Miami-Dade County Public Schools
Curriculum and Instruction (Science)

Essential Labs
(Minimum Required Laboratory Activities)

For High School
Chemistry

June 2007
The School Board of Miami-Dade County Public Schools

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Introduction

The purpose of this packet is to provide the Chemistry teachers with a list of basic laboratories and hands-on activities that students in an 11th-grade Chemistry class should experience. Each activity is aligned with the Chemistry Curriculum Pacing Guide and the Sunshine State Standards (SSS). Emphasis should be placed on those activities that are aligned to the Annually Assessed Benchmarks which are consistently assessed in the grade 11 Science Florida Comprehensive Assessment Test (FCAT). As a result, included in the packet you will find several lab activities that are not usual to a Chemistry class but are specifically related to the grade 11 Annually Assessed Benchmarks.

All the hands-on activities were designed to cover the most important concepts found in Chemistry. In some cases, more than one lab was included to cover a specific benchmark. In most cases, the activities were designed as simple as possible without the use of advanced technological equipment to make it possible for all teachers to use these activities. However, it is highly recommended that technology, such as ExploreLearning’s Gizmos and the hand-held data collection equipment from Vernier, Texas Instruments, and Pasco, is implemented in the science classrooms.

This document is intended to be used by secondary science departments in M-DCPS so that all science teachers can work together, plan together, and rotate lab materials among classrooms. Through this practice, all students and teachers will have the same opportunities to participate in these experiences and promote discourse among learners, which are the building blocks of authentic learning communities.

Acknowledgement:

M-DCPS Curriculum and Instruction (Science) would like to acknowledge the efforts of the teachers who worked arduously and diligently on the preparation of this document.
Materials List

Each list corresponds to the amount of materials needed per station (whether one student or a group of students uses the station). Lab aprons and goggles should be assigned to each student at all labs requiring mixtures of chemicals.

Lab 1: Laboratory Techniques
- table salt
- sand
- distilled water
- 100-mL graduated cylinder
- 250-mL beakers (2)
- 50-mL beakers (2)
- balance
- ring stand
- ring
- funnel
- scoops (2)
- stirring rod
- filter paper
- weighing paper
- water bottle
- watch glass

Lab 2: “The Alka-Popper”
- 2 Alka-Seltzer tablet broken in 1/8, 1/4, and 1/2 pieces.
- film canister
- water
- 10-mL graduated cylinder
- Timer (seconds)

Lab 3: Density
- 100-mL graduated cylinder
- 2-L graduated cylinder (plastic)
- balance (500g capacity)
- tap water
- rubber stopper (#2 solid)
- can of non-diet soft drink
- can of diet soft drink
- dropper

Lab 4: Changes of State
- hot plate
- 250-mL beaker
- water
- temperature probe
- CBL/CBL2/or LabPro
- TI Graphing Calculator with ChemBio or Datamate
- balance
- *thermometer (instead of probe/CBL/TI calculator)
- *timer (instead of probe/CBL/TI calculator)
- *graph paper (instead of probe/CBL/TI calculator)
- hot mitt

Lab 5: Physical and Chemical Changes
- safety goggles
- lab apron
- modeling clay
- wax paper
- plastic teaspoon
- 2 plastic cups
- 1 antacid tablet
- tape measure
- cold water
- balance
Lab 6: Models of Atomic Structure
- white paper
- glue
- colored dots (can be made using a hole-puncher to create dots from peel off labels)
- atom template

Lab 7: Isotopes
- 100 pennies
- balance large
- box top
- 2 plastic cups

Lab 8: Half-Life
- 100 pennies
- plastic cup
- container with lid such as a shoebox
- timer or clock with second hand

Lab 9: Flame Tests
- lab goggles
- lab apron
- Bunsen burner
- nichrome wire loop
- wash bottle with distilled water
- well plate
- hydrochloric acid (HCl) 6.0 M in a dropper bottle
- solutions of metallic salts:
  - calcium nitrate (Ca(NO₃)₂)
  - copper nitrate (Cu(NO₃)₂)
  - lithium nitrate (LiNO₃)
  - potassium nitrate (KNO₃)
  - strontium nitrate (Sr(NO₃)₂)
  - sodium chloride (NaCl)
  - unknown solution

Lab 10: Periodic Trends
- Previous models from Models of Atomic Structure Lab
- Colored construction paper (8 different colors for class activity)
- Glue
- Colored markers
- String or wire

Lab 11: Bonding
Conductivity Apparatus:
A) tape
B) 9-V battery
C) battery clip
D) bare wire leads
E) resistor
F) LED or buzzer
G) wood backing.(tongue depressor)

Tools to make Conductivity Apparatus:
- Scissors
- Wire stripper
- Tape solder
- 1 nail
- Pliers
- Matches, Candle or soldering iron

Other:
- beakers or cups (100 mL or less)
- 1 cm² of Al foil
- distilled water (~100 mL)
- a penny
- wash bottle with distilled water
- rubbing alcohol (isopropyl alcohol) (10mL)
- stirring rod
Approximately 1 g of:
- sucrose (table sugar)
- NaCl
- SiO₂ (sand)
- paraffin wax shavings
- CaCl₂ (calcium chloride)
- CuSO₄ (copper II sulfate)

