GENERAL CHEMISTRY INQUIRY LAB

INSTRUCTOR EDITION

FIELD GUIDE AND SOLUTIONS

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PREFACE

The philosophy behind these labs is based on how students learn and the progression of scientific understanding. The methodology is supported by research. The instructor has a lot of scope in making the students active thinkers in the discipline. The important features of these labs are summarized below.

- Experimental data is the key to scientific understanding. It moves students away from fact-verification as the main purpose of doing a lab.
- The lab procedure is somewhat of an outline and it is NOT step-by-step instructions that result in a cookbook experience with no-learning.
- Experiments are trying to answer an important question. Every experiment is framed as a question that may be very relevant to what we see every day.
- Experiments can and will fail. In order to do successful experiments, silly mistakes can derail the outcome. The key is to make observations that can help correct mistakes. Students are given the scope to make mistakes and get them corrected under the instructor's supervision.
- The process of scientific inquiry - observation, prediction, experimentation, and data analysis to support/re-evaluate prediction & formulate a theory to explain data are emphasized by each lab experiment.
- A close connection to content from General Chemistry lecture is maintained, but these labs do not expect students to have learned specific content in lecture BEFORE coming to the lab. The nature of discussions amongst students and instructor in the lab setting brings out the learning objectives and provides support for the lecture material.
- The instructor plays a major role in the student's inquiry process by directing their queries in a way to help them solve their own questions without providing direct answers.
- Since it is the first lab for students in a college level course, safety is an important issue in the lab. The instructor can be very specific about safety rules and students are expected to come up with their own written procedures, but get it approved by the instructor.
- Whether they work within their group or interact with the entire class, it is important to emphasize communication skills. Many times students realize that they know the material but keep second-guessing their own ability. It is the repeated opportunity provided in a group setting that helps them realize their potential in asking and answering questions.
- By designing these labs as an inquiry-based group-activity laboratory, instructors can help model the student behavior in learning - form learning communities, ask questions without inhibition, read for comprehension, listen for understanding, and speak properly to gain acceptance. This can have an improvement in their entire academic performance beyond a one-semester chemistry course.

Suggested Lab Schedule

Week	Experiment #	Lab Activity
1		Check-in, Safety Orientation
2	1	Density Lab
3	1	Density Lab Week 2
4	2*	Calcium Lab
5	2*	Calcium Lab Week 2
6	3*	Copper lab
7	3*	Copper lab Week 2
8	4	Household Antiseptic Lab
9	5	Gas Laws
10	5	Gas Laws Week 2
11	6	Acid Spill Lab
12	6	Acid Spill Lab Week 2
13	7	Calorimetry Lab
14		Check -out

*: It is preferable to do Calcium Supplement Lab before Copper Lab. We tried both versions and found that this works better as it gives a lot of learning time for writing ionic compounds.

Experiment #1: Density

FOR THE INSTRUCTOR

Chemicals required: none.

Instruments/material: 4 different size polymer pieces, balance, graduated cylinder, ruler, Erlenmeyer flask, and paper towel.

Information for stockroom personnel

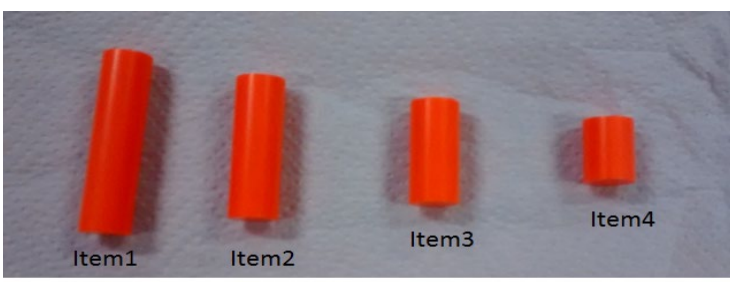
COVERS 2 LAB PERIODS

Students are grouped in units of 4 students and hence a minimum of 6 instruments required per class. Certain units are required in extra quantities as noted below.

Chemicals required: none

Instruments/material:

1. Polymer pieces of 4 different sizes – Six units required



2. Balances that read up to 2 decimals – six units required
3. 100 mL graduated cylinders – 12 units required
4. 1 ft rulers – 12 units required
5. 250 mL ground-glass joint flask with lid – 6 units required
6. Paper towel – 3 pack for the entire lab.

Guide for Instructors

The idea of these labs is to allow students to make mistakes and then figure out by discussing with each other what may cause those errors. It is a way for them to analyze and clear their misconceptions with the help of fellow students in their group and in class. Questions are framed in a way to help them achieve that. Along the way, you can help them without giving answers directly. This guide also helps you to focus on some difficulties experienced by students when we tried the labs.

Experiment #1: Density

While this experiment is simple, it is very important that this lab runs smoothly and that students get a sense of what is expected of them and how the next labs will run. You must set the tone in this session.

Learning Objectives:

Content objectives
Students will be able to:

- Associate the instrument name with the physical quantity measured and its units
- Determine mass, length, volume as extensive and density as intensive property for a solid material
- Recognize that precision is based on the instrument used and the digits reported in a measurement tells about the error in a recorded measurement
- Calculate how errors propagate during data analysis/calculations that was derived from various measurements
- Understand the importance of performing multiple trials to verify the results.

Process objectives
Students will be able to:

- Design different experimental methods for determining the density
- Analyze the data for density from different methods and compare their advantages and disadvantages
- Discover the level of precision from different methods to the requirements set out by the management
- Develop logical oral and written communications necessary to convince non-scientists about the best method selected
- Compare the accuracy and efficiency of the 3 different techniques used to find density.

Overview

The experiment is divided into three components, each demonstrating a method by which a student can obtain the density of a regularly shaped object with different levels of precision. The mass of any experimental object needs to be taken only once as the mass already has enough precision. Therefore, each method focuses on obtaining the volumes of each object with different levels of precision.

Some experimental and theoretical issues that can come up are listed below. The goal is to allow students to go through the discovery process by themselves and own their learning. Our guidance comes from initiating discussions within a group with probing questions. Sometimes it is advantageous to have the entire class discuss and offer solutions that can enlighten everyone.

While the labs may seem simple and straightforward to you, students often come to class unsure of what to do and what the goal of the lab is. It is important to spend 5-15 minutes at the start of each class to go over important content and give students a general outline of the lab. Students are enrolled in different lecture courses and may not have covered the content needed to understand the lab.

Goal: Determine the best method to verify the density of the plastic polymer.

Things to look out for:

1. When circulating and questioning groups, pose questions to the student whom you feel is lost/relying too much on their group.
2. Students are often general in their procedure. Stating "The volume was found by finding the length and diameter." is not enough. Have students explain how they found the diameter and length. Did they stand up the cylinder and align it with the ruler? Did they start at the end of the ruler or at the zero mark? These details are important to include so that anyone could repeat the same procedure and obtain very similar results.
3. The third technique is always confusing for students and even to instructors who have not used this technique themselves. A detailed explanation of the pycnometer technique:

A sample lesson outline is provided below:

The balances students will use measure up to the 100th place. Thus, using mass to find the volume of the cylinder is a way of increasing the precision of the measurement. The difference in mass between the stoppered Erlenmeyer flask completely filled with water and the stoppered Erlenmeyer flask with the cylinder inside is measured. The mass of the displaced water can be converted to volume of displaced water with the same precision. This volume is equal to the volume of the cylinder.

The first set of questions deals with measuring the mass of an object to the correct number of digits. It is important to check if they are all doing the measurement. The difference in mass between the stoppered Erlenmeyer flask completely filled with water and the stoppered Erlenmeyer flask with the cylinder inside is measured. The mass of the displaced water can be converted to volume of displaced water with the same precision. This volume is equal to the volume of the cylinder.

- **Density through Fluid Displacement in a graduated cylinder. [**During lab 1, randomly pick groups to use different size blocks if you wish. Since it is discovery process during lab 1, tell them to stick with one piece and make sure they agree on their readings and procedure. *Note that if there is not adequate time, you can use lab 2 to complete all 4 pieces.*]
    - Requires a graduated cylinder (50 or 100 mL)
    - Students should fill the graduated cylinders with 30 mL of tap water and insert the object and record the changes in volume.
    - In both steps, they are likely to record the numbers incorrectly and also subtract forgetting the rules of significant digits.

- An experimental error also includes water splashing out when they drop the object. It may be worth making them think if this water loss while dropping the object increase or decrease the measured density over true value. [This is posed as an instant question, allow the groups to write the answer and you can decide to discuss it if they are all wrong or if there are some incorrect answers. Like in all inquiry-based activities, the answer is not directly given. Probing questions are given to make them arrive at the answer OR make the students explain their own right and wrong answers and allow them to see their own errors]
- Depending on the precision of mass and volume, they are expected to have appropriate values for density. If there are errors, bring it through discussions and allow them to make corrections
- Review the written procedure from each group and see if they have the necessary steps in the correct order, and that they've taken care of mistakes that were revealed during the discussion.

[Performed in lab period 2]

- **Density through Fluid Displacement in a Pycnometer**
  - *Background: A pycnometer is a device that allows for the determination of an object's density when it is suspended within the solution in a closed environment. A pycnometer functions by the concept of relative density, as the density of the suspension fluid is used to calculate the volume and density of the suspended object. This may be a bit difficult for the student, but it is important for them to finally get the idea that the precision of the above two methods is compromised because of poor precision in measuring the volume and not the mass. This method allows them to get a measure of volume with the same precision as the mass because the volume is indirectly measured using the mass of the fluid displaced.*

  - For the purposes of our experiment, a flask with a ground glass stopper (or a rubber cork with a capillary tube) is filled with water. Whenever the flask is filled, it is important to avoid the formation of air bubbles:
    - Pour slowly down the sides of the flask
    - Fill to the brim of the flask before inserting the stopper and then slowly insert it.
      - Excess water is expected to spill out.
      - Students should have access to trays (or work near the sink) and paper towels to contain spillage.
  - The mass of the **Dry Flask** is obtained before adding any water.
  - Empty the flask, insert an object and refill the flask as described above.
  - Reweigh the flask making sure it is dry outside. [They can reconfirm by keeping it in the air for air-drying and also wiping it again with a paper towel and reweighting.
  - Remove object and repeat the previous two steps for all pieces.

- Expect 3-4 significant figures for the final density, depending on the mass of the measured object.

Review the procedure.

In lab 1, they are expected to complete the first method and have procedures written for the second method. Some students may be able to complete data for the second method, but do not rush them to the pycnometer method until lab period 2.

During lab 2, they complete method 3, [illegible] repeat any experiment necessary by taking care of errors, digits, etc. Calculate the density of all 4 objects. Write them on the board using a format as shown below. Lead a class discussion on the number of digits reported by each method by different groups, their consistency, etc. and eventually pointing towards an idea where the best method for the required precision may be achieved and what errors may have to be avoided. As mentioned earlier, the discussions center on bringing the answers from students and making them critically think.

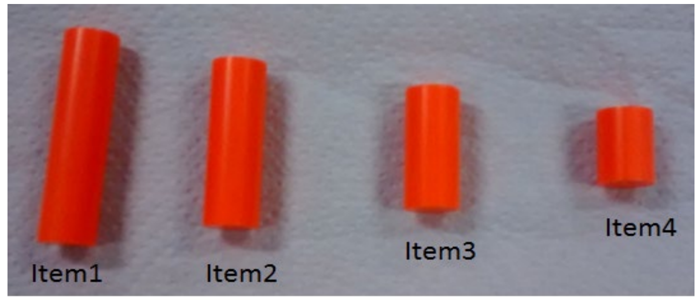
Make sure in the final discussion they are able to figure out what size piece is ideal, what technique is ideal, what are the time requirements for each technique.

The table below should be provided for entire the class to view and can be put on the whiteboard OR projected on a screen while discussions take place.

Group	Item 1 (largest piece)	Item 2	Item 3	Item 4 (smallest piece)	Average time per piece
A Method 1 Method 2 Method 3					
B Method 1 Method 2 Method 3					
C Method 1 Method 2 Method 3					
D Method 1 Method 2 Method 3					

Experiment #1: Density

Your Project:



You are an analytical consultant for a plastics company. To produce a consistent product for the marketplace, the plastic should have a density of 1.38 $\pm$ 0.01 g/mL. Management is interested to know what sample size and what method will work the best for validating their product before shipping out to customers. You are given 4 samples that are part of the same batch.

Management believes something that is easy and inexpensive like density should be used to validate the purity of the product. Prepare a report to the management outlining the different methods available to measure density and your recommended choice based on your results. The given materials in the lab allow you to measure density by three different methods. Show complete procedure, actual results, its reproducibility, how to minimize errors, estimated labor time, material cost, etc., so that you are able to recommend your method to management.

📖 **Recommended Readings:** Units and measurements

Background:

You are doing a guided activity laboratory with your group members and sharing the results with the entire class as and when required to complete the lab. The final laboratory report is submitted at lab 3. There will also be a laboratory quiz during the following week (see schedule).

In order to prepare for this lab:

1. Review the recommended readings listed above.
2. Print this document. **Answer the pre-lab questions before you come to the lab.** You are expected to submit pre-lab before the lab starts.
3. Since you are expected to answer in-lab questions with your group, you are expected to stop reading after the description of the activity.

PreLab – Exp #1: Density **Name: ______________________________**

1. What are the SI units for mass, length, and temperature?

   Mass:___________ Length: ___________Temperature: ___________

2. The length of a rectangular field is 250 ft. Express this length in cm units. [1 ft = 30.54 cm]

3. The length and width of a rectangular field are 250 ft and 150 ft. What is its area in SI units?

4. A cylinder has a diameter of 10 cm and a height of 20 cm. What is the volume of the cylinder? [Volume of cylinder = $\pi r^2 h$]

5. A piece of metal weighed 25.0 g when placed on a balance. When placed on another balance, the same metal weighed 25.5 g. How can you ascertain what is the true mass of the metal?

6. Label the following [choices: stoppered flask, ruler, balance, graduated cylinder}

Goal: Determine the best method to verify the density of the plastic polymer.

Safety	**Materials**
Always wear <u>goggles</u>, a <u>lab coat</u>, and <u>gloves</u> inside the lab. You will be penalized if you fail to do so.	**Chemicals needed:** Water **Instruments needed:** 50-mL graduated cylinder, balance, ruler, stopper flask.

[**INSTRUCTORS NOTES:** Review the following with the class:

1. Today you are working for a plastic making company. You need to verify the purity of the plastic you made. There are several different ways to measure the purity of a product. Today you will use density to determine purity.
2. Go over density formula: density = mass / volume.
3. You will calculate density using three different techniques and compare how accurate/efficient each technique is. The mass is measured the same way for each technique (using a balance). Thus, each method only differs in the technique used to measure the volume.
4. Discuss the difference between accuracy and precision.
5. Remind students to keep track of how long each technique takes as efficiency is an important factor they must take into account.]

Experimentation

Focus Question 1: Consider volume, mass, density, and predict each one as intensive or extensive property.

Make your prediction:

One of the important requirements of lab work is prediction and experimentation. Since this is your first lab, you will be guided through some activities which will familiarize you with the instruments and measurements before you can experiment correctly with them.

(i) Observe the materials given and state some of the physical properties you observe - shape, color, physical state, and smell.

(ii) Measure the length and diameter of the largest sample piece using the appropriate equipment from the lab. Allow each of your group member to make the measurement before comparing with each other.

Length:__________ Diameter:__________

(iii) Compare the number with each other. If your numbers differ from each other, discuss if they are acceptable or not acceptable. If they are not acceptable, discuss how to solve it and write the appropriate number. Your group's accepted value for:

[**INSTRUCTORS NOTES**: The first set of questions deals with measuring the mass of an object to the correct number of digits. It is important to check if they are all doing correctly and if needed, stop the whole class and make students explain their point of view both for right and wrong answers. For example, they may get values within 0.01 unit of each other in a group and wonder who is correct. Some may forget to look for zero reading on the balance before taking the measurement. Some may not even know that the balance may be broken. We can purposely mess the calibration of one of the balance to bring out the systematic error. It is important to make students realize their source(s) of error and make it come through their own words.]

Length:_____________ Diameter:__________ Volume of the piece: ___________

(iv) ✎ Write a few sentences to indicate how the volume should be measured so that everyone will do it correctly without errors.⦸ See instructor for suggestions and approval before proceeding.

[**INSTRUCTORS NOTES:** Students should use a ruler to obtain the object's length, width, and height. When using a cylinder the diameter and radius should also be recorded.

    - Students may find it helpful to trace the length of the cylinder for an accurate length reading. However, students should **not** trace a circumference to measure the diameter. This will increase the diameter and increase overall volume, thus reducing calculated densities. [Make sure this comes out of discussions]
    - They may use different units like inches or cm. [If no one chose inches, you can probe them with questions to make sure they consciously chose 'cm' for a reason]
    - As mentioned earlier, student's may have errors in recording the value along with the incorrect number of digits. Suitable intervention as described earlier is done to make them realize those errors on their own if they are unable to solve it within their group.
    - A discussion may also be necessary after they calculate the volume as they may not use the correct formula or may not have correct digits
  - Depending on the precision of mass and volume, they are expected to have appropriate values for density. If there are errors, bring it through discussions and allow them to make corrections.
  - Review the written procedure from each group and see if they have the necessary steps in the correct order, and that they've taken care of mistakes that were revealed during the discussion.
  - You can also allow procedures read out loud by different groups with probing discussions for the need for a step OR lack of a step so that they all gather the pieces necessary to write a thorough procedure.]

(v) Now measure the volume of each piece as per your written procedure. Compare your values with your group members and make sure you reach a consensus for the volume of each piece.

	Your Value	Group Consensus
Item 1		
Item 2		
Item 3		
Item 4		

Scrap space for calculations:

(vi) Is volume the same for each object? Is volume intensive or extensive property of a material?

Volume is an extensive property.

