<table>
<thead>
<tr>
<th>Lab #</th>
<th>CHEMISTRY LAB - ACTIVITY TITLES</th>
<th>Page</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Lab Safety</td>
<td>5</td>
</tr>
<tr>
<td>2</td>
<td>Matter Classification</td>
<td>7</td>
</tr>
<tr>
<td>3</td>
<td>Classifying Physical and Chemical Changes</td>
<td>9</td>
</tr>
<tr>
<td>4</td>
<td>Chemical Properties of Four Liquids</td>
<td>11</td>
</tr>
<tr>
<td>5</td>
<td>Density of Pennies</td>
<td>13</td>
</tr>
<tr>
<td>6</td>
<td>Percent Cu in Pennies</td>
<td>15</td>
</tr>
<tr>
<td>7</td>
<td>Law of Definite Composition</td>
<td>17</td>
</tr>
<tr>
<td>8</td>
<td>Particle Size Probability</td>
<td>19</td>
</tr>
<tr>
<td>9</td>
<td>Emission Spectroscopy</td>
<td>21</td>
</tr>
<tr>
<td>10</td>
<td>Flame Test</td>
<td>23</td>
</tr>
<tr>
<td>11</td>
<td>Electron Probability – An Analogy</td>
<td>25</td>
</tr>
<tr>
<td>12</td>
<td>Mendeleev for a Day</td>
<td>27</td>
</tr>
<tr>
<td>13</td>
<td>Periodic Trends: Alkali Metals</td>
<td>29</td>
</tr>
<tr>
<td>14</td>
<td>Ionic and Molecular Compounds</td>
<td>31</td>
</tr>
<tr>
<td>15</td>
<td>Models of Covalent Compounds</td>
<td>33</td>
</tr>
<tr>
<td>16</td>
<td>Chemical Name and Formula Writing</td>
<td>35</td>
</tr>
<tr>
<td>17</td>
<td>Composition of Hydrates</td>
<td>37</td>
</tr>
<tr>
<td>18</td>
<td>Aluminum Foil</td>
<td>39</td>
</tr>
<tr>
<td>19</td>
<td>Empirical Formula</td>
<td>41</td>
</tr>
<tr>
<td>20</td>
<td>Single Replacement Reactions</td>
<td>43</td>
</tr>
<tr>
<td>21</td>
<td>Double Replacement Reactions</td>
<td>45</td>
</tr>
<tr>
<td>22</td>
<td>Four Solution Problem</td>
<td>47</td>
</tr>
<tr>
<td>23</td>
<td>Types of Reactions</td>
<td>49</td>
</tr>
<tr>
<td>24</td>
<td>Moles Fe and Cu</td>
<td>51</td>
</tr>
<tr>
<td>25</td>
<td>Mole-Mass Relationship</td>
<td>53</td>
</tr>
<tr>
<td>26</td>
<td>Changes in Physical State</td>
<td>55</td>
</tr>
<tr>
<td>27</td>
<td>Boyle's Law</td>
<td>57</td>
</tr>
<tr>
<td>28</td>
<td>Determination of Absolute Zero</td>
<td>59</td>
</tr>
<tr>
<td>29</td>
<td>Molar Volume of Gas</td>
<td>61</td>
</tr>
<tr>
<td>30</td>
<td>Paper Chromatography</td>
<td>63</td>
</tr>
<tr>
<td>31</td>
<td>Ice Cream Lab</td>
<td>65</td>
</tr>
<tr>
<td>32</td>
<td>Rate of a Reaction</td>
<td>67</td>
</tr>
<tr>
<td>33</td>
<td>Investigation of Chemical Equilibrium</td>
<td>69</td>
</tr>
<tr>
<td>34</td>
<td>Change in Enthalpy of a Reaction</td>
<td>71</td>
</tr>
<tr>
<td>35</td>
<td>pH and Indicators</td>
<td>73</td>
</tr>
<tr>
<td>36</td>
<td>Acid Base Titration</td>
<td>75</td>
</tr>
<tr>
<td>37</td>
<td>Understanding Half-life</td>
<td>77</td>
</tr>
<tr>
<td>38</td>
<td>Determining the Half-life of Ba-137m</td>
<td>79</td>
</tr>
<tr>
<td>39</td>
<td>Back to Chernobyl</td>
<td>81</td>
</tr>
</tbody>
</table>
1. LAB SAFETY

Welcome to Chemistry. The first order of business is to familiarize yourself with the lab room you will be working in. This safety activity will review rules you are familiar with and introduce you to new ones rules which are unique to Chemistry. At the culmination of this lab, you and your parent/guardian will read and sign a safety contact. You will be held responsible for performing and behaving in a safe manor throughout the year. You will be required to pass the safety quiz before performing any other labs.

PRE-LAB QUESTIONS

1. Set up your lab notebook as specified by your teacher.
2. What is the purpose of this lab activity?

OBJECTIVE

1. Familiarize yourself with safety equipment in the classroom.
2. Review safety rules.
3. Conduct safety orientation activities
4. Formulate safety rules

MATERIAL

7 stations with posted questions

PROCEDURES

1. Start at the station where you are seated. Stations can be done in any order.
2. Stay with your lab table group. Read the question on the card.
3. Perform the activity as stated and record information on the data table.
4. When group is ready, switch to the next station.

STATION QUESTIONS/ACTIVITIES

1. What is the effect of acid on clothing? GOGGLES AND GLOVES ARE REQUIRED
2. What is the effect on acid on skin or eyes? GOGGLES AND GLOVES ARE REQUIRED
3. How do you safely light a Bunsen burner? GOGGLES REQUIRED
4. Can all these materials be used to run an experiment?
5. Why is this not a safety station?
6. What is wrong with this lab station?
*7. Identify each piece of equipment. Give brief description of their use. Record.

CONCLUSION

1. Complete data table.
2. State at least one safety rule associated with each station (6 rules).
2. Classification of Matter

Introduction:
There are many ways to classify or group matter. Chemists have agreed upon a general system of classification. After you become familiar with the terms you will examine different samples and classify them.

DO NOT OPEN VIALS

PRELAB

1. What is the purpose of this activity
2. Define the vocabulary words on the chart provided.

Procedures:
1. Working in lab groups as assigned by your teacher.
2. Exam the sample in the vials. Do not open the bottles.
3. In the chart provided, describe some physical characteristics (color, shape, phase…), then indicate the state of matter as: S = solid, L = liquid, or G = gas.
4. Sort the substances into similar groups. (Example: all the elements ….).

Conclusion:
1. Fill in the conclusion chart, by listing the substance in one of five groups.
3. CLASSIFYING PHYSICAL AND CHEMICAL CHANGES

There are two types of changes that occur in nature: physical and chemical. In a physical change, one or more physical properties of a substance are altered but the composition remains the same. No new substances are formed. A chemical change results in the formation of one or more “new” substances. These new substances differ in chemical properties and composition forms the original substance. In this experiment you will observe a number of changes and will classify the change as physical or chemical based on your observations.

PRELAB
1. State the purpose of this activity in your own words.
2. List three examples of a physical change.
3. Give three observable indications that a chemical change has occurred.

GOGGLES MUST BE WORN DURING ALL PARTS OF THIS LAB.

PROCEDURE

Record your observations after each step.

1. Experiment 1
   a) Attach a small candle to a glass plate. Light the candle and move it out of the way and proceed with b). Observe the glass plate at the end of the lab.

   b) Obtain a test tube with wax. Heat water in a beaker on a hot plate. Immerse a test tube with wax in hot water. As soon as you see a change, remove the test tube to a rack and allow it to cool.

2. Experiment 2
   a) Place 2 sugar cubes in a mortar and grind the cubes. You will use a portion of this ground sugar for each of the following experiments.

   b) Place half in a small test tube and add enough water to dissolve. Shake the tube to dissolve.

   c) Place the remaining sugar in a deflagrating spoon and heat over a Bunsen burner. When cool wash contents down the sink.

3. Experiment 3
   a) Add a small scoop of baking soda, sodium bicarbonate, to a test tube and add water to height of about 1 centimeter. Stir.

   b) Add a few drops of silver nitrate solution to the baking soda solution.

   c) In another test tube place the same amount of baking soda and dissolve as above. Add hydrochloric acid. CAUTION: ACID CAN CAUSE SEVERE BURNS. WASH IMMEDIATELY IF ANY ACID COMES IN CONTACT WITH YOUR SKIN. Carefully touch the bottom of the test tube.
4. Experiment 4
Place a small scoop of copper sulfate in a test tube. Heat gently over a Bunsen burner. Refer to experiment 1b) for proper heating technique. Observe any material accumulating at the mouth of the test tube.

5. Experiment 5
a) Place about 25 mL of water in a 100 mL beaker. Add a level teaspoon (plastic spoon) of copper chloride to the beaker. Stir until the solid disappears. Measure the temperature.

b) When the solid has dissolved, place a piece of aluminum in the solution. Make sure the solution covers the aluminum. Continue to observe the mixture and measure the temperature after a minute and again after 5 minutes.

Demonstration if time permits.
6. Experiment 6
a) Mix iron and sulfur and add to a test tube. Run a magnet along the side of a test tube.

b) Heat the contents of the test tube in the fume hood. When cool run the magnet along the side of a test tube.

CONCLUSION

1. For each experiment indicate whether the change was chemical or physical. Briefly explain each. Summarize your findings in a chart.

2. Does the formation of bubbles always indicate a chemical change? Explain.

3. Does heating a substance always cause a chemical change? Explain using your data.

4. Does burning (combustion) always indicate a chemical change? Explain.
4. CHEMICAL PROPERTIES OF FOUR LIQUIDS

The first task of a chemist is to identify substances so that one can be distinguished from another. Sometimes this is simple. If, at room conditions, substance A is a green solid, substance B a red liquid, and substance C a colorless, odorless gas, one can tell them apart by simply looking at them. Color, odor, and physical state at room conditions are examples of physical properties. Physical properties are those that can be tested without changing the chemical formula of the substance. Chemical properties are revealed when, for example, a solid substance is placed in water and rapid bubbling occurs, or when an iron nail is placed in an unknown liquid and a reddish substance forms on it. When chemical properties are tested, substances with new chemical formulas are formed. Every substance has both physical and chemical properties.

In this experiment, you will collect data to determine if any of four liquids are the same or if all four are different. All of the liquids have similar physical properties. They are all clear and odorless at room temperature. You will compare their chemical properties. You will decide how to record your data and then reason logically to determine whether any liquids are the same or whether they are all different.

PRE LAB

1. State the purpose of the experiment in your own words.
2. List three chemical and five physical characteristic properties of pure substances.
3. How many characteristic properties of two substances must be alike for the two substances to be the same?
4. How many characteristic properties of two substances must be different for the two substances to be the different?
5. List three examples of properties that can be observed to distinguish between substances
6. What does the formation of bubbles in a liquid indicate?

OBJECTIVE

1. Observe visible changes.
2. Compare and contrast chemical properties of the four liquids.
3. Interpret experimental results and determine which liquids may be the same.
4. Reason logically to formulate correct conclusions.

MATERIALS

Apparatus
- 24 well micro plate
- Pipet with unknown
- Wash bottle with distilled water
- Forceps
- Laboratory apron
- Safety goggles
- Small straw scoop

Reagents
- blue litmus paper
- red litmus paper
- universal indicator
- aluminum pellets
- zinc pellets
- manganese (IV) oxide powder
- liquids A, B, C, & D
PROCEDURES

1. Put on your laboratory apron and safety goggles.
2. Place 10 drops of liquid A into each of the six top wells of the 24 well microspot plate. Place 10 drops of liquid B into each of the six wells of the next row. Continue with 10 drops of liquid C in the third row, and 10 drops of liquid D in the bottom row.

3. Tear a piece of red litmus paper into four pieces and one piece in each well of column 1 of the microplate. Record your observations on your data table.
4. Repeat step 3 using blue litmus paper placed into each well of column 2. Record your observations on your data table.
5. Be sure there are no flames in the laboratory before you add one drop of universal indicator to each well of column 3. Record your observations on your data table.
6. Add one piece of zinc to each well of column 4. Wait at least two minutes and then record your observations on your data table.
7. Add one piece of aluminum to each well of column 5. Wait at least two minutes and then record your observations on your data table.
8. Add a small amount of manganese (IV) oxide to each well of column 6. Wait at least two minutes and then record your observations on your data table.
9. Obtain a bottle of an unknown liquid and add three drops to each of the six top row wells of your micro plate.
10. Repeat Procedures 3 through 8 for your unknown.
11. Using forceps remove the pieces of litmus paper from the micro plate and place them on a paper towel. Rinse and dry the forceps then remove the zinc and aluminum and place them on the paper towel also. Since the litmus paper, zinc, and aluminum are quite harmless, throw the paper towel containing them into the trash. Empty the liquids remaining in the micro plate into the sink. Rinse the micro plate with tap water by gently pouring water over the wells then turning the micro plate over to drain it into the sink. Rinse and dry the forceps.
12. Before leaving the laboratory, clean up all other materials and wash your hands thoroughly.

CONCLUSION

1. Do any liquids share the same properties? If so, explain.
2. Could any two of the liquids be the same? Explain.
3. What other chemical and/or physical properties might be used to identify each of the liquids?
4. What is the identity of your unknown? Explain in detail how you identified it.
5. DENSITY OF PENNIES

Today's penny is quite different from the penny of a decade ago. Before 1982, pennies were made of an alloy of copper. Since then, they have been made with an outside coating of copper and an inner core of a different metal. The difference in composition of older and more recently minted pennies have resulted the penny’s characteristics; including it’s density, or mass per unit of volume. In this experiment, you will determine and compare the densities of pennies minted before 1982 and after 1983. You will use this data to try to identify the metal used in the core of pennies minted after 1983.

PRE LAB

1. What is the purpose of this lab activity?
2. Define density.
3. What is the reason for measuring five pennies at a time rather than an individual penny?
4. How is the volume of the pennies determined?
5. Why is it important to work with dry pennies?
6. Look up the densities for the following metals: ALUMINUM, COPPER, SILVER, ZINC

OBJECTIVE

1. Make accurate measurements using a balance and graduated cylinder
2. Collect data and graph
3. Identify unknown
4. Perform density and other calculations

MATERIAL

<table>
<thead>
<tr>
<th>Pre-83 Pennies</th>
<th>Post-83 Pennies</th>
</tr>
</thead>
<tbody>
<tr>
<td>Vial</td>
<td>Vial</td>
</tr>
<tr>
<td>25 pennies</td>
<td>25 pennies</td>
</tr>
<tr>
<td>100 mL graduated cylinder</td>
<td>Balance</td>
</tr>
<tr>
<td>paper towel</td>
<td></td>
</tr>
</tbody>
</table>

PROCEDURE

1. Work with one set of pennies at a time; either pre-1982 pennies or post-1983 pennies. Find the mass of 5 pennies from one set. Record the mass in the appropriate Data table.
2. Add 5 more pennies to the first group and obtain the mass of these 10 pennies. Record the mass.
3. Repeat step 2, each time adding more pennies to those already on the Balance, until you have used all 25 pennies
4. Fill a 100mL graduated cylinder to the 20-mL mark with water. Be sure to use the bottom of the meniscus to measure the water level.
5. Still working with the same set of 25 pennies; gently drop 5 of the pennies into the graduated cylinder. Record the new water level in the appropriate table.
6. Add 5 more pennies to the graduated cylinder, making a total of 10 pennies. Record the water level in the table.
7. Add 5 more pennies to the cylinder and record the water level.
8. Repeat step 7 until you have added all 25 pennies to the cylinder. Record the volume after each addition.
9. Discard the water. Dry the pennies with a paper towel and either pass them to another group to use or give them to your teacher.
10. Repeat step 1-9 using the 25 pennies in the other set of coin. Record your data in the other table.
11. Complete your data tables. Find the net volume of each group of pennies by subtracting 20 mL from the total volume recorded for each group (column 3). Enter the net volume for each group in column 4 of the data tables.