Lab 12: Activities of Metals
- 1.0M Zn(NO₃)₂
- 1.0M Al(NO₃)₃
- 1.0M Cu(NO₃)₂
- 1.0M Mg(NO₃)₂
- pipettes (4)
- wire cutters
- Cu wire
- Al wire
- Mg ribbon
- Zn metal strips (4)
- emery cloth or fine sandpaper
- 24-well microscale reaction plate

Lab 13: Determining Reaction Rates
Solution A: Dissolve 4.3 g of potassium iodate (KIO₃) per liter of water.
Solution B: Make a paste of 4 g soluble starch in a small amount of water. Slowly add paste to 900 mL of boiling water. Boil for several minutes and allow solution to cool. Just before using, add 0.2 g of sodium thiosulfate Na₂S₂O₅ and 5 mL of 1 molar H₂SO₄, and add water to bring the final volume to 1 liter of solution.
Materials per group:
- Three 250-ml beakers
- Two graduated cylinders
- Stopwatch
- Stirring rods
- Distilled water at room temperature
- 120 ml of Solution A
- 90 ml of Solution B

Lab 14: Hydrated Crystals
- hotplate
- balance (preferably 2 decimal place)
- hydrated MgSO₄ (Epsom salts)
- 2 beakers (50 or 100 mL and 400 or 600 mL)
- hot mitts or beaker tongs

Lab 15: A Bagged Chemical Reaction
- Safety goggles
- Lab apron
- Calcium chloride pellets (CaCl₂)
- Baking soda, sodium bicarbonate (NaHCO₃)
- Phenol red solution
- Measuring cup or graduated cylinder
- 2 plastic teaspoons
- Plastic cup
- Scissors
- 1-gallon Ziploc-type bag
- Water
Lab 16: Mole Ratio
- iron metal filings, 20 mesh
- copper(II) sulfate pentahydrate (CuSO₄ · 5H₂O)
- 400-mL and 150mL beakers
- 100-mL graduated cylinder
- weighing paper (filter paper can be used)
- balance
- hot plate
- beaker tongs
- distilled water
- stirring rod

Lab 17: Energy Content of Foods
- Lab Pro or CBL 2 system*
- TI graphing calculator*
- Temperature probe*
- DataMate program
- 2 Food samples
- food holder (paper clip + foil)
- wooden splint
- candle (large diameter)
- aluminum foil squares
- utility clamp and slit stopper
- 2 Stirring rods
- ring stand and 4-inch ring
- 100-mL Graduated cylinder
- soda can
- cold water
- matches
- goggles

Lab 18: Rates of Evaporation
- distilled water
- ethanol
- isopropyl alcohol
- acetone
- household
- ammonia
- droppers (5)
- small plastic cups (5)
- grease pencil or marking pen
- masking tape
- paper towel
- square of waxed paper
- stopwatch

Lab 19: Boyle’s Law
- safety goggles
- Boyle’s law apparatus
- ring stand clamp
- 5 chemistry textbooks
- 2 pens or pencils of different colors

Lab 20: Solubility Curves
- potassium chloride (KCl)
- distilled water
- balance
- evaporating dish (or 100-mL beaker)
- 25-mL graduated cylinder
- watch glass
- 250 or 400mL beaker
- hot plate or burner with ring stand, 2 rings & wire gauze
- test tube (18X150 mm)
- utility clamp
- glass stirring rod
- thermometer
- funnel
- cotton wadding
- tongs or hot mitts
Lab 21: Percentage of Acetic Acid in Vinegar

- apron
- goggles
- 10 mL graduated cylinder
- 24-well plate or 3 small beakers
- 2 thin-stemmed pipets or droppers

- phenolphthalein indicator
- 2.0 mL standardized 0.6 M NaOH
- 2.0 mL white vinegar
- stirrer
Parts of a Lab Report
A Step-by-Step Checklist

Good scientists reflect on their work by writing a lab report. A lab report is a recap of what a scientist investigated. It is made up of the following parts.

**Title (underlined and on the top center of the page)**

**Benchmarks Covered:**
- Your teacher should provide this information for you. It is a summary of the main concepts that you will learn about by carrying out the experiment.

**Problem Statement:**
- Identify the research question/problem and state it clearly.

**Potential Hypothesis(es):**
- State the hypothesis carefully. Do not just guess; instead try to arrive at the hypothesis logically and, if appropriate, with a calculation.
- Write down your prediction as to how the independent variable will affect the dependent variable using an “if” and “then” statement.
  - If (state the independent variable) is (choose an action), then (state the dependent variable) will (choose an action).

**Materials:**
- Record precise details of all equipment used.
  - For example: a balance that measures with an accuracy of +/- 0.001 g.
- Record precise details of any chemicals used.
  - For example: (5 g of CuSO₄·5H₂O or 5 g of copper (II) sulfate pentahydrate).

**Procedure:**
- Do not copy the procedures from the lab manual or handout.
- Summarize the procedures; be sure to include critical steps.
- Give accurate and concise details about the apparatus and materials used.