(vii) Measure the mass of the largest piece with the appropriate instrument. If there is more than one instrument in the lab, each of you uses a different instrument to measure and come back to your group to discuss. Are your values the same, different but acceptable, or different and not acceptable? Discuss with your group members and come to a conclusion. Once you are clear, note the value obtained by you and the consensus. Clearly, indicate the reasons if your consensus value is different from your own measurement.

Mass of item 1: Your value:____________ Consensus: ___________

(viii) ✎Now, write a clear procedure for measuring the mass of the piece.

⦸ See instructor before proceeding.

(ix) Now measure the mass of each piece as per your written procedure. Compare your values with your group members and make sure you reach a consensus for the mass of each piece.

	Your Value	Group Consensus
Item 1		
Item 2		
Item 3		
Item 4		

(x) Is the mass the same for each object? Is mass an intensive or extensive property of a material?

Mass is an extensive property.

(xi) Density is a measure of mass-volume ratio. Use your group's consensus values for mass and volume and calculate density independently. Then, compare the value with your group. Discuss if your values are different and come to a consensus. For each piece, measure the mass-volume ratio and compare it with your group members. Is density an intensive or extensive property?

	Your Value	Group Consensus
Item 1		
Item 2		
Item 3		
Item 4		

If your value and the consensus were different, write the reasons that led to the consensus. Make sure this is also incorporated in your final lab write-up.

(xii) Is density an intensive or extensive property of a material?

Density is an intensive property.

✎[Now take a moment to write all the steps clearly that will help your client measure the density of the regular polymer sample you are supplying. *Use a separate piece of paper*. Only one procedure is required per group now. When you write a procedure, make sure it is clear enough so that another person may perform the same procedure after reading your notes. Call this procedure, "Density Measurement, Method 1: Ruler Method". This rough hand-written work can be properly typed at home by each group member for their own lab reports.]

⦸ See instructor for suggestions and approval.

STOP for a class discussion concerning significant figures. Ask each group for the average density of their plastic and write them on the board. Ask the students what differences they see between the values on the board. There is usually at least one group who excludes the last digit of uncertainty in their measurements. If not, ask how the students knew how many decimal places to record. Use this as a starting point for discussing the last digit of uncertainty, as well as accuracy and precision. Make sure you have this discussion with the class before they move onto Focus question 2 so they are able to accurately record their data moving forward.

Focus Question 2: How will you find the density of an irregular shaped object?

Make your suggestion:

(i) Take the 50 mL graduated cylinder and fill it with water so that it is somewhere between 30-35 mL in the cylinder. There is no need to fill it to any specific value. Each of you looks at the water level and note the value independently. DO NOT ASK FOR GUIDANCE and do not consult with your friend at this point. After everyone in your group is finished, compare numbers, discuss, and come to a consensus. If there are any discrepancies in your answers, discuss why you recorded your answer the way you did; if possible come to a consensus. Write your reasons for keeping or changing the answer.

Your value:_______________ Final after discussion:______________

(ii) Take one of the objects whose density you are measuring and drop it in the graduated cylinder. Does it sink or float? What does this mean?

(iii) Record the volume of water in the graduated cylinder. Again, compare with your group members and come to a consensus for the recorded value. Write reasons if you change the value after discussions.

Your value:_______________ Final after discussion:______________

(iv) Calculate the volume of the object. Compare with your group members and come to a consensus for the recorded value. Write reasons if you change the value after discussions.

Your value:_______________ Final after discussion:______________

(v) ✎ Write clearly the procedure for finding the volume of an object using the water displacement method.

⦸ See instructor before proceeding

(vi) Using the procedure written, find the volume of the four items. It is important that each of you do it individually and compare your numerical values along with significant digits.

☈If necessary, repeat your measurements.

Volume - Water Displacement Method

	Your Value	Group Consensus
Item 1		
Item 2		
Item 3		
Item 4		

(vii) Calculate the density of the object using this water-displacement method. Use group consensus value for mass (recorded earlier) and volume (water-displacement method). Compare with your group members and come to a consensus if your values differ with each other. Write reasons if you change the value after discussions.

Density - Water Displacement Method

	Your Value	Group Consensus
Item 1		
Item 2		

Item 3		
Item 4		

⦸ See instructor for suggestions and approval.

✎ [At this point, your instructor will review your answers and there will be a class discussion. Make a note of errors, and write a proper procedure to measure density. Use a separate piece of paper. Title this procedure "Method Two: Finding Density by Water Displacement." When you write a procedure, make sure it is clear enough so that another person may perform the same procedure after reading your notes. It is important to write the advantages of this method over the previous one so that the end user is aware when this method may be needed.]

Experimentation Week 2

⟐Focus Question 3: So far, you investigated two methods to measure density. Compare the precision. Is it possible to get volume with higher precision? Each one in your group may have 'yes' or 'no' prediction. Discuss reasons amongst yourselves and also with the class if prompted by your instructor.

(i) Now it is time to experiment. Take the Erlenmeyer flask with the appropriate glass stopper. For this class, we will call the stoppered flask, a pycnometer (Note: A pycnometer is constructed slightly differently, but this unit can do the same job for the materials we are measuring.) Measure the mass of the empty flask with stopper.

Mass of empty flask with stopper:____________

(ii) Fill the flask with water so that when the stopper is inserted the water will overflow. Make sure the water is at room temperature and there are no air bubbles. Close the flask with the stopper. Wipe all the excess water on the sides of the flask with a paper towel. Try to keep the outside dry and then take the mass again.

Mass of flask with stopper and filled with water: ____________

Mass of the water filling the entire flask: _______

(iii) Remove the stopper from the flask and drop item 1 in the flask (pycnometer) with water. Close the pycnometer with a stopper. Wipe the excess water outside with a paper towel. Keep the outside as dry as possible. Put it on a balance and weigh it.

Mass of the flask with stopper plus item 1 and water: _________

(iv) After subtracting the mass of item 1 and empty flask you had previously determined, calculate just the mass of water now filling the flask in the presence of the object.

Mass of the water filling the flask in the presence of item 1: _______

What is the mass of water displaced because of the presence of item 1: ___________

(v) The density of water at room temperature (25 °C) is 0.997044 g/mL. What is the volume of water displaced in the flask based on the mass determined above?

Volume of water displaced = ________

What is the relationship between the volume of the water displaced and the volume of the object?

(vi) Compare the precision in volume obtained by this method to previous methods. Does it agree with your prediction?

(vii) ✎Make sure you understand the steps clearly and write a proper procedure on a separate piece of paper. Call this as "Density Measurement, Method 3 Pycnometer method". ⦸ See instructor before proceeding.

(viii) Repeat the volume measurement for all items and make the density table.

Density - Pycnometer Method

	Volume	Density
Item 1		

Item 2		
Item 3		
Item 4		

(ix) Now that you have 3 methods to measure density, make a table with density values obtained for each item. [Indicate the consensus mass for your group. Indicate volume and density obtained by each of the methods.]

Density - All Methods

	Mass	Method 1	Method 2	Method 3
Item 1				
Item 2				
Item 3				
Item 4				

(x) Take a moment to compare the volume of the same material obtained by the three different methods. Explain which method has the highest precision and why.

⦸ See instructor for suggestions and approval before proceeding.

✎ [At this point, your instructor will review your answers and there will be a class discussion. You want to discuss things like the shape and size of the object suitable for easy measurement, reproducibility, advantages/ disadvantages of a particular method, etc. Make note of these as they are important for writing your final lab report.]

FINAL LAB REPORT:

After you are done with the lab, write a report detailing the experiments you performed and analyzing the data you gathered. Although you worked with your group to conduct the experiment, each of you must submit your *own* lab report. Note that data and procedure can be similar to your group, but the rest of the write-up should show your individual work. Plagiarism will get a zero for the lab and may lead to further serious actions. Your lab report must include the following items at the minimum.

Abstract: A brief summary of your finding that answers the question posed at the beginning of the lab.

Introduction:

Procedure(s): Include the procedures for the 3 methods used to find density. Make sure your procedure is detailed.

Raw Data:

Description of results and analysis:

Conclusion: Give the overall result of your experiment in a sentence or two. Explain any sources of error that may have impacted your calculated density.

[Note that your lab instructor may have detailed guidelines on how your report should be done and how it is graded]

POST LAB QUESTIONS

[Note that for numerical answers, there will be a penalty for the incorrect number of significant digits]

1. An object is dropped in a liquid having the same density. Will it sink or float? Explain.
2. Which method will you use for finding the density of an irregularly shaped rock?
3. Will density of a substance increase, decrease, or stay the same with an increase in temperature? Explain.
4. What is the mass of 2.50 mL of mercury? [density of mercury = 13.6 g/mL]
5. What is the volume occupied by a 50.00 g irregularly shaped rock piece that has a density of 2.45 g/mL?

Experiment #1: Density – **10 points**

1. Mass: ___kg___. Length: ___m___.______Temperature: __K___.[1 pt each. TOTAL= 3 pt]

2. (250 ft) x (30.54 cm / 1 ft) = 7635 cm [no penalty for sig digits. 1 pt]

3. (250 ft) x (0.3054 m / 1 ft) = 76.35 m

   (150 ft) x (0.3054 m / 1 ft) = 45.81 m

   (76.35 m) x (45.81 m) = 3498 m^2 [no penalty for sig digits. 1.5 pts. 0.5 penalty if m^2 units are not written. 0 points if the answer is not in SI units.]

4. Volume of cylinder = $\pi r^2 h$; 1570 cm^3 or 1.57E-3 m^3 [no penalty for sig digits. 1.5 pts]

5. You can ascertain the true mass of the metal by anyone valid reason – calibrate, standardize with a known weight, etc. [1 pt]

6. Graduated cylinder, ruler, balance, and stoppered flask. [0.5 pt each. TOTAL= 2 pt]

Quiz (Experiment #1: Density)

10 points - 15 minutes **Name:______________________________**

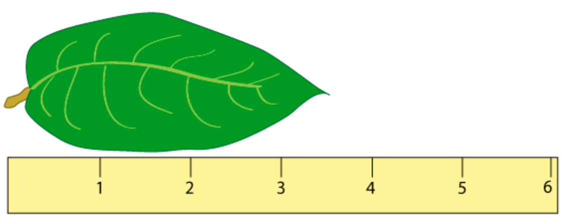
Reference data: Density of water at 20.0 °C = 0.99784 g/mL

Metal	Color	Density (g/cm^3)
Aluminum-	Silvery	2.7
Zinc -	Grey	6.9
Copper -	Red	8.9
Iron -	grey	7.9
Gold –	yellow	19.4

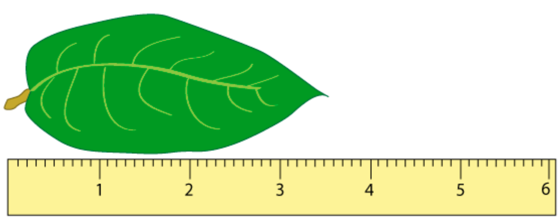
1. Joe has discovered an unknown rectangular solid metal from his great grandfather's closet which he suspects to be either Al, Zn, Cu, Fe, or Au based on his family's business. He observed that it was greyish weighing about 273 g and measured 2.0 cm x 2.0 cm x 10.0 cm. With this information, he was able to identify the metal. Show work and identify the metal. [show work. No work no credit.]

   The metal is:__________.

2. Two students measured the length of a leaf using two different 6 cm rulers. Write the readings that should be properly recorded in their scientific notebooks.



Length of leaf is: _______.



Length of leaf is: _______.

3. An instructor wanted to find the correct volume of a 250 mL flask with a lid. He asked two of his Gen Chem students to figure it out and report it back to him. Each used a different method.

Student A put an empty Erlenmeyer flask with the stopper on a lab balance and tared the reading to zero. Then he filled it with water [measured temperature 20.0 ^{o}C] all the way to the top and capped it with the lid. After the overflowing water outside the flask was thoroughly dried, it weighed 272.56 g. Using this information, he reported the volume of flask as __________.

Show calculation:

Student B filled the flask immediately with water, put the lid to remove the excess water and then transferred it to a 500 mL graduated cylinder. He reported back with the volume as ______.

Note: 500 mL graduated cylinder with water from the Erlenmeyer flask. [The expanded section with the meniscus of water is shown with a line for clarity]

Which student's reading has more precision? Explain

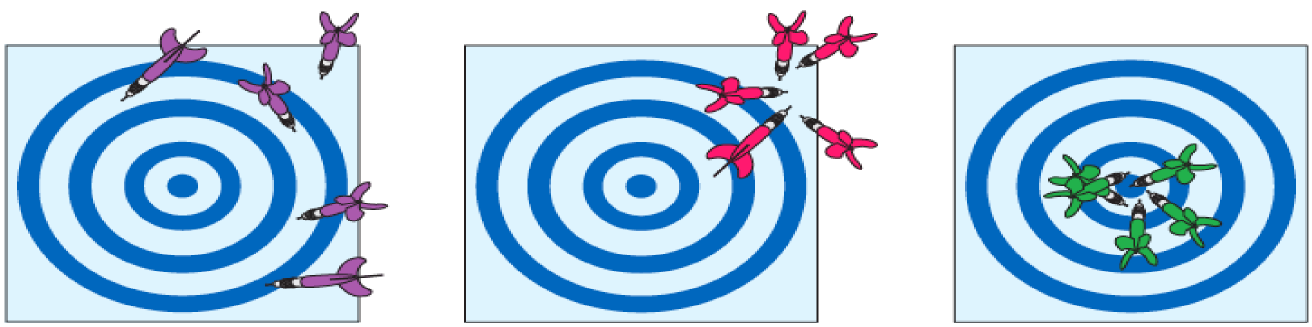
QUIZ SOLUTION

Experiment #1: Density - **10 points**

1. Calculating the density as $273g/40cm^3 = 6.825\ g/cm^3$ according to the given table, zinc corresponds the closest. Color matching alone isn't enough reason and won't get points as both iron and zinc have greyish color. [Choosing zinc 1 pt, reasoning 2 pt. TOTAL= 3 pts]

2. (a) 3.6 cm (b) 3.58 cm [Correct significant figs digits are necessary. Actual value can vary within rules of sig digits. 1 pt each. Mentioning units, 0.5 pt each. TOTAL= 3 pts]

3. Student A: 272.56 / 0.99784 = 273.15 mL [If the student has assumed density is 1, they lose 0.5 pt. Correct significant fig digits and units are needed to get full 1.5 pts]
   Student B: 266 mL [Correct sig digits and units are needed to get full 1.5 pts]
   Student A is more precise because his measurement has more digits OR the student used the more precise mass measurement to indirectly obtain volume. [1 pt]
   [TOTAL= 4 pts]



a Neither accurate nor precise.

b Precise but not accurate.

c Accurate and precise.

Rules for Counting Significant Figures

1. *Nonzero integers.* Nonzero integers always count as significant figures.
2. *Zeros.* There are three classes of zeros:
   a. *Leading zeros* are zeros that *precede* all the nonzero digits. These do not count as significant figures. In the number 0.0025, the three zeros simply indicate the position of the decimal point. This number has only two significant figures.
   b. *Captive zeros* are zeros *between* nonzero digits. These always count as significant figures. The number 1.008 has four significant figures.
   c. *Trailing zeros* are zeros at the *right end* of the number. They are significant only if the number contains a decimal point. The number 100 has only one significant figure, whereas the number 1.00×10^2 has three significant figures. The number one hundred written as 100. also has three significant figures.
3. *Exact numbers.* Many times calculations involve numbers that were not obtained using measuring devices but were determined by counting: 10 experiments, 3 apples, 8 molecules. Such numbers are called *exact numbers*. They can be assumed to have an infinite number of significant figures. Other examples of exact numbers are the 2 in $2\pi r$ (the circumference of a circle) and the 4 and the 3 in $\frac{4}{3}\pi r^3$ (the volume of a sphere). Exact numbers also can arise from definitions. For example, 1 inch is defined as *exactly* 2.54 centimeters. Thus, in the statement 1 in = 2.54 cm, neither the 2.54 nor the 1 limits the number of significant figures when used in a calculation.

Experiment #2: The Creation of a Calcium Supplement

FOR THE INSTRUCTOR

Chemicals required: Solid calcium carbonate, pH paper, phenolphthalein indicator solution, 1M of the following: sodium acetate, sodium bromide, sodium carbonate, sodium phosphate, sodium hydroxide, sodium iodide, sodium chloride, sodium sulfate, calcium nitrate, HCl, and NaOH.

Instruments/material: Balance, thermometer, beakers, burette, burette stand, and clamps.

Information for stockroom personnel

COVERS 2 LAB PERIODS

Chemicals required: Solid calcium carbonate, pH paper, phenolphthalein indicator solution, 1M solutions of each of the following:

- Sodium Acetate
- Sodium Bromide
- Sodium Carbonate
- Sodium Phosphate
- Sodium Hydroxide
- Sodium Iodide
- Sodium Chloride
- Sodium Sulfate
- Calcium nitrate
- HCl
- NaOH

Instruments/material: Burette, burette clamp, ring stand.

Experiment #2: The Creation of a Calcium Supplement

Learning Objectives

Content objectives
Students will be able to:

- Write covalent, ionic, and net ionic equations.
- Use percent composition to determine which soluble calcium salt is the most ideal to use.
- Predict the solubility of an ionic compound based on their experimental results.
- Perform mole calculations
- Determine soluble and insoluble ionic compounds by experiment
- Recognize chemical reactions visually and write chemical equations on paper
- Calculate purity of starting materials
- Quantitative calculations to carry out a chemical reaction.

Process objectives
Students will be able to:

- Perform a solubility test.
- Perform a titration.