**GRAPH ANALYSIS**

1. Construct a graph of your results. Let the y-axis reflect the mass of the pennies. Plot the data for the pre-1982 pennies first. Then draw the best-fitting straight line (the straight line that connects as many point. a. possible).
2. On the same graph, plot the data for the pennies minted after 1983. Draw the best-fitting straight line.

**CALCULATIONS**

Find the slope of each line:

**CONCLUSION**

1. What do the values you obtained for the slopes of the lines represent?
2. Compare the density of copper obtained for the pre-1982 pennies.
3. Calculate the percent error for the pre-1982 pennies.
4. What is the density of the post-1983 pennies?
5. Compare this value to the density of the metals listed in the pre lab.
6. Which metal could be inside the post-1983 pennies?

**Useful Formulas**

<table>
<thead>
<tr>
<th>Density = Mass (g)</th>
<th>Slope = ( \frac{Y_2 - Y_1}{X_2 - X_1} )</th>
</tr>
</thead>
<tbody>
<tr>
<td>Volume (mL)</td>
<td></td>
</tr>
<tr>
<td>% Error = ( \frac{\text{Theoretical} - \text{Experimental}}{\text{Theoretical}} \times 100 )</td>
<td></td>
</tr>
</tbody>
</table>
6. PERCENT COPPER IN A PENNY

You have determined that pennies made after 1983 contain other metals electroplated with copper to appear the same as pre-1982 pennies. If you take a metal file and score the surface of a post-1983 penny, you will note the silvery color of metal beneath the surface.

In this activity you will remove the other metal by reacting the penny with hydrochloric acid, leaving only copper shell. By carefully massing before and after the reaction with acid you will be able to determine the percent of copper in post-1983 pennies.

PRELAB
1. State the purpose of the lab in your own words.

GOGGLES MUST BE WORN IN THIS EXPERIMENT

MATERIALS

One post-1983 penny  
tongs
balance  
small beaker
hydrochloric acid (HCl)  
triangular file
acetone
watch glass

PROCEDURE

1. Mass the penny.
2. Make four heavy scratches at 45° angles on the edges of the penny using the file as demonstrated.
3. Label the beaker with your name, so you can identify it tomorrow.
4. Place the penny in the beaker and carefully pour about 25 mL of hydrochloric acid into the beaker. Cover with a watch glass.

CAUTION: THIS ACID IS CONCENTRATED AND VERY CORROSIVE. AVOID CONTACT WITH SKIN AND INHALING THE FUMES!

5. Observe the reaction and record your observations. Place your beaker on the tray provided. This reaction will continue overnight.

NEXT DAY

6. Observe your beaker and some of the other beakers. Record your observations.
7. With your goggles on carefully add about 25 mL of water to the beaker to dilute any unreacted acid. Pour off (decant) as much liquid as you can, leaving the penny in the beaker. Wash the penny in the beaker with water several times.
8. Do a final rinse it with acetone.
9. Remove the penny carefully from the beaker and set it on the paper towel. Let stand for several minutes.
10. Mass the penny when dry.
CONCLUSION

1. Calculate the percent of copper in the penny. Show set up.
2. Some of the pennies were floating the next day. What was the cause of this?
3. What can you conclude about the effect of hydrochloric acid on copper?
Elements are a kind of matter that cannot be decomposed by ordinary chemical means. Compounds are chemical combinations of elements. What is a chemical combination as compared to just any combination of elements? It would help to give an example of what is not a compound. The combination of oxygen and nitrogen in the air is a mixture of the gases. This mixture can be in all proportions of oxygen and nitrogen. The properties of the mixture depend upon the properties of the individual components of the mixture. Compounds, on the other hand, are special combinations of the elements. A compound has its own properties, distinct from the properties of the elements from which it was made.

The Law of Definite Composition states that the elements forming a compound always combine in the same proportion by mass. The compound water, H$_2$O, is always a chemical combination of hydrogen and oxygen in a 1:8 ratio by mass. If a mixture of hydrogen and oxygen were reacted in some other mass ratio, for example, 1:2, water would be formed but some hydrogen would remain unreacted. Water forms only in the 1:8 ratio by mass.

In this experiment you will examine the reaction between magnesium metal, Mg, and oxygen gas, O$_2$. When heated, magnesium reacts readily with oxygen in the air. The magnesium will be heated strongly in an open crucible for several minutes. You will measure the mass of the magnesium that reacts and the mass of magnesium oxide that is formed, and use your data to calculate the mass of the oxygen that combined. You can then find the ratio of mass of magnesium to mass of oxygen and compare your experimental ratio with the ratios of your classmates and with the actual ratio. By calculating the percent error you can then evaluate how accurately you performed the experiment and look for sources of error.

**PRE LAB QUESTIONS**

1. State the purpose of the experiment in your own words.
2. Why is it important to start the experiment with a clean and dry crucible?
3. What is the purpose of making sure the surface of the magnesium ribbon is clean and shiny?
4. With what element or elements does the magnesium combine when it is heated in the crucible?
5. In the procedure you are asked to reheat the crucible repeatedly until the last two masses agree to within 0.03 gram. What is the purpose of this reheating?
6. Suppose a compound of sodium and chlorine is formed in a ratio of 1.54 grams of chlorine for each 1.00 gram of sodium. How much sodium would you need to completely react with 45.0 grams of chlorine?

**OBJECTIVE**

1. Observe a reaction between magnesium and oxygen.
2. Calculate a ratio of mass of magnesium to mass of oxygen.
3. Measure masses carefully to obtain accurate results.

**MATERIALS**

<table>
<thead>
<tr>
<th><strong>Apparatus</strong></th>
<th><strong>Reagents</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>crucible &amp; lid</td>
<td>laboratory apron and safety goggles</td>
</tr>
<tr>
<td>crucible tongs</td>
<td>pipe stem triangle</td>
</tr>
<tr>
<td>dropper pipette</td>
<td>electronic balance</td>
</tr>
<tr>
<td>ring stand &amp; ring</td>
<td>ceramic tile</td>
</tr>
<tr>
<td>Bunsen burner &amp; striker</td>
<td></td>
</tr>
</tbody>
</table>
PROCEDURE

1. Put on your laboratory apron and safety goggles.
2. Obtain a piece of magnesium ribbon from your teacher. If the surface of the ribbon is not shiny, use a piece of sandpaper or steel wool to shine the surface.
3. Obtain a clean, dry crucible and cover. Find the combined mass of the crucible and cover and record it on the Report Page.
4. Break the magnesium ribbon into small pieces and place them into the crucible. Find and record the combined mass of the crucible, cover, and magnesium.
5. Set up the ring stand, ring, Bunsen burner, and pipe stem triangle as shown in Figure A. Place the crucible on the pipe stem triangle and begin heating the crucible gently with the lid completely on. Heat slowly by moving the flame around underneath the crucible. Remove the heat temporarily if a large amount of smoke comes out of the crucible.
6. After about four minutes of direct heating with no smoke, remove the lid slightly and heat the crucible to redness for four more minutes. Finally, remove the lid completely and heat strongly for four more minutes. (Be careful where you place the lid – it is very hot!) Total of 12 minutes of heating.
7. Turn off the Bunsen burner and put the lid back on the crucible. Allow the crucible and cover to cool to a temperature low enough so that you can touch it comfortably. Minimum of 7 minutes wait time. Find the combined mass of the crucible, cover, and contents and record it on your Report Page.
8. Repeat heating of the same sample for about 5 minutes. This is the second heating. Record mass
9. Compare the masses found in steps 7 and 8. If the masses do not agree to within 0.030 gram, reheat the crucible for four minutes, cool, and find the mass again. This is the third heating. Repeat the process until you get two consecutive readings that agree to within 0.030 gram. This is known as “heating to constant mass”. This step may not be necessary.
10. Before leaving the laboratory, clean up all other materials and wash your hands thoroughly.

CONCLUSION

1. How would your results be affected if some of your magnesium did not react?
2. Use your textbook to help you determine the formula for the magnesium oxide that formed in this experiment.
3. Use the accepted ratio to determine the mass of magnesium that would combine with 16.0 grams of oxygen.
PARTICLE SIZE FROM COLLISION PROBABILITIES

Most of the information we have about the structure of the atom was derived by indirect means because the technology was not available to make direct measurements. In this activity, we are going to assume that for some reason we are not able to measure the diameter of a marble directly. We will use an indirect method to determine the size of the diameter by counting how often we succeed in hitting a marble when we shoot another marble at a group of them. The experimental arrangement is shown in the figure below.

The probability of a “hit” is related to a number of variables: the size of the marbles, the distance between the walls of the enclosure, the number of trials, and the number of target marbles. An equation can be derived using these variables. You will collect the necessary data and use the equation to determine the diameter of a marble. After determining the diameter by this method, you can check its success by direct measurement of the diameter of the marbles and find out how well the indirect method worked.

PRE LAB

1. State the purpose of this lab activity
2. What is probability?
3. How are the target marbles to be lined up?
4. Why will 3 meter sticks be needed?
5. What is the equation for percent error?

OBJECTIVE

1. collect data
2. use probability equation to indirectly determine marble diameter
3. directly measure marble diameter
4. calculate percent error

MATERIALS

11 marbles  
tacky material  
3 meter sticks
PROCEDURE

1. Obtain ten target marbles and one marble with which to bombard them. Arrange the target marbles inside a three sided enclosure made of meter sticks. The target marbles may be placed in any random position EXCEPT that there must always be room for the bombarding marble to get through between any two target marbles and no marble should be shielded from being hit. That is, each target marble must be a potential target. The target marbles will probably have to be rearranged after each hit to achieve this. See figure above.

2. Roll the bombarding marble toward the target marbles. Attempt to roll the bombarding marble along a path parallel to the walls at all times. Also, try to roll the bombarding marble randomly, that is DO NOT AIM. It may help to close your eyes.

3. Continue rolling the bombarding marble toward the target marbles, counting both the total number of rolls and the number of times a roll results in a hit. Remember that if the bombarding marble collides with any number of target marbles before hitting the wall, this is ONE hit. You must have at least 200 rolls for good results. More than 500 do not increase the accuracy of the results. Record the number of marbles rolled and the number of hits on the Data Table. One partner should roll and the other record. You may switch rolls half way.

4. To get a direct measurement, line up your ten target marbles so that they touch each other in a straight line between two meter sticks and measure the length in centimeters to 2 decimal places.

CALCULATIONS

1. Calculate the diameter using the equation:  \[ D = \frac{H \times d}{2N \times Tr} \]
   where:
   \( Tr \) = number of trials
   \( N \) = number of target marbles
   \( H \) = number of hits
   \( d \) = distance between the walls
   \( D \) = diameter

2. Calculate the diameter of one marble using the direct measurement.

3. Calculate the percent error between the calculated and the measured diameter, assuming the measured diameter is the true value.

CONCLUSION

1. What are some sources of error?

2. How did Rutherford describe the atom as a result of his experiment?

3. In what ways does the situation in this experiment (which is a model of Rutherford’s experiment) differ from Rutherford’s experiment in which alpha particles were used to bombard a foil composed of gold atoms? (Hint: consider the charges on the subatomic particles.)
9. EMISSION SPECTROSCOPY

All atoms give off electromagnetic radiation if their gases or ions are energized by heating or by high voltage electric discharge. If the light emitted by a gas is passed through a spectroscope, a pattern of narrow lines of light is produced. Each element produces its own distinct pattern that differs from the pattern of every other element. The particular pattern of frequencies of light emitted by an atom is referred to as its emission spectrum or bright-line spectra. The emission spectrum of an element can be used as a means of identification, just as fingerprints (or DNA) can be used to identify a human being.

The unique pattern of frequencies of light emitted by an atom corresponds to the set energy given off as electrons drop from higher allowed energy states or energy levels to lower energy levels. The energy of each transition is given by Max Planck’s equation, \( E = h \nu \), where \( E \) = energy in Joules, \( \nu \) = frequency, and \( h \) = a constant (Planck’s constant). An electron can be raised from its lowest allowed energy level (ground state) to other higher allowed energy levels by absorbing certain set amounts of energy. This “excited” electron cannot remain at any of these higher allowed energy levels if there are unoccupied lower energy levels closer to the nucleus. The electron is attracted back to one of the lower allowed energy levels. As the electron drops back it emits a photon, a set amount of energy, in the form of electromagnetic radiation. The amount of energy that the electron emits is equal to the energy difference between the higher energy level of the excited state and the lower energy level to which the electron dropped.

PRELAB

1. State the purpose of this experiment.
2. According the modern theory of the atom, where may an atom’s electrons be found?
3. How do electrons become “excited”?
4. What form of energy emission accompanies the return of excited electrons to ground state?
5. State two equations that are used to relate the energy content of a packet of light and its wavelength.
6. A line spectrum is sometimes called a “fingerprint” of an element. What do you think is meant by this term?

OBJECTIVE

1. Use a spectroscope
2. Observe continuous spectra
3. Observe bright line spectra from gaseous elements.

WARNING: DO NOT TOUCH POWER SUPPLIES OR TUBES.
MATERIALS

Spectroscope Tube or spectral glasses  incandescent light source  fluorescent light source
High-voltage power supplies containing spectrum tubes of:
  argon, helium, hydrogen, mercury, and neon

PROCEDURE

1. Examine spectrum of incandescent and fluorescent light sources around the room.  
   Record observations of the continuous spectra in the field provided.
2. Aim spectroscopes at the spectrum tubes set up around the room.  Record your observations
   for each element.  (Draw the specific spectral lines)

CONCLUSION

1. Compare the spectra produced by incandescent and fluorescent sources.
2. How are these spectra different from those observed in the discharge tube?
3. What causes the differences in the bright-line spectra of different elements?
4. Which element showed more lines in the red region?  The blue-violet region?
5. Prior to its discovery on Earth, the existence of helium was first confirmed in the sun.
   Explain how this can be possible.
6. Calculate the energy of the red line of the hydrogen spectrum.
10. FLAME TESTS

When elements are heated to high temperatures they may enter an “excited state”, which occurs when electrons absorb energy and move to higher energy levels. When the electrons return to “ground state”, (the original state), they release the amount of energy that had been absorbed. This energy can be seen in the form of visible light. Each element gives off a unique set of wavelengths, which our eyes recognize as a particular color.