**Variables and Control Test:**
- Identify the variables in the experiment. State those over which you have control. There are three types of variables:
  1. **Independent variable** (also known as the manipulated variable): The factor that can be changed by the investigator (the cause).
  2. **Dependent variable** (also known as the responding variable): The observable factor of an investigation that is the result or what happened when the independent variable was changed.
  3. **Constant variable**: The other identified independent variables in the investigation that are kept or remain the same during the investigation.
- Identify the control test. A control test is the separate experiment that serves as the standard for comparison to identify experimental effects and changes of the dependent variable resulting from changes made to the independent variable.
Data:
- Ensure that all data is recorded.
  - Pay particular attention to significant figures and make sure that all units are stated.
- Present your results clearly. Often it is better to use a table or a graph.
  - If using a graph, make sure that the graph has a title, both axes are labeled clearly, and that the correct scale is chosen to utilize most of the graph space.
- Record all observations.
  - Include color changes, solubility changes, whether heat was evolved or taken in, etc.

Results:
- Ensure that you have used your data correctly to produce the required result.
- Include any other errors or uncertainties that may affect the validity of your result.

Conclusion and Evaluation:
A conclusion statement answers the following seven questions in at least three paragraphs.

I. First Paragraph: Introduction
1. What was investigated?
   a) Describe the problem.
2. Was the hypothesis supported by the data?
   a) Compare your actual result to the expected result (either from the literature, textbook, or your hypothesis).
   b) Include a valid conclusion that relates to the initial problem or hypothesis.
3. What were your major findings?
   a) Did the findings support or not support the hypothesis as the solution to the restated problem?
   b) Calculate the percentage error from the expected value.

II. Middle Paragraphs: These paragraphs answer question 4 and discuss the major findings of the experiment, using data.
4. How did your findings compare with other researchers?
   a) Compare your result to other students’ results in the class.
     - The body paragraphs support the introductory paragraph by elaborating on the different pieces of information that were collected as data that either supported or did not support the original hypothesis.
     - Each finding needs its own sentence and relates back to supporting or not supporting the hypothesis.
     - The number of body paragraphs you have will depend on how many different types of data were collected. They will always refer back to the findings in the first paragraph.

III. Last Paragraph: Conclusion
5. What possible explanations can you offer for your findings?
   a) Evaluate your method.
   b) State any assumptions that were made which may affect the result.
6. What recommendations do you have for further study and for improving the experiment?
   a) Comment on the limitations of the method chosen.
   b) Suggest how the method chosen could be improved to obtain more accurate and reliable results.
7. What are some possible applications of the experiment?
   a) How can this experiment or the findings of this experiment be used in the real world for the benefit of society?
Lab Roles and Their Descriptions

Cooperative learning activities are made up of four parts: group accountability, positive interdependence, individual responsibility, and face-to-face interaction. The key to making cooperative learning activities work successfully in the classroom is to have clearly defined tasks for all members of the group. An individual science experiment can be transformed into a cooperative learning activity by using these lab roles.

<table>
<thead>
<tr>
<th>Project Director (PD)</th>
<th>Materials Manager (MM)</th>
</tr>
</thead>
<tbody>
<tr>
<td>The project director is responsible for the group.</td>
<td>The materials manager is responsible for obtaining all necessary materials and/or equipment for the lab.</td>
</tr>
<tr>
<td><strong>Roles and responsibilities:</strong></td>
<td><strong>Roles and responsibilities:</strong></td>
</tr>
<tr>
<td>- Reads directions to the group</td>
<td>- The only person allowed to be out of his or her seat to pick up needed materials</td>
</tr>
<tr>
<td>- Keeps group on task</td>
<td>- Organizes materials and/or equipment in the work space</td>
</tr>
<tr>
<td>- Is the only group member allowed to talk to the teacher</td>
<td>- Facilitates the use of materials during the investigation</td>
</tr>
<tr>
<td>- Shares summary of group work and results with the class</td>
<td>- Assists with conducting lab procedures</td>
</tr>
<tr>
<td></td>
<td>- Returns all materials at the end of the lab to the designated area</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Technical Manager (TM)</th>
<th>Safety Director (SD)</th>
</tr>
</thead>
<tbody>
<tr>
<td>The technical manager is in charge of recording all data.</td>
<td>The safety director is responsible for enforcing all safety rules and conducting the lab.</td>
</tr>
<tr>
<td><strong>Roles and responsibilities:</strong></td>
<td><strong>Roles and responsibilities:</strong></td>
</tr>
<tr>
<td>- Records data in tables and/or graphs</td>
<td>- Assists the PD with keeping the group on-task</td>
</tr>
<tr>
<td>- Completes conclusions and final summaries</td>
<td>- Conducts lab procedures</td>
</tr>
<tr>
<td>- Assists with conducting the lab procedures</td>
<td>- Reports any accident to the teacher</td>
</tr>
<tr>
<td>- Assists with the cleanup</td>
<td>- Keeps track of time</td>
</tr>
<tr>
<td></td>
<td>- Assists the MM as needed.</td>
</tr>
</tbody>
</table>

When assigning lab groups, various factors need to be taken in consideration:

- Always assign the group members, preferably trying to combine in each group a variety of skills. For example, you can place an “A” student with a “B”, “C”, and a “D” and or “F” student.
- Evaluate the groups constantly and observe if they are on task and if the members of the group support each other in a positive way. Once you realize that a group is dysfunctional, re-assign the members to another group.
Grade 11 Annually Assessed Benchmarks

The following lists the seventeen Annually Assessed Benchmarks that will be tested each year of the Grade 11 Science FCAT. It should be noted that within specific benchmarks, other benchmarks are embedded and could be tested annually.