Overview

The experiment is divided into two weeks. First, start by explaining what a double replacement reaction is. Make clear the purpose of a solubility test. Explain that to make a calcium supplement, the salt you produce must have a calcium ion in it (a source of calcium) and an anion. The purpose of the lab is to determine what calcium salt to synthesize and how to synthesize it. The solubility tests are performed to determine which calcium salts are soluble. Based on the results of the solubility test, students will select the calcium salt they think is the best choice (*calcium chloride*). Using impure calcium carbonate, students will determine how to synthesize the calcium salt. Give examples of a covalent, ionic, and net ionic equation. At the end of the first lab session, you should inform students that the solubility rules were developed after a series of solubility tests were performed.

Before each group begins a titration make sure to explain proper technique.

1. Rinse the buret with water. Make sure the stopcock works well.
2. Rinse the buret with 5 mL of NaOH.
3. Rinse the burette with 5 mL of NaOH.
4. Fill the burette with NaOH.

Goal: Create a liquid calcium supplement.

Things to look out for:

1. When circulating and questioning groups, pose questions to the student whom you feel is lost/relying too much on their group.
2. Students often are unsure of why a titration must be performed; a titration must be done since the calcium carbonate is impure. Thus in a 10g sample, only a fraction of the 10g is actually calcium carbonate while the rest of the weight is due to impurities.

Experiment #2: The Creation of a Calcium Supplement

Your Project:

A chemical manufacturing plant recently opened up a vitamin supplement division in which you are employed as a chemist. Your supervisor has requested the creation of a liquid calcium supplement as an alternative to chewable ones. The product will be marketed to the kitchens of schools, nursing homes as well as hospitals to ensure the oral consumption of nutritional calcium.

Your own goal is to determine which calcium salts in solution will return a pure aqueous calcium supplement and come up with a procedure to make one from a cheap source like chalk (calcium carbonate), make it into a concentrated solution, and ship it to food industries and pharmacies with appropriate instructions. In Lab period 1, you will identify a suitable soluble calcium compound and suitable procedure for making it. In lab period 2, you will synthesize it in pure solid form or in solution form.

📖 **Recommended Readings:** Writing formulas for ionic compounds and precipitation reactions.

Background:

Metals like calcium form ionic compounds. Some ionic compounds like those with group 1 metals or those with nitrate as the counterion are very soluble but the solubility for other combinations vary significantly. A solubility table can be used as a reference to know which combination may result in insoluble compounds before you set out to prepare one. You are expected to predict, before testing, based on whatever you can infer from your daily life.

Once you have identified soluble calcium compounds, it is important to find cheap starting materials and make the conversion. Limestone (calcium carbonate) is a natural source for calcium. You can decide on the counterion based on what chemicals are available in the lab for this experiment and then make the pure compound by doing a chemical reaction with limestone.

[**NOTE TO INSTRUCTOR:** Determine the correct concentrations of 1 M HCl and 1 M NaOH using potassium hydrogen phthalate (KHP) standard. Put this number on the containers. Students will use this number to calculate the volume of HCl needed for the reaction]

↯Goal: Determine which calcium salts in solution will return a pure aqueous calcium supplement (week 1). Determine the most cost-effective procedure to extract the pure salt (week 2).

Safety	**Materials**
Both acids and bases are CORROSIVE, meaning that skin contact with either one will result in severe burns. Be very careful when handling acids and bases. Do not directly inhale the chemicals. Always wear <u>goggles</u>, a <u>lab coat</u>, and <u>gloves</u> inside the lab. You will be penalized if you fail to do so.	**Chemicals needed:** Solid calcium carbonate, pH paper, phenolphthalein indicator solution, ● 1M solutions of each of the following: ● Sodium Acetate ● Sodium Bromide ● Sodium Carbonate ● Sodium Phosphate ● Sodium Hydroxide ● Sodium Iodide ● Sodium Chloride ● Sodium Sulfate ● Calcium nitrate ● HCl ● NaOH

Experimentation

⟐Focus Question 1: Which calcium compounds are soluble in water?

Make your prediction. Consult with your group members and write one or more reasons that you have found convincing. Two examples are given below. Predict the others.

Compound	Soluble	Insoluble	Reason(s)
Calcium Nitrate	✓		Given in the background information above
Calcium Carbonate		✓	Marble stone is insoluble, chalk is insoluble, background information provided above also suggests that, …
Calcium Acetate			
Calcium Bromide			
Calcium Phosphate			
Calcium Hydroxide			
Calcium Iodide			
Calcium Chloride			
Calcium Sulfate			

[**NOTE TO INSTRUCTOR:** Tell students that they need not be right about their predictions and there is no need to refer to any solubility table. Prediction can come from anything that they have to come to know based on their daily life. For example, they can say things like, calcium phosphate is what makes up the bones and it can't be soluble in water. Calcium chloride sounds similar to common salt and may be soluble in water. All halides are similar and therefore if I know one of them is soluble, others may be soluble too. Calcium deposits in pipes and things of that nature. If they use the solubility table already, it is fine too, but tell them that we are not expecting people to memorize but learn to experiment and organize knowledge. Prompt with questions like 'think where you may come across such compounds – in soil, body, man-made items, etc. Think of similar compounds that you may have known to make a prediction for things you don't know.]

Locate the stock solutions of calcium nitrate and other solutions for counterions in the form of their sodium salts. Locate the test tubes needed for this experiment and make sure they are clean. To make the solution visible in the test tube, you may need to add the solution to about half inch height. Estimate this volume using water in the test tube and a graduated cylinder. All groups members can do it once, average it and round it off to the nearest mL.

[**NOTE TO INSTRUCTOR:** Anything of the order of 2 -3 mL is fine. 1 mL may be too little and more than 3 mL may be too high. The point of this exercise is to make them realize that if they use too little they won't be able to make clear observations and using too much means they will have no space for mixing the second reactant and/or wasting chemical unnecessarily. Probing questions like, why you shouldn't add too much, what may happen if you fill half the test tube with a solution, etc. can make them realize the reasons for this without giving such answers directly. Finally, make them realize that it is a qualitative test (no precise measurements are made) and hence volumes are for a matter of convenience to see what we want to observe.]

(i) There are about 8 reactions you are testing. Estimate the total amount of calcium nitrate solution you need. If you don't have your own supply of calcium nitrate solution at your bench, you can bring this estimated amount to your bench in a small clean beaker and label it as calcium nitrate solution.

Amount of calcium nitrate needed:_____ mL

(ii) ⦸✎Now think about the reactions you are testing to experimentally observe your predictions. Write a couple of sentences about how you will do it. Discuss with your group first and ⦸show the procedure to your instructor before proceeding.

[**NOTE TO INSTRUCTOR:** Guide the students to think about what they are trying to see without giving the answer (formation of a solid from solution), the procedure in terms of minimizing time, minimizing the use of chemicals, and still be able to observe what is needed to make the confirmation. Experimentation is a key feature of how science evolves. Students should change their mindsets from memorizing or verifying facts to learn how to predict, experiment, and develop new knowledge

Students are expected to have one of the 2 procedures - Approximately 2 to 3 mLs of calcium nitrate solution is taken in each of the 8 test tubes. Different sodium salt solutions are added in drops to each one and see if any precipitate forms. Even after adding equal amounts, no precipitate forms, then it can be inferred as soluble.

Alternately, 2 to 3 mLs of different sodium salts be taken in each test tube and calcium nitrate solution can be added in drops to each one. If no precipitate forms after adding about the 2 to 3mLs of the solution, that calcium salt can be taken as soluble.

Verify if their methods are clearly articulated. Allow them to experiment. After experimentation, you can pool the experience and decide if one method is better than the other in terms of efficiency (time and chemicals waste), did the results differ one way to other, etc. Encourage the group members to divide the work for adding solutions and keeping track. Observations can be by consensus. If there are some issues with any particular observation, they can repeat only that reaction.]

The expected results are shown below:

Anion	Soluble	Precipitate
Acetate	✓	
Bromide	✓	
Carbonate		✓
Phosphate		✓
Hydroxide		✓
Iodide	✓	
Chloride	✓	
Sulfate		✓

Salt	Soluble	Insoluble	Does it match prediction?
Calcium Acetate			
Calcium Bromide			
Calcium Carbonate			
Calcium Phosphate			
Calcium Hydroxide			
Calcium Iodide			
Calcium Chloride			
Calcium Sulfate			

Calcium Nitrate	soluble		This is already given as a solution.

(iii) For each of the compounds identified above as <u>soluble</u>, write a balanced equation showing the dissociation of that compound into ions in solution. For example, calcium nitrate is soluble. This dissociates completely as follows: $Ca(NO_3)_2\ (aq) \rightarrow Ca^{2+}\ (aq) + 2\ (NO_3)^-\ (aq)$

(iv) For each of the compounds identified as insoluble, write a balanced equation showing all the reactants and products formed in the reaction. Write molecular, ionic, and net ionic equations for these reactions. For example, if you added two reactants to get the insoluble compound, they must be clearly shown as reactants in (*aq*) state and then show appropriate products along with their physical state.

(v) Predict an outcome based on your experimental results: Will $BaSO_4$ be soluble? Give a reason for your answer.

(vi) Choosing the right compound to make the calcium supplement: You may have discovered some compounds (calcium with different anions) are soluble. Write their formulas correctly.

(vii) List some criteria that you will look for to choose one from the above list. (A business is for profit while drugs are for human consumption and they both must be considered when you think about suitable criteria.)

[**NOTE TO INSTRUCTOR:** Percent calcium in each of the compounds, toxicity, acidity, basicity, cost of production, etc. are all possible criteria. Make sure all these are listed, but students can come up with even more possible answers. Then ask them which are the ones that they can readily obtain now without any outside reference except the periodic table. It may be worthwhile to point to them that Google is a good source for other information like toxicity, but the authentic reference sources are CRC handbook, Merck Index, etc. and also MSDS supplied by the manufacturer of such chemicals. Probe students in such a way that they can identify percent composition, acidity/basicity as readily inferable with basic chemistry knowledge (applying what they have learned in class, students should be able to identify H^+ in the formula is acidic, and hydroxides are basic.]

(viii) Calculate the percent calcium in each of the compounds.

Formula for Compound	% Ca

[**NOTE TO INSTRUCTOR:** They can calculate percent composition of Ca in all the soluble compounds using periodic table and figure which has got the highest % Ca which will lead them to choose $CaCl_2$. Calculated data is shown for soluble compounds including the calcium nitrate. They can be off in the last digit because of rounding of atomic mass.]

Compound	% Ca
Calcium nitrate	24.4
Calcium acetate	27.0
Calcium chloride	36.0
Calcium bromide	20.0
Calcium iodide	13.5

(ix) What is your choice compound for making calcium supplement? Clearly identify the reason.

[**NOTE TO INSTRUCTOR:** It should be obvious to the students that % Ca is the highest in $CaCl_2$ in the above list and hence will make sense to use that one. Students may ask about toxicity data. You can mention that is not part of this lab investigation, but that is required by FDA before it can be approved. Typically company makes pure material first and give it to the next team for toxicity studies.]

⟐Focus Question 2: A cheap source of calcium ion in solid form is limestone ($CaCO_3$). How much of this compound is required to make 2.00 grams of pure $CaCl_2$?

Make a prediction and show your method.

[**NOTE TO INSTRUCTOR:** Students should be able to see from the formula that 1 mole of $CaCO_3$ can make 1 mole of $CaCl_2$. They can calculate based on this information. You can probe 'why it is a prediction rather than a certain amount' to make them realize things like the purity of the starting material, efficiency of the reaction itself, loss while purifying, etc. can contribute to lower yield than theoretical yield. The expected answer is 1.80 g.]

(i) Consider the two sources of chloride in the reactant list available for the lab – NaCl and HCl. Write the possible reactions with calcium carbonate in the form of balanced chemical equations. Which one of the two is appropriate for making calcium chloride and why?

[**NOTE TO INSTRUCTOR:** Probe students to think about products of the reaction, not just the product of interest. Any similar example that can get them to think about isolating a pure product will depend on what else is made along with that.

Consider the two reactions,

$CaCO_3$ (s) + 2 HCl (aq) → $CaCl_2$ (aq) + H_2O + CO_2 (g) &

$CaCO_3$ (s) + 2NaCl (aq) → $CaCl_2$ (aq) + Na_2CO_3 (aq).

In the first one, gas is formed which will escape out of solution and the other product is water which can be removed by evaporation. In the second case, all ions will be in solution and it will be difficult to make the reaction proceed to the right side. In other words, the reaction will spontaneously go from right to left and not the other way around.

Learning Goal: Students should be able to reason out how reactions happen in aqueous solutions. The reaction will proceed in the direction to form a precipitate or gas. Students may already know Acid-Base reactions produces salt and water in aqueous solution. If not, they are seeing it in the next step in this lab. In this case, it is the formation of molecular compound water which drives the reaction. See that students get this concept clear in writing in their notebook – Reactions will happen in aq solutions only if there is a possibility to form a solid, gas, or a molecular compound. Thus, examining the possible combinations of the opposite ions in reactants, and having some idea about the physical state of compounds that can form, we can infer if the reaction can go or not.]

(ii) Once you have settled on the reactants, calculate the amount of the second reactant needed to convert all of the calcium carbonate (number of grams you identified in your prediction) to calcium chloride. How can you be sure that you have added enough to convert all of the calcium carbonate? If you have added an excess of the second reagent, how will you figure out the amount of excess? Don't forget that there may be litmus paper in the lab. ✎Think about this and discuss with your group and write a procedure to identify the exact amount needed. ⦸ See instructor for suggestions and approval before proceeding.

[**NOTE TO INSTRUCTOR:** To ensure complete conversion of calcium carbonate and the uncertainty in the purity of the starting material, students may have added more reactant than what is needed and figure out how much excess reactant (in this case HCl) remains by titrating it with NaOH which is also given to them in the chemicals kit available for this lab. Probing questions like 'what happens if you add more than what you need may make them realize that acid will be left over and then you can probe them how to figure out that excess. On the other hand, if they add too little HCl which will leave a lot of $CaCO_3$ unused, we are wasting starting material. But that can be inferred by simple observation of undissolved starting material. Thus, you can probe them of suitable scenarios to think about identifying the need to know the right amount to add.

Allow them to write the procedure in such a way that they calculate either the stoichiometric amount or a slight excess and check the acidity of the solution with a pH paper, then take NaOH in a burette, and titrate to see the end point and figure out how much excess was added. For this purpose, you may want to mention them that the instructor has personally verified the concentration of NaOH using a standard and can be considered accurate. They are expected to do titration at this point to figure out the excess added and then subtract it to know the exact amount that reacted.

You can also alert them about indicators like phenolphthalein instead of pH paper while doing the titration which makes it practically useful because it is always in solution and takes away the need for testing it after each addition.

The procedure has to be repeated 1 or 2 more times if needed. The exact volume of HCl that reacted will give the amount of $CaCl_2$]

(iii) Based on your experimental data, did you add the correct or less than the stoichiometric amount of HCl? Based on the amount of HCl added, what is the calculated yield of $CaCl_2$ in the reaction? Does it match your prediction? Explain.

[**NOTE TO INSTRUCTOR:** Probe them to think about the purity of starting material. They should be able to calculate the percent purity of the starting material based on the above data. A natural resource like limestone is not completely pure calcium carbonate. We expect answers in the range of 80 to 95% purity of $CaCO_3$.

Note that if there is no more time for the lab, you can stop right here. If you have more time, you can continue to next step of doing a fresh reaction and isolating pure solid $CaCl_2$.]

Experimentation Week 2

Goal: Isolate Solid $CaCl_2$.

(iv) Using your procedure, carry out a new reaction in a beaker. Use the calculated amount of HCl based on the purity of your starting material. The next step of the process is to figure out the amount of pure calcium chloride formed in the reaction. How will you remove any solid unreacted material, how will you remove the solution? ✎Think about these and add those steps to your procedure. *[Note: Hotplate may be available in the hood. Know when to weigh your beakers. Your lab should have filter paper and funnel.]*

[**NOTE TO INSTRUCTOR:** You may want to add your notes to the following procedure to make it complete.

Start with a clean and dry 100 mL beaker. Weigh it.

Estimate 2.00 g of CaCO3 using a weighing paper and then add this to the beaker and reweigh.

Calculate the amount of 1.0 M HCl needed and then add this to the beaker. There will be some fizzing.

At this point, we have $CaCl_2$ in solution along with other products. Note that unreacted solids can be removed by filtration with a filter paper and water can be removed by placing the beaker on a hot plate at 100 - 125 ºC in the hood. All weighings must be done when the beaker is cold.

If filtration is needed, another clean beaker must be weighed and the filtrate must be transferred to that beaker which is then placed on a hot plate for evaporating the water.

Ask students how they will know that they have evaporated all the water. It is important to put it back on the hotplate for 3 to 5 minutes and reweigh. The weights should be very close to each other and not differ by more than 0.1 g. If so, they continue this process until the weights are within 0.1 g]

(v) Carry out your experiment. Show calculations and report the final mass of solid $CaCl_2$ formed.

FINAL LAB REPORT:

Your final lab report should clearly include the solubility data and clearly explain why you chose $CaCl_2$. In addition to providing the experimental details include the procedure for preparing pure solid $CaCl_2$, include in the procedure to make 500 mL of a 1.00 M solution of $CaCl_2$. This section must provide the number of grams of $CaCl_2$ to be synthesized, and how to make 500 mL of a 1.00 M solution.

POST LAB QUIZ:

Make sure you practice writing net ionic equations. There are stoichiometric calculations using solids and ions in solution, hence you need to know how to calculate concentration and moles of ions. You must be able to carry out all types of calculations that are used in this lab.

PreLab – Exp #2: The Creation of a Calcium Supplement **Name: __________________________**

This assignment is based on nomenclature, balancing equations, and stoichiometry from your textbook. These must be completed before you come to the lab.

1. Write the formulas for the following compounds and if they are ionic compound, show the ions that make up that compound.

