). Each solvent should be in its own test tube.
3. Put a few crystals of unknown solid in each test tube. Put Parafilm over the opening of the tube. Hold the Parafilm in place and shake the tube side-to-side. Try to keep the Parafilm dry. *(Be careful not to dissolve the Parafilm.)*
4. Observe whether or not the unknown solid dissolves. If the solid dissolves, it is soluble in that solvent. If the solid does not change at all, it is insoluble in that solvent. If the solid changes slightly and the solvent becomes cloudy, it is slightly soluble in that solvent. Record your observations for each solvent.

(E) Determination of Density; Pennies

Perform this experiment three times; Trial 1, Trial 2, and Trial 3.
Do one trial at a time. Use 10 pennies that are either pre-1982 or post-1982.

1. Determine the mass of your pennies.
   f) Use an empty, clean, dry 50 mL beaker.
   g) Place the beaker on the lab balance and record the mass of the empty beaker.
   h) Remove the beaker from the balance and add the ten pennies to the beaker. Place the beaker back on the balance and record the mass of the beaker and pennies; record all of the digits for the mass that the balance shows.
   i) Return to your lab bench so other students can use the balance.
2. Determine the volume of your pennies.
   e) Obtain an empty 50 mL or 100 mL graduated cylinder.
   f) Add approximately 20 mL of DI water into the cylinder. Record the exact volume of water with the correct number of significant figures from the calibration marks on the cylinder.
   g) Add the pennies to the graduated cylinder. Use all of the sample that gave the recorded mass in step 1. Make sure all of the pennies are under the surface of the water. Shake gently to remove air bubbles.
   h) Use the calibration marks on the cylinder to record the volume of water and pennies now in the graduated cylinder.
3. Pour the water and pennies from the graduated cylinder, then repeat steps 1 and 2 for your other two trials. Dry the pennies after each trial.

Calculations: Calculate the density and average density for the unknown liquid in Part B and the pennies in Part E.

$$\text{density} = \frac{\text{mass}}{\text{volume}} \qquad d = \frac{m}{v}$$

a) Calculate the density of your sample for each trial using the equation:
(Do not include the mass of the graduated cylinder or mass of the beaker when calculating the densities. Do not include the volume of the water with the pennies; just use the volume of the pennies.)

b) Calculate the average density value for the liquid and for the pennies.
Average = (sum of density values) / (number of trials)

Identification of the Unknown Liquid:
Use Table 1 in Blackboard (Experiment 2 module), along with your density and solubility results to determine the identity of the unknown liquid.

Identification of the Unknown Solid:
Use Table 2 in Blackboard (Experiment 2 module), along with your solubility results to determine the identity of the unknown solid.

Calculate the percent composition of the pennies as % Cu and % Zn.

d) Copper and zinc are the only elements in pennies. So, the % Cu and % Zn will add to 100%.
e) Use the average density you calculated for your sample of pennies. Use literature values for the density of copper and zinc *(https://pubchem.ncbi.nlm.nih.gov)*
f) Follow the math equations below.

$$\text{density}_{\text{penny}} = \frac{p\, d_{Cu} + q\, d_{Zn}}{100}$$

p = the % Cu by mass and q = the % Zn by mass

You need to solve for p and q.

Rearrange to solve for one unknown at a time.

$$p + q = 100\ \%$$

$$q = (100 - p)$$

rewrite as

$$\text{density}_{\text{penny}} = \frac{p\, d_{Cu} + (100 - p)\, d_{Zn}}{100}$$ Solve for p

$q = (100 - p)$ Solve for q

p is the % Cu

q is the % Zn

Name: _____

RCBC CHE116 Notebook Worksheet for Experiment 2: Laboratory Measurements

Section 1
Experiment work was performed on _____.
Lab Partner: _____.

Section 2 Raw Data

Part B Determination of Density: Unknown Liquid
Trial 1
Mass of empty 10 mL graduated cylinder: _____

Volume of unknown liquid in the cylinder: _____

Mass of cylinder with liquid: _____

Trial 2
Mass of empty 10 mL graduated cylinder: _____

Volume of unknown liquid in the cylinder: _____

Mass of cylinder with liquid: _____

Trial 3
Mass of empty 10 mL graduated cylinder: _____

Volume of unknown liquid in the cylinder: _____

Mass of cylinder with liquid: _____

Part C Determination of Solubility: Unknown Liquid

Observation for test tube with water + Unknown: _____

Observation for test tube with ethanol + Unknown: _____

Observation for test tube with cyclohexane + Unknown: _____

Part D Determination of Solubility: Unknown Solid
Observation for test tube with water + Unknown: _____

Observation for test tube with ethanol + Unknown: _____

Observation for test tube with cyclohexane + Unknown: _____

Part E Determination of Density: Pennies

Type of pennies; pre or post 1982 _____

Trial 1

Mass of empty 50 mL beaker _____

Mass of beaker + 10-pennies _____

Volume of water in the graduated cylinder _____

Volume of water + pennies in cylinder _____

Trial 2

Mass of empty 50 mL beaker _____

Mass of beaker + 10-pennies _____

Volume of water in the graduated cylinder _____

Volume of water + pennies in cylinder _____

Trial 3

Mass of empty 50 mL beaker _____

Mass of beaker + 10-pennies _____

Volume of water in the graduated cylinder _____

Volume of water + pennies in cylinder _____

Section 3: Experimental Procedure

Part B: The mass of a ____ mL graduated cylinder was determined using the lab _____. Then, _____ mL of the unknown liquid was added to the cylinder. The mass of the cylinder and liquid was then determined. These steps were repeated for trial _____ and trial _____.

Part C: The solubility of the unknown liquid was determined using three solvents: _____, _____, and _____. Approximately _____mL of each solvent was put into its own test tube. Approximately _____ drops of the unknown liquid was put into each solvent, and observations were recorded.

Part D: The solubility of the unknown solid was determined using three solvents: _____, _____, and _____. Approximately _____mL of each solvent was put into its own test tube. A few crystals of the unknown solid were put into each solvent, and observations were recorded.

Part E: The mass of an empty _____ mL beaker was recorded. Ten pennies were then put into the beaker, and the mass of the pennies and beaker was recorded. The volume of the pennies was determined by volume-by-difference, using a _____ mL graduated cylinder.

Section 4: Data Tables

Part B (show units)

Trial Number	Mass of empty cylinder	Mass of cylinder + liquid	**Mass of liquid**	**Volume of liquid**
1				
2				
3				

Part C

Solvent	Observation of unknown liquid in solvent
Water	
Ethanol	
Cyclohexane	

Part D

Solvent	Observation of unknown solid in solvent
Water	
Ethanol	
Cyclohexane	

Part E (show units)

Trial Number	Mass of empty beaker	Mass of beaker + pennies	**Mass of pennies**	Volume of only water in cylinder	Volume of water + pennies	**Volume of pennies**
1						
2						
3						

(Density of pennies **= Mass of pennies / Volume of pennies)** calculate this on page 21

Section 5: Calculations & Results

Part B: Show the calculation for the density of the unknown liquid for each trial, and then calculate the average density

Density calculation, Trial 1
Density calculation, Trial 2
Density calculation, Trial 3
Average Density of unknown liquid

Part C: What is the result for the solubility of the unknown liquid in each solvent?
Soluble, Sparingly Soluble, or Insoluble?

Solvent: Water
Solvent: Ethanol
Solvent: Cyclohexane

Part D: What is the result for the solubility of the unknown solid in each solvent?
Soluble, Sparingly Soluble, or Insoluble?

Solvent: Water
Solvent: Ethanol
Solvent: Cyclohexane

Part E: Show the calculation for the density of the pennies for each trial, and then the average density. Use the average density to calculate the % composition of the pennies. (Show calcs.)

Density calculation, Trial 1
Density calculation, Trial 2
Density calculation, Trial 3
Average Density of pennies
% composition

Section 6: Conclusion

Part B & C: Use the given information in Blackboard and your results to identify the unknown liquid.

Part D: Use the given information in Blackboard and your results to identify the unknown solid.

Part E: Compare your calculated % composition of pennies to literature values.

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Experiment 3: Density of Aqueous Sodium Chloride Solutions

(This experiment is adapted from Ross S. Nord and Stephen E. Schullery, Eastern Michigan University.)

Purpose: The purpose is to determine the concentration of an unknown sodium chloride solution by calculating its density. The density of this unknown solution will be experimentally determined, and compared to a standard curve prepared from solutions of known density (a graph of density vs. concentration).

Background: The density of a solution is usually directly related to the solution's concentration. Because the density of a solution is easy to measure, this provides a convenient method of determining the concentration of the unknown sample.

$Density = mass \,/\, volume$

For a solution, the volume is readily determined using accurately graduated glassware. The mass is determined using a laboratory balance.