In this lab you will observe the colors given off by pure solutions of metal ions, and one mixture. You will then use this property of atoms to identify an unknown metal ion by comparing the color observed with the colors of known metal ions.

**GOGGLES MUST BE KEPT ON DURING THE ENTIRE EXPERIMENT.**

**PRE LAB**

1. What is the purpose of this experiment?
2. What is meant by ground state?
3. How do electrons become “excited” in this lab?
4. When is the energy absorbed by electrons released?
5. What is the form of this energy?
6. How should the burner flame be adjusted for best results?
7. Why is this lab done without room or outdoor lighting?

**OBJECTIVE**

1. observe the colors given off by a group of metal ions when heated in a flame.
2. identify an unknown metal by the color in a flame test.

**MATERIALS**

<table>
<thead>
<tr>
<th>Goggles</th>
<th>splints saturated with metal ion solutions</th>
<th>forceps or tongs</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bunsen burner</td>
<td>cobalt glass</td>
<td>unknown sample</td>
</tr>
</tbody>
</table>

**PROCEDURE**

1. You will be working in pairs to make observations of known metals. Each person will have his/her unknown sample to identify. **NO TWO STUDENTS IN THE CLASS SHOULD HAVE THE SAME UNKNOWN SAMPLE.**

2. One person from each pair will obtain wooden splints that have been soaked in solutions of each metal ion to be tested. Take two splints with sodium, two with potassium and two with the mixture. Each person will then be given an unknown sample. Immediately record your own unknown number.

3. Light the Bunsen burner. Be sure to have a good blue flame with the inner cone clearly visible. Place the end of the splint that has been soaking in the solution in the flame of the burner. Record the color of the flame.
4. Test all samples including the mixture of sodium and potassium and your unknown.

5. * Repeat the flame test with samples of potassium, sodium, and the mixture, this time, viewing the flame through the blue cobalt glass.

*Optional

CONCLUSION

1. What metal ion was in your unknown sample?
2. Explain how you identified your unknown using your data. Be specific.
3. Give two reasons why the flame test is not always valid to identify a metal.
4. What would be observed if a spectroscope were used during a flame test?
5. What color was the flame for the mixture of sodium and potassium? Explain.
6. Is it possible to use the flame tests to identify each individual metal in any mixture? Explain.

*7. What is the purpose of the cobalt glass for viewing the flame?

*optional
11. ELECTRON PROBABILITY – AN ANALOGY

The Bohr model states that the electron travels in a specific orbit around the nucleus in a Planetary Model and therefore its position is space and its momentum can be readily determined. However, Werner Heisenberg in his Wave Mechanical Model showed that this is impossible since both position and momentum cannot be known at the same time. This is the famed Heisenberg Uncertainty Principle. We must therefore speak in terms of probable location of the electron in the atom.

The region of space in which an electron can be found 95% of the time is called the space orbital. The most probable distance of the electron from the nucleus is called the radius of maximum probability.

OBJECTIVE
1. use a model to determine the probable location of an electron in a space orbital.
2. calculate the region of space in which an electron can be found within the space orbital.

MATERIALS
Bulls eye target  masking tape
Ball point pen  cushioning material

PROCEDURES
1. You will be working individually.
2. Place paper as cushioning under the target, and secure to floor using masking tape.
3. Stand above target, in such a way so that your hand is 2 meters above the bull’s eye.
4. Aim for the bull’s eye, and drop the pen…a total of 100 times.
PRE LAB QUESTIONS

1. Why must the target be cushioned?
2. Describe Bohr’s model of the atom. (Key feature of the electron position).
3. What is your prediction about the outcome of the activity? In what region will the electron most likely be located?
4. What is the purpose of this lab activity?

CALCULATIONS

1. Location the 95 dots closest to the bulls-eye. Draw smooth circle around the pen dots to represent the space orbitals.
2. For each circle preprinted on the target, count the pen dots, and record.
3. Graph the number of dots in a region vs. average radius. Draw a smooth curve.
4. Divide the total number dots in each circular region, by the area of that region according to the values given on the data table. (This is the number of dots per square centimeter). Record these values on the data table.
5. Create another graph; plot the dots/cm² value vs. average radius.

CONCLUSION QUESTIONS

1. When you through a dart at a target, can you predict precisely where it will strike?
2. The s (spherical) orbital is non directional. Does your orbital appear to be this way? Explain.
3. Interpreting the graph: what is the area of maximum probability (of electron location).
4. Make certain to attach your target to your lab notebook before submission.
12. MENDELEEV FOR A DAY

As more and more elements were discovered during the 1800’s it became important to categorize them according to similarities in chemical and physical properties. A Russian chemist, Dmitri Mendeleev (1834-1907) was the first to do so successfully. He arranged the elements in vertical columns in order of increasing atomic mass. The columns were then arranged so that elements with similar chemical properties were placed side by side.

Mendeleev left numerous spaces in his table because there were no known elements with the appropriate properties to fill the spaces. He then predicted the properties of these unknown elements based on the properties of the elements next to them on his periodic table. When these “missing elements” were discovered, they were found to have properties very similar to those he predicted.

As a model of Mendeleev’s work, in this investigation you will group unknown compounds (instead of elements) according to their chemical behaviors. You will see if a precipitate forms when two solutions are mixed together. You then will observe which precipitates are dissolved by another solution. Finally you will record any color changes that occur when two solutions are mixed together. After analyzing your data, you will place each unknown solution into a group with other solutions of similar characteristics.

OBJECTIVE

To model Mendeleev’s ideas
To group nine unknown compounds based on their chemical behaviors.

PRELAB

1. How does Mendeleev’s periodic table differ from the modern periodic table?
2. Why did Mendeleev leave blank spaces on his periodic table?
3. What is the reason (unknown to Mendeleev) that elements in the same group have similar properties?
4. Why is it necessary to rinse the stirring rod between test tubes?

MATERIALS:

3 test tubes labeled A, B, C
10-mL graduate cylinder
Stirring rod
Unknown solutions 1-9
4 dropper pipets containing test solutions A, B, C & D

PROCEDURE:

GOGGLES AND APRONS ARE NEEDED FOR THIS LAB. CAUTION – SOLUTION D IS VERY CORROSIVE. WASH YOUR HANDS WITH SOAP AND WATER IMMEDIATELY IF ANY SOLUTIONS COME IN CONTACT WITH YOUR SKIN.

1. Each group will work with one of the unknown solutions. These unknown solutions are identified by numbers 1-9. The data will then be shared with the class at the end of the experiment.
2. Put on goggles and aprons.
3. Label three test tubes A, B, C.
4. Obtain one unknown solution. Note the number. Place 1 mL of the unknown solution in each of the three labeled test tubes.

5. To test tube labeled A, add 12 drops of test solution A.
   To test tube labeled B, add 12 drops of test solution B.
   To test tube labeled C, add 12 drops of test solution C.

6. Any test tube that does not show an immediate change should be stirred for about 10 seconds. Make sure you rinse the stirring rod after each use to avoid contaminating the next solution. Record the observations on the Data Table. Be sure to record in the space for your unknown sample.

7. To any test tube that contains a precipitate, add 20 drops of test solution D and stir for at least 15 seconds. Again rinse the stirring rod after each use. Record the results.

8. Dispose of all solutions in the waste container provided, then rinse out the test tubes and graduated cylinder in the sink.

9. Clean up and wash your hands. Leave materials out on the bench.

**CONCLUSION QUESTIONS**

1. Why do you think this investigation is titled “Mendeleev for a Day”?

2. Based on your observations, group the unknown solutions into “Families” according to similarities in chemical behavior. Arrange your groups in a chart.

3. Justify your arrangements by referring to your data.
4. Some members of the same “family” that show similar but not identical reactions. Give an example of this.

5. No members of Group 8A (18) of the modern periodic table can be found on Mendeleev’s version of the periodic table. Suggest an explanation for their absence.
13. PERIODIC TABLE I:
A STUDY OF REACTIVITY OF METALS

OBJECTIVE
1. Compare reactivity of metals in two Groups and in a Period by observing reactions.
2. Identify an unknown Group II metal in a compound.
3. Determine reactivity trends in families and in periods.

PRE LAB
1. What is the purpose of this lab?
   Read the textbook on reactivity of metals and the procedure to answer the following:
2. What is the trend in activity in a Group as the atomic number increases?
3. What is the trend across a Period as atomic number increases?

GOGGLES AND APRONS MUST BE WORN FOR THESE EXPERIMENTS.

PART I

MATERIALS
test tubes, test tube rack, calcium piece, magnesium ribbon, aluminum pellet, phenolphthalein solution, splint, match, forceps, steel wool, Bunsen burner, test tube holder

PROCEDURE
1. The reactivity of Alkali metals will be observed by demonstration.
2. Pour about 5 mL of distilled water into a clean, dry test tube and place it in the test tube rack. Using the forceps, add a piece of calcium to the water. DO NOT TOUCH THE CALCIUM WITH YOUR HANDS. IT IS CORROSIVE TO SKIN. Observe the reaction. Collect the gas being released by holding an inverted dry test tube over the reactant tube.
3. Test for hydrogen gas by inserting a burning splint into the mouth of the upper test tube.
4. Add a drop of phenolphthalein solution to the reactant tube. Record your observations.
5. Clean the surface of the magnesium ribbon with steel wool. Break it into small pieces and place the pieces in a test tube. Repeat the above procedure.
6. If there is no visible reaction, gently heat the water in the test tube just to boiling. BE SURE TO POINT THE TUBE AWAY FROM YOURSELF AND OTHERS WHILE HEATING.
7. Once it is boiling, turn off the burner, return the test tube to the rack and collect the gas. After a few seconds, test for hydrogen as above. Add phenolphthalein. Note any color change in the test tube.
8. Repeat the procedure with one aluminum pellet.
PART II
The more active a metal, the more reactions will be observed. Formation of a precipitate will indicate that a reaction has occurred.

SALTS OF BARIUM AND STRONTIUM ARE EXTREMELY TOXIC. AVOID CONTACT WITH THESE CHEMICALS AND WASH YOUR HANDS THOROUGHLY AFTER USE.

MATERIALS
spot plate, dropper pipets containing solutions of Mg(NO$_3$)$_2$, Ca(NO$_3$)$_2$, Sr(NO$_3$)$_2$, Ba(NO$_3$)$_2$, H$_2$SO$_4$, Na$_2$CO$_3$, Na$_2$CrO$_4$.

PROCEDURE
1. To each well of a spot plate place about 10 drops of one of the following solutions containing alkaline earth metals: Mg(NO$_3$)$_2$, Ca(NO$_3$)$_2$, Sr(NO$_3$)$_2$, Ba(NO$_3$)$_2$. To each solution add two drops of H$_2$SO$_4$. Record whether a precipitate appears, the relative amount and color. If no precipitate is formed leave the box blank.
2. Repeat the procedure but add a few drops Na$_2$CO$_3$(aq) to each solution.
3. Repeat the procedure again but add Na$_2$CrO$_4$(aq) to each solution.
4. Obtain an unknown solution. The solution will contain ions of one of the alkaline earth elements: Mg$^{2+}$, Ca$^{2+}$, Sr$^{2+}$, or Ba$^{2+}$. Test with the solutions used above (steps 2 – 3) in order to identify the unknown ion.

CONCLUSION
1. Did this experiment verify the predicted reactivity trend of Group 1A and 2A elements? Explain using your observations.
2. Predict what you would have observed if the following elements had been tested: Beryllium, Strontium, Francium. Explain.
3. You observed three elements in period 3 in this lab. Do your observations support the predicted reactivity trend? Describe.
4. If zinc and iron had been tested, how would their reactivity compare to the metals in the same PERIOD that you observed in this lab?
5. Describe any relationship that you can determine between the number of precipitates formed by each compound and the location of the alkaline earth metal on the periodic table.
6. What is the identity of your unknown compound? Support your conclusion by referring to your data.
14. IONIC AND MOLECULAR COMPOUNDS: COMPARING PROPERTIES

Compounds can be classified by the types of bonds that hold their atoms together. Ions are held together by ionic bonds in ionic compounds; atoms are held together by covalent bonds in molecular compounds. You cannot tell whether a compound is ionic or molecular simply by looking at a sample because both types can look similar, when they are solids. In this activity some simple tests will be done to classify compounds by type because each type has a set of characteristic properties shared by most members. Ionic compounds are usually hard, water-soluble, have high melting points and can conduct electricity when dissolved in water. Molecular compounds can be soft or hard, are usually less water-soluble, tend to have lower melting points, and cannot conduct electricity when dissolved in water. In addition, ionic compounds are always solids at room temperature, while molecular compounds can be solids, liquids or gases.

OBJECTIVE
1. observe some characteristics of six different solid substances
2. classify the solids as ionic or molecular

MATERIAL

<table>
<thead>
<tr>
<th>Item</th>
<th>Quantity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hot plate</td>
<td></td>
</tr>
<tr>
<td>glass slides or watch glasses</td>
<td></td>
</tr>
<tr>
<td>7 samples in vials</td>
<td></td>
</tr>
<tr>
<td>Spot plate</td>
<td></td>
</tr>
<tr>
<td>conduction apparatus with battery</td>
<td></td>
</tr>
</tbody>
</table>

PROCEDURE
1. Obtain seven solid samples vials.
2. Using a clean scoopula or wood splint for each sample, and place several grains on a glass surface to observe melting. Keep on hot surface no more than 3 minutes. (VAPOR HAZARD) Observe which samples melt.
3. Place several grains of each sample into their own spot plate well. Add de-ionized water. Observe which samples dissolve.
4. Using the solutions you made in step 3, and place the conduction apparatus electrodes in each well. Clean apparatus electrodes with de-ionized water before each testing. Observe which mixtures cause the light to glow or flash the fastest. Compare results with de-ionized or distilled water, used as a control.
PRE LAB
1. What is the purpose of this lab?
2. Describe some characteristics of ionic compounds.
3. Describe some characteristics of molecular compound.

CONCLUSION
1. Complete the statements:
   a) The following substances ______are ionic because ______
   b) The following substances _____are molecular because _____

2. If you know the formula of the compound can you predict whether it is ionic or molecular?
   Explain.

3. Predict the following, based on the patterns established in this experiment:
   a) Solubility of sodium iodide in water(high/low)
   b) Relative melting point of benzoic acid (C_6H_5COOH) (high/low)
   c) Electrical conductivity of glucose in water
   d) Electrical conductivity of magnesium chloride in water
15. MODELS OF COVALENT COMPOUNDS

Why should people care about the shapes of molecules? Consider that the properties of molecules depend not only on their molecular composition and structure, but on their shape as well. Molecular shape determines a compound’s boiling point, freezing point and the nature of its reactions.