Name	**Formula**	**If ionic, write a balanced equation when it dissolves in water to show the ions. If they are not ionic, just write molecular.**
Potassium sulfide		
Silicon tetrachloride		
Iron(II) phosphide		
Phosphorus pentoxide		
Calcium sulfate		
Dichlorine heptoxide		
Magnesium nitrate		
Iron(II) acetate		
Magnesium perchlorate		
Sodium phosphate		

2. How many moles of magnesium ion are present in 5.25 g $Mg_3(PO_4)_2$?

3. How many moles of phosphate ion are present in 5.25 g $Mg_3(PO_4)_2$?

4. How many grams of elemental phosphorus is present in 5.25 g $Mg_3(PO_4)_2$?

5. If all the phosphate available in 5.25 g $Mg_3(PO_4)_2$ is converted to $AlPO_4$ using chemical reactions, what is the maximum amount (theoretical yield) of $AlPO_4$ that can be obtained? In the lab, the yield of $AlPO_4$ was only 75%. How many grams of $AlPO_4$ was actually obtained?

Name	**Formula**	**If ionic, write a balanced equation when it dissolves in water to show the ions. If they are not ionic, just write molecular.**
Potassium sulfide	K_2S	$K_2S \rightarrow 2K^{+} + 2S^{-}$
Silicon tetrachloride	$SiCl_4$	NOT IONIC
Iron(II) phosphide	Fe_3P_2	$Fe_3P_2 \rightarrow 3Fe^{2+} + 2P^{3-}$
Phosphorus pentoxide	P_2O_5 or P_5O_{10}	NOT IONIC
Calcium sulfate	$CaSO_4$	$CaSO_4 \rightarrow Ca^{2+} + SO_4^{2-}$
Dichlorine heptoxide	Cl_2O_7	NOT IONIC
Magnesium nitrate	$Mg(NO_3)_2$	$Mg(NO_3)_2 \rightarrow Mg^{2+} + 2NO^{3-}$
Iron(II) acetate	$Fe(C_2H_3O_2)_2$	$Fe(C_2H_3O_2)_2 \rightarrow Fe^{2+} +2(C_2H_3O_2)^{-}$
Magnesium perchlorate	$Mg(ClO_4)_2$	$Mg(ClO_4)_2 \rightarrow Mg^{2+} + 2ClO_4^{-}$
Sodium phosphate	$Na_3(PO_4)_2$	$Na_3(PO_4)_2 \rightarrow 3Na^{+} + PO_4^{3-}$

2. 5.99×10^{-2} mol Mg^{2+}

3. 3.99×10^{-2} mol $(PO_4)^{3-}$

4. 1.24g P

5. Theoretical yield = 4.87g $AlPO_4$

Actual yield: 3.65g $AlPO_4$

Quiz (Experiment #2: The Creation of a Calcium Supplement)

10 points - 15 minutes **Name:______________________________**

Use a copy of the periodic table. All ionic compounds that dissolve break apart into their constituent ions. When compounds react with each other, the balanced equation is necessary to know their stoichiometry.

MOLECULAR WEIGHTS:

Mg: 24.305 g/mol P: 30.974 g/mol O: 15.999 g/mol

N: 14.007 g/mol Na: 22.990 g/mol

1. In a solution of 0.50 M $Mg(NO_3)_2$.write the species present and their concentrations.

2. In a solution of 0.50 M Na_3PO_4, write the species present and their concentrations.

3. When 100 mL solution 0.50 M $Mg(NO_3)_2$ is added to a 100 mL solution of 0.50 M Na_3PO_4 a precipitate of Magnesium phosphate is obtained.
   (a) Write a balanced net ionic equation for the formation of magnesium phosphate precipitate.

   (b) How many grams of magnesium phosphate can be obtained in the above reaction?

   (c) List all the ions that are present in solution after the precipitation.

(d) What is the concentration of nitrate ion in solution after the precipitation of magnesium phosphate?

Experiment #2: The Creation of a Calcium Supplement **- 10 points**

1. **0.5M Mg^{2+} , 1M NO_3^-**
2. **1.5M Na^+ , 0.5M PO_4^{3-}**
3. **a) $3Mg^{+2}_{(aq)} + 2PO_4^{3-}_{(aq)} \rightarrow Mg_3(PO_4)_2$**

    b) 4.38g $Mg_3(PO_4)_2$

    c) Na+ and NO_3^-

    d) The nitrate ion is a spectator ion since it is not involved in a net ionic reaction. Its initial concentration was 1M (see the answer to question 1), but now since an additional 100 ml of water was added (from 100 ml of sodium phosphate solution), now the new concentration is 0.5M

Experiment #3: What Can I Bid For This Chemical Stockpile?

FOR THE INSTRUCTOR

Chemicals required: samples of two compounds copper(II) sulfate and copper(II) chloride, aluminum foil, zinc granules, water, and 6M HCl.

Instruments/material: hot plate, beakers, stirrers, balance.

Information for stockroom personnel

COVERS 2 LAB PERIODS

Make sure the chemicals do **not** have the complete formula written on the outside of the bottle. Call them copper sulfate and copper chloride. NO FORMULAS. Other chemicals can have formulas.

Amounts of chemicals needed (approximate amounts needed):

Copper(II) sulfate: 60 grams for each section

Copper(II) chloride: 60 grams for each section

Zinc granules: 60 grams for each section

Aluminum foil: About 50 sq. ft. Reynolds foil. Weight about 60 grams for each section

1 L of 6 M HCl for each section. [Unlikely to use all of it. May cover all 3 sections]

Experiment #3: What Can I Bid For This Chemical Stockpile?

Learning Objectives:

Content objectives
Students will be able to:

- Predict the products of single replacement reactions
- Determine the relative reactivity of metals
- Predict the solubility of an ionic compound based on their experimental results
- Perform mole calculations
- Understand why sodium chloride must be added to activate the reaction between copper sulfate and aluminum
- Write ionic equations for chemical reactions.

Process objectives
Students will be able to:

- Carry out quantitative calculations for a chemical reaction.
- Connect the importance of quantitative chemical reaction to commerce.

Overview

The experiment is divided into two weeks. First, start by explaining single replacement reactions briefly; be sure to mention that if the element is less active than no reaction occurs. Do not explain the activity series yet. Most single replacement reactions are redox reactions. Briefly explain what a redox reaction is.

At the end of lab session 1, be sure to inform students that the activity series was developed by performing a series of single replacement reactions.

↯*Goal: Determine if aluminum and zinc can displace copper (week 1). Determine the percent yield of the displacement of copper for their selected metal (week 2).*

Things to look out for:

1. When circulating and questioning groups, pose questions to the student whom you feel is lost/relying too much on their group.
2. Students will likely try to search the activities of the metals online to answer the first focus questions. Urge them not to search for information online. Probe them to think about what they know about the periodic trends from the periodic table.

Experiment #3: What Can I Bid for This Chemical Stockpile

Your Project:

A manufacturing company wants to get rid of one of their divisions in a corporate reshuffle. This division has about 500,000 kg each of two copper compounds – copper(II) sulfate and copper(II) chloride which is not in the pure state because they may have absorbed some moisture due to long storage. The company wants to get rid of them as soon as possible in its current form. As a trader in copper, your employer, XYZ metal corporation wishes to get hold of these at the cheapest price possible, purify the copper, and sell it at a **profit**. They made a preliminary offer to acquire these copper compounds at $0.10 per kg. Even though XYZ metal corp. estimated making a profit, they want to analyze the compounds and make sure the deal was worth. Samples of both compounds were obtained by XYZ corp. and given to you for analysis. Your assignment is to estimate the value of copper in these compounds.

📖 **Recommended Readings: Writing chemical formulas of ionic compounds, calculating percent composition, writing chemical equations, calculating the mass of reactants and products (stoichiometry).**

Background:

Copper is a very valuable element used in making coins, utensils, electrical wires, etc. It is the first one in the group of noble metals. Many other metals are so reactive that they can readily displace copper. The periodic table is a useful source to look for reactivity patterns and pick what may be suitable metals to displace copper.

During lab 1, you will figure out whether Al or Zn can displace copper from its compounds, the conditions necessary for the reaction to occur, and practical consideration for completing the reaction and isolate pure copper. To do the reaction on a quantitative scale, you will calculate the amounts necessary to carry out the reactions on a 2-g scale and write a procedure that can help you accomplish isolating pure copper and weigh it. One procedure is enough for each group. ⦸Your procedure must be approved by your instructor.

During lab 2, you will carry out the reaction as per your procedure, get the dry copper and weigh it to find the mass percent. You are expected to write your individual final reports to your employer addressing total mass of copper that can be recovered, and its value. Feel free to include your thoughts on the profitability of the deal and any other issues to consider.

Complete the pre-lab exercise at home and submit to your instructor as soon as you enter the lab.

Goal: Determine if aluminum and zinc can displace copper.

Safety	**Materials**
ACIDS are corrosive. Handle with care and avoid skin contact. Do not inhale the chemicals. DO NOT REMOVE THE ACID CONTAINER FROM UNDER THE HOOD. Hydrogen gas is flammable and reactions of acids with metals are vigorous, produce heat, and may release hydrogen gas. Use hood wherever needed. All waste should be properly disposed of in the appropriate containers as directed by your instructor. Always wear goggles, a lab coat, and gloves inside the lab. You will be penalized if you fail to do so.	**Chemicals needed:** samples of two compounds copper(II) sulfate and copper(II) chloride, aluminum foil, zinc granules, water, and 6M HCl **Instruments needed:** hot plate, beakers, stirrers, balance

Experimentation

Focus Question 1: Predict if zinc granules can displace copper from its compounds. Explain.

[**NOTE TO INSTRUCTOR**: Students can use the periodic table and make them think about I metallic nature which changes across the table. A reference table of activity series can also be added after probing. You can probe things like which metals are used as such and which metals are not known to be used as such. Students are likely to say things like aluminum, gold, silver, copper, iron, titanium, etc. *More reactive elements are found as compounds and they are difficult to keep them in their metallic state.* So, you can probe them like – explain why you see sodium chloride but not sodium. These kinds of probing will help them realize that *active metals give up electrons to become ions. Less active metals can be kept as such without losing their neutral state.* Once you get past this idea, ask them to think about the relative activity of the stable metals. This question is about the relative reactivity of Zn Vs Cu. They can be wrong about their predictions and we don't have to correct them. They are going to experiment and find out.]

It is time to experiment. Do a *qualitative* test to see what conditions and observations are necessary to check if Zn can displace copper ions. Take a small sample of copper sulfate and copper chloride in 2 separate beakers and dissolve them in lesser amounts of water until everything dissolves. The amount of sample to use and the amount of water to use are up to you, but to be employed in a company, you will be expected to minimize the amount. So, plan to use just enough to make visual observations. Throw in a few bits of zinc into each and stir.

(i) Explain what you observed. Are there any indications that a chemical reaction occurred and if so, any indication that the reaction is complete and if not complete, what needs to be done to make it go to completion?

(ii) Use chemical equations to describe your observation. Show both molecular and net ionic equations. **Do not throw away the solution yet.** Space is given below for work.

(iii) From what you have experimented, can Zinc displace copper from both compounds? Show your work to your instructor.

Focus Question 2: Predict if Aluminum foil can displace copper from its compounds.

Explain.

[**NOTE TO INSTRUCTOR**: This question is about the relative reactivity of Al Vs Cu. They can be wrong about their predictions and we don't have to correct them now. However, you can ask some students to explain to the entire class (both Yes and No answers). They are going to experiment and find out. You can also probe their reasoning using the same type of questions on relative activities like you did for the previous question.]

[**NOTE TO INSTRUCTOR**: The blue color of the copper solution should slowly disappear first. Students are likely to conclude either it doesn't work OR it works slowly. A class discussion on slow reaction needs to be had. Theoretical activity suggests that Aluminum should be more active than Zinc, but why it is not observable. Some students may know that Aluminum metal is protected by some surface covering, but you can bring it out slowly by hinting about things like rust, etc. To scratch the coating, one can scratch the surface with sandpaper or in this case, adding a bit of salt which will do the same job in solution.]

It is time to experiment. Do a qualitative test to see what conditions are necessary and what observations are necessary to check if Al can displace copper ions. Take a small sample of copper sulfate and copper chloride in 2 separate beakers and dissolve it in a small amount of water until everything dissolves (like you did in the previous step). Throw in tiny bits of aluminum foil and stir.

(i) Explain what you observed. Are there any indication that there was a reaction and if so, any indication that the reaction is complete and if not complete, what needs to be done to make it go to completion?

(ii) Use chemical equations to describe your observation. Show both molecular and net ionic equations. **Do not throw away the solution yet.** Space is given below for work.

(iii) If there was a reaction, is it faster or slower compared to zinc? Refer to the table of activity amongst Zn and Al, and see if your results agree. If you agree, go to next step. If not, explain what are the possible reasons and suggest some solutions that you wish to try. **Do not throw away the solution yet.** Talk to your instructor first.

(iv) From what you have experimented, can aluminum displace copper from its compounds? Show your work to your instructor. Make sure you write molecular and net ionic equations wherever appropriate.

(v) Having done both experiments, it is time to choose one of the two metals. If both Al and Zn worked for both copper compounds, then which one will you choose? Explain your choice.

(vi) If only one of them worked for both compounds, choose that metal for carrying out the quantitative determination of copper. Go back to your experiments and check the solutions in the beaker.

Focus Question 3: If you have excess Al or Zn in your solutions, they need to be removed before you can obtain dry copper and weigh it.

Predict what can be done to get copper only and not the unreacted solid Al or Zn?

[**NOTE TO INSTRUCTOR**: Students may struggle with this part as they are not used to worrying about excess. They generally think that we need to add more to get more yield! Here, adding more of one reagent (metal) is a problem because they need to be removed. It is important to get that emphasized that stoichiometric calculations are needed precisely for this purpose and set them to calculate correctly for the next step where they are going to calculate reagents on a 2 g scale.]

(i) There are some acids available in your lab. One of them is hydrochloric acid. Add in tiny amounts of acid into the beakers containing unreacted Al or Zn and stir to see if the metals react. Write your observations and provide chemical equations. Make sure you write all the visual clues that can help you confirm if everything other than copper is not in solid form. Write molecular and net ionic equations for the reaction between Al or Zn with acid.

✎BEFORE YOU SET OUT TO WRITE A COMPLETE PROCEDURE:

- Review the experimental results for your 3 focus questions and ascertain the following:
- What is the physical state of copper salt before it can be displaced?
- What can displace copper?
- How can you remove any excess of the metal after all copper is removed?
- What are the appropriately balanced chemical equations?

⦸✎Now you are ready to write the procedure for this lab to quantitatively estimate the percent of copper on a 2 grams scale. Carry out all the calculations needed, write the procedure, and ⦸show it to your instructor. *You will use this procedure to do the experiment in the following lab.*

- Make sure you use appropriate containers for measuring the 2 grams of material correctly.
- Assume copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$) as the formula of copper sulfate and copper(II) chloride dihydrate ($CuCl_2 \cdot 2H_2O$) as the formula of copper chloride to calculate the amount of metal needed for the chemical reaction. Your balanced equation will help you get the correct numbers.
- As you may have seen in your earlier write-ups, every small thing like 'zeroing the balance', using a weighing paper/plastic weighing boat, dry beaker, etc. should be explicitly stated. Or else, no one can follow the steps correctly.
- Give a rough estimate of water needed to dissolve and visual clues to know if it has dissolved. Be sure to document the steps and visual clues for every step to know the completion of the reaction and for removing excess Zn or Al.
- At the point where you need to remove the dry copper metal in the beaker, you can simply

decant the excess liquid from the beaker, wash it 2 times with water, decant the water, and put the beaker with wet copper on a hot plate to dry.

- Based on the last step, make sure your procedure clearly identifies the steps where you get the weight of beaker, the weight of beaker + copper compound, and weight of beaker + copper metal at the appropriate step in your procedure.

    ⦸ See instructor for suggestions and approval before leaving the lab.

Mass of the beaker	Mass of the beaker + copper compound	Mass of the beaker + copper metal

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Experimentation Week 2

Goal: Determine the percent yield of the displacement of copper to see which compound is more lucrative to extract copper.

Focus Question 4: Predict the maximum mass percent of copper that you may be able to retrieve experimentally from copper(II) sulfate pentahydrate and from copper(II) chloride dihydrate. Explain.

[**NOTE TO INSTRUCTOR**: Students can use the formula mass in each compound to calculate % of Cu. But the important thing is for them to realize that this is maximum %. Anyone giving smaller than that number saying that they are only 90% efficient in the lab OR reactants may be contaminated are all giving GOOD answers. You can pose them a question like this 'I say I am going to get 80% copper from copper chloride sample. Comment on my prediction.]

Use your procedure developed in the previous lab. Use two separate clean dry beakers with no cracks. Use one for copper(II) sulfate and the other one for copper(II) chloride. Do both experiments simultaneously. Keep track of what you are doing.

(i) From your masses of the beaker, beaker + copper compound, and beaker + copper metal, calculate the mass percent of copper in each of the copper compounds. Does it agree with your prediction? Explain.

	Mass of beaker	**Mass of beaker + copper compound**	**Mass of beaker + copper metal**
copper(II) sulfate			
copper(II) chloride			

FINAL LAB REPORT:

At this point, discuss with your instructor on how to write your final lab report. Your experiment should lead you to find the mass of copper metal in the entire stockpile. Make your reasoning if Al or Zn should be used for displacing the copper.

Use today's metal prices to find the monetary value of copper based on this mass. You can use the following link as a source: https://markets.businessinsider.com/commodities/copper-price. Based on your employer’s bid at \$0.10 per kg, will the company make a profit on the extracted copper?

PreLab – Exp #3: What can I bid for this chemical stockpile? **Name:____________________**

1. (a) Write the formula for sodium sulfate decahydrate. (b) What is the mass percent of sodium in sodium sulfate decahydrate? (c) How many grams of sodium are present in 500 g of sodium sulfate decahydrate?

2. Solid magnesium reacts with dilute sulfuric acid to form magnesium sulfate solution and hydrogen gas.
   Write a balanced equation to represent the above statement. Be sure to indicate the physical state of each of the reactants and products.