The concentration of an aqueous solution can be expressed in many different ways, but the simplest is mass percent (also called weight percent).

$Mass\ \% = (\ mass\ solute\ /\ (\ mass\ solute + mass\ H_2O\ )\ )\ x\ 100\%$

The density of aqueous NaCl solutions is a nearly-linear function of the NaCl concentration (in mass percent). A linear relationship permits a reliable standard curve to be constructed. The best-fit line is added using linear regression. This line shows the relationship between density and % mass concentration NaCl.

In this experiment, five standard solutions of known mass percent will be prepared and their densities determined. A standard curve will be constructed by plotting the density vs. mass % and the best-fit line will be determined via linear regression (use Excel or Google Sheets). The density of an unknown solution will be calculated and its mass % concentration will be determined with the equation of the best-fit line.

Record measured and calculated values with correct units and the appropriate number of significant figures. All calculations should be clearly organized, make proper use of significant figures, and include the units.

Chemicals:

Deionized water
Sodium Chloride

Equipment:

Beakers (50 mL, 100 mL)	10 mL pipet & pipettor	glass stirring rod
Laboratory Balance	spatula	thermometer
Erlenmeyer flasks (50 mL)	transfer pipets (plastic)	Parafilm

Procedure:

Part A: Preparing the Standard Solutions

1. Weigh out the mass of NaCl for solution number 1 (given in Table 1 below).
   a) Place a clean, dry, empty 50-mL (or 100 mL) beaker on a balance, and record its mass.
   b) Use a spatula to add solid NaCl to the beaker. The mass of solid does not need to be exactly the mass given in the table, but should be close. However, you do need to know exactly how much you measured out, so record the mass of the (beaker+ NaCl) from the lab balance. Record all the digits on the display.
   c) Repeat steps 1a and 1b for the other solutions in Table 1 (solutions 2 – 5).

Table 1: Mass of NaCl for Standard Solutions

Standard Soln	**1**	**2**	**3**	**4**	**5**
Mass NaCl, (g)	1.0	2.0	3.0	4.5	6.0

2. Return to your lab bench with the five beakers of salt from steps 1 (a, b, c). Add 30 mL of distilled water to each of the beakers. (Use a graduated cylinder to measure the amount of liquid.) Bring your five solutions back to the same balance used in step 1. Measure the mass of each (beaker+salt+water) solution.

3. Bring your five solutions back to your lab bench. Mix the contents of each beaker until the salt dissolves. Rinse and dry your glass stirring rod in between beakers so you don't contaminate any of them. There should now be a homogeneous mixture (a solution) in each beaker.

Part B: Determining the Density of the Standard Solutions and Unknown Solution

1. Bring six dry, clean, empty small beakers or small Erlenmeyer flasks to the balance. These beakers should already be labeled as solution 1 – 5, and unknown.
2. Measure and record the mass of each labeled beaker or flask. Record all digits in the balance display.
3. Return to your lab bench. Pipet 10.0 mL of each solution (1 – 5, unknown) into its labeled beaker or flask. (You will need to get approximately 20 mL of unknown solution from the bottle in the hood to pipet 10.0 mL of this solution. Record the ID number of the unknown.)
4. Bring these beakers back to the balance used in step 1. Measure and record the mass of each (beaker+10.0 mL solution). You should now have the mass of all solutions (1 – 5, unknown).

Calculations

1. Calculate the mass % NaCl in each of your five standard solutions. Show your calculations. Do not include the mass of the beaker in these calculations; subtract out the mass of the empty beaker from the mass NaCl and the (mass NaCl + mass H_2O)

$Mass$ % = ($mass$ NaCl / ($mass$ NaCl + $mass$ H_2O)) x 100%

2. Calculate the density for each of your five standard solutions and the unknown solution. Show your calculations. Use 10.0 mL as the significant figures for the volume.
Subtract out the mass of the empty beaker before doing these calculations.

Density = ((mass of 10.0 mL solution) / (10.0 mL))

3. Generate your standard curve and best-fit line using Excel or Google Sheets. *(There is a video in Blackboard to show you how to use these programs.)* Density values are on the Y-axis and % mass NaCl values are on the X-axis. Put the best-fit line on the graph. Print the graph with the best-fit line, equation of the line, and R^2 value. **Attach a printout of this graph with your notebook worksheet.**

4. Calculate the concentration (% mass) of your unknown using its density and the equation of the best-fit line: y=mx+b

y is the density of your unknown, and you are solving for x, which is the concentration. Show your calculations.

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Name: _____

RCBC CHE116 Notebook Worksheet for Experiment 3: Density of NaCl Solutions

Section 1

Experiment work was performed on _____.

Lab Partner: _____.

Section 2 Raw Data Temperature of the lab: _____

Part A Preparing the Standard Solutions (show units)

Standard Solution Number	**1**	**2**	**3**	**4**	**5**
Mass of empty beaker					
Mass of beaker + NaCl					
Mass of beaker + NaCl + water					

Part B Determination of the Density of the Solutions (show units)

Solution Number	**1**	**2**	**3**	**4**	**5**	**Unknown**
Mass of empty beaker						
Mass of beaker + 10.0 mL solution						

Section 3: Experimental Procedure

Part A: The mass of an empty beaker was measured with the lab balance. Solid _____ was added to the beaker, and then the mass of the _____ and beaker was obtained. This was repeated with four more beakers, adding the appropriate amount of _____ to each beaker, according to Table 1 on page 24 of the experiment document.

To each of these five beakers with NaCl, _____ mL of DI H_2O was added. The five beakers with solution were brought back to the same balance used previously, and the mass of each was determined. Each mixture was stirred to dissolve the NaCl.

Part B: The density of each standard solution and the _____ solution was determined by pipetting exactly _____ mL of each solution into a previously weighed empty beaker or Erlenmeyer flask. The mass of the beaker + solution was measured. The mass of the solution was determined by mass-by-difference.

Section 4: Data Tables

Part A (show units)

Standard Solution Number	**1**	**2**	**3**	**4**	**5**
Mass of empty beaker					
Mass of beaker + NaCl					
Mass of beaker + NaCl + water					
Mass of NaCl					
Mass of (NaCl + H2O)					

Part B (show units)

Solution Number	**1**	**2**	**3**	**4**	**5**	**Unknown**
Mass of empty beaker (or flask)						
Mass of beaker (or flask) + 10.0 mL solution						
Mass of 10.0 mL solution						

Section 5: Calculations & Results

Part A: Show the calculation for the % mass NaCl, for each standard solution.

Solution 1
Solution 2
Solution 3
Solution 4
Solution 5

Part B: Show the calculation of the density for each solution.

Solution 1
Solution 2
Solution 3
Solution 4
Solution 5
Unknown

Summary of Results

Solution	1	2	3	4	5	Unknown
% mass NaCl						**???????**
Density						

Do not graph the unknown.

Calibration Curve: Printout, Equation of the Line, R^2

The equation of the best-fit-line: _____

The R^2 value: _____

Calculated %mass NaCl in the unknown solution: _____

(show calculation here)

Attach a printout of your calibration curve as the last page of this worksheet.

Section 6: Conclusion (Use complete sentences.)

Did the calibration curve using the 5 standard solutions yield a straight line? Was the R^2 value close to 1? What does the R^2 and b value mean with respect to solution preparation?

Does the calculated % mass NaCl for the unknown make sense, based on its density? Explain your answer.

Experiment 4: Gravimetry

(This experiment is from CC-BY Torres & González-Urbina, CUNY.)

Purpose

The purpose of this experiment is to determine the percent yield of the reaction between barium nitrate ($Ba(NO_3)_2$) and sulfamic acid (NH_2SO_3H) that yields solid barium sulfate ($BaSO_4$).

Background

Sulfamic acid (NH_2SO_3H) is a very common chemical used to remove grout and mortar haze, as well as rust and mineral deposits [1]. It can be found in many home improvement retail stores [2]. The hydrolysis of sulfamic acid to give sulfates can be used to precipitate barium in the form of barium sulfate [3].

$$Ba(NO_3)_{2(aq)} + NH_2SO_3H_{(aq)} + H_2O_{(l)} \rightarrow BaSO_{4(s)} + NH_4NO_{3(aq)} + HNO_{3(aq)}$$

In this experiment, you will react sulfamic acid with barium nitrate in hot water to produce a precipitate of barium sulfate. The solid material will be isolated by gravity filtration. When the barium nitrate is the limiting reactant and the amount of barium nitrate used is known, that information, along with the mass of precipitate isolated, will allow for the percent yield to be calculated. The advantage of using sulfamic acid instead of sulfuric acid, is that sulfamic acid produces a coarse, crystalline precipitate with fewer impurities [4].