- **SC.A.1.4.3**- The student knows that a change from one phase of matter to another involves a gain or loss of energy. (Also assesses B.1.4.3)

- **SC.A.1.4.4**- The student experiments and determines that the rates of reaction among atoms and molecules depend on the concentration, pressure, and temperature of the reactants and the presence or absence of catalysts.

- **SC.A.2.4.5**- The student knows that elements are arranged into groups and families based on similarities in electron structure and that their physical and chemical properties can be predicted.

- **SC.B.1.4.1**- The student understands how knowledge of energy is fundamental to all the scientific disciplines (e.g., the energy required for biological processes in living organisms and the energy required for the building, erosion, and rebuilding of the Earth).

- **SC.C.1.4.1**- The student knows that all motion is relative to whatever frame of reference is chosen and that there is no absolute frame of reference from which to observe all motion. (Also assesses C.1.4.2 and C.2.4.6)

- **SC.C.2.4.1**- The student knows that acceleration due to gravitational force is proportional to mass and inversely proportional to the square of the distance between the objects.

- **SC.D.1.4.1**- The student knows how climatic patterns on Earth result from an interplay of many factors (Earth's topography, its rotation on its axis, solar radiation, the transfer of heat energy where the atmosphere interfaces with lands and oceans, and wind and ocean currents).

- **SC.D.1.4.2**- The student knows that the solid crust of Earth consists of slow-moving, separate plates that float on a denser, molten layer of Earth and that these plates interact with each other, changing the Earth's surface in many ways (e.g., forming mountain ranges and rift valleys, causing earthquake and volcanic activity, and forming undersea mountains that can become ocean islands).
Grade 11 Annually Assessed Benchmarks (cont.)

- **SC.D.2.4.1** - The student understands the interconnectedness of the systems on Earth and the quality of life. (Also assesses SC.G.2.4.4)

- **SC.E.1.4.1** - The student understands the relationships between events on Earth and the movements of the Earth, its moon, the other planets, and the sun. (Also assesses SC.E.1.4.2 and SC.E.1.4.3)

- **SC.F.1.4.1** - The student knows that the body processes involve specific biochemical reactions governed by biochemical principles. (Also assesses SC.F.1.4.3 and SC.F.1.4.5)

- **SC.F.2.4.3** - The student understands the mechanisms of change (e.g., mutation and natural selection) that lead to adaptations in a species and their ability to survive naturally in changing conditions and to increase species diversity. (Also assesses SC.D.1.4.4 and SC.F.1.4.2)

- **SC.G.1.4.1** - The student knows of the great diversity and interdependence of living things. (Also assesses SC.G.1.4.2)

- **SC.G.2.4.2** - The student knows that changes in a component of an ecosystem will have unpredictable effects on the entire system but that the components of the system tend to react in a way that will restore the ecosystem to its original condition. (Also assesses SC.B.1.4.5 and SC.G.2.4.5)

- **SC.H.1.4.1** - The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.H.1.2.1, SC.H.1.2.2, SC.H.2.4.2, SC.E.2.4.6, and SC.E.2.4.7)

- **SC.H.2.4.1** - The student knows that scientists control conditions in order to obtain evidence, but when that is not possible for practical or ethical reasons, they try to observe a wide range of natural occurrences to discern patterns.

- **SC.H.3.4.2** - The student knows that technological problems often create a demand for new scientific knowledge and that new technologies make it possible for scientists to extend their research in a way that advances science. (Also assesses SC.H.3.4.5 and SC.H.3.4.6)
Laboratory Safety

Rules:

• Know the primary and secondary exit routes from the classroom.
• Know the location of and how to use the safety equipment in the classroom.
• Work at your assigned seat unless obtaining equipment and chemicals.
• Do not handle equipment or chemicals without the teacher’s permission.
• Follow laboratory procedures as explained and do not perform unauthorized experiments.
• Work as quietly as possible and cooperate with your lab partner.
• Wear appropriate clothing, proper footwear, and eye protection.
• Report to the teachers all accidents and possible hazards.
• Remove all unnecessary materials from the work area and completely clean up the work area after the experiment.
• Always make safety your first consideration in the laboratory.

Safety Contract:

I will:

• Follow all instructions given by the teacher.
• Protect eyes, face and hands, and body while conducting class activities.
• Carry out good housekeeping practices.
• Know where to get help fast.
• Know the location of the first aid and fire fighting equipment.
• Conduct myself in a responsible manner at all times in a laboratory situation.

I, _______________________, have read and agree to abide by the safety regulations as set forth above and also any additional printed instructions provided by the teacher. I further agree to follow all other written and verbal instructions given in class.