The geometry (shape) of a small molecule can be predicted by examining the central atom and identifying the number of atoms bonded to it and the number of unshared electron pairs surrounding it. The names of the shapes you will be able to identify are: linear, bent, trigonal pyramidal, tetrahedral, and trigonal planar.

Covalent bonds are formed when electrons are shared by two atoms. If the two atoms are alike, that is have similar electronegativity values, the bond is said to be nonpolar covalent. If the atoms are not alike, that is there is a large difference between their electronegativity values, the bond is polar covalent. More than one pair of electrons can be shared. This results in a double or triple bonds.

Molecules can also be polar or nonpolar. If all the covalent bonds are nonpolar, the molecule will be nonpolar. If the covalent bonds are polar, the molecule can still be nonpolar if the polar bonds are symmetrically arranged around a central atom, so that their charges cancel each other out. On the other hand, if the arrangement of the polar bonds is asymmetrical, the electrons will be attracted more to one end of the molecule and a polar molecule or dipole will result. Polarity of the molecule depends on the bonds and the shape.

A model in chemistry is a way of thinking about the structure of matter. In this laboratory, you will build molecular models to show the three-dimensional nature of matter, identify the shape, and associate the electron dot diagram with the shapes.

To give additional illustrations, this lab will also introduce you to organic compounds. These substances contain carbon chains. (examples: C₂H₂, C₂H₄, C₂H₆)

PRELAB
1. State the purpose of this lab activity
2. What properties of compounds are determined by the shape of the molecule?
3. What shapes can molecules have?
4. How can you determine if BOND is polar or nonpolar?
5. Predict the bond type (ionic, polar covalent, nonpolar covalent) for the following. Show the electronegativities differences for each pair.
   a. Na and Cl
   b. C and H
   c. S and O
   d. N and N
6. What factors determine if a MOLECULE is polar or nonpolar?
OBJECTIVE
1. *build* molecular models to show the three-dimensional nature of matter
2. *observe* symmetrical and asymmetrical molecules
3. *identify* molecular shapes, and associate them to electron dot diagrams
4. *examine* several organic compound structures
5. *determine* covalent bonding type, and molecule type

MATERIALS
Molecular models kits  model color coded key

PROCEDURE
1. Assemble models for all formulas in the data table.
2. Complete the Data Table.

CONCLUSION
1. Which shapes always produce polar molecules?

2. List the NONPOLAR MOLECULES which have POLAR BONDS.

3. Both water and carbon dioxide are molecules composed of three atoms and 2 bonding clouds. One is polar and the other is nonpolar. Explain why.

4. Why is CH₃Cl polar, while CCl₄ is not?

5. The polarity of a substance can have a great effect on its solubility. A rule of thumb for solubility is “like dissolves like”. Knowing this general rule, what can you predict about the polarity of alcohol if you know that alcohol dissolves in water?

6. Classify each of the compounds using one of the following: (Refer to the electronegativity chart to determine bond type.)
   - *ionic bonding*
   - *polar covalent bonds in polar molecules*
   - *polar covalent bonds in nonpolar molecules*
   - *nonpolar covalent bonds in nonpolar molecules.*

   a) I₂  
   b) CBr₄  
   c) H₂S  
   d) NaF
16. CHEMICAL NAMES AND FORMULAS

Chemical substances are described not only by unique names but also by unique chemical formulas. We use this language to communicate about chemistry. In this lab activity you will combine certain ions, and produce new compounds. This type of formation occurs when a cation is able to combine with an anion. The formation of these new compounds will be visible to the unaided eye. These specific compounds will then be named and their formulas written. Some of these ionic compounds will need roman numerals in their names. You will need to refer to your textbook, reference charts and class notes for additional assistance.

SAFETY: goggles and aprons must be worn at all times

PRE LAB

1. What is the purpose of this lab activity?
2. What is a precipitate?
3. When is a roman numeral used in a name?
4. Which cations require the use of roman numerals?
5. What is a cation? List the cations in this experiment.
6. What is an anion? List the anions in this lab activity.

OBJECTIVE

1. write formulas for ionic compounds
2. write names for ionic compounds
3. observe precipitation of these compounds
4. identify cations and anions

MATERIALS

9 solutions in dropper pipettes
ionized water
spot plate or plastic reaction surface

PROCEDURE

PART I:
1. Obtain a set of vials for each lab table. In this activity you may work in 3-4 person groups. Make observations and write names and/or formulas on data table # 2.
2. After your group is done you may rotate to another table.

PART II:
1. Put on goggles and apron.
2. Clean spot plate or plastic reaction surface thoroughly, and rinse with ionized water.
3. Collect solutions from indicated area.
4. Add three drops of each indicated solutions into the indicated spot.
5. Record observations in the numbered data table #1 boxes using the following symbols:
   ppt = precipitate formation (cloudy or grainy appearance)
6. Follow instructions for safe disposal of solutions from your spot plate or reaction surface.
7. Clean up your lab area thoroughly and return all materials and equipment to their original location.
8. Remove and put away your goggles and aprons
CONCLUSION

1. In each box in **table # 2**, write the name and formula of the precipitate formed. 
   **Reaction #1 has been done for you.**
   **Reminder:** Some names will require Roman Numerals.

2. Write the formulas for the following compounds: lead (II) chloride, lead (II) chlorate, zinc phosphate, zinc phosphide

3. Using your answers in question #2, explain how you distinguish between the binary ionic compounds and ionic compounds with polyatomic ions
Hydrates are ionic compounds (salts) that have a definite amount of water (water of hydration) as part of their structure. The water is chemically combined with the salt in a definite ratio. Ratios vary in different hydrates, but are specific for any given hydrate. The formula of a hydrate is represented in a special manner. The hydrate of Cobalt (II) chloride has the formula CoCl$_2 \cdot x$H$_2$O. The unit formula for the salt appears first, and the water formula is last. The dot means that water is loosely bonded to the salt. The coefficient $x$ stands for the number of molecules of water bonded to one unit of salt. This special formula illustrates the law of definite composition.

When hydrates are heated, the "water of hydration" is released as vapor. The remaining solid is known as the anhydrous salt. The general reaction for heating a hydrate is

$$\text{hydrate} \rightarrow \text{anhydrous salt} + \text{water}$$

The percent of water in a hydrate can be found experimentally by accurately determining the mass of the hydrate and the mass of the anhydrous salt. The difference in mass is due to the water lost by the hydrate. The percentage of water in the original hydrate can easily be calculated.

1. **What is the purpose of this lab activity?**
2. **Calculate the theoretical percent of water in the following hydrates.**

   (HINT: One of these will be your unknown!)
   
   MgSO$_4 \cdot 7$ H$_2$O
   CuSO$_4 \cdot 5$ H$_2$O
   BaCl$_2 \cdot 2$ H$_2$O
   Na$_2$SO$_4 \cdot 10$ H$_2$O
   CaCl$_2 \cdot 2$ H$_2$O

### OBJECTIVE

1. **Determine** the % of water in an unknown hydrate
2. **Identify** an unknown hydrate sample from a given list of hydrates

### MATERIALS

- Evaporating dish
- Scoopula
- Balance, ring stand
- Crucible tongs or mitten
- Iron ring
- Wire gauze
- Burner
- Vial containing unknown hydrate sample
PROCEDURE

1. **YOU ARE WORKING INDIVIDUALLY.** Set on ring stand, ring with wire gauze. Wash and dry an evaporating dish then heat it in the hottest part of the flame for about one minute. Use tongs/or mitten to handle dish. First allow dish it to cool thoroughly. Mass the evaporating dish.

2. Obtain a sample of unknown hydrate. _Be sure to record the number on the vial._
   Mass about 2 grams of the hydrate into an evaporating dish. Record the masses accurately. Do not waste time trying to get _exactly_ 2 grams. Break up any lumps.

3. Adjust Bunsen burned so there is a double blue cone. Set ring on stand, about 8 -10 cm from the barrel.

4. Heat dish gently by moving the burner flame back and forth along the bottom, for 5 minutes. Increase the heat gradually and set burner down on stand. _Avoid any popping and splattering._

5. Heat strongly for an additional 5 minutes. If the edges of the hydrate start to turn brown, remove the heat momentarily and resume heating at a gentler rate.

6. Remove the dish and allow it to cool for at least 5 minutes. Find the mass as soon as it is cool. _Massing dish while warm will cause experimental error._

7. Repeat heating (steps 4 and 5 to obtain the second heating value. Cool and mass to make sure all the water has been removed. If the mass changes by 0.1g or more, heat a third time. Repeat until there is essentially no change in the mass. You can then assume that all the water has been driven off.

CONCLUSION

1. Write the number and complete formula of your unknown hydrate?
   Why did you select it from the list? Use your calculations to support your choice.

2. Why do you think it is necessary to measure the mass of the anhydrous salt immediately after cooling?

3. A hydrate compound has a mass of 1.632g before heating and 1.008g after heating. Compute the experimental percentage of water in the hydrate.
18. ALUMINUM FOIL

This activity will be done cooperatively. All members of the group must be involved and one grade will be given to the group. (If however, I observe that a member of the group is not participating, that person will receive a lower grade.)

OBJECTIVE:
To determine the thickness of a piece of aluminum foil to three significant digits.
To determine the thickness in atoms.
To determine the number of atoms of aluminum in the piece of foil.
To answer the question stated below (at the end.)

SOME DIRECTIONS:
1. Write the names of the group members.
2. You may use any equipment that is presently in the room.
3. Decide on a procedure and write it.
4. Prepare a data table.
5. Show calculations clearly.

SOME THINGS THAT YOU ARE NOT PERMITTED TO DO:
1. You may not destroy, crumble, tear or fold your piece of foil. It must be returned in exactly the same condition that you received it.
2. You can only confer with members of your group. You are not permitted to spy or eavesdrop on other groups.
3. You cannot remove any papers from this lab. All materials must be turned in by the end of the period.

SOME INFORMATION THAT MAY BE HELPFUL:
1. The density of aluminum is 2.69 g/cm$^3$.
2. The diameter of one atom is $2.50 \times 10^{-8}$ cm. (Assume the atom is a sphere.)

WHAT IS TO BE SUBMITTED BY THE END OF THE PERIOD
1. The names of the group members.
2. The aluminum foil
3. The procedure and data table.
4. The calculations with the results of the objectives clearly marked.
5. The answer to the following questions with calculations shown:

If the population of the world is about $5.6 \times 10^9$ people, how many atoms of aluminum from your foil could you distribute to each person in the world?
19. **EMPIRICAL FORMULAS**

You will be given an unknown binary compound to analyze. The general formula of each compound will be AxBy. Heating any of the compounds will cause them to decompose according to the following equation:

\[ \text{AxBy} \rightarrow xA(\text{s}) + yB(\text{g}) \]

Since “B” is released as a gas, strong heating will leave only the "A" of the unknown in the evaporating dish.

**GOGGLES AND APRON MUST BE WORN**

**PRE LAB**

Show all work, and circle final answer:

1. What is the purpose of this lab?
2. A 0.750 g sample of tin is oxidized with nitric acid to form tin oxide. Calculate the empirical formula of tin oxide if the original tin sample gained 0.201 g of oxygen.
3. Excess sulfur reacts with 0.565 g of cobalt to give 1.027 g of cobalt sulfide. Find the empirical formula of the product.
4. If 1.164 g of iron filings reacts with chlorine gas to give 3.384 g of iron chloride, what is the empirical formula of the compound?
5. Why should you wait at least 5 minutes before massing the evaporating dish?

**Precaution:** THE SOLIDS ARE SAFE IF HANDLED PROPERLY. IF ANY SOLID COMES IN CONTACT WITH YOUR SKIN, WASH WITH SOAP AND WATER.

**Cleanup:** WASH RESIDUE IN THE EVAPORATING DISH DOWN THE DRAIN WITH PLENTY OF WATER. RETURN VIAL.

**OBJECTIVE**

1. *analyze* a binary compound
2. *determine* formula of the unknown

**MATERIAL**

- evaporating dish
- wire screen
- balance
- tongs or mitten
- Bunsen burner
- iron ring with stand
- striker

**PROCEDURE**

1. You will be working individually. Obtain a sample and record the unknown number.
2. Clean an evaporating dish with water. Dry the dish by placing the evaporating dish on a wire screen and iron ring on a ring stand over a laboratory burner and heat for several minutes.
3. Carefully remove the dish with tongs and allow it to cool on a tile until you no longer can feel heat radiating from it (min. 5 minutes).
4. When the dish is cool (min. 5 minutes), measure the mass of the dish to the nearest 0.01 g. Record the mass. **HOT DISHES SHOULD NOT BE PLACED ON THE BALANCE.**
5. Add about 3 grams of the unknown to the evaporating dish and measure the mass to the nearest 0.01 g. Record mass.
6. Begin heating slowly. (The flame should be a double blue cone, but keep farther from the stand). Increase the heat until you have heated the dish strongly for about 10 minutes. You are driving off all of the "B" in your unknown in this step.

7. Remove the dish form the ring stand; let it cool on the tile. Measure the mass of the dish and contents (now just the "A" in your unknown). Record the mass.

8. Reheat with a hot flame for a couple of minutes, cool on the tile and measure the mass again. If the mass has changed (in the 0.) place) heat again and mass until there is no change. These steps confirm that all the "B" has been driven off.

CONCLUSION

1. What is the number and empirical formula of your compound?
2. What is an empirical formula?
3. What additional information would you need to calculate the molecular formula?