Reaction yield

Stoichiometric calculations can predict the amount of product that should be formed in a chemical reaction. However, chemical reactions are not perfect, and often times reactions do not proceed to full completion, and the amount of product predicted is not obtained in the experiment. The *yield* of a chemical reaction refers to the amount of product actually obtained in the experiment with respect to the expected quantity calculated with stoichiometric calculations.

%Yield = ((Actual amount obtained in lab) / (Expected amount from calculations)) x 100%

Example Calculation:

We mix 2.0 moles of $Ba(NO_3)_2$ with an excess of NH_2SO_3H to produce a $BaSO_4$ precipitate. We isolate only 1.0 mole of $BaSO_4$ in the lab.

$$Ba(NO_3)_{2(aq)} + NH_2SO_3H_{(aq)} + H_2O_{(l)} \rightarrow BaSO_{4(s)} + NH_4NO_{3(aq)} + HNO_{3(aq)}$$

Calculate the reaction yield.

Answer: According to stoichiometric calculations, 2 moles of $Ba(NO_3)_2$ should produce 2 moles of $BaSO_4$. However, the actual number of moles of $BaSO_4$ obtained in lab is 1.0 mole. Hence the yield is

%Yield = ((1.0 mole obtained in lab) / (2.0 moles expected from calculations)) x 100%

%Yield = 50%

Limiting Reactant

In a chemical reaction, the reaction stops when one of the reactants is completely used. This reactant is referred to as the limiting reactant, and this reactant limits the amount of product generated. After all limiting reactant has reacted, some of the other reactant will remain unused. Typically, one needs to do some basic calculations to identify the limiting reactant.

Filtration

Filtration is a technique employed in chemistry to separate a solid compound from the liquid. The mixture is poured onto a filter and gravity makes the liquid go through the filter, while the solid remains on the filter. When the mixture containing the solid is very hot, we call the filtration procedure "hot gravity filtration". Filter paper must be folded before proceeding with the filtration, forming a cone. In order to do this one needs to first fold the filter in half, and then in half again, as shown in Figure 1. Do not press the tip of the cone while folding, because it will weaken the paper.

Figure 1

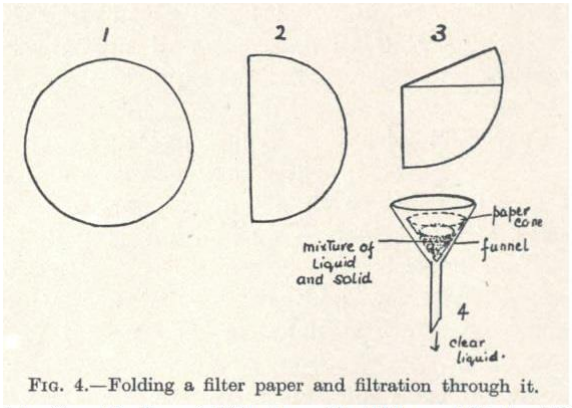

FIG. 4.—Folding a filter paper and filtration through it.

https://en.wikipedia.org/wiki/File:Paper_filter_folding_and_filtration.JPG

Chemicals

(Barium compounds are toxic. Make sure you wear gloves and wash your hands after handling the barium solids and solution.)

0.060 M $Ba(NO_3)_2$ solid NH_2SO_3H (MW = 97.1 g/mole)

Equipment

250 mL beaker	50 mL or 100 mL graduated cylinder
Stirring rod	Filter paper and funnel
250 mL Erlenmeyer flask (to support the funnel)	
400 mL beaker (for the ice bath)	
Tongs	Spatula
Watch glass	Drying Oven

Procedure
The Reaction
(1) You will need 2.3 g to 2.5 g NH_2SO_3H. This reactant is located near the lab balance.

(2) Bring your lab notebook and a clean, dry 250 mL beaker to the balance.

(3) Tare a piece of weighing paper and weigh out 2.5 _ _ g of NH_2SO_3H. Record the exact mass in your notebook, and then transfer all of this sulfamic acid into the 250 mL beaker. Return to your lab table.

(4) Do this step in the hood, then return to your table. Use a 100 mL or 50 mL graduated cylinder to measure 100.0 mL of the 0.060 M barium nitrate solution. Transfer all of this barium solution into the 250 mL beaker that contains the sulfamic acid. Return to your table.

(5) At your table, add 25 mL of deionized water to the graduated cylinder as a rinse, and add this to the beaker. Mix with a glass stirring rod.

(6) **IN THE HOOD**, use a hotplate to heat the solution to boiling. Allow the solution to boil GENTLY for approximately 30 minutes. Cover the beaker with a watch-glass, and add small amounts of deionized water to make sure the amount of liquid does not change much. Add small amounts of the water so you don't stop the boiling. Stir the solution occasionally. Do not lose sample on the glass stir rod.

(7) Use the 250 Erlenmeyer flask to support the funnel for the filtration.

(8) After 30 minutes of boiling, turn off the hot plate. Use several layers of paper towel to make potholders so you can remove the beaker from the hotplate and onto your lab bench. When the beaker has cooled a bit, place the very warm beaker into an ice bath. Be careful not to contaminate your reaction mixture with the ice water. Cool the beaker to almost room temperature.

(9) Obtain a piece of filter paper, write you name on it with pencil, and then record its mass.

(10) When the reaction mixture containing the precipitate is cooled, proceed to filter the mixture. Make sure no trace of precipitate remains in the beaker. Use a wash bottle to rinse any remaining product out of the beaker and onto the filter paper.

(11) Transfer the filter paper and its contents to a watch glass. Make sure the precipitate does not fall off of the filter paper.

(12) Place the watch glass with filter paper in an oven at 105°C for 15 minutes.

(13) Remove the dry filter paper with dry precipitate from the oven after the 15 minutes. Be careful not to burn yourself. Put the hot watch glass on the lab bench, and slide the filter paper and precipitate onto the lab bench to cool. Once the filter paper and precipitate are cool, weigh them and record the mass into your notebook. Then return the filter paper with precipitate back to the watch glass.

(14) Return the filter paper/precipitate and watch glass to the hot over for another 15 minutes. Then repeat step 13 to find the mass of the filter paper and precipitate. Record this mass in your notebook. Use this second mass for your calculations. We are assuming this second mass is the dry precipitate. Was this second mass much less than the first 'dry' mass?

(15) Calculate the % Yield.

a) Convert the liters of barium nitrate used to moles of barium nitrate. (M x L=moles)
b) Convert the mass of sulfamic acid to moles of sulfamic acid. (g / g/mole = mole)
c) Confirm that barium nitrate is the limiting reactant. Determine the amount of excess sulfamic acid remaining after the reaction stopped, assume all of the barium nitrate was used. (moles start – moles used = moles left over)
d) Base your calculations on the number of moles of limiting reactant, barium nitrate, that was used. How many moles of barium sulfate should have been made, based on the moles of barium nitrate used? (1:1 ratio)
e) Convert the moles of barium sulfate you should have made to grams. (mole x g/mole = g)
f) Calculate the % Yield, using the grams of barium sulfate precipitate obtained in lab and the grams of barium sulfate you expected to make *(calculation step e)*.

References

(1) pubchem., https://pubchem.ncbi.nlm.nih.gov, Accessed: 2017-04-22.
(2) www.homedepot.com/., http://www.homedepot.com/, Accessed: 2017-04-22.
(3) Wagner,W. F.;Wuellner, J. A. *Analytical Chemistry* **1952**, *6*, 1031–1032.
(4) Notley, J.M. *Journal of Applied Chemistry and Biotechnology* **1973**, *10*, 717–723.

Procedure

Name: _____

RCBC CHE116 Notebook Worksheet for Experiment 4: Gravimetry

Section 1

Experiment work was performed on _____.

Lab Partner: _____.

Section 2 Raw Data

Volume of 0.060 M $BaNO_3$ used: _____

Mass of NH_2SO_3H used: _____

Mass of empty filter paper (with name on it): _____

Mass of filter paper (with name) + precipitate after 1st drying: _____

Mass of filter paper (with name) + precipitate after 2nd drying: _____

Section 3: Experimental Procedure

The reaction between sulfamic acid and barium nitrate was accomplished by using _____ g of sulfamic acid and _____ mL of 0.060 M barium nitrate solution. The chemicals were combined in a 250 mL _____. The chemicals were mixed. The contents of the beaker were covered with a watch glass and heated to boiling, on a hotplate in the fume hood. The mixture was boiled for _____ minutes. Deionized water was added to the beaker in small increments so the volume would not decrease and the mixture would not stop boiling. After 30 minutes, the mixture was cooled slightly and then placed in an _____. Once at about room temperature, the mixture was _____. (The mass of the empty filter paper was determined before filtering.) The mass of dry precipitate and filter paper was determined after _____ minutes of drying. The mass of dry precipitate and filter paper was determined again after an additional _____ minutes of drying in the 105°C drying oven.