Signature: _______________________________  Date: ___________________
Laboratory Techniques and Lab Safety
(Adapted from Glencoe Lab Manual p. 1-4)

Benchmark:
SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.E.2.4.6, and SC.E.2.4.7, SC.H.1.2.1, SC.H.1.2.2, SC.H.2.4.2,)

SC.H.1.4.4 -The student knows that scientists in any one research group tend to see things alike and that therefore scientific teams are expected to seek out the possible sources of bias in the design of their investigations and in their data analysis. CS

SC.H.2.4.1- The student knows that scientists control conditions in order to obtain evidence, but when that is not possible for practical or ethical reasons, they try to observe a wide range of natural occurrences to discern patterns.

Objective:
• Measure the mass of solid substances.
• Measure a volume of water.
• Separate components of a mixture through filtration.

Background Information:
Chemistry has been developed largely through experimentation. Chemistry courses use laboratory experiences to demonstrate, clarify, and develop principles of chemistry. Behavior in the laboratory is more structured than in the classroom. Certain rules of conduct pertaining to safety and keeping a clean work environment must be followed at all times. You must also adopt correct procedures for using glassware and other pieces of equipment. General safety rules are summarized at the beginning of this lab manual. However, there often will be more specific safety rules or special procedures to follow when performing an experiment. Your teacher will provide these added instructions before you perform any lab activity. If you are unsure of any procedure, always ask your teacher before proceeding.

In this activity, you will practice some laboratory techniques and apply laboratory safety rules. You will determine the mass of different solid materials, measure the volume of a liquid, and separate mixtures of chemicals. You will also review specific safety rules.

Lesson Lead:
How can the mass of an object be measured? How can the volume of a liquid be measured? How can a mixture be separated?

Materials:
• table salt
• sand
• distilled water
• 100-mL graduated cylinder
• 250-mL beakers (2)
• 50-mL beakers (2)
• balance
• ring stand
• ring
• funnel
• scoops (2)
• stirring rod
• filter paper
• weighing paper
• water bottle
• watch glass
Pre-Lab:
1. What is the safety rule concerning working alone in the laboratory?
2. What is the safety rule concerning the handling of excess chemicals?
3. What should you do if you spill a chemical?
4. Read the entire laboratory activity. State the safety precautions which will be needed to handle the different chemicals and lab equipment in this experiment. Record your answer below.

Safety Precautions: [Student writes.]

Procedure:
1. Using a scoop, transfer a small amount of table salt to a 50-mL beaker.
2. Measure the mass of a piece of weighing paper to 0.1 g or 0.10 g depending on your laboratory balance. Record this mass in Data Table 1.
3. Add about 5.0 g of table salt from the 50-mL beaker to the weighing paper. Record the mass of the weighing paper and table salt to 0.1 g in Data Table 1.
4. Transfer the table salt to the 250-mL beaker and place all excess table salt into an appropriate waste container, as indicated by your teacher.
5. Using another scoop, transfer a small amount of sand to the second 50-mL beaker. Using the techniques described in steps 2 and 3, measure out about 5.0 g of sand. Then transfer the sand to the 250-mL beaker containing the table salt.
6. Using a 100-mL graduated cylinder, measure out 80 mL of distilled water. Measure the volume of the water to 0.1 mL by reading at the bottom of the meniscus, as illustrated in Figure A. Record the volume of water measured in Data Table 1.
7. Pour the water into the 250-mL beaker containing the table salt and sand. Using the stirring rod, gently stir the mixture for 1 minute. Record your observations in Data Table 2.
8. Place a clean 250-mL beaker on the base of the ring stand. Attach the ring to the ring stand and set the funnel in the ring so that the stem of the funnel is in the beaker. Adjust the height of the ring so that the bottom of the funnel stem is approximately halfway up the beaker. Fold a piece of filter paper as illustrated in Figure B. Place the folded filter cone in the funnel. Note: Label your filter paper with your period and groups name using a pencil.
9. To avoid splashing and to maintain control, you will pour the liquid down a stirring rod. Place the stirring rod across the top of the 250-mL beaker that contains the mixture, as shown in Figure B. The stirring rod should rest in the spout and extend several inches beyond the spout. Grasp the beaker with your hand and place your index finger over the stirring rod to keep it in place. Slowly pour the contents of the beaker into the filter cone, allowing the liquid to pass through the filter paper and collect in the beaker.
10. While holding the beaker at an angle, use the water bottle to rinse the beaker and wash any remaining solid from the beaker into the filter cone. Record your observations in Data Table 2.

Note: If you wish to continue and completely isolate and analyze the components of the mixture, complete the additional procedure steps listed under the Extensions on the next page, instead of steps 11-14.
11. Allow the filter cone to drain. Then remove the filter cone and carefully unfold the filter paper. Place the filter paper on a watch glass to and record your observations in Data Table 2.
12. Place all chemicals in the appropriately labeled waste container.
13. Return all lab equipment to its proper place.
14. Clean up your work area

Data Table 1:

<table>
<thead>
<tr>
<th>Amounts</th>
<th>Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of table salt + weighing paper (g)</td>
<td></td>
</tr>
<tr>
<td>Mass of weighing paper (g)</td>
<td></td>
</tr>
<tr>
<td>Initial Mass of table salt (g)</td>
<td></td>
</tr>
<tr>
<td>Mass of sand + weighing paper (g)</td>
<td></td>
</tr>
<tr>
<td>Mass of weighing paper (g)</td>
<td></td>
</tr>
<tr>
<td>Initial Mass of sand (g)</td>
<td></td>
</tr>
<tr>
<td>Volume of water (mL)</td>
<td></td>
</tr>
</tbody>
</table>

Data Analysis:
- To find the "Mass of table salt," subtract the "Mass of weighing paper" from the "Mass of table salt + weighing paper."
- To find the "Mass of sand," subtract the "Mass of weighing paper" from the "Mass of sand + weighing paper."