THE ATOMIC MASSES OF "A" AND “B" CAN DETERMINE FROM THIS TABLE

<table>
<thead>
<tr>
<th>Unknown Number</th>
<th>Atomic Mass of &quot;A&quot;</th>
<th>Atomic Mass of &quot;B&quot;</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>106</td>
<td>20.7</td>
</tr>
<tr>
<td>2</td>
<td>35.3</td>
<td>62.0</td>
</tr>
<tr>
<td>3</td>
<td>89.6</td>
<td>35.0</td>
</tr>
<tr>
<td>4</td>
<td>46.6</td>
<td>27.3</td>
</tr>
<tr>
<td>5</td>
<td>68.9</td>
<td>13.5</td>
</tr>
<tr>
<td>6</td>
<td>30.5</td>
<td>71.3</td>
</tr>
<tr>
<td>7</td>
<td>122</td>
<td>35.7</td>
</tr>
<tr>
<td>8</td>
<td>46.1</td>
<td>53.9</td>
</tr>
<tr>
<td>9</td>
<td>23.0</td>
<td>40.3</td>
</tr>
<tr>
<td>10</td>
<td>92.2</td>
<td>27.0</td>
</tr>
<tr>
<td>11</td>
<td>47.7</td>
<td>41.9</td>
</tr>
<tr>
<td>12</td>
<td>78.5</td>
<td>11.5</td>
</tr>
<tr>
<td>13</td>
<td>34.1</td>
<td>39.9</td>
</tr>
<tr>
<td>14</td>
<td>63.3</td>
<td>24.8</td>
</tr>
<tr>
<td>15</td>
<td>212</td>
<td>41.4</td>
</tr>
<tr>
<td>16</td>
<td>68.2</td>
<td>79.8</td>
</tr>
<tr>
<td>17</td>
<td>70.6</td>
<td>124</td>
</tr>
<tr>
<td>18</td>
<td>61.0</td>
<td>142.6</td>
</tr>
<tr>
<td>19</td>
<td>46.0</td>
<td>80.6</td>
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<tr>
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<td>92.2</td>
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<td>106</td>
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<tr>
<td>22</td>
<td>35.3</td>
<td>62.0</td>
</tr>
<tr>
<td>23</td>
<td>89.6</td>
<td>35.0</td>
</tr>
<tr>
<td>24</td>
<td>46.6</td>
<td>27.3</td>
</tr>
</tbody>
</table>
In a single replacement reaction one element replaces another element in a compound as indicated by the general equations:

\[ A + BX \rightarrow AX + B \quad \text{or} \quad Y + BX \rightarrow BY + X \]

In the first, A replaces B and B is liberated. In the second, Y replaces X and X is liberated. In single replacement reactions, an element reacts with a compound, displacing an element from that compound. An example of the first type of single replacement is the reaction of zinc with hydrochloric acid in which zinc replaces hydrogen, liberating hydrogen gas:

\[ \text{Zn} + 2 \text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2. \]

An example of the second case is the replacement of bromine in potassium bromide by chlorine:

\[ \text{Cl}_2 + 2 \text{KBr} \rightarrow 2 \text{KCl} + \text{Br}_2. \]

Most single replacement reactions occur in aqueous solutions. Metals can be organized according to the ease with which they undergo certain chemical reactions (generally single replacement reactions). These Tables are called activity series, which list elements in order of their ability to replace other elements. The most active element, which is placed at the top of the series, can replace every element below it.

In this experiment you will determine the order of activity of four metals: iron, copper, zinc and magnesium by studying the single replacement reactions that you perform.

**PRE LAB**

1. What is the purpose of this lab?
2. Why must the spot plate be cleaned after each trial?
3. During a single replacement reaction what happens to a metal that is least reactive?

**OBJECTIVE**

1. *perform* single replacement reactions
2. *study* single replacement reactions
3. *interpret* observations
4. *determine* the order of activity of four metals

**MATERIALS**

- Spot plate
- 6 solutions in dropper pipets
- Dionized water
- 4 types of metals in vials (Pb, Cu, Zn, Mg)
PROCEDURE

1. Clean spot plate wells. Do a final rinse with dionized water.
2. Pour 5 drops of Cu(NO$_3$)$_2$ solution into well 1, 5 drops of Pb(NO$_3$)$_2$ solution into well 2, 10 drops of Zn(NO$_3$)$_2$ solution into well 3, 5 drops of AgNO$_3$ solution into well 4, 5 drops of HCl into well 5, and 5 drops of Mg(NO$_3$)$_2$ into well 6.
3. To each well add a clean piece of Pb.
4. Note and record the results.
5. Clean spot plate. **Dispose of metal into solid waste can.**
6. Repeat steps 1 – 4, three more times. (Using Cu, Zn, and Mg instead of Pb).

CONCLUSION

1. Write balanced equations for each reaction that occurred.
2. Based on your observations, list the order of decreasing activity. (Most active first).
   Justify your answer.
21. DOUBLE REPLACEMENT REACTIONS

Precipitation occurs in solution when two chemicals join together to form a product that is insoluble in water and falls out of solution like rain or snow. A precipitate is a solid substance that separates from solution during a chemical reaction, and can be identified by the cloudy, milky, gelatinous or grainy appearance it gives to the mixture. In a test tube the, the solid might settle to the bottom of the container. The reactions are usually double replacement reactions.

In this laboratory activity you will mix solutions of lead, silver, and calcium compounds with other compounds in solution. You will observe and describe the precipitates that are formed. You will then write and balance complete chemical equations to describe precipitation reactions.

GOGGLES MUST BE WORN. WASH YOUR HANDS AFTER CLEANING UP.

PRE LAB

1. State the purpose of the experiment in your own words
2. What are cations? Give two examples.
3. What are anions? Give two examples
4. Where are anions and cations found on the periodic table?
5. What is a solution? What is a precipitate?

OBJECTIVE

1. observe precipitation formation
2. describe precipitation formation
3. write balanced double replacement reactions
4. identify the precipitate in each reaction that occurs

MATERIALS

11 solutions in pipettes
plastic reaction surface
dionized water

PROCEDURE

1. Place lab table in acetate sheet. Make certain it is clean by rinsing with dionized water. Dry with paper towel.
2. Place one drop of each solution as shown in lab table. (Combine solutions as shown in the table).
3. Record observations on data table. Write “ppt” if a precipitate forms.
   Do not let the tip of the pipette touch the drop to avoid contamination and false reactions.
3. Wash chemicals down the sink. Flush with water and wash the plastic.
CONCLUSION

1. How do you know a double replacement reaction occurs?
2. What two CATIONS commonly form precipitates?
3. Which of the mixtures give no visible reaction?
4. Write balanced chemical equations for all the precipitation reactions of AgNO₃ that you observed. Be sure to combine ions in formulas according to their charges. Mark the precipitate with an (s).
5. Write balanced chemical equations for all the precipitation reactions of CaCl₂ that you observed.
22. FOUR SOLUTION PROBLEM

You are provided with 8 pipets containing aqueous solutions. Four are identified and labeled AgNO₃, FeCl₃, KSCN and NaOH. The other four pipets contain the same solutions but labeled only A, B, C and D. Note: SCN⁻ is a polyatomic ion called the “thiocyanide” ion. You are expected to identify what solution each pipet contains. You will use your observations of double replacement reactions as a tool to help you solve this problem.

OBJECTIVE

Identify the contents of pipets A, B, C and D.

MATERIALS

- AgNO₃, FeCl₃, KSCN and NaOH in labeled pipets.
- AgNO₃, FeCl₃, KSCN and NaOH in pipets labeled A, B, C and D.
- Plastic Sheet Protector (reaction surface).
- Distilled water bottle.
- Cup (to hold pipets).

SAFETY PRECAUTIONS

- Wear safety goggles and aprons.
- Wash hands after the lab.
- AgNO₃ will stain skin and clothing.

PROCEDURE

1. React each of the “known solution” with each other. Record your observations for each reaction in Data Table 1. If a precipitate formed, use the solubility chart to predict which compound was the precipitate and its color and record the formula and color of the precipitate in the Table 1. Use colored pencils to indicate the color.
2. Once you have reacted all known reactants, react each unknown with each known. Record your observations in Data Table 2. Observe the colors of the precipitates that formed in order to determine the identity of the unknown solutions. Use colored pencils to indicate the color.
CONCLUSION

1. Write the chemical and net ionic equation for each reaction in which a precipitate was formed from Table 1.

2. What is the identity of the unknown pipets?
23. TYPES OF CHEMICAL REACTIONS

There are five basic reaction types that we can observe in this course: Single replacement, double replacement, decomposition, synthesis, combustion (complete or incomplete). In this lab activity you will carefully perform each of the reactions, making careful observations. Based on class discussion, other lab activities, demonstrations, your textbook and class notes, you will determine the reaction types for parts A – F and write balance equations.

Goggles and Aprons must be worn throughout this lab.

PRE LAB

1. What is the purpose of this lab?
2. Write balanced chemical equations to describe the following chemical reactions:
   a) Aluminum reacts with the oxygen in the air to form aluminum oxide.
   b) Hydrogen peroxide, H$_2$O$_2$, decomposes into water and oxygen gas.
   c) Hydrochloric acid reacts with magnesium to produce hydrogen gas and magnesium chloride.
   d) Hydrochloric acid also reacts with sodium hydroxide to produce table salt and water.
3. Now, identify each type of reaction.

OBJECTIVE

1. observe different reaction types  
2. write word equations for each reaction  
3. identify reaction types  
4. write balance equations for each reaction.

MATERIALS

right-angle glass bend     large test tubes (3)     splints
burner, large beaker       small beaker            forceps
small test tubes (2)       solid stoppers        test tube rack
ring stand                 one-hole stopper       rubber tubing
test tube holder

PROCEDURE

REACTION A
1. Place the CuCO$_3$ in a small test tube to a height of about 1 cm.
2. Note the appearance of the sample.
3. Using the test tube holder heat the tube strongly for about 3 minutes. Extinguish the flame and then hold a burning splint into the test tube. Observe the result and note any changes in the appearance of the residue in the test tube.

REACTION B
1. Obtain a piece of magnesium ribbon. Polish with steel wool.
2. Using crucible tongs ignite the magnesium in the Bunsen burner and hold it over a watch glass.

DO NOT LOOK DIRECTLY AT THE BURNING MAGNESIUM!
3. Examine the material in the watch glass.
4. Wash the contents of the watch glass down the sink.
REACTION C
1. In the small test tube mix a few drops of AgNO₃ and HCl. Record the results.
2. Dispose of liquids in the hood

REACTION D (See fig. 1)
1. Place 10 ml of hydrochloric acid in one test tube. Insert the one- hole stopper with the glass tube into the test tube.
2. Clamp on a ring stand.
3. Fill a test tube with tap water, and invert the tube in a large beaker of tap water as shown in figure 2. Keep the mouth under the water.
4. Run the rubber tubing to the collection test tube so that the free end of the test tube is just inside the mouth of the tube.
5. Obtain some pieces of mossy zinc. Open the test tube containing the HCl and carefully drop in the zinc. Record your observation.
6. Allow the collection tube to fill with gas. Collect another if necessary.
7. Stopper the tubes of gas with solid stoppers and remove from the water. Keep them inverted.

fig. 1
Zinc + HCl

REACTION E
1. Dry the tubes on the out side. Remove the stopper from one of the test tubes and carefully bring a burning splint to the mouth of the gas-filled tube. Notice any change in tube.
2. Record observation.
3. Carefully dispose of the acid in the beaker labeled “acid” in the hood.

BE SURE TO KEEP GOGGLES ON DURING CLEANUP.

CONCLUSION
1. Study the observations you recorded and then identify the type of reaction that took place in each case (A-E).
3. Write balanced equations for each reaction.
24. A CHEMICAL REACTION: MOLES OF IRON AND COPPER

PRE LAB

1. What is the purpose of this lab activity?
2. What are the two possible formulas for iron chloride?
3. Write the two balanced equations that could occur in this lab when iron reacts copper II chloride. (In one assume FeCl₂ is formed and in the other FeCl₃).
4. What does it mean to decant in the context of the lab activity?
5. Why must the product of this reaction be completely dry before massing?

OBJECTIVE

1. determine the number of moles of copper produced and the number of moles of iron used in the reaction.
2. calculate the ratio of moles of iron to moles of copper.
3. determine the formula and the name of the iron chloride produced.

MATERIALS

- 100 Ml beaker
- 250 Ml beaker
- graduated cylinder
- stirring rod
- wash bottle w/de-ionized water
- steel wool
- forceps
- balance
- copper(II) chloride
- drying oven or heat lamp
- iron nail
- 1.0M HCl solution

PROCEDURE

1. Weigh the 100 Ml beaker.
2. Add about 2.50 g of copper chloride to the beaker. Add 15 Ml of de-ionized water and stir until all the copper chloride dissolves.
3. Use the steel wool to make the nail shiny (if not new) and weigh the nail. Record. Place the nail in the beaker containing the solution. Leave it for about 20 minutes.
4. Use tongs to pick up the nail. Any copper that is adhering to the nail should be added to the beaker by scraping and washing it off with water in the wash bottle. Set the nail aside to dry on a paper towel.
5. Carefully decant* the liquid from the copper. Pour the liquid into the large beaker so if some the copper is poured, off you can retrieve it. Discard the liquid in the sink.
6. Rinse the copper again with 15 Ml de-ionized water. Decant and repeat this step three more times.
7. Wash the copper with 10 Ml 1.0M HCl, decant and wash one more time with the water.
8. Label the beaker and place under drying lamp or in drying oven. Dry.
* See diagram on the next page.

It may be necessary to wait until the next day.

10. Weigh the beaker with the dry copper. Record. At this point the beaker with the Cu must dry completely.
11. Clean out the copper into the waste basket. Rinse the beaker and return to the side of the room.

CALCULATIONS: Show all set-ups on the table. Make certain to use units.

1. Calculate the mass of the iron that reacted. (nail before reaction – nail after reaction)
2. Calculate the moles of iron that reacted.
3. Calculate the mass of copper that formed. (beaker with Cu – beaker)
4. Calculate the moles of copper that formed.
5. How many moles of Fe reacted and how many moles of Cu were formed in your lab? Is the mole ratio of iron to copper closer to 1:1 or 2:3?
6. Use the ratio of Fe to Cu you calculated to determine which equation best describes the reaction that occurred in this lab. (Refer to prelab question #3.)

CONCLUSION

1. What type of reaction is this?
2. Can a reaction occur with copper nails and iron chloride? Explain.
In a balanced chemical equation, all reactants and products must be represented by symbols or formulas. The total number of atoms of each element must be the same on each side of the equation to satisfy the law of conservation of mass.

A calculation of the formula mass of a reactant or product enables a researcher to convert from grams of a particular substance, to moles of that substance. The mole relationship, given by coefficients of the balanced equations, then allows the researcher to calculate how many moles of every substance will take part in the reaction.

In this experiment, you will investigate the qualitative relationships in the reaction:

\[ \text{NaHCO}_3(s) + \text{HCl(aq)} \leftrightarrow \text{NaCl(aq)} + \text{CO}_2(g) + \text{H}_2\text{O(g)} \]

A known mass of sodium hydrogen carbonate will be reacted with excess hydrochloric acid. Knowing the mass of NaHCO\textsubscript{3}(s) that reacts, you can determine from the balanced equation the mass of NaCl that should be produced. You can compare this theoretical value with the actual experimental mass of NaCl produced.

**PRELAB**

1. What is the purpose of this lab?
2. What information is given by the coefficients in chemical reactions?
3. What is a ratio? Give an everyday example of a ratio.
4. How are ratios related to chemical reactions?
5. How can you show that mass is conserved in a chemical reaction?

**OBJECTIVE**

1. understand mole-mass relationships in chemical reactions.
2. demonstrate the difference between the experimental mass of the product of a chemical reaction and the theoretical mass predicted for that product by calculation.

**EQUIPMENT**

Balance  
burner  
spatula  
evaporating dish  
Ring stand  
iron ring  
wire gauze  
watch glass  
safety goggles/apron

**MATERIALS**

6M hydrochloric acid (HCl)  
sodium hydrogen carbonate (NaHCO\textsubscript{3})

**SAFETY**

GOGGLES AND APRONS MUST BE WORN. 
BE CAUTIOUS IN HANDLING HOT GLASSWARE.