Section 4: Data Table

Volume of 0.060 M $Ba(NO_3)_2$	
Mass of NH_2SO_3H	
Mass of empty filter paper with name	
Mass of filter paper with name + ppt, 1st drying	
Mass of filter paper with name + ppt, 2nd drying	
Mass of precipitate after 2nd drying *(This is the experimental amount of barium sulfate.)*	

Section 5: Calculations & Results

Calculate the moles of barium nitrate used. M x L = moles
Calculate the moles of sulfamic acid at the start: g / (g/mole) = moles
Determine amount of sulfamic acid left over: (moles s. acid start – moles used = moles left)
Moles of barium sulfate that should have been made: (1:1 ratio)
Convert moles barium sulfate to grams barium sulfate: (mole x g/mole = g) *(This is the expected amount of barium sulfate.)*
% yield of barium sulfate

Section 6: Conclusion

What was the % yield? Were you able to get 100%? If not, was the % yield too high or too low? Explain your answer.

Experiment 5: Chemical Reactions

Part A: Precipitation and acid-base reactions

(This part of the experiment was adapted from CC-BY Torres & González-Urbina Experiment: Precipitation and Acid-Base Reactions)

Purpose

The goal of this experiment is to perform a set of simple acid-base and precipitation reactions in aqueous solutions. For the precipitation reactions you will use the solubility rules to identify the solid and aqueous products of the reaction. For the acid-base reactions, you will identify the acidic/basic characteristic of each reactant chemical.

Chemicals

Solutions of	NaCl	NaBr	Na_2SO_4	KCl
Solutions of	$Ba(NO_3)_2$,	$AgNO_3$	*(Barium nitrate is toxic!!!)*	
Solutions of	HCl	CH_3COOH		
Solutions of	NaOH	Na_2CO_3		

Equipment

Red and blue litmus paper
4 Small test tubes (You will clean and reuse these test tubes during this experiment.)
Parafilm Test tube rack

Background

Many chemical reactions take place in aqueous solution, and most of these reactions involve ions. Let's consider as an example what happens when mixing a colorless solution of silver nitrate ($AgNO_3$) with a colorless solution of sodium chloride (NaCl). The solution of silver nitrate contains Ag^{1+} cations and NO_3^{1-} anions, whereas the solution of sodium chloride contains Na^{1+} cations and Cl^{1-} anions. When we mix these two aqueous solutions, a white precipitate (AgCl) forms immediately due to the ion exchange process.

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

Acid-base reactions are aqueous reaction also. Acids and bases react by means of a neutralization reaction. An example would be:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

This experiment addresses these two important types of chemical reactions.

Precipitation reactions

Some ionic compounds are soluble in water whereas others are not. In a precipitation reaction, two strong-electrolyte solutions are mixed to produce an insoluble solid called a precipitate.

Solubility Guidelines:

Ions that form soluble compounds (aq)…….…….. except when combined with …….	
Group I ions (Na^+, Li^+, K^+, etc)	no exceptions
Ammonium (NH_4^+)	no exceptions
Nitrate (NO_3^-)	no exceptions
Acetate (CH_3COO^-)	no exceptions
Hydrogen carbonate (HCO_3^-)	no exceptions
Chlorate (ClO_3^-)	no exceptions
Halides (F^- , Cl^- , Br^-)	Pb^{2+}, Ag^+ and Hg_2^{2+}
Sulfate (SO_4^{2-})	Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Hg_2^{2+} and Pb^{2+}

Ions that form insoluble compounds (s)…….except when combined with….	
Carbonate (CO_3^{2-})	group I ions (Na^+, Li^+, etc) or (NH_4^+)
Chromate (CrO_4^{2-}) (NH_4^+)	group I ions (Na^+, Li^+, etc) or Ca^{2+}, Mg^{2+} or
Phosphate (PO_4^{3-})	group I ions (Na^+, Li^+, etc) or (NH_4^+)
Sulfide (S^{2-})	group I ions (Na^+, Li^+, etc) or (NH_4^+)
Hydroxide (OH^-)	group I ions (Na^+, Li^+, etc) or Ca^{2+}, Mg^{2+}, Sr^{2+} or NH_4^+

Acids and Bases

Acids are chemicals that produce hydrogen ions (H^{1+}) in water. Bases produce hydroxide ions (OH^{1-}) that accept hydrogen ions. H^{1+} and OH^{1-} combine to form water. Acids and bases change the color of certain chemicals called indicators, and litmus is a well-known acid-base indicator. Acids and bases can be classified as strong or weak according to the extent to which they ionize/dissociate in solution. A strong acid is completely ionized in solution, whereas a weak acid is only slightly ionized. The same can be applied to bases.

Strong Acids	**Strong Bases**
HCl, HBr, HI	NaOH, KOH
$HClO_4$, $HClO_3$	$Sr(OH)_2$, $Ba(OH)_2$
H_2SO_4	Na_2O, BaO
Weak Acids	**Weak Bases**
HF, CH_3COOH, H_2SO_3	NH_3

Acid-Base Neutralization Reactions

A neutralization reaction between a strong acid and a strong base yields a salt and water:

Acid (aq) + Base (aq) $\rightarrow$ Salt (aq) + $H_2O(l)$

Litmus paper

Litmus paper is a quick test to identify whether a solution is acidic or basic. There are two variants for litmus paper; red litmus paper and blue litmus paper. Blue litmus paper turns pink under acidic conditions whereas red litmus paper turns blue under basic conditions. Regardless of which litmus you start with, remember that ***bases turn litmus blue and acids turn litmus red.***

	Red litmus paper	**Blue litmus paper**
Acidic solution	Stays Red	Turns Red
Basic solution	Turns Blue	Stays Blue

Ionic and Net ionic equations

The ionic equation for a precipitation reaction or acid base reaction shows all the species as they actually exist in solution. Because dissolved ionic compounds exist as separate aqueous ions, the ions should be shown separately. Some of these ions appear the same as both reactants and products. This means that they play no role in the reaction: they are spectator ions. In the ionic equation, you can simplify the chemical equation by canceling the spectators out on each side of the arrow; this produces the net ionic equation.

The ionic and net ionic equations for the reaction between aqueous solutions of silver nitrate ($AgNO_3$) and potassium chromate (K_2CrO_4) to give a precipitate of silver chromate (Ag_2CrO_4) and a solution of potassium nitrate (KNO_3) are:

$2Ag^{1+}_{(aq)} + 2NO_3^{1-}_{(aq)} + 2K^{1+}_{(aq)} + CrO_4^{2-}_{(aq)} \rightarrow Ag_2CrO_{4(s)} + 2K^{1+}_{(aq)} + 2NO_3^{1-}_{(aq)}$

$2Ag^{1+}_{(aq)} + CrO_4^{2-}_{(aq)} \rightarrow Ag_2CrO_{4(s)}$ *(no spectator ions shown)*

Summary Checklist of Chemical Reactions for Part A, Steps 1-5:

	NaCl	NaBr	Na_2SO_4	KCl
$Ba(NO_3)_2$	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt
$AgNO_3$	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt

Cloudy means a ppt, clear means no ppt

****Do not put your observations on this paper. Your observations go in your notebook worksheet.*

Summary Checklist of Chemical Reactions for Part A, Steps 6-8:

	NaOH	Na_2CO_3	HCl	CH_3COOH
Red Litmus	color	color	color	color
Blue Litmus	color	color	color	color

The color will be blue or red.

****Do not put your observations on this paper. Your observations go in your notebook worksheet.*

Summary Checklist of Chemical Reactions for Part A, Steps 9-11:

	NaOH	Na_2CO_3
HCl	Heat?	Gas or heat?
CH_3COOH	Heat?	Gas or heat?

These reactions do take place, even if you don't feel heat or see gas formation.

****Do not put your observations on this paper. Your observations go in your notebook worksheet.*

Procedure

Part A: Precipitation reactions

1) Arrange in the following order the following set of reactants:
NaCl (aq) NaBr (aq) Na_2SO_4(aq) KCl (aq)

2) Arrange in the following order the following set of reactants:
$Ba(NO_3)_2$(aq) $AgNO_3$(aq)

3) Use NaCl (aq) and add 10 drops to a set of two test tubes. Each tube should have now 10 drops of NaCl. Repeat this step with two test tubes for NaBr, then two test tubes for Na_2SO_4, then two test tubes for KCl.

4) Use $Ba(NO_3)_2$(aq). Add 10 drops to the first test tube containing the NaCl. Then add 10 drops of $Ba(NO_3)_2$(aq) to each first test tube for NaBr, Na_2SO_4, and KCl. Write down your observation:
Indicate soluble as (aq) for aqueous or insoluble as (s) for solid.

5) Use $AgNO_3$(aq). Add 10 drops to the second test tube containing the NaCl. Then add 10 drops of $AgNO_3$(aq) to each second test tube for NaBr, Na_2SO_4, and KCl. Write down your observation:
Indicate soluble as (aq) for aqueous or insoluble as (s) for solid.