Data Table 2:

<table>
<thead>
<tr>
<th>Step</th>
<th>Observations</th>
</tr>
</thead>
<tbody>
<tr>
<td>Procedure Step 7</td>
<td></td>
</tr>
<tr>
<td>Procedure Step 10</td>
<td></td>
</tr>
<tr>
<td>Procedure Step 11</td>
<td></td>
</tr>
</tbody>
</table>

Analysis and Conclusions:
1. Why were the excess reagents not put back into the original reagent bottle?
2. What differences were observed between the mixture of salt and sand in the 250-mL beaker and the same materials after the water was added?
3. Why were the samples of table salt and sand placed into 50-mL beakers prior to weighing?
4. If one of the pieces of glassware is dropped and breaks, why is it necessary to clean up and tell the teacher about broken glass immediately?
5. Why is it necessary to wear safety goggles and a lab apron while performing experiments in the lab?
6. Why is eating, drinking, or chewing gum not allowed in a laboratory?
7. Why must you always wash your hands after working in a laboratory?
8. Why do you never work alone in a chemical laboratory?
Extension: Separation of a Mixture

Procedure:
1. Allow the filter cone to drain. Then remove the filter cone and carefully unfold the filter paper. Place wet filter paper and collected material A on a watch glass to dry. Once it is completely dry, record the mass of the filter paper plus collected material A in Data Table 3.
2. Heat the beaker containing the filtered water till all the water is gone.
3. Allow the beaker to cool and record the mass of the beaker and material B.
4. Place all chemicals in the appropriately labeled waste container.
5. Return all lab equipment to its proper place.
6. Clean up your work area

Data Table 3:

<table>
<thead>
<tr>
<th>Amounts</th>
<th>Mass</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of filter paper (g)</td>
<td></td>
</tr>
<tr>
<td>Mass of filter paper + collected material A(g)</td>
<td></td>
</tr>
<tr>
<td>Mass of collected material A(g)</td>
<td></td>
</tr>
<tr>
<td>Final mass of beaker + material B (g)</td>
<td></td>
</tr>
<tr>
<td>Mass of material B (g)</td>
<td></td>
</tr>
</tbody>
</table>

Analysis and Conclusions:
1. Using the information recorded in Tables 2 and 3, determine the identity of materials A and B.
2. Compare the mass of material A and material B with the original mass of sand and salt. Add the initial mass of sand and salt and compare with the final mass of material A + B.
3. What are some possible sources of error in this activity?
“The Alka-Popper”

Benchmark:
SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.E.2.4.6, and SC.E.2.4.7, SC.H.1.2.1, SC.H.1.2.2, SC.H.2.4.2)

SC.H.1.4.4 - The student knows that scientists in any one research group tend to see things alike and that therefore scientific teams are expected to seek out the possible sources of bias in the design of their investigations and in their data analysis. CS

SC.H.2.4.1 - The student knows that scientists control conditions in order to obtain evidence, but when that is not possible for practical or ethical reasons, they try to observe a wide range of natural occurrences to discern patterns.

Objectives:
• Use the scientific method to conduct an investigation.
• Determine the dependent and independent variables of an experiment.
• Analyze the number of trials required.
• Identify the control.
• Graph and analyze data.

Background Information:
Chemistry is a basic science which deals with matter and energy, the composition of matter and the changes that matter undergoes. Much like other science disciplines, chemistry relies on the scientific method, a systematic way to find answers to problems through experimentation. Chemists make observations using their senses and take measurement using tools available in the laboratory. The collected data is analyzed in order to formulate conclusions.

A key part of the scientific method is that you have a testable hypothesis which clearly states the relationship between the dependent and independent variables. This lab will focus on using the scientific method to learn about chemical reactions.

Lesson Lead:
How can you increase the rate of “popping’’?

Hypothesis: [Student writes.]

Materials:
• 2 Alka-Seltzer tablets broken in 1/8, 1/4, and 1/2 pieces.
• film canister
• water
• 10-mL graduated cylinder
• Timer (seconds)

Safety Precautions:
• Always wear safety goggles and a lab apron
Procedure:
1. Break tablet into three pieces: 1/2, 1/4, and 1/8. Hint: break tablet in half, place one half aside. Take 2nd half and break it in half again, now you have two quarters. Set one aside. Take the final quarter and break it in half now you have an eighth. You will have an eight left over.
2. Measure 10ml of water and place it into the film canister.
3. Place 1/8 piece of an Alka-Seltzer tablet into a film canister with the water.
4. Place the lid on the canister quickly!
5. Measure the amount of time in seconds for the lid to “pop” and place in the data table.
6. Repeat for 1/4 and 1/2 tablet and record data as Trial 1.
7. Repeat steps 1-6 and record data as Trial 2.
8. Average your data from your two trials.
9. Compare your data with the other groups in your class.