Handle hydrochloric acid with care. Flush any spills with cold water and dilute solution of sodium bicarbonate and report to your teacher immediately. Never lean over an apparatus, or set up apparatus near lab top edge.
PROCEDURE

1. Flame dry a clean evaporating dish by heating it in the hot part of a burner flame for about 5 minutes. Allow the dish to cool.

2. Find the combined mass of the evaporating dish plus a watch glass. This is mass (a) in your list of data.

3. Leaving the watch glass and evaporating dish on the balance, add 2.50 g sodium hydrogen carbonate (NaHCO₃) to the evaporating dish until the scale balances. Record this mass as (b) in your list of data.

4. Set up the ring stand, ring and wire gauze as shown in Fig.16.1. Place the watch glass on top of the evaporating dish and place the dish on the wire gauze.

5. Obtain about 5 ml of 6M hydrochloric acid (HCl). CAUTION: Handle this acid carefully. It can cause painful burns if it touches your skin. Slowly add HCl to the NaHCO₃ in the evaporating dish, a few drops at a time (see Fig.16.2). Continue adding acid until the reaction (bubbling) stops. Carefully tilt the evaporating dish back and forth a couple of times to make sure that the acid has contacted all the NaHCO₃. After making sure that all bubbling has stopped, remove the watch glass and place it curved side up on the lab bench.

6. Holding the burner in your hand, gently heat the evaporating dish. Use a low flame and move the burner back and forth to avoid spattering. When almost all the liquid is gone remove the burner and replace the watch glass on the evaporating dish, leaving the small opening for vapor to escape. Heat gently again until no liquid remains. Allow the dish to cool.

7. Find the combined mass of the watch glass, evaporating dish and contents (NaCl). Record this mass, I, in your list of data.

CONCLUSION

1. According to the balanced equation for the reaction used in this experiment, what is the ratio of moles of NaHCO₃ to moles of NaCl?

2. Using your data, calculate, the moles of NaHCO₃ you used and the moles of NaCl that formed in this experiment? What is the mole ratio? Is it the same as the ratio, obtained from the balanced equation?

3. What is the % yield? How do the theoretical yield and the actual yield compare? What might be the source of error if yield is less than 100%? …if it is greater than 100%?

4. If the masses of all but one of the substances that take part in a chemical reaction are known, explain, why is it possible to determine the unknown mass by subtraction?
26. CHANGES IN PHYSICAL STATE

A pure substance can exist in three physical states: solid, liquid, and vapor (gaseous). Changes in physical state occur at constant, discrete temperatures that are characteristic of the substance. Changes in physical state include solids melting, liquids freezing and boiling, and gases condensing. In this experiment you will examine what happens when a pure substance, an organic compound called Lauric acid or CH₃(CH₂)₁₀COOH, undergoes a change in state. This substance is a waxy solid at room temperature.

In Part I, the substance will be melted at a temperature above its normal freezing point. As it cools at a constant rate, temperature readings will be made at regular intervals until the substance changes back to its solid phase at a temperature below its freezing point.

In Part II the procedure will be reversed. Starting with the solid, heat will be added and temperature readings taken until the substance returns to the liquid phase.

PRECAUTIONS: GOGGLES AND APRONS MUST BE WORN.
AVOID INHALING IRRITATING FUMES, THEREFORE DO NOT ALLOW THE LAURIC ACID TO BE HEATED ABOVE 60°C.
DO NOT DISPOSE OS LAURIC ACID IN SINK

PRE LAB

1. What is the purpose of the lab?
2. List the four changes a substance can undergo and state whether each is endothermic or exothermic.
3. Which phase changes will be observed in this lab activity?
4. What do you predict will be the relationship between the freezing point and melting point?
5. In what phase will the Lauric acid be at the start of the experiment?.
6. How will you determine the freezing point and melting point from your data?

OBJECTIVE

1. Observe melting and freezing of an organic substance.
2. Plot two sets of data
2. Determine freezing and melting of the substance by interpretation of graph data

MATERIALS

Hot plate
beaker for hot water bath
ring stand with clamps
test tube containing Lauric acid
cool water bath
gloves (mittens)
2 thermometers
stop watch
test tube rack
PROCEDURE

PART I:  COOLING CURVE  ** {Procedure 1 – 4 may be done by your instructor.}

One partner will be responsible for reading the temperature and the other for recording and timing.

1. Place water in a beaker. Begin heating the water on a hot plate.
   Keep one thermometer for taking temperature readings of the water bath ONLY.

2. Remove the stopper from the test tube and place it in the water bath. Make sure there is enough water to reach the level of Lauric acid in the test tube. Keep test tube from touching bottom by clamping. **LAURIC ACID WILL MELT. DO NOT ALLOW ANY WATER TO GET INTO THE TEST TUBE, AS IT WILL CONTAMINATE THE LAURIC ACID.**

3. When water bath reaches about 55°C, remove the melted Lauric Acid test tube and place it in the test tube rack. Insert a dry thermometer into test tube and swirl gently to get an even temperature reading.

4. Begin recording temperatures every 30 seconds. Leave the thermometer in it. You may gently stir with the thermometer. Continue taking temperature reading every 30 seconds and record until the temperature has fallen to 40°C. **Leave the thermometer in the acid at all times.**

5. Take test tube and place in cool water bath set at 30°C, and then add one or two ice cubes to the cool water. **The thermometer in now stuck in the acid! Do not remove. Go to part II.** (or return test tube with thermometer to the front desk)

PART II:  HEATING CURVE (optional).

1. Place cool water bath (at approx. 30°C) with the frozen Lauric acid t.t. on hot plate.
2. Begin recording the temperature of the Lauric acid every 30 seconds. Gradually warm until Lauric acid has liquefied.
3. **Do not heat Lauric acid beyond 60°C.**
4. Remove the thermometer and clean with paper towel while Lauric acid is still a liquid. Replace the stopper on the t.t. Return everything to original locations

CONCLUSION

1. Construct graphs with time on the x-axis and temperature on the y-axis. Use your data to construct two lines. Use the same set of axes for both graphs. Draw a smooth curve through the points.
2. Does the temperature of the substance vary while it is freezing or melting?
3. Using your cooling curve, determine the freezing point of lauric acid.
4. ’’Determine the melting point of lauric acid. How do both temperatures compare?’’
5. Explain the diagonal parts of the cooling curve in terms of changes in kinetic and potential energy. Do the same for the horizontal portions of the curve.
6. Explain how an increase in the amount of lauric acid used would affect the shape of the curves.
7. Explain in your own words what is going on at the molecular level as the liquid Lauric acid cools and freezes?
27. BOYLE’S LAW

Gases in a closed system are affected by change in pressure, while temperature remains constant. This relationship between gas volume and change in pressure is readily observed when the gas is in a closed system. In the lab activity air will be the gas studied. A plunger/cylinder apparatus will provide the closed system. To increase or decrease pressure in the column standard weight will be added or removed from the plunger platform. Pressure is measured most commonly in units of \( \text{atm} \), or \( \text{kPa} \). Careful calculations will be done to convert to Newtons (N).

### PRE LAB

1. State the purpose of this lab activity.
2. State Boyle’s law. What is the constant?
3. What is the conversion from grams to kilograms
4. What does the Newton (N) measurement represent?

### OBJECTIVE

1. *observe* the effect of pressure on the volume of a confined gas.
2. *graph* the pressure-volume relationship and describe the relationship.
3. *determine* the Boyle’s Law constant.

### MATERIAL

Plunger/cylinder apparatus  
4 weights totaling 2000 grams

### PROCEDURE

1. Obtain a plunger and cylinder apparatus. Work in assigned groups. Push the plunger all the way down and then pull it up to the 30.0 Ml mark. Cover the tip of the syringe with the red cap to confine the gas.

2. Push down a few times and when it comes to rest record the volume to 0.1 Ml. This is the initial volume for zero weight.

3. Place 500 grams of weight on the platform above the plunger so that the center of the weight is over the center of the platform. Record volume to the nearest 0.1 Ml.

4. Continue to add weights in 500 gram increments to a maximum of 2000 grams and recording the volume change each time.

5. All the volume measurements you made are slightly too large because of friction. You will now make a second series of the measurements while the weights are being removed from the platform. With 2000 grams on the platform, gently push the plunger down a few Ml. Gradually release your hand and when the plunger comes to rest, record the volume to the nearest 0.1 Ml.

6. Remove each weight one at a time and record each new volume. Record the volume with zero weight.
CALCULATIONS

1. Calculate weight: mass x 9.8

2. Calculate the pressure: weight/ 4.5 x 10^-4 m^2. This gives the pressure in Pa.

3. Calculate the average volume: V_{av}

4. The pressure that acts on the plunger with no added weights can be found by reading the barometric pressure and converting to Pa.
   Barometric reading x 133.3

5. Find total pressure: P_t = P + P_{atmospheres}

6. Calculate the Boyle’s Law constant: V_{avg} x P_t (Round to 2 sig fig.)

7. Plot a graph of P_t on the x-axis and the V_{avg} on the y-axis.

CONCLUSION

1. State a generalization about the effect of pressure on the volume of gases at constant temperature.

2. What kind of a relationship is shown by the graph?

3. When no weights are on the plunger, there is still pressure being exerted on the gas in the cylinder. This is the P_{atm}. What causes that pressure?

4. Does this experiment verify Boyle’s Law, within experimental error? Explain how it does.
28. DETERMINATION OF ABSOLUTE ZERO

A gas has different characteristics at varying temperatures. In this laboratory activity you will use air as the gas. Temperature is measured in the laboratory in Celsius. However, in the gas laws this value is changed to the Kelvin scale. You will utilize Charles’s Law in an effort to determine the volume of a gas at Absolute zero.

GOGGLES MUST BE WORN. CAUTION WITH HOT PLATES.

PRE LAB

1. What is the purpose of this lab activity?
3. What does extrapolate mean in respect to graphing.
4. What is the equation for percent error?

OBJECTIVES

1. Record air volume change under varied temperature conditions
2. Compare experimental with theoretical values for volume
3. Determine absolute zero by graphical analysis
4. Calculate % error

MATERIALS

- hot-plate with 1000mL beaker (HOT BATH)
- thermometer
- 250 mL Erlenmeyer flask
- mitten and tongs
- Styrofoam container with water and ice (ICE BATH)
- 1 or 2 hole stopper
- Tap water
- 100 ml graduated cylinder

PROCEDURE

1. Each student will perform the experiment independently. Two people will share the hot water bath and ice bath.

2. Set a large beaker of water (3/4 full) on a hot plate. Turn the hot plate to high so the water boils as soon as possible.

2. Prepare a cold water bath with ice and cold water. The temperature should be less than 10°C or as close to zero as possible.

3. Tightly stopper a DRY 250-mL flask with a one-hole rubber stopper and submerge it in the hot water bath (neck up) so the air in the flask is completely submerged in the water. The flask should be submerged up to the rubber stopper. Make certain no water gets into the flask. You may use tongs to keep the flask down.
4. As soon as the water appears to boil, measure the temperature of the water. It should be approx. 100°C. Record the temperature you read on your thermometer. Record this as $T_1$.

5. Leave the flask in the boiling water for at least one minute, then remove it and quickly invert it in the ice bath (neck down). A mitten may help in this tricky maneuver. Push the flask down so it is completely submerged under the ice water. Allow it to remain there until the flask has reached the temperature of the ice bath water.

6. Record the temperature of the ice water at this time. This is $T_2$.

7. Remove the flask by placing your finger over the hole(s) so no more water can enter or leave the flask.

8. Measure the volume of water drawn into the flask. Record.

9. Now, completely fill the flask with water then insert the rubber stopper. Measure this volume using the graduate cylinder. (It will be more than 100 mL.) This measurement gives you $V_1$ or the volume the air occupies at the high temperature stopper.

10. Clean up.

**CONCLUSION**

1. Prepare a graph with the volume on the y-axis and the temperature on the x-axis. Label the points.

2. Extrapolate the line to the point where it crosses the x-axis (the x intercept). Record this point.


4. You calculated the value for $V_2$ (calculation 5) and determined it experimentally (calculation 4). How do the two values compare?

5. Use your graph to determine what would be the volume of the sample of gas in the flask when the temperature is 50°C?
29. MOLAR VOLUME OF A GAS

You will experimentally determine the molar volume of a gas by reacting a known mass of magnesium metal with an excess of hydrochloric acid to produce hydrogen gas, which will be collected by water displacement in a gas collection tube. Whenever a gas is collected over water, the result is a mixture of the collected gas and water vapor. According to Dalton’s law of Partial Pressures the vapor pressure of the water can be subtracted from the total pressure to give the pressure of just hydrogen gas present in the mixture. We will assume that the total pressure is equal to atmospheric pressure and the vapor pressure of water can be obtained from a reference table. Having determined the volume and pressure of the hydrogen at the temperature of the lab, you can use the combined gas laws to find the volume this sample of gas would occupy at STP. The number of moles of hydrogen can be determined from the balanced chemical equation and the mass of magnesium used.

GOGGLES AND APRONS ARE REQUIRED DURING THIS ACTIVITY.
HANDLE THE HYDROCHLORIC ACID WITH CAUTION
WASH WITH SOAP AND WATER IF ANY ACID IS SPILLED ON YOUR SKIN.

PRE LAB

1. What is the purpose of this lab activity?
2. Write the balanced equation for the reaction of magnesium and hydrochloric acid.
3. What is the ratio of magnesium used to moles of hydrogen produced in the reaction?
4. What is meant by STP?
5. What is a eudiometer? What two gases will be collected in the eudiometer?
6. What piece of information, from a reference table will you need in order to complete the calculations in this experiment?

OBJECTIVES

1. Observe the reaction of magnesium and hydrochloric acid.
2. Apply Dalton’s law to find the partial pressure of the hydrogen gas collected.
3. Determine the volume this gas would occupy at STP.
4. Calculate the volume of one mole of this gas at STP.
5. Compare your experimental value to the accepted value for molar volume.

MATERIALS

Beaker (either 400 Ml or 600 Ml), Ring stand and clamp, Eudiometer (gas collection tube) and stopper with copper wire (from cart), Thermometer, Magnesium ribbon, ruler, 6M HCl, Barometer (located in the front of the room), water at room temperature.
PROCEDURE

1. THIS EXPERIMENT WILL BE DONE INDIVIDUALLY.
   Observe the demonstration set-up.

2. Obtain a piece of magnesium ribbon, carefully mass it and record to the nearest 0.1 cm.

3. Add water that is close to room temperature to a 400 Ml or 600 Ml beaker. (approx. 22°C)

4. Obtain a eudiometer (gas collection tube) and stopper fitted with copper wire hook. Wrap the magnesium ribbon around the wire.

5. Using the markings on the gas collection tube to measure the acid, **pour about 8 mL of HCl** in the tube. Then fill the collection tube to the top with the room temperature water.