Acid Base Neutralization Reactions

6) Arrange in the following order these reactants: NaOH (aq) Na_2CO_3(aq)

7) Arrange in the following order these reactants: HCl (aq) CH_3COOH (aq)

8) Put the litmus papers on a watch glass while doing this step. For each of the four reactants, put one drop of solution on its own piece of red litmus paper. Then put one drop of each reactant on its own piece of blue litmus paper. Record all of your observations and identify each chemical as an acid or base.

9) Use NaOH (aq) and add 10 drops to a set of two test tubes. Each tube will have now 10 drops of a NaOH. Repeat this step with two test tubes for Na_2CO_3.

10) Use HCl. Add 10 drops to the first test tube containing the NaOH. Then add 10 drops to the first test tube containing Na_2CO_3. Observe if you feel any heat generated. Sometimes the acid and base concentrations are too low to feel heat. However, a reaction has still taken place.

11) Use CH_3COOH. Add 10 drops to the second test tube containing the NaOH. Then add 10 drops to the second test tube containing Na_2CO_3. Observe if you feel any heat generated. Sometimes the acid and base concentrations are too low to feel heat. However, a reaction has still taken place.

References
(1) Fetzer-Gislason, P. R.D.W. S., Lab Experiments in Introductory Chemistry; Freeman: 2003.
(2) Murov, S., Experiments in General Chemistry; Cengage: 2013.
(3) Ebbing, R. W.D. D., Experiments in General Chemistry; Houghton Mifflin Harcourt Publishing Company: 2004.

Part B: The Scientific Method

This experiment was adapted from Torres & González-Urbina, CUNY, The Scientific Method.

Purpose

The goal of this experiment is to understand how the scientific method works. In order to do this, you will make observations, look for patterns and come up with a hypothesis in order to identify an unknown compound. You will do this by comparing the chemical reactions of the unknown compound with a set of known compounds. Observe changes that indicate a chemical reaction has occurred, such as gas evolution, the formation of a precipitate, a change of color, or the generation of heat.

Chemicals

Caution: Solutions of barium are toxic.

Known solutions: $NaCl_{(aq)}$ $NaI_{(aq)}$ $Na_2CO_{3(aq)}$ $Na_2HPO_{4(aq)}$ $Na_2SO_{4(aq)}$,
Testing solutions: $Ba(NO_3)_{2(aq)}$ $AgNO_{3(aq)}$ thymol blue
Unknown compound$_{(aq)}$
$NH_{3(aq)}$, which is also written as $NH_4OH_{(aq)}$

Equipment

Six test tubes (you will clean and reuse these test tubes during this experiment)
Test tube rack Parafilm

Background The scientific method is a sequence of logical steps.
The first step is to collect data by making observations and/or measurements of a sample. The next step consists of looking for patterns and trends in the data. When a pattern is observed, scientists develop a hypothesis, that is, a feasible explanation of the observations. After formulating a hypothesis, scientists think of experiments to test the hypothesis. If the results of repeated experiments support the hypothesis, scientists formulate a theory that explains the observations.
This experiment was adapted from the literature [1, 2].

Summary Checklist of Chemical Reactions for Part B:

Reagents to add to Test Tubes	Known Chemicals and Unknown in the Test Tubes (6 test tubes, all solutions)					
	NaCl	NaI	Na_2CO_3	Na_2HPO_4	Na_2SO_4	Unknown
$Ba(NO_3)_2$	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt
$AgNO_3$	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt
thymol blue	color	color	color	color	color	color

****Do not put your observations on this paper. Your observations go in your worksheet.*

Procedure

1) Obtain the unknown solution and the known solutions.

Testing for a barium precipitate

2) Using 6 test tubes, add 10 drops of each known solution to its own test tube. Each test tube should have a different known solution, and the last test tube should have the unknown solution.

Add 2 drops of ammonia (NH_3) solution to each test tube and mix. (NH_3 will not be part of the chemical equation.)

3) Add five drops of the $Ba(NO_3)_2$ solution to each of the test tubes. Shake gently to obtain homogeneity. Examine each test tube carefully and look at the results. Record your observations: precipitate/color or no precipitate.

4) Discard the solutions from the test tubes into the waste container. Wash the test tubes, and rinse them with distilled water. You will reuse these test tubes for the following steps.

Testing for a silver precipitate

5) Using the cleaned 6 test tubes, add 10 drops of each known solution to its own test tube. Each test tube should have a different known solution, and the last test tube should have the unknown solution.

6) Add five drops of the silver nitrate $AgNO_3$ solution to each test tube. Record your observations: precipitate/color or no precipitate.

7) Discard the solutions from the test tubes into the waste container. Wash the test tubes, and rinse them with distilled water. You will reuse these test tubes for the following steps.

Testing with thymol blue

8) Using the cleaned 6 test tubes, add 10 drops of each known solution to its own test tube. Each test tube should have a different known compound solution, and the last test tube should have the unknown solution.

9) Add one drop of the thymol blue solution to each test tube. Shake each test tube gently before recording your observations: write down the colors of the resulting solutions. (If the color is too faint, add one more drop of thymol blue.)

10) Discard the solutions from the test tubes into the waste container. Wash the test tubes with tap water and then DI water. Then put them in the collection bin.

Identifying the unknown compound

11) The unknown is one of these compounds:

NaCl NaI Na_2CO_3 Na_2HPO_4 Na_2SO_4

You should be able to identify your unknown compound by matching its chemical reactions with those of one of the known compounds. Explain your logic when you identify your unknown.

References
(1) Ebbing, R. W.D. D., Experiments in General Chemistry; Houghton Mifflin Harcourt Publishing Company: 2004.
(2) Beran, J. A., Laboratory manual for principles of general chemistry; John Wiley and Sons: 2010.

Name: _____

RCBC CHE116 Notebook Worksheet for Experiment 5: Chemical Reactions

Section 1

Experiment work was performed on _____.

Lab Partner: _____.

Section 2 Raw Data

Chemical Reactions for Part A, Steps 1-5. If there is a ppt, indicate its color.

	NaCl	NaBr	Na_2SO_4	KCl
$Ba(NO_3)_2$	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt
$AgNO_3$	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt

Chemical Reactions for Part A, Steps 6-8:

	NaOH	Na_2CO_3	HCl	CH_3COOH
Red Litmus	color	color	color	color
Blue Litmus	color	color	color	color

Chemical Reactions for Part A, Steps 9-11:

	NaOH	Na_2CO_3
HCl	Heat?	Gas or heat?
CH_3COOH	Heat?	Gas or heat?

Chemical Reactions for Part B: If there is a ppt, indicate its color.

Reagents to add to Test Tubes	Known Chemicals and Unknown in the Test Tubes (6 test tubes)					
	NaCl	NaI	Na_2CO_3	Na_2HPO_4	Na_2SO_4	Unknown
$Ba(NO_3)_2$	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt
$AgNO_3$	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt	ppt or no ppt
thymol blue	color	color	color	color	color	color

Section 3: Experimental Procedure

The chemical reactions listed in the experiment document were performed. Approximately 1 mL (_____ drops) of each solution was used for each chemical test. One drop of each solution was put on the _____ litmus paper and the _____ litmus paper when testing for acids and bases. For Part B, _____ test tubes were used, five for the knowns and one for the _____.

Section 4: Data Table

This section contains the observations and it is combined with Section 5: Results

Section 5: Results

Part A: Record your observation for each chemical reaction listed below and write the balanced chemical reaction using the molecular equation format.

Chemicals mixed together (the reactants)	**Observation (Section 4)**	**Balanced Chemical Equation (molecular equation format)**
$Ba(NO_3)_2$ and Na_2SO_4		
$AgNO_3$ and NaCl		
$AgNO_3$ and NaBr		
HCl and NaOH		
CH_3COOH and NaOH		
HCl and Na_2CO_3		

Part A: Results for Acid, Base Tests

Chemical	**NaOH**	**Na_2CO_3**	**HCl**	**CH_3COOH**
Red Litmus Observation				
Blue Litmus Observation				
Result: Acid or Base?				

Part B: Identification of Unknown Chemical

Chemical Test	**Known Chemicals that Matched with Unknown**
Barium nitrate	
Silver nitrate	
Thymol blue	

Section 6: Conclusion Use complete sentences.

Part A: For the litmus paper tests, were you able to identify whether the chemicals were acids or bases? Explain your answer.

Part B: What is the identity of the unknown compound? Explain the logic used to make the decision.

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Experiment 6: Chemical Reactions and the Activity Series

(This experiment was adapted from two experiments from Santa Monica College: CH10 Chemical Reactivity and Types of Chemical Reactions.)