3. Consider the reaction between magnesium metal and hydrochloric acid to form magnesium chloride and hydrogen gas.
   (a) Write a balanced equation.

   (b) If 1.50 grams of Mg is used, how many moles of HCl are required for complete reaction?

4. Sodium metal reacts with cold water to liberate hydrogen, magnesium metal will not react with cold water but reacts with steam while Aluminum will not react with water or steam. Use this information to arrange the metals based on their activity. What is the general periodic trend that can be expected based on this information?

PRELAB SOLUTION

Experiment #3: What can I bid for this chemical stockpile? – **10 points**

1. (a) $Na_2SO_4.10H_2O$ (b) 14.3% (c) 71.4 g [1 pt each. TOTAL= 3 pts]
2. (a) Mg (s) + H_2SO_4 (aq) $MgSO_4$ (aq) + H_2 (g) [0.5 pt for each formula, TOTAL= 2 pts]
3. Mg (s) + 2 HCl (aq) $MgCl_2$ (aq) + H_2 (g) [1.5 pts]
   Mol Mg: 1.50/24.3 = 0.0617 mol Mg. It needs 0.0617 x 2 mol HCl = 0.123 mol HCl [1 pt each, TOTAL= 2 pts]

4. Na > Mg > Al. As we go across the period from left to right, activity decreases. [1.5 pts]

Quiz (Experiment #3: What Can I Bid for this Chemical Stockpile?)

10 points - 15 minutes **Name:______________________________**

1. (a) Write a balanced equation for the reaction between solid aluminum (Al) and copper(II) sulfate ($CuSO_4$) solution to form Aluminum sulfate ($Al_2(SO_4)_3$) solution and solid copper (Cu).

   (b) How many grams of solid aluminum is required to produce 5.25 g of solid copper? Assume you have excess copper(II) sulfate solution for the reaction with aluminum.

2. (a) Copper metal can displace silver ion from its compounds to form silver metal. Silver metal can displace gold(III) ion from its compounds to form gold metal. Predict the periodic trend based on this information.

   (b) Write Yes if the reaction will proceed from left to right. Write NO if it will not happen as written.

   i. $2\ Au\ (s) + 3\ Cu(NO_3)_2\ (aq) \rightarrow 3\ Cu\ (s) + 2\ Au(NO_3)_3\ (aq)$______

   ii. $2\ Ag\ (s) + Cu(NO_3)_2\ (aq) \rightarrow Cu\ (s) + 2\ Ag(NO_3)\ (aq)$ ______

   iii. $3\ Cu\ (s) + 2\ AuCl_3\ (aq) \rightarrow 2\ Au\ (s) + 3\ CuCl_2\ (aq)$ ______

QUIZ SOLUTION

Experiment #3: What Can I Bid for this Chemical Stockpile? **- 10 points**

1. (a) 2 Al (s) + 3 $Cu(SO_4)$ (*aq*) 3 Cu (s) + $Al_2(SO_4)_3$ (*aq*) [2 pts for balanced equation either ionic or molecular. 1 pt if physical states are clearly shown for each. TOTAL= 3 pts]

   (b) 5.25/63.546 mol Cu = 0.0826 mol Cu [Give that point if you see the math 5.25/63.546 as many students do everything in one step, 1 pt]

Needs 0.0826 x 2/3 mol Al = 0.0551 Al [Give the point if you see the math step, 1 pt]

Which is 0.0551 x 26.9815 g = 1.49 g of Al [1 pt. TOTAL= 3 pts]

2. (a) The activity of metals in the group IX-B or column 11 (coinage metals or noble metals) decreases as we go down the group. It's also okay if they write Cu>Ag>Au [1 pt]

   (b) (i) NO, (ii) NO, (iii) YES [1 pt each. TOTAL= 3 pts]

Experiment #4: Safety of household antiseptic

FOR THE INSTRUCTOR

Chemicals required: 0.02 M potassium permanganate, household hydrogen peroxide, 1 M or 3 M H_2SO_4

Instruments/material: Burette, 2 to 5 mL graduated pipette, 10 mL graduated cylinder, 50 or 100 mL graduated cylinder, 250 mL Beaker or Erlenmeyer flask

Information for stockroom personnel

COVERS 1 LAB PERIOD

Chemicals needed:

- 0.02 M potassium permanganate
- household hydrogen peroxide
- 1 M or 3 M H_2SO_4

Experiment #4: Safety of household antiseptic

Learning Objectives

Content objectives
Students will be able to:

- Interconvert percent by mass and molar concentration.
- Balance a redox reaction using oxidation numbers
- Understand why acid is added to the hydrogen peroxide.
- Write a titration procedure with clear instructions for materials and methods

Process objectives
Students will be able to:

- Perform a titration
- Establish proficiency in experimental observations and reproducibility of results
- Quantitative calculations following all rules for scientific reporting

Overview

Explain percent by mass and molar concentration. Students were introduced to these two units of concentration in the prelab. After all the students have handed in the prelab, balance the redox equation with the class. Students should be familiar with titration techniques from the previous labs. If necessary, guide students towards the proper techniques.

Goal: Verify the concentration of the hydrogen peroxide sample.

Things to look out for:

1. When circulating and questioning groups, pose questions to the student whom you feel is lost/relying too much on their group.
2. Since this is a one-week lab, be sure to remind students to take meticulous data by following titration protocol performed in Experiment #2.

Experiment #4: Safety of a Household Antiseptic

Your project:

As a quality control chemist, your job is to ascertain the purity of hydrogen peroxide before it is packed. For peroxide marketed for a household antiseptic purpose, your company wants to be sure the concentration is within the range 3.0 $\pm$ 0.2 % by mass.

📖 **Recommended Readings: Redox reactions and titration.**

Background:

Hydrogen peroxide is an unstable compound which slowly decomposes to water and oxygen. However, in the presence of catalase, the reaction proceeds fast giving out nascent oxygen and killing bacteria on the wounds. It will also kill normal cells and hence should be used with caution. The concentration of peroxide is important to minimize scarring.

In the presence of acid, hydrogen peroxide can be oxidized by potassium permanganate. The purple-colored manganate ion is reduced to colorless Mn^{2+} ion in the process. This reaction can be used to calculate the exact concentration of hydrogen peroxide by the titration method.

Goal: Determine the concentration of household antiseptic.

Safety	Materials
ACIDS are corrosive. Handle with care and avoid skin contact. Do not inhale the chemicals. DO NOT REMOVE THE ACID CONTAINER FROM UNDER THE HOOD. Dispose of chemicals in the proper waste container. Potassium permanganate is a strong oxidizer, therefore it can irritate skin. Always wear goggles, a lab coat, and gloves inside the lab. You will be penalized if you fail to do so.	**Chemicals:** Household hydrogen peroxide, household hydrogen peroxide, 0.02 M potassium permanganate, 1 M or 3 M H_2SO_4 **Instruments:** 10-mL and 50-mL graduated cylinder, balance, ruler, stopper flask.

Experimentation

Focus Question 1: Predict how many mLs of 0.020 M $KMnO_4$ are needed to completely react with 2.0 mLs of 3% hydrogen peroxide solution. Explain your reasoning.

[**NOTE TO INSTRUCTOR**: Since the concentration of peroxide solution is 0.88 M and the stoichiometric ratio is 2:5, mLs required is 0.88 x 2.0 mL x 2/5 = 0.704 millimole = 0.704/0.020 = 35.2 mL. Even though this is expected, students need not be correct about this. They can fix their error after seeing results.]

(i) Write a procedure for doing the titration and verify your prediction.

See instructor for suggestions and approval before proceeding.

[Hint: Put $KMnO_4$ in the burette. 2 mLs may be too small to see in a 125 or 250 mL flask and you may have to add some water to make the solution visible.]

[**NOTE TO INSTRUCTOR**: Probe them with questions using the balanced reaction. Ask how much acid is needed and how they will introduce that. As per balanced equation, you can allow them to calculate that and put in a slight excess as we don't monitor that for calculation. They can easily estimate that more than 2.1 mL of 1 M H^+ is needed and adding 3 to 5 mL of 1.0 M H_2SO_4 is not an issue. Sample procedure below. Make sure they include cleaning and lining burette before filling. This is the second titration experiment they are supposed to be doing and hence they must be familiar with titration requirements.

1. Clean the burette and line the burette with $KMnO_4$ using 2 to 5 mL of KMnO4 solution and discard the waste.
2. Fill burette with $KMnO_4$ to the 0 mark.
3. Record the initial volume of KMnO4 to the correct number of significant figures.

4. Transfer 2.0 mL of the 3% hydrogen peroxide using a clean pipette into a 250 mL Erlenmeyer flask. Add about 20 to 25 mL water.
5. Add about 5.0 mL of 1 M H_2SO_4 into the flask. [OR 1 mL of 3 M H_2SO_4]
6. Swirl the contents of the flask and Titrate the hydrogen peroxide solution with the $KMnO_4$ solution. The disappearance of the purple color indicates reaction is proceeding. Add slowly and stop when the color stays in solution with the addition of just one drop. Record the final volume of $KMnO_4$.
7. Place the contents of the flask into the waste beaker and rinse the flask. Note down the volume
8. Repeat the titration two more times to check for consistency.
   NOTE: All waste must be properly disposed of in the waste container.
9. Take the average of consistent trials.]

(ii) Based on your experiment, are the mLs of $KMnO_4$ required to agree with your prediction?

(iii) What is the experimental concentration of hydrogen peroxide household antiseptic?

(iv) Is it within the accepted error range ($\pm$ 0.2 % by mass)?

FINAL LAB REPORT:

- Include your procedure and calculations.
- Explain why you need to add H_2SO_4 in the titration.
- In the conclusion, make sure you answered the question that your employer wanted you to find out.

PreLab – Exp #4: Safety of household antiseptic **Name: ________________________**

1. Calculate the molar concentration of H_2O_2 in a 3.0 % by mass hydrogen peroxide solution.

   [3.0 % by mass means 3.0 grams H_2O_2 per 100 grams of solution. At this low concentration, the density of solution is 1.0 g/mL]

2. Balance the following redox reaction that happens in an acidic medium.

   MnO_4^- (aq) + H_2O_2 (aq) → Mn^{2+} (aq) + O_2 (g)

- 0.88 M
- $2MnO_4^-{}_{(aq)} + 5H_2O_{2\,(aq)} + 6H+_{(aq)} \rightarrow 2Mn_2^+{}_{(aq)} + 5O_{2\,(g)} + 8H_2O_{(l)}$

Experiment #5: Gas Law

FOR THE INSTRUCTOR

Chemicals required: Methanol, acetone, isopropanol, hexane, water.

Instruments/material: Hotplate, ring stand, clamps, aluminum foil, rubber bands, 500 mL beaker, 100 mL graduated cylinder, boiling chips, a pin needed for making a hole in the middle of the aluminum foil, waste container.

Information for stockroom personnel

COVERS 2 LAB PERIODS

Chemicals Needed: Methanol, acetone, isopropanol, and hexane.

Experiment #5: Gas Law

Learning Objectives

Content objectives
Students will be able to:

- Understand each postulate of the Kinetic Molecular Theory.
- Apply the Ideal Gas Law equation.
- Understand the relationship between boiling point and deviation from the ideal gas law.
- Analyze the relationship between pressure, volume, and temperature to the number of gas-particles.
- Determine the molar masses of several volatile substances to understand the limitations of Ideal Gas Law and Kinetic Molecular Theory.

Process objectives
Students will be able to:

- Analyze data for patterns and draw appropriate conclusions.
- Explain the reasoning behind the experimental setup.

Overview

This experiment is divided into two weeks. Firstly, review the Kinetic Molecular Theory (KMT) postulates and the ideal gas law equation. Show students the experimental set-up. Do not explain any part of the setup. Have the students figure it out. In lab 1 the students decide the right procedure for measuring the molar mass of each of the liquids assuming they follow ideal gas law at the boiling temperature of water. In lab period 2, they will complete the molar mass determination for all four compounds.

A typical procedure should have the following steps:

1. Fill in a ----- mL beaker with ------ mL water. Put some boiling chips, Assemble the stand and hot plate, keep the beaker on the hot plate, and bring it to boiling.
2. Take a clean, dry, 250 mL Erlenmeyer, cut a square piece of Al foil that can cover it and a rubber band and weigh them to two decimal places.
3. Add 3 mL of volatile liquid. Close with the foil and put the rubber band. Make a pinhole in the foil.
4. Slowly lower the clamped flask into hot water until the neck is inside hot water and clamp it.
5. Check if the liquid inside Erlenmeyer flask has completely evaporated and then remove it from hot water. Measure the temperature of boiling water.
6. Wipe the outside of the cooled Erlenmeyer flask and then take its mass.
7. Empty the liquid and waste container. Fill the flask with water. Measure the volume of water using a 100 mL graduated cylinder.
8. Note down the atmospheric pressure.

9. Calculate the molar mass of the volatile liquid. n = PV/RT and MW = mass/n

Goal: Determine the molecular weight of 4 volatile liquids. Determine which gasses behave closest to an ideal gas.

Things to look out for:

1. When circulating and questioning groups, pose questions to the student whom you feel is lost/relying too much on their group.
2. Leave a common 10 mL graduated cylinder and a small beaker for each of the chemical. Students can bring their flask to the hood and transfer the right amount using the common 10 mL graduated cylinder. This way, they will not waste chemicals by taking too much chemical it to their bench. In addition, they need not keep rinsing their own beaker or graduated cylinder for each chemical and contribute more waste. Keep the chemicals spread out in the hood so that there is no possibility of contamination.

Your Project:

In this lab, you are required to examine four different volatile liquids – methanol (CH_3OH), isopropanol (C_3H_7OH), hexane (C_6H_{14}), and acetone (C_3H_6O) and determine if they follow the ideal gas law at the water's boiling point temperature.

You will be working in groups as decided by your instructor. In lab period 1, you will make calculations and a sample experiment to decide the right procedure for measuring the molar mass of each of the liquids, assuming they follow ideal gas law at the boiling temperature of water. In lab period 2, you will complete the molar mass determination for all four compounds. Since you will use ideal gas law to calculate the molar mass, the correlation between their experimental value and actual value gives you an indication of their ideal nature at a specific temperature. In this case, all four compounds will exist in the gas phase at the boiling point of water, 100^{o}C. Observe patterns and explain your observations by developing scientific ideas you have learned about molecular behavior.

📖 **Recommended Readings**: **Kinetic molecular theory (KMT) and the Ideal Gas law.**

Background:

The study of gases is extremely important for the fundamental understanding of molecular behavior. Amongst the three common phases of matter, gases contain individual particles that move freely, while solids and liquids tie the molecules together and exhibit bulk properties. Volatile liquids vaporize easily and convert to gas phase completely when placed in boiling water. No matter which substance is involved, when their phase changes to gas, they behave as individual gas particles. Boyle's Law, Charles' Law, and Avogadro's law that are combined together as ideal gas law summarizes the observed behavior of gas molecules. The Kinetic Molecular Theory (KMT) of gases is the theoretical basis to explain the ideal gas behavior.

⁴⁄₄Goal: Determine the molecular weight of 4 volatile liquids. Determine which gasses behave closest to an ideal gas.

Safety	**Materials**
All chemicals listed below are FLAMMABLE. Hence, the hot plate is preferred over a Bunsen burner. Do not bring chemicals to your bench. Leave them in some common area or hood. Your instructor will tell you how to get the appropriate amounts to your bench. Dispose of chemicals in the proper waste container. Always wear <u>goggles</u>, a <u>lab coat</u>, and <u>gloves</u> inside the lab. You will be penalized if you fail to do so.	**Chemicals needed:** Methanol [Boiling Point: 65 ºC] Acetone [Boiling Point: 56 ºC] Isopropanol [Boiling Point: 83 ºC] Hexane [Boiling Point: 68 ºC] **Instruments Needed:** 100 mL Graduated cylinder, boiling chips, a pin needed for making a hole in the middle of the aluminum foil, waste container. See the diagram below for hardware.

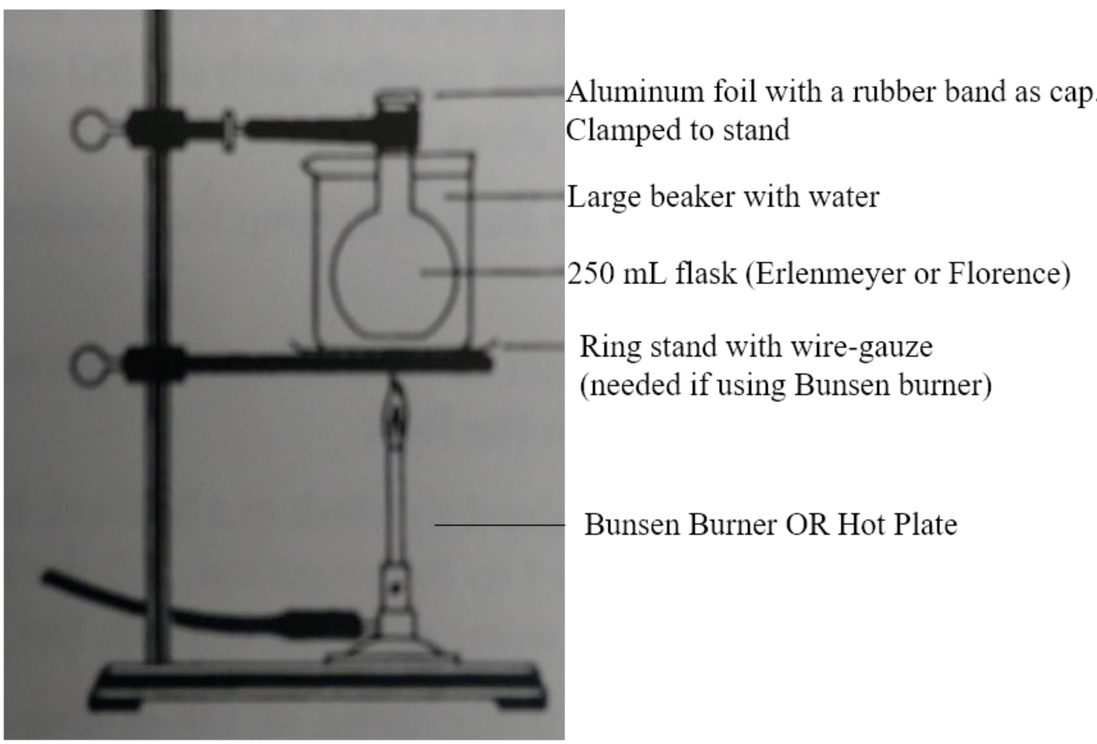


Figure: Set-up for molar mass determination of a volatile liquid

Experimentation

⟐Focus Question 1: Calculate the number of grams and the volume of isopropanol (l) required to fill a 250 mL flask with vapors of isopropanol at 100 ºC and the normal atmospheric pressure of 1 atm. *[density of isopropanol = 0.786 g/cm³]*

[**NOTE TO INSTRUCTOR:** Probe students with statements like, take a moment to look at the setup for this experiment. Where are you going to put the volatile liquid and how do you plan to convert it to gas and maintain its temperature?