Data Table:

<table>
<thead>
<tr>
<th>Size of Tablet</th>
<th>Time (Seconds)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Trial 1</td>
</tr>
<tr>
<td>1/8</td>
<td></td>
</tr>
<tr>
<td>1/4</td>
<td></td>
</tr>
<tr>
<td>1/2</td>
<td></td>
</tr>
</tbody>
</table>

Analysis and Conclusions:
1. What are the independent and dependent variables?
2. Graph your average data and the class average data using different symbols or colors.
3. Remember to title and label graph.
4. Explain the construction of your graph:
   a. Why did you choose this type of a graph?
   b. How did you decide where to place the dependent and independent variable?
   c. How did you choose the distance for your intervals on the x and y axes?
5. Which size of tablet caused the fastest reaction rate? Which one had the slowest reaction? Explain.
6. What is the trend?
7. What is the relationship between the variables?
8. Did you get the same results as other groups? Why or why not?
9. How many trials did you do? Is this appropriate?
10. What was your control?
Density
(Adapted from Glencoe Lab Manual p. 1- 4)

Benchmarks:
SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.H.1.2.1, SC.H.1.2.2, SC.H.2.4.2, SC.E.2.4.6, and SC.E.2.4.7)

SC.H.2.4.1- The student knows that scientists control conditions in order to obtain evidence, but when that is not possible for practical or ethical reasons, they try to observe a wide range of natural occurrences to discern patterns.

Objective:
• Measure the mass and volume of several different objects.
• Calculate the density of objects by using their measured mass and volume.
• Compare the densities of various objects.
• Calculate percentage error.

Background Information:
Density is a physical property of a substance and is often used to identify what the substance is. Density is the ratio of the mass of a substance to its volume. Density can be computed by using the equation

\[
\text{Density} = \frac{\text{Mass}}{\text{Volume}}
\]

Mass and volume measurements can be made in the laboratory. Mass can be determined by using a balance. If the object has a regular shape, such as a cube or a cylinder, volume can be calculated from length measurements. However, most objects have irregular shapes, and the volume must be determined indirectly. One way to measure the volume of an irregularly shaped item that does not dissolve in or react with water is by water displacement. An item that is entirely submerged in water will displace a volume of water equal to its volume. It is necessary to use the proper units when calculating the density of a substance. Densities of liquids and solids are usually expressed in terms of g/mL or g/cm³. Densities of gases are usually expressed in g/L.

The accuracy of your methods can be reported in terms of percent error. The percent error of measurements is a comparison of the differences between experimental results and theoretical or standard values, expressed as a percentage. The equation for percent error is:

\[
\% \text{ error} = \frac{|\text{experimental value} - \text{theoretical value}|}{\text{theoretical value}}
\]

Lesson Lead:
How can you find the densities of objects by using water displacement to measure their volumes?

Pre-Lab:
1. Define density.
2. Write the mathematical expression of density. What units are associated with density?
3. Read the entire laboratory activity. Form a hypothesis that compares the density of a rubber stopper to the density of water. Form a second hypothesis that compares the densities of a non-diet soft drink and a diet soft drink to water. Record your hypotheses.
4. The density of aluminum is 2.70 g/cm³. What volume will 13.5 grams of aluminum occupy?

Hypotheses: [Student answers Pre-Lab question #3.]
   a) __________________________________________________________________________
   b) __________________________________________________________________________

Materials:
   • 100-mL graduated cylinder
   • 2-L graduated cylinder (plastic)
   • balance (500g capacity)
   • tap water
   • rubber stopper (#2 solid)
   • can of non-diet soft drink
   • can of diet soft drink
   • dropper

Safety Precautions:
   • Always wear safety goggles and a lab apron.
   • Clean up any spills immediately.
   • Do not eat or drink anything in a laboratory.

Procedure:
Part A: Density of Water
   1. Find the mass of a clean, dry 100-mL graduated cylinder. Record this mass in Data Table 1.
   2. Fill the cylinder with distilled water. Use a dropper to adjust the bottom of the meniscus exactly to the 100.0-mL mark.
   3. Find and record the mass of the graduated cylinder and water.
   4. Calculate and record the mass of the water.

Part B: Density of a Rubber Stopper
   1. Find the mass of a solid #2 rubber stopper. Record this mass in Data Table 2.
   2. Pour about 50 mL of tap water into the 100-mL graduated cylinder. Read and record the exact volume.
   3. Place the rubber stopper into the graduated cylinder. Make sure that it is completely submerged. (Try using a pencil to hold the stopper just under the surface of the water.)
   4. Read and record the exact volume.

Part C: Density of a Can of Non-Diet Soft Drink
   1. Find the mass of an unopened can of non-diet soft drink. Record this mass in Data Table 3.
   2. Pour about 1000 mL of tap water into the 2000-mL graduated cylinder. Read and record the exact volume.
   3. Place the can of soft drink into the graduated cylinder, making sure that it is completely submerged.
   4. Read and record the exact volume.