Purpose:

a) To perform and observe the results of a variety of single and double displacement reactions
b) To become familiar with some of the observable signs of these reactions
c) To identify the products formed in each of these reactions
d) To write balanced chemical equations for each single and double displacement reaction studied

Background

During a chemical reaction both the form and composition of matter are changed. Old substances are converted to new substances, which have unique physical and chemical properties of their own. Some of the observable signs that a chemical reaction has occurred include the following:

- A metallic deposit appears
- Bubbles appear
- A temperature change occurs
- A color change occurs
- A precipitate appears (cloudy, tiny particles)

Single Displacement Reactions

All single displacement reactions have the general form: **A + BC → B + AC**
Here, A is an element and BC is usually an aqueous ionic compound or an acid (consisting of B^+ and C^- aqueous ions). A displaces B in BC, resulting in the formation of a new element B and a new ionic compound or acid, AC. If the new element B is a metal, it will appear as a solid. If it is a gas, it will appear as bubbles.
An *Activity Series* of elements is often used to determine if A will displace B in a single displacement reaction. An *Activity Series* is provided on page 52. As a rule, if A has a higher activity that B, a single displacement reaction will occur. However, if A has lower activity than B, a single displacement reaction will <u>not</u> occur.

Double Displacement Reactions

All double displacement reactions have the general form: **AB + CD → AD + CB**
Reactions that can be classified as double displacements include precipitation reactions, neutralization reactions, and gas forming reactions.

Precipitation Reactions

Here AB and CD are usually aqueous ionic compounds (or acids) consisting of aqueous ions (A^+ and B^-, C^+ and D^-). When a double displacement reaction occurs, the cations and anions switch partners, resulting in the formation of two new ionic compounds AD and CB, one of which is in the solid state. This solid product is an insoluble ionic compound called a precipitate. To determine whether a product compound will be soluble (aq) or insoluble (s), consult the *Solubility Guidelines*. Note that if both of the predicted products are soluble, a precipitation reaction will not occur.

Neutralization Reactions

Here AB is an acid (consisting of H^+ and X^- aqueous ions) and BC is a base (consisting of M^+ and OH^- ions). When a double displacement reaction occurs, the cations and anions switch partners, resulting in the formation of water and a new ionic compound (a salt), which is usually soluble. Neutralization reactions are exothermic, and are generally accompanied by a noticeable release of heat.

Gas Forming Reactions

In these reactions one of the products after the double displacement is in the gaseous state (AD or CB). One such example is hydrogen sulfide (H_2S). However, one of the products could also be carbonic acid (H_2CO_3) or sulfurous acid (H_2SO_3). Both carbonic acid and sulfurous acid are unstable and will decompose to form carbon dioxide and sulfur dioxide gases, respectively:

Carbonic acid
H_2CO_3 (*aq*) → H_2O (*l*) + CO_2 (*g*)

Sulfurous Acid
H_2SO_3 (*aq*) → H_2O (*l*) + SO_2 (*g*)

Writing Equations for Reactions

• Write the correct formulas for each reactant and place a yield arrow (→) after the last reactant. • Identify the reaction type (single or double displacement)
• If you determine that a reaction will occur, write the correct formula(s) of the products after the arrow. If you determine that a reaction will not occur, simply write “no reaction” after the arrow
• Balance the equation
• Be sure to include the physical states of all reactants and products in your final equation.

Safety

Be especially cautious when using the 6 M acid and base solutions as they can burn your skin. Also be aware that skin discoloration will result from contact with $AgNO_3$.

Chemicals
Metals: Aluminum, Copper, Zinc, Magnesium
Solutions: 6M sodium hydroxide 6M hydrochloric acid

All of the following solutions are 0.1 M: barium chloride, copper(II) sulfate, iron(III) chloride, silver nitrate, sodium carbonate, sodium hydroxide, sodium phosphate, sodium sulfate, zinc nitrate.

Equipment
10 small test tubes test tube rack Parafilm
plastic droppers

Solubility Guidelines:

Ions that form soluble compounds (aq)…….…… except when combined with …….	
Group I ions (Na^+, Li^+, K^+, etc)	no exceptions
Ammonium (NH_4^+)	no exceptions
Nitrate (NO_3^-)	no exceptions
Acetate (CH_3COO^-)	no exceptions
Hydrogen carbonate (HCO_3^-)	no exceptions
Chlorate (ClO_3^-)	no exceptions
Halides (F^- , Cl^- , Br^-)	Pb^{2+}, Ag^+ and Hg_2^{2+}
Sulfate (SO_4^{2-})	Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Hg_2^{2+} and Pb^{2+}

Ions that form insoluble compounds (s)…….except when combined with….	
Carbonate (CO_3^{2-})	group I ions (Na^+, Li^+, etc) or NH_4^+
Chromate (CrO_4^{2-}) NH_4^+	group I ions (Na^+, Li^+, etc) or Ca^{2+}, Mg^{2+} or
Phosphate (PO_4^{3-})	group I ions (Na^+, Li^+, etc) or NH_4^+
Sulfide (S^{2-})	group I ions (Na^+, Li^+, etc) or NH_4^+
Hydroxide (OH^-)	group I ions (Na^+, Li^+, etc) or Ca^{2+}, Mg^{2+}, Sr^{2+} or NH_4^+

Activity Series Table

$Li_{(s)} \rightarrow Li^{1+}_{(aq)} + 1e^-$ **Highest Activity (easily oxidized)**
$K_{(s)} \rightarrow K^{1+}_{(aq)} + 1e^-$
$Ba_{(s)} \rightarrow Ba^{2+}_{(aq)} + 2e^-$
$Ca_{(s)} \rightarrow Ca^{2+}_{(aq)} + 2e^-$
$Na_{(s)} \rightarrow Na^{1+}_{(aq)} + 1e^-$
$Mg_{(s)} \rightarrow Mg^{2+}_{(aq)} + 2e^-$
$Al_{(s)} \rightarrow Al^{3+}_{(aq)} + 3e^-$
$Mn_{(s)} \rightarrow Mn^{2+}_{(aq)} + 2e^-$
$Zn_{(s)} \rightarrow Zn^{2+}_{(aq)} + 2e^-$
$Cr_{(s)} \rightarrow Cr^{3+}_{(aq)} + 3e^-$
$Fe_{(s)} \rightarrow Fe^{2+}_{(aq)} + 2e^-$ / $Fe_{(s)} \rightarrow Fe^{3+}_{(aq)} + 3e^-$
$Cd_{(s)} \rightarrow Cd^{2+}_{(aq)} + 2e^-$
$Ni_{(s)} \rightarrow Ni^{2+}_{(aq)} + 2e^-$
$Sn_{(s)} \rightarrow Sn^{2+}_{(aq)} + 2e^-$
$Pb_{(s)} \rightarrow Pb^{2+}_{(aq)} + 2e^-$
$H_{2\,(g)} \rightarrow 2\,H^{1+}_{(aq)} + 2e^-$
$Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$
$Ag_{(s)} \rightarrow Ag^{1+}_{(aq)} + 1e^-$
$Hg_{(l)} \rightarrow Hg^{2+}_{(aq)} + 2e^-$
$Au_{(s)} \rightarrow Au^{3+}_{(aq)} + 3e^-$ **Lowest Activity (not easily oxidized)**

Procedure

- Always reuse clean test tubes that have been rinsed with ***distilled water***. The test tubes do not have to be dry. At the end of this experiment, do a final rinse of your test tubes and return them to the test tube bin.
- Use approximately 2 mL quantities of all solutions.
- For reactions involving metals, use just 1-2 pieces of each metal. Place the metal in the test tube first, and then add the solution. The metal should be completely immersed in the solution used.
- Perform the following reactions and record your observations in your notebook worksheet, in the order shown below. Note that some redox reactions take longer than others. If results are not obtained immediately, give the reaction some time. You may have to warm the test tube contents in the hot water bath on the side bench.

All waste is to be disposed of in the waste container in the hood.

1. Aqueous barium chloride + aqueous sodium sulfate

2. Aqueous sodium phosphate + aqueous barium chloride
3. Aqueous sodium phosphate + aqueous copper(II) sulfate
4. Aqueous sodium phosphate + aqueous iron(III) chloride
5. Aqueous sodium phosphate + aqueous zinc nitrate

6. Aqueous sodium hydroxide + aqueous barium chloride
7. Aqueous sodium hydroxide + aqueous copper(II) sulfate
8. Aqueous sodium hydroxide + aqueous iron(III) chloride
9. Aqueous sodium hydroxide + aqueous zinc nitrate

Reactions 6-9 use 0.1M NaOH

10. Hydrochloric acid (6 M) + aqueous sodium hydroxide (6 M)

11. Aqueous sodium carbonate + aqueous barium chloride
12. Aqueous sodium carbonate + aqueous copper(II) sulfate
13. Aqueous sodium carbonate + aqueous iron(III) chloride
14. Aqueous sodium carbonate + aqueous zinc nitrate
15. Aqueous sodium carbonate + aqueous sodium sulfate

16. Aluminum metal + hydrochloric acid
17. Copper metal + hydrochloric acid
18. Magnesium metal + hydrochloric acid
19. Zinc metal + hydrochloric acid

Caution with the Mg and Zn metals. Use large test tubes and only 1 piece of metal.