The following calculations help them apply the ideal gas law to the current situation. Without giving answers, probe them to think about it and make some predictions.

Using PV = nRT, R = 0.0821, T = 373, P = 1, V = 0.250 , we can calculate that n = 0.00816 mol. In grams of isopropanol (molar mass = 60.1), it is 0.49 g. In volume, it is 0.62 mL. Both mass and volume are important for this experiment. The mass is important because the balance that they use should be able to measure this correctly. Volume is important because they need to add more than this volume to the flask, take it to the boiling temperature of water, and allow excess to evaporate. Their reasoning for deviation could be experimental errors in not being able to know if all liquid has evaporated, whether all the gas have the same temperature inside flask, whether water will boil at 100 in today's atmospheric pressure which may or may not be 1 atm, flask's volume may be more than 250, etc. You can probe them with questions on these issues if they are stuck on not being able to make any guesses. There may be water sticking to the aluminum foil. Almost all of these errors can cause the weight to be more than it should be.] add densities of each metal

Predict if the experimental determination of this mass will be more or less than this amount. Write your reasoning and discuss with your group.

If the gases are actually attracted to each other irl, then there will be more molecules in there than the kmt would lead us to believe… making the mass higher.

Experimental verification: From your calculation of the volume of isopropanol needed, pick an amount much higher than that. Typically, 3 mL should be excess enough for an experiment. Weigh a clean, dry 250 mL Erlenmeyer flask along with Aluminum foil cap and a rubber band. Put the 3 mL liquid inside the flask, close it with aluminum foil and put a rubber band to hold the foil in its place. Make a pinhole at the center of the foil so that excess vapor can escape. Have the set up using hot plate and beaker with water as shown in the diagram (or as shown by your instructor). Choose the appropriate size beaker and figure out how much water you need to put in the beaker and write that amount. The Erlenmeyer flask with the volatile liquid can be lowered into the beaker in such a way that the water is touching the neck of the flask, but not the aluminum foil. Put some boiling chips in water and bring the water to a boil.

[**NOTE TO INSTRUCTOR:** If the students made a mistake of having long Aluminum foil as a cap, you can tell them to cut a small square and use it. However, if they still used a long protruding one, it may be touching the water and gathering some water vapor and it could affect their results. Since this is the first trial, they can afford to be wrong, but allow them to correct it and redo the experiment if needed.]

(i) Note down the amount of time needed to see all the volatile liquid evaporate and become gas. Visually confirm if all the liquid has evaporated and only gas is occupying the flask.

(ii) Take a moment and think about how you will you record the temperature and pressure of the gas inside Erlenmeyer flask. ␥Consult with your instructor after you agree on the correct method.

(iii) Take the Erlenmeyer flask out of the hot water and allow it to cool down to near room temperature. You can wipe the outside of the flask with a paper towel and make sure you can see the condensed liquid inside the flask. If the outside is cool and dry, you can weigh it and see how much liquid has condensed. Does it agree with your prediction?

(iv) Think of some of the errors that are possible and discuss how they can be overcome. How can you measure the actual volume of the flask? Why is a pinhole is needed instead of closing the flask completely OR keeping it wide open? As you think about these questions, you can refine your experiment to minimize errors.

(v) Using the experimental volume, calculate the molar mass of isopropanol.

Note that the condensed liquid must be disposed of in the waste container.

(vi) ␥Write your experimental steps clearly and show it to your instructor.

[**NOTE TO INSTRUCTOR:** Atmospheric pressure is the pressure of the gas inside flask because the pinhole made it an open system with the atmosphere. Therefore, noting today's atmospheric pressure using barometer will be enough. If the hole is too big, vapor will escape easily and air will fill in. If there is no hole, the pressure will build up due to vaporizing liquid and the flask may explode. The temperature of boiling water is the temperature of the gas because they are in contact with each other for a sufficient amount of time. Thus, we do not worry about putting the thermometer inside the flask that would have made the gas escape easily. It is also important to stop the evaporation process as soon as the liquid has evaporated. Prolonged heating can cause the more vapor to escape and replaced by air. The experiment can be improved by bringing the water to boil first and then lowering the Erlenmeyer flask. To measure the volume of Erlenmeyer flask, they can fill it with water and measure the amount using a 100 mL graduated cylinder.]

Experimentation Week 2

⟐Focus Question 2: Predict which of the four volatile liquids – methanol, isopropanol, acetone, and hexane – will deviate the most from ideal gas behavior at 100 °C.

Write your reasoning and discuss with your group.

(i) With your written lab procedure from the previous week, prepare your dry Erlenmeyer flasks, and carry out the molar mass determination of all four compounds. Fill in the table as shown below.

Compound	Boiling Point (°C)	Expected Molar Mass	Experimental Molar Mass
Methanol			
Isopropanol			
Acetone			
Hexane			

Discuss the patterns you see from your data amongst your group members. If needed, your instructor may initiate a class discussion to pool different views of students. Based on your results, write your final lab report. Your report must answer the question which is (are) most ideal and which one deviates the most. Explain your reasoning. Make sure your lab report contains everything mentioned by your instructor.

[**NOTE TO INSTRUCTOR:** Students may see some correlation between boiling points and molar mass to the ideal gas nature. Boiling points is an indicator of intermolecular attractions that must be negligible as per KMT. Thus, higher boiling points may show more deviations from ideality. In general, intermolecular interactions go up with the size of the molecule unless special forces like H-bonding are present. In addition, the molecular volume is neglected in KMT. As the size of the molecule increases, this becomes a factor, as the empty flask volume is lower than that of smaller molecules. All gases behave ideally if the following conditions are met –high temperatures or very low pressures. The post-lab questions bring out these concepts.]

POST LAB QUESTIONS

1. If the atmospheric pressure is 754 torr, what is the value in atm units?
2. Assume that all the four liquids in today's experiment are heated to 1000 ^{O}C using a heat chamber; will they behave more ideal or less ideal gas compared to 100 ^{o}C? Explain.
3. Which of the three gases -He, N_2, and CO_2 you will expect to be the most ideal at STP? Explain.
4. Explain why most gases behave ideally at lower pressures than at higher pressures.
5. Explain what simple adjustments were made by van der Waals to 'P' and 'V' terms in the ideal gas law to make it applicable to real gases.

PreLab – Experiment #5: Which is the most ideal? **Name: ____________________**

R = 0.0821 L•atm/mol•K

1. State the three postulates (principle statements) of Kinetic Molecular Theory.

2. As per KMT, which one will have a higher volume at the same temperature and pressure – one mole of Helium OR one mole of Neon?

3. 2.50 g of CO_2 is taken in a 0.400 L rigid container at 25 ^{o}C. What will be the pressure in 'atm' units inside the container?

4. The barometric pressure is 753 Torr. What is the barometric pressure in 'atm' units?

5. At 0.95 atm and 98 ^{o}C, how many moles of an ideal gas can occupy a 250 mL flask? If the mass of the gas filling the flask is 0.745 g, what is the molar mass of that gas?

Experiment #5: Gas Law – **10 points**

1. Volume: volume of individual particles = 0

   Motion: Particles are in constant motion

   Collisions: Particles exert no force on each other, only on the walls of the container

   [3 pts for at least 3 of the ideas]

2. Same volume [1 pt]

3. 3.47 atm [2 pts]

4. 0.991 atm [2 pts]

5. 0.0078 mol, 96 g/mol [2 pts]

Quiz (Experiment #5: Gas Law)

10 points - 15 minutes **Name:______________________________**

Ideal Gas Law: $PV = nRT$, density = PM/RT Van der Waals Eqn: $(P+a/n^2)(V\text{-}nb) = RT$
$R = 0.0821$ L·atm/mol·K 1 atm = 760 torr = 760 mm (millimeters of mercury)

1. The barometric pressure is 780 torr. What is the barometric pressure in 'atm' units?

2. Calculate the density in g/L units for CO_2 at 780 Torr and 298 K.

3. The boiling point of liquid A is 40 ^{o}C and liquid B is 90 ^{o}C. While both become gases at the boiling point of water, which one will deviate the most from ideal gas law at that temperature?

4. Gas 'A' has a molar mass of 32 g/mol while gas B has a molar mass of 90 g/mol. Both are gases at room temperature. Which one will deviate the most from ideal gas law at room temperature?

5. The Van der Waals equation corrects the ideal gas law equation by adjusting the measured pressure upwards, making it $P + a/n^2$ ['a' is a constant and it depends on the nature of gas]. Explain why the measured pressure may be lower that warrants this adjustment for the gas law equation.

6. The Van der Waals equation corrects the ideal gas law equation by adjusting the measured volume downwards, making it V-nb ['b' is a constant and it depends on the nature of gas]. Explain why the measured volume may be larger that warrants this adjustment for the gas law equation.

1. 1.03 atm
2. 1.96 g/L
3. Liquid B because of stronger intermolecular forces
4. Gas B will deviate the most because it is the heaviest.
5. The measured pressure is lower because the particles are actually colliding into each other. The KMT states that **all** of the force exerted by the particles is on the walls of the container. The measured value doesn't take the collisions between particles into account, thus reporting a smaller value.
6. The measured volume will be higher because the particles actually have volume. The KMT states that the particles have no volume, so the actual volume of the particles is being factored into the measured value.

Experiment #6: Are You Prepared to Counter an Acid Spill?

FOR THE INSTRUCTOR

Chemicals required: 1 M HCl, 1 M NaOH, solid potassium hydrogenphthalate (KHP), solid sodium bicarbonate, phenolphthalein solution.

Instruments/material: Balance, calorimeter, thermometer, beakers, Erlenmeyer flask, burette, 20 mL pipettes (if available), burette stand and clamps.

Information for stockroom personnel

COVERS 2 LAB PERIOD

Make sure to provide a **full** bottle of HCl and NaOH solution enough for two lab periods. The contents of the chemicals must remain unaltered and not refilled for the 2nd week, therefore enough must be provided for the lab. Failure to maintain the solution will defeat the purpose of the 2nd lab.

-or-

During the first week of the lab, students will standardize both the HCl and NaOH solutions. It is imperative that a full bottle of each solution be designated for each specific class and retained for the following week. Failure to maintain the solution will render the first week's experiment useless.

Chemicals required: 1 M HCl, 1 M NaOH, solid potassium hydrogenphthalate (KHP), solid sodium bicarbonate, phenolphthalein solution.

Instruments/material: Balance, calorimeter, thermometer, beakers, Erlenmeyer flask, burette, 20 mL pipettes (if available), burette stand and clamps.

Experiment #6: Are You Prepared to Counter an Acid Spill?

Learning Objectives

Content objectives
Students will be able to:

- Predict the products of neutralization reactions.
- Write the ionic equation for acid-base neutralization reactions.
- Differentiate weak and strong acid/base and the meaning of pH
- Understand how “concentrated” differs from “strong.”
- Associate the heat energy from chemical reactions to the chemical bonds that are broken and formed.
- Determine which base is most appropriate to clean an acid spill.

Process objectives
Students will be able to:

- Analyze data for patterns and draw appropriate conclusions.
- Perform a titration.

Overview

In the first week, students will perform a titration to verify the concentrations of HCl and NaOH. In the second week, students will use the standardized acid and base to perform neutralization reactions.

Explain what acids and bases are and what a neutralization reaction is. Go over the products of neutralization reactions. Students must verify the concentrations of the acid and base (first week) in order to perform a complete neutralization reaction (second week).

Review the steps of preparing a buret for titration. Make sure to tell students to wash the buret at least two times with the basic titrant.

Goal: Determine if a weak base or strong base is more appropriate to clean an acid spill.

Things to look out for:

1.When circulating and questioning groups, pose questions to the student whom you feel is lost/relying too much on their group.

2. If a new batch of NaOH and HCl is given the second week, students may use the standardized concentration given in the modification portion of the lab.

Experiment #6: Are you prepared to counter an acid-spill?

Your Project:

You are an intern working in a chemistry laboratory with concentrated acids. Your supervisor wants to develop a protocol for handling accidental acid spills. In the cabinet for bases, there are two bottles available: sodium hydroxide and sodium bicarbonate or sodium hydrogencarbonate (baking soda). Your goal is to determine which chemical is more suitable to clean up the acid spill and write a protocol. You must also explain your scientific reasoning very clearly in your report.

This is a 2-week lab. In lab 1, you need to figure out the exact concentrations of 1 M HCl and 1 M NaOH. To do the standardization, you are given an acid called potassium hydrogenphthalate (KHP) which serves as a primary standard. Once the exact concentrations are known, you plan your experiment to figure out the relationship between the heat of reaction and the amount of substance used for neutralization.

In lab period 2, you will carry out the experiment to compare the heats of reaction of HCl with the two different bases and explain which one may be a better choice.

📖 **Recommended Readings: Acid-base reactions, stoichiometry, balancing ionic equations, enthalpy, and calorimetry, titration.**

Background:

An acid is a compound which releases hydrogen ions in solution. Hydrobromic acid, HBr, is an example of an acid. A base is a compound which accepts hydrogen ions. Potassium hydroxide, KOH, is an example of a base. A reaction between an acid and a base is called a neutralization reaction. An example of a neutralization reaction is:

$$HBr + KOH \rightarrow H_2O + KBr$$

Some acids and bases are found in the typical kitchen pantry; for example, acetic acid is vinegar and sodium bicarbonate is baking soda. However, other acids and bases are only found in laboratories and can cause serious injury if they make contact with skin. It is clear that some acids and bases can be considered “weak” (such as vinegar and baking soda) while others are “strong” (such as HBr and KOH). What makes an acid/base strong or weak? *A strong acid/base completely dissociates in water while a weak acid/base only partially dissociates in water.*

The pH scale is used to measure how acidic or basic something is in aqueous solution. pH (power of hydrogen) is equal to the -log[H^+]. Neutral substances have a pH of 7, implying that they have equal amounts (& in a very small concentration of 10^{-7} M) of H^+ and OH^-. pH less than 7 is acidic and more than 7 is a basic solution.

Heat of Reactions

Any reaction can be classified as endothermic or exothermic. An exothermic reaction releases heat while an endothermic reaction absorbs heat. In determining if heat is absorbed or released, it is important to define the system and surroundings. For example, a cup of water is placed in the freezer which will freeze the water. If the system is the cup of water then we say that heat is released by the system as water becomes ice. The heat released is picked up by the freezer and it gains that energy. *Heat always flows from a hot body to cold body.* Liquid water is hot compared to the freezer. The opposite situation applies when you hold a piece of ice on your hand. Your hand is hot compared to ice and energy flows from your hand to ice which melts it. Thus, ice to water is endothermic as far as ice is concerned because it took energy from your hand to melt it. For your hand, it is exothermic because heat is released from your hand. It is important to know that *energy is conserved.* Thus, exothermicity or endothermicity of a chemical reaction is whether energy is released or absorbed by the reaction at room temperature. Just like the ice to water example, if a reaction releases heat (exothermic), it will warm up the surroundings, and if takes heat from surroundings (endothermic), it will feel chilly in the immediate vicinity.

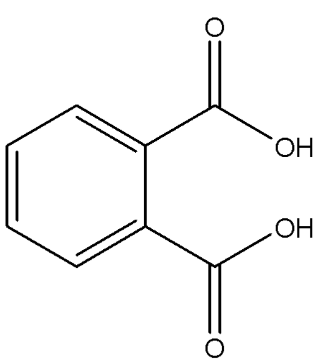
The heat absorbed or evolved in a reaction can be represented by q. Since energy is conserved, if the surrounding is a defined material like water or metal, we can calculate the change in temperature of that substance using the relation, $q=ms\Delta t$. 'm' represents the mass of the substance. Δt represents the change in temperature of the substance. 's' represents the specific heat capacity of the substance. *The specific heat capacity is the heat energy needed to raise the temperature of 1g of substance by 1 degree Celsius.*

Goal: Determine if a weak base or strong base is more appropriate to clean an acid spill.

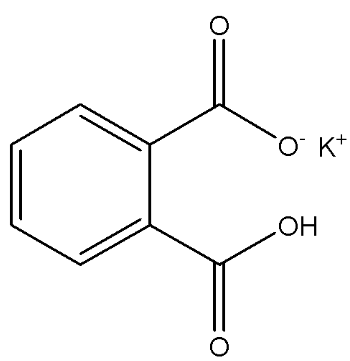
Safety	**Materials**
Both acids and bases are CORROSIVE, meaning that skin contact with either one will result in severe burns. Be very careful when handling acids and bases. Do not directly inhale the chemicals. Always wear <u>goggles</u>, a <u>lab coat</u>, and <u>gloves</u> inside the lab. You will be penalized if you fail to do so.	**Chemicals needed:** 1 M HCl, 1 M NaOH, solid potassium hydrogenphthalate (KHP), solid sodium bicarbonate, phenolphthalein solution. **Instruments needed:** Balance, calorimeter, thermometer, beakers, Erlenmeyer flask, burette, 20 mL pipettes (if available), burette stand and clamps.