6. Place the rubber stopper with the Mg in the tube. There should be no air bubbles in the tube. If there are remove the stopper and add more water.

7. Place your finger over the stopper hole and quickly **invert** the tube into the large beaker of water. Clamp the tube to the ring stand.

8. Observe the reaction. All the Mg will react in time. When the reaction stops, **allow about 5 minutes for the solution to cool to room temperature**.

9. While you are waiting, record the atmospheric pressure (in mm Hg) in the room from the barometer.

10. Measure and record the temperature of the water in the beaker. Assume that the collected gas is at the same temperature.

11. Gently tap the tube on the bottom of the beaker to release any bubbles of hydrogen on the bottom. Read the volume of the gas to the nearest 0.05 mL. Be sure to read the bottom of the meniscus.

12. Empty the collection tube and repeat the experiment if there is time. It is not necessary to dry the collection tube or change the water in the beaker.

**When you are finished, return all equipment to its original location. Dry your bench and wash your hands.**

CONCLUSION

**Show all work for calculations**

1. What are some sources of error in this experiment?
2. How would the volume of the hydrogen collected differ if twice as much Mg had been used?
3. What happens to the other product of the reaction formed in this experiment?
4. A volume of 35.0 mL of nitrogen gas was collected over water at 23°C and 753 mm Hg. The vapor pressure of water is 21.1 mm Hg.
   a) What is the pressure of nitrogen gas alone?
   b) What would the volume of the gas be at STP?
   c) What should be the volume of one mole of this gas at STP?
5. Find the volume of the following masses of gases at STP:
   a) 80 g O₂   b) 66 g CO₂
6. How many Liters would the following moles of any gas occupy at STP?
   a) 0.25 mole   b) 3.50 mole
30. PAPER CHROMATOGRAPHY

Chromatography is a technique for separating and identifying mixtures of compounds. The name stems from the Greek word chromos, meaning color. It was first applied to separate small quantities of colored compounds. All types of chromatography employ two different immiscible phases in contact with each other: the mobile phase and the stationary phase.

A small amount of the mixture to be separated (in this case ink and dyes) is placed near the edge of the paper, the stationary phase. The substances in the mixture are more or less tightly adsorbed on the surface of the paper. The bottom of the paper is then placed in a solvent, the mobile phase. As the solvent travels up the paper, by capillary action, the substances more attracted to the solvent will travel with it; substances tightly bound to the paper will not move at all; substances weakly attracted to the paper will move more slowly with the solvent. Separation occurs because different chemicals in the mixture travel different distances. When the solvent has moved the entire length of the paper, the paper is removed from the solvent and dried. Once developed, the paper, called a chromatogram, will contain different chemicals located at different positions on the paper.

The color and location of unknown components separated under specific chromatographic conditions can be matched with the color and location of other compounds subjected to the same conditions. This can be done by calculating R_f values:

\[ R_f = \frac{D_s}{D_f} \]

Where, \( D_s \) = distance traveled by a spot, and \( D_f \) = distance traveled by the solvent.

PRE LAB

1. What is the purpose of this lab activity?
2. What is a solvent? What is the solvent used in this activity?
3. Where is the mixture placed on the chromatography paper before placement in the solvent?
4. Why does the mixture travel up the paper once placed in the solvent?
5. What does separation of the mixture indicate?
6. What is a chromatogram?

OBJECTIVE

1. separate mixtures of compounds found in various pens.
2. compare components of inks in markers by calculating \( R_f \) values.

MATERIALS

- 2 pieces of chromatography paper (cut as directed)
- Pencil and ruler
- Bottle or plastic cup
- markers (or food coloring)
PROCEDURE
1. The lab will be done individually.
2. Draw a line with pencil about 2.5 cm from the edge of the strip of paper. This will mark the point of origin. Place a spot from the marker on the line. Let the ink dry.
3. Pour water in the bottle as demonstrated. Place the paper (spot down) in the bottle, allowing the bottom of the paper to touch the solution. DO NOT LET THE SPOT TOUCH THE SOLUTION.
4. Observe, and when the solution has nearly reached 2 cm from the top, remove the paper and mark the distance the solvent has traveled.
5. Mark the distance each spot has traveled by placing a dot in the center of each spot.
6. Measure the distance traveled by solvent for each chromatogram (Df). Measure the distances traveled by each spot (Ds) on each chromatogram. In all cases, measure from the origin (not the end of the paper), to the pencil line. Calculate the Rf for each spot.

CALCULATIONS
Calculate the Rf value for each spot. Where one substance has separated into two or more colors, calculate Rf for each colored spot. Place all measurement and calculations (show set-up) on the data table.

CONCLUSION
1. What is the purpose of the process of chromatography?
2. What causes the components to separate?
3. Which markers appear to contain only one compound? Explain.
5. Were there any colors which seemed to be the same in different markers?
   (Compare with others at your table.)
6. Why is it important to mark the chromatography paper with pencil, not pen?
31. ICE CREAM LAB

A PRACTICLE APPLICATION OF COLLIGATIVE PROPERTIES

Colligative properties of solutions are determined only by the number of dissolved particles and the nature of the solvent used. The “nature” includes osmotic pressure, boiling point elevation and freezing point depression (the property applied in this activity). Sodium chloride dissociates into one mole of sodium ions and one mole of chloride ions, while dissolving in ice. This does not occur in the ice cream mix being created in the separate compartment.

One solution freezes while the other melts!

ICE CREAM PRODUCTION

MATERIALS

<table>
<thead>
<tr>
<th>One plastic sphere</th>
<th>Plastic spoon</th>
<th>small paper cups</th>
<th>spoons</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 cup Rock salt (NaCl) or ½ cup CaCl₂</td>
<td>Ice</td>
<td>Towels</td>
<td>measuring cups/spoons</td>
</tr>
</tbody>
</table>

Recipe: Ingredients for 1 pint sphere
2 tablespoons sugar ¼ cup half and half or cream 1 tablespoon vanilla extract
1½ cups milk 1 Tablespoon flavored syrup (chocolate, strawberry, vanilla)

PROCEDURE

1. Clean lab bench and cover with paper. Work in groups of four.

2. Open (use special too;) and fill 1-pint sphere with 1 cup salt and ice. Record the temperature. Seal carefully.

3. Open other end that goes into the metal cylinder. Add milk, sugar, and vanilla. Seal carefully.

4. Now begin to shake, roll and pass sphere around to your other lab partners.

5. After 10 minutes open the ice cream end and use the plastic tool included to scrape the frozen ice cream away from the walls. Reseal.

6. Open the ice end and record the temperature. Add more ice if most is melted. Continue to shake for an additional 5 minutes

7. OK. You’ve worked hard enough.

8. Open the ice cream end, using the plastic scoop to remove the ice cream and place in a small paper cups. Get a spoon. Be nice and share.

9. Add extra flavoring if you like and enjoy!

Clean and dry all equipment
CONCLUSION QUESTIONS
Submit answers in full sentences on a separate sheet of paper at the end of class.

1. What were the temperatures of the pure ice and the ice with salt mixture after 10 minutes?
2. Why was salt added to the ice?
3. What phase change occurred in the milk?
4. Is the change in the milk exothermic or endothermic?
5. Describe what happened to the heat in the reaction.
6. If you did not add sugar would the ice cream have frozen faster? Why?
7. Why did the outside of the sphere begin the sweat? No it’s not a leak!
8. Why is salt spread on the roads during a winter storm?
32. RATE OF A REACTION

The rate of a chemical reaction is the time required for a given quantity of reactant(s) to be changed to product(s). The rate is affected by several factors including the type of reactants, concentration, temperature, pressure (in gases), and catalysts. Some reactions are fast and some reactions are slow. The rate of a specific reaction can be found only by experiment.

In this laboratory activity, you will study the effects of varying concentrations of a reactant on the rate. In a second experiment, you will study the effect of temperature on the rate. The reaction occurs when aqueous solutions of iodate (IO$_3^-$) and sulfite ion (SO$_3^{2-}$) are mixed. The sulfite solution also contains starch. The reaction is known as the Iodine Clock reaction, because a measurable amount of time elapses before the reaction reaches completion. The end of the reaction is recognized when a dark blue-black color appears. This color results from a complex which forms between starch and I$_2$, which is one of the products.

You will run several experiments in which the concentration of the iodate ion is varied and accurately time the reactions. You will then measure the time of reaction run at another temperature.

PRE LAB

Read the Introduction and Procedure and answer these questions.

1. In your own words, state the purpose of the experiment.
2. What factor is varied in Part I? What factor is constant?
3. What factor is varied in Part II? What factor is constant?
4. What is the purpose of the starch in the experiment?
5. Why is it necessary to have two 10 mL graduated cylinders for this activity?

MATERIALS

2 graduated cylinder (10 mL labeled A and B) 2 large test tubes
thermometer watch with second hand
beaker for ice water bath de-ionized water.
Solution A (with IO$_3^-$) Solution B (with SO$_3^{2-}$ and starch)

PROCEDURE

You will be working in pairs, one partner will time and the other will mix the reactants and note the color change. **Be sure to have a watch that can time seconds.**

PART I

1. Obtain solutions A (iodate ion) and B (sulfite- starch). Clean two 10 mL graduates and mark them "A" and "B". Clean two large test tubes.
2. Pour 6 mL of solution A into the graduated cylinder and add 1 mL of water.
3. Pour 3 mL of solution B into the other graduated cylinder.
4. Prepare to time the reaction. While one lab partner pours solution A into solution B, the other partner should immediately start timing the reaction. Pour the solutions back and forth several times from one test tube to the other to ensure thorough mixing. Then allow the mixture to stand. At the instant a color change occurs, the partner timing the reaction should note the elapsed time. Record this in your data table. Rinse the test tubes and shake out the water.

5. Prepare different concentrations of solution A as shown in the data table, and add the water to solution A. Repeat the experiment with each concentration.

6. Calculate average of the times for each trial and calculate rate for each trial.

7. Measure the temperature of the room and record.

**PART II**

1. Place ice and water in the beaker so the temperature is between 0°C and 10°C. Place 4 mL of solution A and 3 mL of water in one test tube and 3 mL of solution B in the other. Put the test tubes in the water baths making sure the solutions are totally submerged below the water. Place a thermometer in solution A, and wait until the solution reaches the temperature of the water baths (within a few degrees). Record the actual temperature of the solution.

2. Follow the same procedure as in PART I to run the reaction. Run two trials at the colder temperature using the concentrations of experiment 3. Average your trials.

**ANALYSIS**

Prepare two graphs: time vs. volume of A, and rate vs. volume of A.

**CONCLUSION**

1. Based on your experimental data, make a general statement about the effect of concentration of reactants on time and reaction rate.

2. Make a similar statement about the effect of temperature on reaction rate.

3. What other factors affect the rate of a reaction?
33. INVESTIGATION OF CHEMICAL EQUILIBRIUM

Le Chatelier’s Principle states that when a system at equilibrium is disturbed by an external factor, the system will react to the change and reach a new equilibrium. These changes can be studied in systems where shifts in colors can be observed. In this lab activity four different systems at equilibrium will be disturbed. In each case, you will predict the change then observe the result after the reaction is completed.

PRE LAB
1. State the purpose of this activity.
2. State the ways a system at equilibrium can be disturbed?
3. What type of systems will be affected by pressure?
4. What is a spectator ion?

PART I: Demonstration OBJECTIVE
1. Observe carefully as four separate activities are performed.
2. Predict the change and record your prediction on the data table before the experiment is done.

PART II: Activity OBJECTIVE
1. Observe reaction as you perform lab activities
2. Collect data and explain observations

MATERIALS
3 small test tubes  2 beakers  test tube holder
test tube rack  ice  Copper (II) chloride solution (CuCl₂)
hot plate  spot plate  pipets containing NaCl and AgNO₃

GOGGLES MUST BE WORN

PROCEDURE
1. Prepare a hot water bath by heating water in one beaker on the hot plate. Prepare a cold water bath in the other beaker using cold water and ice.
2. Place about 3 mL of the CuCl₂ solution into the 3 small test tubes.
3. Place one test tube in the hot water bath and the other in the ice bath. Record your prediction on the data table. After the systems have changed, record the results. Do not change the prediction if it does not agree with the observed results.
CONCLUSION

1. Write the equilibrium constant expressions for each system.
2. Using examples from this lab (including demo), describe each of the ways equilibrium systems were disturbed. Be specific about what was added or removed in each example.
3. Which type of change on a system was not done in this lab?
4. List at least four spectator ions in this lab.
5. If heat is added to an exothermic reaction, which increases: products or reactants?
6. If heat were added to an endothermic reaction, which increases: products or reactants?
7. In general, what must be done to remove a reactant or product?
34. CHANGE IN ENTHALPY OF A REACTION

INTRODUCTION: When a chemical reaction takes place, chemical bonds in the reactants are broken and new chemical bonds in the products are formed. Energy is always absorbed in the breaking of bonds and always released when bonds are formed. If the energy required to break old bonds is less than the energy released in forming new bonds, the difference in energy is given off and the reaction is said to be exothermic. The enthalpy change for an exothermic reaction is given a negative sign to indicate that energy flows from the system.

In this experiment you will measure the amount of energy released by the decomposition of hydrogen peroxide, \( \text{H}_2\text{O}_2 \), into water and oxygen and compare the experimentally determined result with a calculated value.

The common disinfectant, 3% hydrogen peroxide (\( \text{H}_2\text{O}_2 \)) solution, will be used with manganese dioxide as a catalyst for the decomposition. The amount of energy released in the reaction can be calculated from the mass of the water in the solution, the change in temperature, and the specific heat of water. The formula is

\[
\text{Energy released} = (\text{mass}) \times (\text{specific heat}) \times (\text{change in temperature})
\]

or

\[
\Delta H = MC\Delta T, \text{ where } \Delta T = T_f - T_i \text{ (final temperature – initial temperature)}
\]

PRELAB QUESTIONS:

1. State the purpose of this experiment.
2. Define enthalpy.
3. Write the equation for the decomposition of hydrogen peroxide into water and oxygen gas.
4. The reaction may occur in two steps. The energy for each step is given. Use this information to calculate the net energy released. This will be the theoretical or calculated value in kJ/mol.

\[
\begin{align*}
\text{H}_2 + \frac{1}{2} \text{O}_2 & \rightarrow \text{H}_2\text{O} \quad \Delta H = -286 \text{ kJ/mol} \\
\text{H}_2\text{O}_2 & \rightarrow \text{H}_2 + \text{O}_2 \quad \Delta H = +191 \text{ kJ/mol}
\end{align*}
\]

5. Calculate the heat released when 31 mL of hydrogen peroxide solution are decomposed and the temperature changes from 21.3\(^\circ\)C to 37.6\(^\circ\)C. Assume the density of the solution is the same as water (1g/mL) and use the specific heat of water 4.18 J/g \(^\circ\)C.

GOGGLES MUST BE WORN AT ALL TIMES

PROCEDURE:

Each person will do the experiment independently. NO PARTNERS.