20. Aluminum metal + iron(III) chloride
21. Copper metal + iron(III) chloride
22. Magnesium metal + iron(III) chloride
23. Zinc metal + iron(III) chloride

24. Aluminum metal + zinc nitrate
25. Copper metal + zinc nitrate
26. Magnesium metal + zinc nitrate

27. Aluminum metal + copper(II) sulfate
28. Magnesium metal + copper(II) sulfate
29. Zinc metal + copper(II) sulfate

30. Copper metal + aqueous silver nitrate

Name: _____

RCBC CHE116 Notebook Worksheet for Exp. 6: Chemical Rxns and the Activity Series

Section 1

Experiment work was performed on _____.

Lab Partner: _____.

Section 2 Raw Data

The data in this section is identified by the reaction numbers listed on pages 53-54.

Reaction Number	Observation
1	
2	
3	
4	
5	
6	
7	
8	
9	
10	
11	
12	
13	
14	
15	
16	
17	
18	
19	
20	
21	
22	
23	
24	
25	
26	
27	
28	
29	
30	

Section 3: Experimental Procedure

The chemical reactions listed in the experiment document were performed. Approximately _____ mL of each solution was used for each chemical test. When a reaction involving a metal was done, just _____ pieces of the metal was used. The reactions were done in the order shown in the experiment document. The reactions are identified by the reaction _____ listed on pages 53-54.

Section 4: Data Table

Reaction Number	Chemical formulas of the two reactants (Show ionic compounds as separate ions.)
1	$Ba^{2+} + 2\ Cl^{1-} + 2\ Na^{1+} + SO_4^{2-}$
2	
3	
4	
5	
6	
7	
8	
9	
10	
11	
12	
13	
14	
15	
16	
17	
18	
19	
20	
21	
22	
23	
24	
25	
26	
27	
28	
29	
30	

Section 5: Results Write a balanced chemical equation for only the reactions listed in this table. Use the net ionic equation format. The equations must be balanced.

Reaction Number	Balanced Chemical Equation (use the Net Ionic Equation format) *(Note: There are only 20 reactions listed here.)*
1	$Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \rightarrow BaSO_{4(s)}$
2	
3	
5	
8	
9	
10	
12	
13	
16	
18	
20	
22	
23	
24	
26	
27	
28	
29	
30	

Section 6: Conclusion

Did the observations made in lab class match with what was listed in the Solubility tables and the Activities Series table? If not, why not? Explain your answer.

Experiment 7: Solution Preparation

Purpose

Scientists in many different fields use solutions for their laboratory work. Solutions have a solute dissolved in a solvent, and the mixture is homogeneous. For our general chemistry laboratory experiments, the solutions are aqueous, meaning that the solute is dissolved in water.

This experiment is an exercise in preparing solutions, starting with the preparation of a stock solution. Serial dilution solutions will then be prepared with the stock solution. These solutions will let you practice solution preparation. A spectrophotometer will be used to measure the amount of light absorbed by your solutions. You will also need to prepare standard solutions (solutions of known concentration) to make a calibration curve that lets you determine the concentration of a red dye solution of unknown concentration. This will let you "see" how well your solutions were prepared, and to check the linear working range of the spectrophotometer.

Background

Solution preparation will involve the dilution of the concentrated stock solution. The dilution equation will be used to determine the concentration of the new, diluted solution.

$(M_c)(V_c) = (M_d)(V_d)$ where M_c is the concentration of the concentrated stock solution, V_c is the pipeted volume of the stock solution, M_d is the concentration of the new diluted solution, and V_d is the volume of the new diluted solution. The concentration unit does not have to be molarity. However, M_c and M_d must be the same concentration unit. V_c and V_d do not have to be in the unit of liters, but they both have to be the same volume unit.

The absorbance of light by the solute in solution will follow the Beer-Lambert Law (Beer's Law), which states that the absorbance of light is directly proportional to its concentration. The mathematical equation is $A = abc$, where A is absorbance, a is the molar absorptivity coefficient, b is the path length, and c is concentration.

Calibration curves are graphs that show the relationship between concentration and absorbance of light by the solute. Absorbance is dependent on the concentration of solute in the solution. Therefore, concentration is on the x-axis and absorbance is on the y-axis. The data points obtained in lab are plotted using Excel or Google Sheets. The computer program will put a best-fit line through the data points. The equation of the linear line is obtained, and this equation shows the mathematical relationship in the

$y = mx + b$ format which translates to Absorbance = m(concentration) + b

(Note: The b in this mathematical equation is not the same b as in Beer's Law. b is the y-intercept, and m is the slope of the best-fit-line.)

The spectrophotometers that are in the general chemistry laboratory cover the wavelength range of visible light. Therefore, in order to measure the absorbance of light by the solute in solution, the solute must absorb visible light. So, we work with colored solutions.

Chemicals

Red food dye #1 (liquid form) Deionized water

Equipment

Pipets and pipettors	Volumetric flasks, 100 mL	Spectrophotometer
Parafilm	Plastic droppers	Test tubes for spectrophotometer
Beakers, small sizes	Laboratory balance	

Procedure

Part A: Preparation of the stock solution

(Save this solution until the end of the experiment.)

1) Bring your laboratory notebook, pen, plastic dropper, and a clean, dry, 150 mL beaker to the laboratory balance.
2) Place the beaker on the balance pan, and tare the balance so the beaker has a mass of zero grams.
3) Use your clean plastic dropper to put 0.25 g of liquid red dye into your beaker. (This will be approximately 6 drops.) Record the exact mass of the red dye in your notebook worksheet. Do not tare the balance again.
4) Remove the beaker from the balance pan, and add approximately 40 mL of deionized water using a graduated cylinder.
5) Return the beaker to the balance pan, and gently add deionized water to the beaker until the total mass in the beaker is 50 g. The mass does not have to be exactly 50 grams, but the exact total mass must be recorded in your notebook worksheet. (Use all digits displayed on the balance.)
6) Return to your lab bench. Use a clean, dry, glass stir rod to mix your stock solution well.
7) Calculate the concentration of the stock solution, in units of % wt.

Part B: Serial Dilutions

1. Use a calibrated pipet to measure and dispense your stock solution for this part of the experiment (*do not use a plastic dropper*). Rinse the pipet with your stock solution, and then dispense the rinse into a waste beaker.

2. Prepare serial dilution (1). Draw up your stock solution into the pipet and adjust the bottom of the meniscus to pipet exactly 5.0 mL. Dispense this volume to a 100 mL volumetric flask. (The flask should already have been rinsed with DI water. The flask does not have to be dry, since you will be adding water to it.) Dilute the contents of the flask with DI water until the bottom of the meniscus sits exactly on the calibration mark on the neck of the flask. The total volume of solution in the flask will now be 100.0 mL. Cover and seal the flask with Parafilm. Invert to mix fifty times.

3. Prepare serial dilution (2). Rinse the pipet with DI water, and then rinse with your serial dilution (1) solution. Draw up your (1) solution into the pipet and adjust the bottom of the meniscus to pipet exactly 5.0 mL. Dispense this volume to a 100 mL volumetric flask. Dilute the contents of the flask with DI water until the bottom of the meniscus sits exactly on the calibration mark on the neck of the flask. The total volume of solution in the flask will now be 100.0 mL. Cover and seal the flask with Parafilm. Invert to mix fifty times.

4. Prepare serial dilution (3). Rinse the pipet with DI water, and then rinse with your serial dilution (2) solution. Draw up your (2) solution into the pipet and adjust the bottom of the meniscus to pipet exactly 5.0 mL. Dispense this volume to a 100 mL volumetric flask. Dilute the contents of the flask with DI water until the bottom of the meniscus sits exactly on the calibration mark on the neck of the flask. The total volume of solution in the flask will now be 100.0 mL. Cover and seal the flask with Parafilm. Invert to mix fifty times.

   Go to Part C, then come back for step 5

5. After you have completed Part C, determine the absorbance for each of the serial dilution solutions.
   a. Make sure the spectrophotometer is set to your λ_{max} for the wavelength.
   b. Use the DI water as the blank to zero the spectrophotometer.
   c. Transfer each serial dilution solution into its own test tube, about ¾ full.
   d. Measure the absorbance of light for each solution, and record these values into your notebook worksheet.