Experimentation

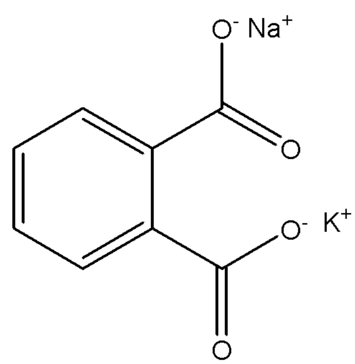
KHP is a monoprotic acid with the formula $K(HC_8H_4O_4)$. Its molar mass is 204.2 g/mol. It is a primary standard because it is a stable solid substance at room temperature. It is very useful to determine the concentrations of other bases like NaOH which are hygroscopic. It is to be noted that phthalic acid is a diprotic acid (see structure below). When one of the acidic hydrogens is neutralized, it forms KHP. The second hydrogen ion can be neutralized by one equivalent of NaOH by the reaction,
$K(HC_8H_4O_4)$ (aq) + NaOH (aq) → $KNa(C_8H_4O_4)$ (aq) + H_2O



phthalic acid



potassium hydrogen phthalate



sodium potassium phthalate

Focus Question 1: Predict how many grams of KHP is needed to neutralize 5 mL of 1.0 M NaOH. Explain the basis for your prediction.

[**NOTE TO INSTRUCTOR:** 5 mL of 1.0 M NaOH is 0.005 moles NaOH. As per the neutralization reaction, 0.02 moles of NaOH requires 0.02 moles of KHP which is 0.005 x 204.2 g = 1.02 g.]

It is time to experiment and verify if the number of grams you calculated that will neutralize exactly 5 mL. Your instructor will demonstrate to you how to line the burette with the NaOH solution you are going to use.

✎Make sure you write the instructions in your experimental procedure. Take 1 M NaOH in a clean burette.

(i) In a clean and dry Erlenmeyer flask, add your estimated amount of KHP. You need to know the exact mass that you are adding to the flask. Use a weighing paper to weigh the solid KHP. Weigh the weighing paper before adding and after adding to make sure of the exact mass that got transferred to the flask. ✎Write the appropriate instructions in your experimental procedure.

(ii) Dissolve the KHP in the Erlenmeyer flask using about 25 mL of water. Is the water volume needs to be precisely 25 mL? Why are you adding water to KHP? Think about this and discuss with your group. ✎Make the appropriate notes in your experimental procedure.

[**NOTE TO INSTRUCTOR:** Amount of water is not important, but dissolving KHP makes the reaction go faster and to completion. Thus, they can add a bit of extra water if all KHP is not in solution.]

(iii) Add one or two drops of phenolphthalein indicator to the flask. Titrate against NaOH. How do you know you have reached the end point of the titration? Note the endpoint of the titration. Calculate the volume of NaOH used and the concentration of NaOH.

(iv) Do a second trial. Decide if you need to refill the burette with more NaOH. Calculate the volume of NaOH used and the concentration of NaOH. If both trials do not agree with each other, do a third trial.

(v) Note the correct concentration of NaOH. You can average the two trials that closely agree with each other.

Focus Question 2: Predict the volume of NaOH needed to neutralize 20 mL of 1M HCl.

[**NOTE TO INSTRUCTOR:** It is based on the number obtained by them for [NaOH]. If it also 1 M, then 20 mL of HCl will need 20 mL of NaOH. If the concentration is ‘x’ M NaOH, then mLs needed will be = 20 x 1.0 /x]

(i) For this experiment, you are measuring 20 mL of 1 M HCl using a pipette. Your instructor will demonstrate how to use a pipette. If a pipette is not available, you can use a smaller size graduated cylinder. Take the 20 mL of 1 M HCl in the flask, add one or two drops of phenolphthalein and carry out the titration. Note the volume of NaOH used and calculate the concentration of HCl.

(ii) Do a second trial. If both trials do not agree, do a third trial.

(iii) Did the concentration of NaOH and HCl come out to be exactly 1 M? If they are not exactly 1 M as indicated, explain what may have caused the numbers to be off.

Focus Question 3: Predict if the amount of heat produced by the neutralization reaction between 1 M HCl and 1 M NaOH will depend on the amount of materials used? Will the temperature change be the same if 50 mL of 1M HCl is added to either 25 mL of 1 M NaOH OR 50 mL of 1 M NaOH? Explain

After you make your prediction, think about carrying out the experiment.

Write the procedure. You can complete the experiment in lab period 2.

Experimentation Week 2

Based on the actual concentration of your HCl and NaOH, calculate how many mLs of NaOH is required to neutralize 50 mLs of HCl. You may want to divide the mLs of NaOH required by three and round it off to the nearest 0.1 mL so that it can be measured using a graduated cylinder. In one experiment, you can add one-third of the total NaOH needed, in a second experiment, add two-thirds, in a third experiment add the correct stoichiometric amount, and in a fourth one add 10 mL more than the stoichiometric amount of NaOH. Thus, you can create 4 different experiments where the amount of HCl is the same, but the added NaOH is different. Each time measure the temperature change in the reaction mixture.

[**NOTE TO INSTRUCTOR:** If a student calculated 48 mL of NaOH is needed for stoichiometric addition, then he/she will need to add 16.0 mL for experiment 1, 32.0 mL for experiment. 2, 48.0 mL for experiment. 3, and 58.0 mL for experiment. 4.]

(i) Weigh an empty calorimeter with lid. Add 50 mLs of HCl. Measure its temperature. Add the appropriate amount of NaOH you calculated for the first experiment. Stir it and measure the temperature change. You may notice the temperature go up and stay constant in about 30 seconds to 1 min. ✎Write the exact details in your procedure.

(ii) Take the mass of the calorimeter with lid again. You will be able to calculate the mass of the water and using the temperature change, calculate the amount of heat energy produced in the reaction. [Use 's' for solution same as that of water which is 4.184 J/g°C]. *Discard the solution into a waste container.*

(iii) Weigh the calorimeter with lid and add 50 mLs of HCl. Measure its temperature.

(iv) This time, add the amount of NaOH you calculated for the second experiment and repeat the process for recording the change in temperature. Did the temperature change the same amount as your previous experiment?

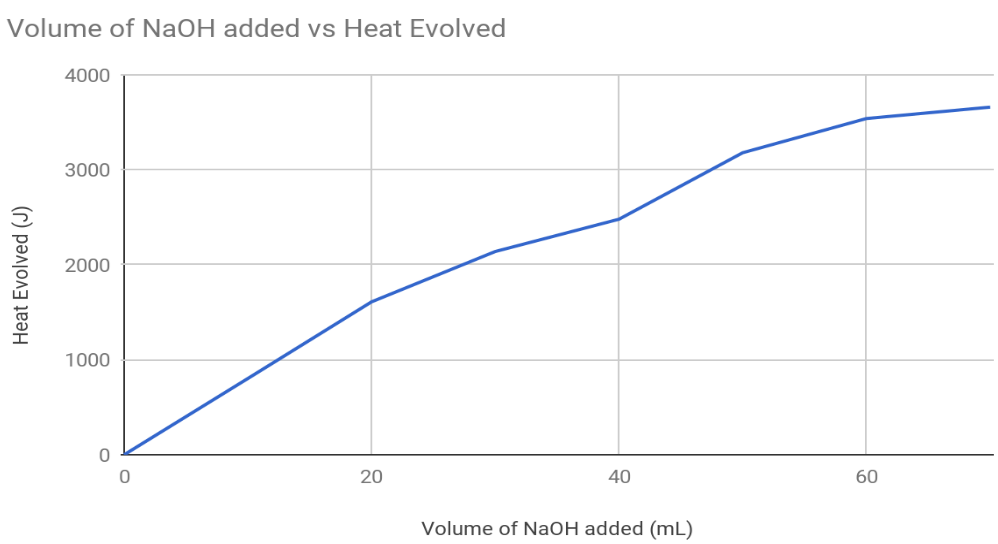
(v) Measure the mass of calorimetry+lid+solution to figure out the exact mass of solution inside. How much heat energy was released in this experiment? Does it match the same as your previous experiment?

(vi) Carry out the third and fourth experiments and calculate the temperature change and heat energy change. Record your data below.

Expt#	mLs of NaOH added	Temp change (°C)	Heat Energy Released by the reaction
1			
2			
3			
4			

(vii) Discuss amongst your group and comment on the results. Make a graph heat energy on 'y' axis and mLs of NaOH added on the 'x' axis. What kind of trend is seen? Does it agree with your prediction?

[**NOTE TO INSTRUCTOR:** A typical graph looks like the one below]



(viii) For each of your experimental results, calculate the number of moles of H_2O formed and the heat energy change per mole of water formed and complete the table below. Explain the significance of the value in the last column for each of the experiment.

Expt #	Moles H_2O formed in the reaction	Heat Energy Released by the reaction	Heat energy per mole of H_2O formed
1			
2			
3			
4			
5			

[**NOTE TO INSTRUCTOR:** The last column is expected to show a value in the range of 55 to 65 kJ/mol]

⎆Focus Question 4: Predict if the heat energy per mole of water produced will be lower or higher if baking soda ($NaHCO_3$) is used to neutralize HCl instead of NaOH. Put your reasoning in words here and also show the balanced chemical equation]

[**NOTE TO INSTRUCTOR:** Predictions do not have to be correct but must have logical reasoning. Make sure they write the balanced reaction here.
In molecular form: $NaHCO_3$ (s) + HCl (aq) → NaCl (aq) + H_2O (l) + CO_2 (g)
If they write it in solution form: $NaHCO_3$ (aq) + HCl (aq) → NaCl (aq) + H_2O (l) + CO_2 (g)
They may also show the net ionic equation: HCO_3^- (aq) + H^+ (aq) → H_2O + CO_2 (g).
If not, drive them to show that. Once they write it, they may be able to reason out that this reaction is different from the strong acid-strong base neutralization. A certain amount of energy may be needed to break the H^+ from HCO_3^-.
Note: If we use sodium carbonate instead of baking soda, the reaction will look like:
Na_2CO_3 (s) + 2 HCl (aq) → NaCl (aq) + 2 H_2O (l) + CO_2 (g).
The net ionic equation (assuming sodium carbonate is in solution) will look like:
CO_3^{2-} (aq) + 2 H^+ (aq) → H_2O + CO_2 (g).
Here again, the reasoning for a different heat of reaction is that the reactant is not with hydroxide]

Now it is time to experiment. Since baking soda ($NaHCO_3$) is a solid, you can calculate the number of grams of baking soda needed to neutralize 50 mL of your HCl.

(ix) ✎Write the procedure clearly outlining how you will carry out this experiment using calorimeter. ⦸ See instructor for suggestions and approval before proceeding.

(x) ☒ You may need to do 2 trials. Calculate the number of moles of water formed and calculate standard heat of neutralization of this reaction per mole of water formed. Don't forget to outline every step needed.

[**NOTE TO INSTRUCTOR**: The molar mass of baking soda is 84 g. For 50 mL of 1 M HCl (0.05 mol HCl), students will need about 4.20 g of baking soda. They can weigh the calorimeter with lid, note the temperature of the HCl solution, add the 4.20 g of baking soda. Since the reaction releases CO_2, there will be fizz. Once the fizz settles, we will know the reaction is complete. This should be fairly quick. There will be initial cooling because of the fizz. Students should be careful not to record this initial cooling effect. After stirring, the temperature will increase or decrease. They need to weigh the solution to get the mass of solution.]

Trial #	Moles H_2O formed in the reaction	Heat Energy Released by the reaction	Heat energy per mole of H_2O formed
1			
2			

FINAL LAB REPORT:

Now that you have done some experiments with the two substances, NaOH and $NaHCO_3$ to neutralize acid spills, explain if one of them is preferable over the other OR it doesn't matter. Your lab report must summarize your experimental data as well as any other relevant information like cost, safety in handling, etc. in reaching conclusions.

PreLab – Exp #6: Are you prepared to counter an acid-spill? Name: ______________________

1. An acid, HA is 90% unionized. Is this a strong or weak acid?

2. (a) NaOH is a strong base. What are the species that will be found in an aqueous solution of NaOH?

   (b) Ammonium hydroxide is a weak base. What are the species that will be found in an aqueous solution of NH_4OH?

3. Write the balanced ionization reaction of sodium bicarbonate (aka sodium hydrogencarbonate or baking soda) in water.

4. Write the balanced neutralization reaction for hydrochloric acid and sodium hydroxide. Show the equation in molecular form and also show the net ionic equation.

5. How many mLs of 0.95 M HCl is required for the complete neutralization of 25.0 mL of 0.85 M NaOH?

6. If a reaction produces 25.0 kJ of heat and if all the heat is transferred to 400.0 grams of water, what will be the change in temperature of the water? [‘c’ for water is 4.184 J/g°C]

7. (a) 0.127 moles of Mg are converted to MgO by complete combustion with excess oxygen. How many moles of MgO will be formed in the reaction?

   (b) If the heat released by the reaction is transferred to 250.0 grams of water surrounding the reaction vessel, the temperature went up by 4.50 °C, how many Joules of heat energy was released by the reaction? [‘c’ for water is 4.184 J/g°C]

   (c) Calculate the standard molar heat of formation of MgO using the above experimental data.

PreLab – Exp #6: Are you prepared to counter an acid-spill?

1. An acid, HA is 90% unionized. Is this a strong or weak acid?

   [1 point] weak acid

2. (a) NaOH is a strong base. What are the species that will be found in an aqueous solution of NaOH?
   [1 point total. 0.5 point each] Na+ (aq) and OH- (aq)
   (b) Ammonium hydroxide is a weak base. What are the species that will be found in an aqueous solution of NH_4OH?

   [1.5 total, 0.5 pt each] NH3, NH4+ (aq), OH- (aq), and NH3 (aq) (NH4OH acceptable)

3. Write the balanced ionization reaction of sodium bicarbonate (aka sodium hydrogencarbonate or baking soda) in water.

   [1 point] NaHCO3 → Na+ (aq) + HCO3- (aq)
   [If the answer is written to include HCO3- + H2O → H2CO3 + OH-, add 0.5 point bonus]

4. Write the balanced neutralization reaction for hydrochloric acid and sodium hydroxide. Show the equation in molecular form and also show the net ionic equation.
   [1 pt] NaOH (aq) + HCl (aq) → NaCl (aq) + H2O (l)

   {1 pt] Net ionic: H+ (aq) + OH- (aq) → H2O

5. How many mLs of 0.95 M HCl is required for the complete neutralization of 25.0 mL of 0.85 M NaOH?
   [1 pt] 22.4 mL

6. If a reaction produces 25.0 kJ of heat and if all the heat is transferred to 400.0 grams of water, what will be the change in temperature of the water? ['c' for water is 4.184 J/g^{o}C]
   [1 pt] 14.9 ^{o}C

7. (a) 0.127 moles of Mg are converted to MgO by complete combustion with excess oxygen. How many moles of MgO will be formed in the reaction?
   (b) If the heat released by the reaction is transferred to 250.0 grams of water surrounding the reaction vessel, the temperature went up by 4.50 ^{o}C, how many Joules of heat energy was released by the reaction? ['c' for water is 4.184 J/g^{o}C]

   (c) Calculate the standard molar heat of formation of MgO using the above experimental data.

QUIZ

An acid, HX is 100% ionized. Is this a strong or weak acid?

$Mg(OH)_2$ is a strong base, but it is not very soluble in water. Only 0.0011 moles of $Mg(OH)_2$ can dissolve in a 1 L solution. What are the concentrations of magnesium and hydroxide ions in a 0.0011 M solution of magnesium hydroxide?

Even though $Mg(OH)_2$ is insoluble in water all the $Mg(OH)_2$ dissolved when it was sprinkled on an aqueous hydrochloric acid spill. Explain why this happened.

Acetic acid (formula: $H(C_2H_3O_2)$) is a weak monoprotic acid. It is soluble in water, but it is a weak acid. What species are present in a solution of acetic acid?

Write the balanced net ionic equation for (a) hydrochloric acid with sodium hydroxide and (b) acetic acid with sodium hydroxide.

Both neutralization reaction is mentioned in question four are exothermic. Even though both are exothermic, one of them generated less heat energy per mole. Identify which one of the two produced less heat energy (a) hydrochloric acid with sodium hydroxide or (b) acetic acid with sodium hydroxide.

Explain the reason for your choice.

QUIZ SOLUTION

Experiment #7: Heat Capacity

FOR THE INSTRUCTOR

Chemicals required: 0.02 M potassium permanganate, household hydrogen peroxide, 1 M or 3 M H_2SO_4

Instruments/material: Burette, 2 to 5 mL graduated pipette, 10 mL graduated cylinder, 50 or 100 mL graduated cylinder, 250 mL Beaker or Erlenmeyer flask

Information for stockroom personnel

COVERS 1 LAB PERIOD

Chemicals needed:

- 0.02 M potassium permanganate
- household hydrogen peroxide
- 1 M or 3 M H_2SO_4

Experiment #7: Which metal is a better choice for making a cooking utensil?

Learning Objectives

Content objectives
Students will be able to:
- Justify the experimental set-up.
- Understand how molecular mass correlates to the specific heat capacity
- Determine which metal is most appropriate for the inner and outer layer of the cooking utensil.
- Predict which metal will have the highest specific heat capacity.

Process objectives
Students will be able to:
- Analyze data for patterns and draw appropriate conclusions.
- Use deductive reasoning to decipher between

Overview

Explain the terms heat capacity and specific heat capacity and the units for each. Explain the difference.
Explain the layering of a pot. Ideally, the metal layer in contact with the heat should heat up quickly (low specific heat) while the layer in contact with the food should stay hot for a long period of time.