1. Place two small scoops of manganese dioxide into the calorimeter.

2. Measure about 45 mL of 3% hydrogen peroxide and record the exact volume on the data table.

3. Measure the temperature of the hydrogen peroxide in the graduated cylinder and record it.
4. Pour the hydrogen peroxide into the calorimeter, put the lid on and push the thermometer through the lid. Swirl gently and note the temperature change. Record the highest temperature.

5. Open the calorimeter and look inside. Record your observation.

6. Empty the calorimeter and rinse it. Return it.

**CALCULATIONS:** *Number the calculations and show set-up for each in your lab book.*
1. Calculate the change in enthalpy for the amount of H$_2$O$_2$ in your data using $MC\Delta T$. Again assume the density of the H$_2$O$_2$ is the same as water.
2. Convert the Joules to kilojoules.
3. Calculate the mass of H$_2$O$_2$ in the solution of 3% hydrogen peroxide.
4. Calculate the moles of H$_2$O$_2$.
5. Calculate the $\Delta H$ in kJ/mol.
6. Calculate what % of the theoretical $\Delta H$ you were able to measure in this experiment. (Refer to question 4 in the Prelab for theoretical value.)

**CONCLUSION QUESTIONS:**
1. How will the mass of the calorimeter and contents change throughout the experiment? Explain.

2. Compare your results to the theoretical value. How do they compare? Refer to your data in your answer.

3. Suggest two reasons for the difference between your results and the theoretical value.

4. The decomposition of hydrogen peroxide is slow. If poured on the table there is no evidence of decomposition. On the other hand when 3% hydrogen peroxide is applied to a cut, the decomposition begins immediately (bubbles appear). Use your knowledge of biology to suggest a reason why this occurs.
35. pH and Indicators

An acid-base indicator is usually a weak acid with a characteristic vivid color. Upon addition of base, this acid is converted into its conjugate base, which is a different color. When the acid and its conjugate base exist in solution in approximately equal concentrations, an intermediated color is evident. For example, the common acid-base indicator used in the equilibrium lab, bromthymol blue (BTB,) is yellow in the acid form and blue in the form of its conjugate base. It changes color around pH 7. When equal concentrations of the acid and base are in solution, the resulting color is green.

PRELAB

1. State the purpose in your own words.
2. Describe the pH scale.
3. What is an indicator?

OBJECTIVES

1. *Observe* the colors of different indicators over a pH range from 1 to 12.
2. *Determine* the pH of an unknown substance.

MATERIALS

Droppers containing indicators;
Droppers containing known solutions of pH from 1 – 12.
Plastic small-scale reaction surface

PROCEDURE

**GOGGLES MUST BE WORN AT ALL TIMES.**

*The known solutions will be done in groups. UNKNOWN SOLUTIONS WILL BE TESTED INDIVIDUALLY.*

1. Obtain a set of 12 solutions of known pH’s.
2. Obtain a set of indicator solutions.
3. Drop 2 drops of solution in the appropriate space across the row and add a drop of indicator down the column. Observe and record the color.
4. Record the number of your unknown and test the unknown solution with each indicator.
CONCLUSION

1. Which would be the best indicator only to distinguish all acids from all bases?
2. Identify your unknown. Justify your answer by referring to your data. Answer in complete sentences.
3. Using the table, list the indicators, the pH range over which the color is changing, and the colors in the acidic range and the basic range. Copy this chart in your lab book to organize your answer.

<table>
<thead>
<tr>
<th>INDICATOR</th>
<th>pH RANGE WHERE COLOR CHANGES</th>
<th>COLOR</th>
<th>ACID</th>
<th>BASE</th>
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36. ACID BASE TITRATION

Titration is a technique for determining the concentration (molarity) of an unknown acid or base by neutralizing it with a known solution of acid or base. An indicator that changes color in the appropriate range is used to show when the neutralization is complete.

OBJECTIVES

1. Perform a titration where an acid is neutralized by a base.
2. Use a titration to determine the concentration of an unknown acid.

MATERIALS

buret, ring stand, buret clamp, 10 mL graduate cylinder,  
Erlenmeyer flask, Beaker for waste liquid, 0.10 M HCl  
0.10 M NaOH, phenolphthalein  
UNKNOWN ACID with a number

PROCEDURE

GOGGLES AND APRONS MUST BE WORN DURING THIS LABORATORY

1. THIS LABORATORY EXERCISE WILL BE DONE INDIVIDUALLY. Solutions, except the unknown, will be shared at a table. All calculations should be done in class and the Report page will be handed in at the end of the period. You will be graded on lab procedure, neatness of the report and on the accuracy of your results.

2. Rinse out the buret with water and make sure the liquid flows through easily. Rinse the buret with a little of the 0.10 M NaOH as demonstrated.

3. Fill the buret with the base. Let a little run out into a waste beaker. Read the volume (the bottom of the meniscus). Record.

4. Add between 5 – 10 mL of 0.10 M HCl to a graduate cylinder and record the exact volume to the nearest 0.1 mL. Pour the HCl into an Erlenmeyer flask and add 1 drop of phenolphthalein.

5. Add base slowly to the flask, as demonstrated, until the indicator just turns A LIGHT PINK COLOR AND THE COLOR REMAINS. This is the end-point or the neutralization point. Read and record the final volume of the buret.

6. Repeat steps 4 and 5 two more times, for a total of three trials. It is not necessary to fill the buret each time. Rinse the flask with water after each trial. It is not necessary to dry the flask.
7. Calculate the molarity of the acid for each trial. If the molarities are close you have mastered the technique. Go on to the unknown.

8. Obtain a bottle containing an unknown concentration of HCl. Record the number of your sample. Repeat the procedure in steps 3-5. Record volumes and calculate the molarity of your unknown.
35. Understanding Half-Life

In any sample of a radioactive isotope, the individual atoms are decaying in a random fashion. It is impossible to predict which atom is the next to decay, yet statistically you can predict how many atoms will decay within a certain period of time. Scientists measure how much time elapses while half the atoms of a given radioactive sample decay. That time is called half-life. For example, the half-life of carbon-14 is 5730 years. This means that if you were to start with 100 grams of carbon-14 today, in 5730 years you would have 50 grams left. After another 5730 years (11,460 years total) you would have 25 grams left. Half-lives of radioactive isotopes vary greatly, from a fraction of a second to billions of years. The half-life is a very important consideration when choosing a radioactive isotope for a specific application such as a medical tracer.

PRE LAB

1. State the purpose of the experiment in your own words.
2. What is meant by the term half-life?
3. What is the half life of carbon-14?
4. How can carbon-14 help in determining the age of a fossil?
5. Suppose you have a radioactive isotope with a half-life of 2 years and you start with 800 grams of this substance today.
   a. How much will you have left 2 years from today?
   b. How much will you have left 8 years from today?
6. Is the quantity of a radioactive isotope ever equal to exactly zero? Explain your answer.

OBJECTIVE

1. Interpret a model of radioactivity and half-life.
2. Demonstrate the connection between half-life and a decay graph.
3. Relate half-life and geologic dating.

MATERIALS

Cardboard box

200 pennies
PROCEDURE

1. Place 200 pennies into a box and put the lid on. Shake the box for several seconds then open the box and remove all the pennies that have the “heads” side up. These are your “decayed atoms”. Carefully count these pennies and record this number on your Report Page. Do not put the pennies back in the box.
2. Shake the box again for several seconds. Open the box and again remove, count, and record the “heads” pennies.
3. Continue this process until you have either one penny or no pennies left. Remember to record the number of pennies removed each time.
4. Put all the materials away and begin the calculations and questions.

CONCLUSION

1. Make a graph of your results: x-axis time, y-axis pennies remaining.
2. Describe the shape of your graph.
3. If you had started with 1000 pennies, would the shape of the graph be different? Explain why or why not.
4. Approximately what percent of the pennies were removed each time?
5. Is it possible to identify which pennies will be “heads” up? Explain.
6. Is it possible to predict approximately how many pennies will be “head” up for each shake? Explain.
38. DETERMINING THE HALF-LIFE OF Ba-137m
(A class experiment)

The half-life of a radioactive isotope (radioisotope) is the time required for the activity (or mass) of a sample to be reduced to one half of its original activity (or mass). Today, the class will determine the half-life of the radioisotope Ba-137m. The m stands for “metastable”, a condition of temporary stability. A Ba-137 generator will be used as the source of the Ba-137m. The generator contains Cs-137, which decays by beta emission to Ba-137m with a half-life of over 30 years. The Ba-137m that is produced possesses more energy than is normally possessed by a stable nucleus and “cools” (decays) to stable Ba-137 by gamma emission with a half-life that you will determine.

PRE LAB
1. What is the purpose of this activity?
2. If you are taking a 5 minute background reading, how do you calculate the counts per minute (cpm)?
3. If your readings are taken for 30 seconds, how do you calculate the cpm (gross)?
4. How do you calculate the cpm (net) from the cpm (gross)?

OBJECTIVES
1. Collect data for the decay of Ba-137m.
2. Construct a decay curve graph of the data.
3. Determine the half-life of the Ba-137m from the decay curve.

MATERIALS
A Cs-137/Ba-137m “minigenerator, Geiger tube and sample holder, and Geiger counter (scalar).

PROCEDURE
1. The Geiger counter will be turned on by your teacher and allowed to run without a sample for 5 minutes so that a background radiation count can be taken. Record the counts.
2. Your teacher will “milk” the minigenerator to wash out a 7 or 8 drop sample of Ba-137m into the small metal dish. A Geiger tube will be placed over the sample in a sample holder.
3. The scalar will be turned on for 30 seconds and then stopped. You will quickly record the counts. The Counter will be set to zero and restarted 30 seconds after it was turned off. In this manner you will collect several 30 second readings with a 30 second “rest” between each reading. This will continue for about 10 minutes.
4. Fill out your data sheet as the experiment proceeds.
5. Construct a graph of the data. Your cpm (net) should be plotted on the y-axis and the time (either in seconds or minutes) on the x-axis. Use the beginning of each interval as your time.
example: for the interval 0-30 sec, plot your activity versus 0 seconds.) Remember to draw the best fit smooth curve.

6. Use your graph to calculate the half-life of Ba-137m. Choose any convenient cpm (net) for your first activity and find the time that corresponds to that activity. Then take half of that first activity (this is your second activity) and find the time that corresponds to that. Repeat using half of the second activity. Be sure to do at least two trials and average the results. Show all your work either on the graph paper or on a separate sheet of paper.

CONCLUSION

1. Based on your graph, what is the half-life of Ba-137m? Remember to show your work.

2. If the “accepted” value for the half-life is 2.6 minutes, what is your percent error? Show your calculation. If you plotted your graph in seconds, remember to convert your units

3. 

4. Some Cs-137 is always mixed into the Ba-137m sample. Does this affect the measurement of the half-life of the Ba-137m sample? Explain why or why not.

5. Why is it safe to throw the remaining Ba-137 down the drain after 10 or 15 minutes?
39. Back to Chernobyl

At 1:23 a.m. on Saturday, April 26, 1986, Reactor #4 at the Chernobyl (in the former Soviet Union) nuclear power plant exploded. Only a shell of the reactor building remained as the reactor fuel core began to burn its way through the concrete floor toward the earth below. The first of the two explosions released a huge plume of radioactive material that spread and settled over most of Europe. Meanwhile, Soviet scientists and firefighters tried to determine how to extinguish the fire. It was the world’s worst nuclear accident.

The heat of the main explosion had carried the plume so high that initial radiation levels near the plant were surprisingly low. Heavy winds carried radiation away from Kiev, a nearby city of 2.5 million people, and into Scandinavia and the rest of Europe. On Monday, April 28, scientists at a power plant north of Stockholm, Sweden detected levels of radiation 4 to 5 times higher than normal. It was not until later that evening that Soviet officials announced that an accident at Chernobyl had occurred.

Following the explosion, contaminated material was removed from the Chernobyl plant and Reactor #4 was sealed. The three other reactors were put back into operation, one only six months after the accident, the other two shortly thereafter. Although the Chernobyl accident prompted significant changes in nuclear policy within the Soviet Union, the nuclear power industry there continues.

Pripyat, located near the Chernobyl plant, was a thriving city of 50,000 before the accident. Although there were no visible effects in Pripyat after the explosion, radioactive material was being scattered over the entire countryside. The city remained deserted for at least three years after the accident.

The film “Back to Chernobyl” looks at the impact of the accident on Pripyat, the Soviet Union, Europe, and the rest of the world three years after the nuclear accident. It also explores the debate that has evolved within the scientific community. Did Chernobyl signal the end of the peaceful use of nuclear power or is it providing lessons that will take the nuclear power industry into the future?

POST LAB QUESTIONS
Answer questions as directed by your teacher
Days before the 17th anniversary on 26 April 2003 of the disaster at Ukraine's Chernobyl nuclear reactor, a Ukrainian intelligence agency declassified several documents that show design flaws, shoddy workmanship, and major violations of safety rules at the plant. The 121 documents released exposed 29 accidents that occurred before the 1982 accident, including a 1982 accident that caused the release of small doses of radiation.

Russia has responded by claiming that post-Soviet Ukraine has not properly maintained the concrete canopy covering the faulty reactor, which was only built to last five years. Atomic Energy Minister Alexander Rumyantsev warned that "nobody has inspected these walls in detail. We do not know what reactions are taking place under the shield." The Ministry has also claimed that there are holes in the so-called sarcophagus through which radiation could leak. A collapse of the structure would likely spread a cloud of radioactive dust and panic throughout the region.

Preceding the disaster's anniversary, 5,000 Chernobyl victims marched in downtown Kiev protesting the continuing lack of funding for cleanup and victim compensation, including housing and healthcare. On April 26 hundreds of people gathered at a small hill with marble plates inscribed with victims’ names, and a bell tolled 17 times for each year that has past since the explosion at the plant initiated a disaster whose effects are still being felt in the region.

The United Nations estimates that 15,000 to 30,000 people living in the contaminated zones have since died due to radiation exposure. Russian activists claim that these figures do not include the 300,000 Russian "liquidators," those who were sent to clean up the disaster and have now fallen victim to radiation related illnesses.

More than 2.45 million people have been hospitalized in Ukraine as of early 2002 with illnesses sparked by the disaster, including 472,400 children, according to the Ukrainian Health Ministry. According to a study published in a July 1999 issue of the journal Cancer, the rate of thyroid cancer among Ukrainians age 15 and younger increased 10-fold in the years following the accident.

Activists and survivors of the accident are not distracted by the blame trading going on between the Russian 'Id Ukrainian governments that are nether compensating victims nor doing enough to protect future I generations. One activist was quoted in the Christian Science Monitor as saying, "What we want is for the (post-Soviet) governments of Russian, Ukraine, and Belarus to work together in dealing with the legacy of Chernobyl, but there is no sign of that happening."

Indian Point 3 – Plant with a containment building