Part C: Determining the λ_{max}

(finding the best wavelength for absorbance of light by red dye)

1) Use the Serial Dilution (1) solution for this part of the experiment. Place some of this solution into a small test tube, about ¾ full. Put DI water into another test tube, also about ¾ full. The DI water is your blank that will be used to zero the spectrophotometer.
2) Red dye looks red because it absorbs light in the blue-green region of visible light (see Figure 1). You need to determine the best absorbance. Turn on the spectrophotometer and set the wavelength to 470 nm.

Figure 1

Coloured absorbed	Visible (Complementary)
Violet	Yellow-Green
Blue	Yellow
Green-Blue	Orange
Blue-Green	Red
Green	Purple
Yellow-green	Violet
Yellow	Blue
Orange	Green-Blue
Red	Blue=Green

https://www.slideshare.net/GranchBerheTseghai/2-fundamentals-of-color

3) Put the test tube with the DI water into the sample compartment of the spectrophotometer, and close the lid. Press the zero Abs button. This sets the zero point on the spectrophotometer, which subtracts out any absorbance of light by the glass test tube and DI water.
4) Put the test tube of red dye solution into the sample compartment and close the lid. Record the absorbance readout in your notebook.
5) Repeat steps 3 and 4 after increasing the nm setting by 10 nm. Continue to do this until you have absorbance readings up to 540 nm. You must set the blank each time the nm setting is changed.
6) Report the best wavelength to use, λ_{max}, when measuring the absorbance of light by the red dye solution.

Return to Part B, step 5

Part D: Preparing Standard Solutions and a Calibration Curve

Solutions of known concentration will be prepared to make a calibration curve.
Use the stock solution you prepared in Part A to prepare these standard solutions.

1) Standard Solution 1: Pipet 1.0 mL of stock solution into a 100 mL volumetric flask and dilute to the mark with DI water. Seal the flask and invert 50 times to mix.
2) Standard Solution 2: Pipet 2.0 mL of stock solution into a 100 mL volumetric flask and dilute to the mark with DI water. Seal the flask and invert 50 times to mix.
3) Standard Solution 3: Pipet 3.0 mL of stock solution into a 100 mL volumetric flask and dilute to the mark with DI water. Seal the flask and invert 50 times to mix.
4) Standard Solution 4: Pipet 4.0 mL of stock solution into a 100 mL volumetric flask and dilute to the mark with DI water. Seal the flask and invert 50 times to mix.

Measure the absorbance of light for each standard solution and the unknown solution.

1) Make sure the spectrophotometer is set to your λ_{max} for the wavelength.
2) Use the DI water as the blank to zero the spectrophotometer.
3) Transfer each standard solution and the unknown solution into its own test tube, about ¾ full.
4) Measure the absorbance of light for each solution, and record these values into your notebook.

Prepare a calibration curve.

Follow the instructions in the Calibration Curve video and use Excel or Google Sheets to prepare a calibration curve. The concentration is the independent variable, so that is your X axis. Absorbance is the dependent value, so that is the Y axis. Plot a scatter plot, linear, and put the best-fit line on the graph. Show the equation of the line and the R^2 value on the graph also. The R^2 value gives some indication as to how well the standard solutions were prepared. **Print this graph and attach it to your worksheet.**

Determine the concentration of your unknown.

Use the equation of the best-fit line to solve for the unknown concentration. The equation of the line is in the y=mx+b format. Y is the absorbance and X is the concentration. You know the absorbance of the unknown solution, so solve for X. Show this calculation in your notebook worksheet, and report the concentration of the unknown with correct units.

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Name: _____

Section 1

Experiment work was performed on _____.

Lab Partner: _____.

Section 2 Raw Data

Part A

Mass of red dye in the tared beaker: _____.

Mass of red dye and water in the tared beaker: _____.

Part B

Size of volumetric flasks used: _____.

Serial dilution (1): volume of stock solution pipetted into volumetric flask: _____.

Serial dilution (2): volume of serial dilution (1) pipetted into vol. flask: _____.

Serial dilution (3): volume of serial dilution (2) pipetted into vol. flask: _____.

(come back here to finish part B after you do part C)

Wavelength set on spectrophotometer, λ_{max}: _____.

Solution	Absorbance
Serial dilution (1)	
Serial dilution (2)	
Serial dilution (3)	

Part C

Solution used for this part: _____ What was used as the blank: _____

Wavelength, nm	Absorbance		Wavelength, nm	Absorbance
470			510	
480			520	
490			530	
500			540	

Part D

Size of volumetric flasks used: _____.

Standard solution 1: volume of stock solution pipetted _____mL, abs. _____.

Standard solution 2: volume of stock solution pipetted _____mL, abs. _____.

Standard solution 3: volume of stock solution pipetted _____mL, abs. _____.

Standard solution 4: volume of stock solution pipetted _____mL, abs. _____.

Unknown solution: abs. _____

Wavelength used: _____, what was used as the blank: _____.

Section 3: Experimental Procedure

Part A: The red food dye stock solution was prepared by combining _____ g of red food dye #1 with enough deionized water in a small beaker to make _____ g of solution. The solution was mixed well and labeled as the stock solution.

Part B: The stock solution from Part A was used to prepare the first of four serial dilution solutions. The volumetric flasks used were _____ mL in size. A _____mL calibrated pipet was used to deliver the solutions into their volumetric flask. Serial dilution (1) was prepared with _____ mL of the stock solution. Serial dilution (2) was prepared with _____ mL of serial dilution (1) solution. Serial dilution (3) was prepared with _____ mL of serial dilution (2) solution. Each solution was diluted to the calibration mark of the volumetric flask with deionized water and mixed well before being used to prepare the next solution.

Part C: The wavelength of maximum absorbance by the red dye was determined using the _____ solution. The wavelength range use was _____ nm to _____ nm. The λ_{max} was determined to be _____ nm.

Part B (continued): The absorbance of light at λ_{max} was determined for each of the serial dilution solutions. The blank used to zero the spectrophotometer was _____. The absorbance values were listed in the raw data section. The blank was used to zero the spectrophotometer only once for Part B.

Part D: The standard solutions were prepared by pipetting the appropriate amount of stock solution into _____ mL volumetric flasks. For standard solution 1, _____ mL of stock solution was used. For standard solution 2, _____ mL of stock solution was used. For standard solution 3, _____ mL of stock solution was used. For standard 4, _____ mL of stock solution was used. Each solution was brought to volume with _____, and mixed well. The absorbance of _____ nm light was measured for each solution. The blank used at the beginning of Part D was _____. The absorbance data was recorded in the raw data section.

Section 4: Data Tables

Part A: Preparation of the Stock Solution

Mass of Red Food Dye #1 used: _____

Mass of Red Food Dye + DI Water: _____

%Mass of the stock solution: _____

Part B: Serial Dilutions

Solution Name	Concentration	Absorbance

Part C: Determining the Wavelength of Maximum Absorbance, λ_{max}

Wavelength, nm	Absorbance		Wavelength, nm	Absorbance
470				

The solution used for Part C was _____

The λ_{max} was determined to be _____ nm.

Part D: Standard Solutions and a Calibration Curve

Solution Name	Concentration	Absorbance
	to be calculated with eqn of the line	

The concentration of the unknown was calculated to be _____%mass.

Section 5: Calculations and Results

Part A: Calculation of the stock solution concentration

Part B: Calculation of the concentration of each serial dilution solution made:

(Show your calculations.) Use $M_c \cdot V_c = M_d \cdot V_d$

Serial Dilution 1	Serial Dilution 2
Serial Dilution 3	

Part C: Determining λ_{max} (no calculations are needed)

Part D: Calculation of the concentration of each standard solution, making the calibration curve, and calculating the concentration of the unknown solution.

(Print out your calibration curve graph, make sure the equation of the line is shown on the graph, along with the R^2 value, and attach the calibration curve graph to this worksheet)

(Show your calculations.)

Standard 1	Standard 2
Standard 3	Standard 4

Equation of the best-fit-line: _____*(R^2 is not part of the eqn.)*

Absorbance of the unknown solution: _____

Calculation of the concentration of the unknown solution:

Section 6: Conclusion
Part A: No conclusion is necessary.

Part B: Was it possible to measure the absorbance of each of the three solutions, or were any solutions too dilute or too concentrated? *If the readout was flashing, the solution was too concentrated. If the readout was unsteady, the solution was too dilute.* Did the absorbance of each solution differ by a factor of twenty, like the concentrations did (5 mL to 100 mL is a dilution by a factor of twenty). Explain your answers.

Part C: No conclusion is necessary.

Part D: Did the absorbance vs. concentration calibration curve graph yield a straight line? What was the R^2 and b values and what does that tell you about the quality of your solution preparation? Does the calculated concentration of the unknown make sense when you compare the %mass and absorbance of the unknown to the %mass and absorbance of the standard solutions? Explain your answers.

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