Goal: Determine which metal is most suitable for the creation of a cooking utensil.

Things to look out for:

1.When circulating and questioning groups, pose questions to the student whom you feel is lost/relying too much on their group.

2. If a new batch of NaOH and HCl is given the second week, students may use the standardized concentration given in the modification portion of the lab.

Experiment #7: Which metal is a better choice for making a cooking utensil?

Your project:

You are an intern working in a utensil manufacturing company. Your company is interested in making utensils that can heat quickly, lower raw-material cost, easy to manufacture, can withstand temperatures of 700 ºC, and durable for repeated use. While it is a handful to tackle all at once, in this project, your manager wants you to focus on the material's ability to heat quickly. You are given 6 different materials, but all have the same size.

You are given 20 mm cubes of aluminum, copper, zinc, iron, lead, and brass. Even though the same size is given, it is important to characterize the ability to heat up based on the material's weight. Thus, your manager wants to know two things from you – for the same mass of material, which of the 6 metals will absorb heat quickly from the stove. Secondly, for the same mass, which metal will hold more heat and release it slowly so that items can slow cook even after a stove has been turned off.

📖 **Recommended Readings: Thermochemistry and Specific heat capacity.**

Background:

Aluminum is a common light-weight metal that is also very cheap. However, it is less desirable to use 100% aluminum in the making of utensils. Usually, layers of different metals are used in the construction to add various properties to the utensil. The most important properties are its ability to absorb heat and cook evenly. Properties like shine after repeated use, ease of maintenance, withstand overheating, etc., adds to the constant search for new materials. Multiclad or polyclad refers to the layers of metals that make up the utensil. One layer is exposed to the heat source while another layer made of a different metal is exposed to the food that is cooking. The layer in contact with heat source must absorb heat quickly while the layer in contact with the food must retain that heat longer so that it is not lost to the surrounding air quickly.

Specific Heat Capacity:

When heat energy is given to a substance, it is absorbed by that substance causing its temperature to go up. Heat flows from the hot body to the cold body until both reach the same temperature. The hotter substance will drop down in temperature as it gives off heat to the colder substance. The amount of heat energy lost by the hot substance equals the amount of heat energy gained by the cold substance. Water can be chosen as a standard to receive the heat so that the amount of heat energy transferred to it can be measured easily by using the standard specific heat capacity of water, 1 cal per gram per ºC (= 4.184 J per gram per ºC.) *The specific heat capacity is a measure of the amount of heat energy required to change one gram of material by one degrees Celsius.* Substances which do not require too much energy to heat up by 1 ºC, will become hot very quickly. By the same token, they also lose that heat quickly. Thus, measuring and analyzing the specific heat capacity is very helpful to choose new materials.

Goal: Determine which metal is most suitable for use as a bottom layer of a cooking utensil.

Safety	Materials
Always wear goggles, a lab coat, and gloves inside the lab. You will be penalized if you fail to do so.	**Chemicals:** 20 mm cubes of aluminum, copper, zinc, iron, lead, and brass. **Instruments:** Hotplate, 500-mL beaker, string, thermometer.

Your experimental set-up will be demonstrated by your instructor. Make sure you are clear with the terms heat capacity and specific heat capacity.

Experimentation

Focus Question 1: You are given 6 metal cubes. Examine them. If each of these cubes is dropped in boiling water and allowed to absorb heat, do you expect them to absorb the same or different amounts of heat energy to reach the boiling water's temperature?

Make your prediction, discuss amongst your group members and summarize your reasons.

[Note: this is a prediction based on your experience and logical arguments within your group. Don't search the internet for the "correct" answers.]

[**NOTE TO INSTRUCTOR:** Probe students with questions if they are completely lost.

Some examples are given below:

S: Are you asking if all metals will reach the same temperature?

Instructor: Read the question. [It is clearly mentioned that they all reach the same temperature]

S: Are you asking for heat capacity or specific heat capacity for each piece?

Instructor: Read the question. You place the entire metal cube in hot water and predicting the heat energy absorbed. [If they are still confused, give them the clue that it is 'q'.]

Note: You're not asking for either. You're trying to convey the idea that the heat energy absorbed is NOT the heat capacity or specific heat capacity.

S: Don't we need to know the specific heat capacity of the metal to answer this question?

Instructor: Use your experiential learning. Let your mind wander and think about instances which may give you some clues. [Examples are their observations in kitchen, metal statues, the relation between the type of metal and its specific heat capacity, etc.]

Some students may have an idea of specific heat of different metals and they may ask the following question

S: Shouldn't we need the mass of each metal cube to answer this question?

Instructor: Maybe. Feel them in your hand and you may be able to make some guesses, if not exact masses at least relative masses.

Most groups may come up with one or two ideas amongst the following –

they will absorb the same amount of heat because they are in the same hot water.

they may have different heat capacities which will affect the heat absorbed

since they are different elements, both mass and specific heat capacity will affect the total heat absorbed, etc.

We are not trying to correct them at this point. At least they should know that it is the total heat absorbed by the metal cube that is in question. And they are given the same heat source.

IF THEY HAVE ANSWERED THAT EACH METAL CUBE WILL ABSORB DIFFERENT AMOUNTS OF HEAT ENERGY, ASK THEM TO PREDICT THE ORDER FROM HIGHEST ABSORBER TO LOWEST ABSORBER.]

It is time to experiment and verify. If you are not used to using the calorimeter, your instructor will show you one and how to record the temperature. Record all values for mass and temperature in table 1.

You need to boil water and put all 6 metal cubes in them. Therefore, choose an appropriate beaker and put enough water in it and raise it to a boil. Weigh each metal cube and note its mass. Tie each metal cube with a separate string, tie them to a stand, and immerse them in the boiling water for about 5 minutes. At this point, what do you expect will be the temperature of the metal cubes? Write this temperature.

While the metal cube is sitting inside boiling water, take a few minutes to discuss what may be needed to measure the amount of heat absorbed by the metal cube or simply to compare the heat energy absorbed by the metal cubes. ✎Write a procedure.⦸ See instructor for suggestions and approval before proceeding.

You have access to a calorimeter, balance, and thermometer. You may recall that a hot substance transfers its heat to a cold substance until they are equal in temperature and that cold water can be used to absorb heat from hot metal.

[**NOTE TO INSTRUCTOR:** If the students realize that they need to keep the same mass of water and put each of the metal cubes and see how much change in temperature is observed for each metal cube. The metal cube which contained the largest amount of heat will cause the largest temperature change in water. If they are unable to get this point, make sure you ask probing questions about heat transfer. Thus, their procedure is all about taking the same mass of water and measuring the temperature change caused by each metal cube.

If 2 sets of calorimeters are available for each group, the experiment will be finished faster. The most important thing is to keep the mass of water nearly the same for each experiment.

Comparing the change in temperature of the water will give the information on the relative amounts of heat absorbed by the metal. It will also have the highest value for the last column which is also the amount of heat released by the metal cube.]

Table-1: Record the data below and calculate the amount of heat absorbed by each metal cube. Does it agree with your prediction?

Block	Mass of metal cube	Boiling water temp	Mass of water in the calorimeter	Initial temp of water in the calorimeter	Final temp of water in the calorimeter	Change in temp of water (^{o}C)	Heat absorbed by the water (J)
Al							
Br (brass)							
Cu							
Fe							
Pb							
Zn							

[**NOTE TO INSTRUCTOR:** Typical data: Table 1

Block	Mass Of metal cube (g)	Boilin g water temp. (oC)	Mass of water in the calorimeter	Initial temp of water in the calorimeter	Final temp of water in the calorimeter	Change in temp of water (oC)	Heat absorbed by the water (J)
Al	22.28	100.1	50.1	20.5	27.5	7.0	1467
Zn	54.3	100.0	50.1	20.5	28.0	7.5	1572
Fe	62.64	100.0	50.1	20.5	30.0	9.5	1991
Br (brass)	67.05	100.0	50.1	20.5	29.5	9.0	1887
Cu	71.05	100.0	50.1	20.5	30.0	9.5	1991
Pb	87.98	100.0	50.1	20.5	24.5	4.0	838

(i) How does your data agree with your prediction? Did you see any surprises? If your answer didn't agree with your prediction, think about possible explanations.

Thinking about the data:

[**NOTE TO INSTRUCTOR:** There are 4 questions based on the data. Depending on the amount of time left for your lab, complete as many questions as possible during the class period. It is preferable to have them answer within their groups and discuss together as a class wherever possible. If any questions remain, ask students to complete them at home and also discuss with peer mentors, if available for your class.]

Q1. Use your data to figure out the specific heat capacity for each metal cube. Use table-2 to arrange the metals based on their atomic mass and compare their specific heat capacities. Use table-3 to describe the pattern observed.

Table -2

Block	Mass Of metal cube (g)	Initial temp of metal	Final temp of metal	Change in temp of metal (°C)	Heat released by the metal (J)	Specific Heat Capacity of metal (J/g•°C)
Al						
Br (brass)						
Cu						
Fe						
Pb						
Zn						

Table -3: Elements arranged by atomic mass

Element	Atomic Mass	Specific Heat Capacity (J/g•°C)

Al		
Iron		
Copper		
Zinc		
Lead		

[**NOTE TO INSTRUCTOR:** Typical data will look like this. Table -2

Block	Mass Of metal cube (g)	Initial temp of metal	Final temp of metal	Change in temp of metal (oC)	Heat released by the metal (J)	Specific heat capacity of the metal
Al	22.28	100.0	27.5	72.5	1467	0.908
Br (brass)	67.05	100.0	30.0	70.0	1887	0.399
Cu	71.05	100.0	30.0	70.0	1991	0.400
Fe	62.64	100.0	30.0	70.0	1991	0.454
Pb	87.98	100.0	24.5	75.5	838	0.126
Zn	54.3	100.0	28.0	72.0	1572	0.402

Table -3: Elements arranged by atomic mass

Element	Atomic Mass	Specific Heat Capacity
Al	26.98	0.908
Iron	55.85	0.454

Copper	63.55	0.400
Zinc	65.39	0.402
Lead	207.2	0.126

Pattern: Based on the atomic masses: Lead with the highest atomic mass has the lowest specific heat capacity and Aluminum with the lowest atomic mass has the highest specific heat capacity. This can be thought of the general trend. Students may not quickly realize the initial and final temperature of metal are obtained from the previous tables.]

Q2. Fitting the data. Dulong and Petit observed that the product of specific heat capacity and atomic mass is constant. Calculate the product of specific heat capacity and atomic mass for the five metals. Fill in table-4. Based on your data, what is the approximate value of this constant? The anticipated value is 3R where 'R' is the gas constant. How does your value compare to the anticipated value?

Table – 4: Product of atomic mass and specific heat capacity.

Block	Atomic mass	Specific Heat Capacity of metal (J/g•°C)	Molar Heat Capacity (J/mol•°C)
Al			
Cu			
Fe			
Pb			
Zn			

[**NOTE TO INSTRUCTOR:** Typical Data: Table 4: Product of atomic mass and specific heat capacity.

Block	Atomic mass (g/mol)	Specific heat capacity	Molar heat capacity, (J/mol•°C)

		of metal (J/g•°C)	
Al	26.98	0.908	24.5
Fe	55.85	0.454	25.4
Cu	63.55	0.400	25.4
Zn	65.39	0.402	26.3
Pb	207.2	0.126	26.2

The average is 25.6 which is closer to the anticipated value of 24.9. The error is about 3%.]

Q3: Using the molar heat capacity data obtained in your experiment, calculate the apparent molar mass of brass. Which pure element is closer to its molar mass? Based on the color of brass, which element you think is the predominant element in the alloy?

[**NOTE TO INSTRUCTOR:** Based on the typical data, the molar mass of brass will be 25.6/0.399 = 64.2 g/mol. If the standard value of 3R is used, it will give 24.9/0.399 = 62.4. There is going to be 3% error anyway because of the way we calculated. As long as they pick an element within this margin, it is fine. Possible answers are copper, zinc, Nickel, etc. Based on the color of brass, it is mainly copper.]

Q4: Your metal cubes are exactly 20 mm cubes. Calculate the density of brass, copper, and zinc from your data. Assume that the brass cube you are given is an alloy (a homogeneous mixture) of only copper and zinc. Find the percentage of copper and zinc in brass using density.

[**NOTE TO INSTRUCTOR:** typical data:

Density of copper: 71.05/8 = 8.88 g/cm^3
Density of zinc: 54.3/8 = 6.79 g/cm^3
Density of brass: 67.05/8 = 8.38 g/cm^3
8.34 = x(8.88) + (1-x)6.79. Solving this, we get, x = 0.74. That is 74% copper and 26% zinc]

FINAL LAB REPORT:

Now that you have established the relationship between the element and its ability to absorb heat and hold heat, answer the following questions in your lab report. Which metal is preferable to be used near the source of heat, that is at the bottom of the utensil? Why layers of metal may be preferable in the construction of utensils rather than using a single metal?

[**NOTE TO INSTRUCTOR:** Make sure they write a procedure which contains the following steps at the minimum. Weigh a dry and empty calorimeter with lid, put 50 or 100 mL water and weigh again to know the mass of water. Measure its temperature. Take the thermometer out, open the lid of the calorimeter, transfer one of the metal cubes by removing it from the stand using the string to handle it and dropping it into cold water in the calorimeter without splashing water. Stir the water for about a minute or two and measure the temperature. Make a note of the change in temperature.

Remove the metal cube and then empty water in the sink. There is no need to reweigh the same calorimeter because it may be slightly wet. Make sure outside of the calorimeter is dry. Add the same amount of fresh water chosen for the previous experiment. Weigh it to get the mass of water in the calorimeter. Record its initial temperature and final temperature after dropping another metal cube and stirring. The same procedure is repeated for each metal cube.]

PreLab – Exp #7: Which metal is a better choice for a cooking utensil?

Name: ________________________________

1. A hot metal at 125 ºC was dropped into 75.0 g of water kept at a room temperature of 20.6 ºC in a calorimeter. Within a couple of minutes, water's temperature increased to 26.8 ºC?
   (a) How much heat energy was absorbed by water? [‘c’ for water is 4.184 J/gºC]

   (b) How much heat energy was released by the metal?

   (c) What information is needed to calculate the specific heat capacity of the metal?

2. Refer to question 1. The metal was taken out of the water and left on the table. After about 5 minutes, the temperature of the water was measured to be 24 ºC. Will the temperature of metal be lower or higher than 24 ºC? Explain your prediction.

3. A block of titanium (‘c’ = 0.54 J/gºC) weighing 65 g and at 105 ºC was dropped into 75.0 g of water kept at a room temperature of 22.6 ºC in a calorimeter. Calculate the final temperature that will be reached by the metal and water?

Experiment #7: Which metal is a better choice for making a cooking utensil? - **10 points**

1. (a) 1946 J [2 pts]
   (b) 1946 J or -1946 J [2 pts]
   (c) mass of metal [1 pt]
2. Metal will be lower. [1 pt] See note below:

NOTE: Some students may think it could be hotter or cooler than water depending on the rate of loss as both mass and its heat conduction may be different.
As long as they make some logical arguments which show that they understood the question and tried to address it, they are good answers.

Explanation [1 pt]: Any reasonable explanation is okay. Some samples are shown below:

I. Metals lose heat faster because they are known to be good conductors of heat compared to water.
II. Metals have lower specific heat capacity compared to water.
III. As the data in the problem implies - metal went down in temperature much more compared to water's increase in temperature showing that metal will lose or gain heat much more than the water in the problem.

Water is still in calorimeter while metal has been taken out and hence will lose heat faster.

3. 30.8 ^{o}C [3 points. Make sure they show work. If work is not clear, reduce 1.5 pts.]

It is also important that they distinguish heat energy and temperature. When the same amount of heat energy is gained or lost, one with the lower specific heat capacity (& lower mass) will show more temperature change.

Quiz (Experiment #7: Which metal is a better choice for making a cooking utensil?)

10 points - 15 minutes **Name:______________________________**

Dulong and Petit Law: Atomic mass x specific heat capacity = 24.9 J/mol•ºC

Reference data

Element/Compound	's' (J/g•ºC)
Cr	0.461
Fe	0.461
Mn	0.477
Ni	0.502
Cu	0.377
Water	4.184

1. True/False: Heat energy spontaneous flows from a body at a high temperature to a body at a low temperature.
2. True/False: Heat will flow faster if the temperature difference is greater between two bodies.
3. True/False: An alloy will have a density that is a weighted average of the densities of metals that make up the alloy.
4. True/False: An alloy is a homogeneous mixture of metals.
5. At room temperature, 2.50 grams of copper, chromium, iron, nickel, manganese were placed on the same hot plate. The hot plate was turned on and the metals were continually heated and their temperatures were recorded. After about 5 minutes, one metal read 126 ºC while another metal read 99 ºC. All other metals had temperature readings in-between these 2 numbers.
   (a) Identify the metal that read 126 ºC.
   (b) True/False: In 5 minutes of heating, the metal nickel absorbed the highest amount of heat energy compared to other metals.
   (c) If you want the water to heat fast, which of the 5 metals will be the best choice?
6. An unknown hot metal weighing 50.0 g at 125 ºC is dropped in 100.0 g of cold water at 25.0 ºC. After a few minutes of stirring, metal and water reached the same temperature of 30.1 ºC. (a) Find the specific heat capacity of the unknown metal.
   (b) What is the atomic mass of the unknown metal?

Experiment #7:Which metal is a better choice for making a cooking utensil? - **10 points**

1. True [1 pt]
2. True [1 pt]
3. True [1 pt]
4. True [1 pt]
5. (a) Copper [1 pt]
   (b) False [1 pt]
   (c) copper [1 pt]
6. (a) 0.450 J/g^{o}C (b) 55.4 g/mol [2 pts for one correct answer and 3 points if both are correct. Note that (b) is 24.9 divided by whatever the answer obtained for (a). Thus, some students may have (a) wrong, but can get (b) correct even if they don't get 55.4.]