Part D: Density of a Can of Diet Soft Drink
   1. Find the mass of an unopened can of diet soft drink. Record this mass in Data Table 4.
   2. Pour about 1000 mL of tap water into the 2000-mL graduated cylinder. Read and record the exact volume.
   3. Place the can of diet soft drink into the graduated cylinder, making sure that it is completely submerged.
   4. Read and record the exact volume.
Cleanup and Disposal:
Return all materials and supplies to their proper place, as directed by your teacher.

Data and Observations:
Data Table 1

<table>
<thead>
<tr>
<th>Part A: Density of Water</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of empty graduated cylinder (g)</td>
</tr>
<tr>
<td>Mass of graduated cylinder and water (g)</td>
</tr>
<tr>
<td>Mass of water (g)</td>
</tr>
<tr>
<td>Volume of water (ml)</td>
</tr>
<tr>
<td>Density of water (g/mL)</td>
</tr>
</tbody>
</table>

Data Table 2

<table>
<thead>
<tr>
<th>Part B: Density of Rubber Stopper</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of rubber stopper (g)</td>
</tr>
<tr>
<td>Initial volume of water in graduated cylinder (mL)</td>
</tr>
<tr>
<td>Final volume of water in graduated cylinder (mL)</td>
</tr>
<tr>
<td>Volume of rubber stopper (mL)</td>
</tr>
<tr>
<td>Density of rubber stopper (g/mL)</td>
</tr>
</tbody>
</table>

Data Table 3

<table>
<thead>
<tr>
<th>Part C: Density of a Can of Non-Diet Soft Drink</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of can of non-diet soft drink (g)</td>
</tr>
<tr>
<td>Initial volume of water in graduated cylinder (mL)</td>
</tr>
<tr>
<td>Final volume of water in graduated cylinder (mL)</td>
</tr>
<tr>
<td>Volume of can of non-diet soft drink (mL)</td>
</tr>
<tr>
<td>Density of can of non-diet soft drink (g/mL)</td>
</tr>
</tbody>
</table>

Data Table 4

<table>
<thead>
<tr>
<th>Part D: Density of a Can of Diet Soft Drink</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of can of diet soft drink (g)</td>
</tr>
<tr>
<td>Initial volume of water in graduated cylinder (mL)</td>
</tr>
<tr>
<td>Final volume of water in graduated cylinder (mL)</td>
</tr>
<tr>
<td>Volume of can of diet soft drink (mL)</td>
</tr>
<tr>
<td>Density of can of diet soft drink (g/mL)</td>
</tr>
</tbody>
</table>

Analysis and Conclusions:
1. Use the mass and volume data to calculate the densities of water, the rubber stopper, a can of non-diet soft drink, and a can of diet soft drink. Show work and then record the answers in the data tables.

2. Did the volume of water change when an object was placed into a graduated cylinder that was half-filled with water?

3. Would you expect the densities of various fruit juices to all be the same? Explain.

4. When you use the terms heavier or lighter to compare objects with the same volume, what property of the objects are you actually comparing?
5. Why do you think the can of non-diet soft drink is denser than the can of diet soft drink?

6. The density of water is 1 g/mL or 1 g/cm³. Use the percent error equation provided in the Background Information section to calculate the percent error of your measurement. What could have you been done to improve the accuracy of your measurements?

Real-World Chemistry:
1. How can the concept of density be used to differentiate between a genuine diamond and an imitation diamond?

2. Explain why a tractor-trailer can be completely filled with one type of merchandise, such as butter, but only partially filled with a second type of material, such as steel.


Change of States

**Benchmarks:**
SC.A.1.4.3- The student knows that a change from one phase of matter to another involves a gain or loss of energy. (Also assesses B.1.4.3)

SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.E.2.4.6, and SC.E.2.4.7, SC.H.1.2.1, SC.H.1.2.2, SC.H.2.4.2,)

SC.H.2.4.1- The student knows that scientists control conditions in order to obtain evidence, but when that is not possible for practical or ethical reasons, they try to observe a wide range of natural occurrences to discern patterns.

**Objective:**
- Determine how the temperature of water changes through a period of time.
- Find the thermal energy needed to boil and evaporate the water.

**Background Information:**
Common phase changes are solid to liquid (melting) and liquid to gas (boiling.). All phase changes involve a change in the internal energy, but **no change in the temperature**.

During the change of phase from ice to liquid water (melting), the heat added is absorbed by the ice. This absorbed heat weakens the forces holding the molecules together causing the ice to melt. As this is occurring, both ice and liquid water are present and the temperature remains at the melting point. The energy needed for this phase change is called the **Heat of Fusion, ΔH_{fus}**.

In a similar manner, the temperature remains at the boiling point until all the water is changed to water vapor. The heat energy added overcomes the forces of attraction holding the molecules of liquid water together, allowing them to become gaseous molecules. The energy needed for this phase change is called the **Heat of Vaporization, ΔH_{vap}**.

Both the Heat of Fusion and the Heat of Vaporization are experimentally determined for each substance and can be reported in Joules or calories per gram. For water, the ΔH_{fus} is 334 J/g (79.7 calories/g) and the ΔH_{vap} is 2260 J/g (537 calories/g). The energy needed for the phase changes of water are represented by “q” (expressed in Joules) and can be calculated from the following equations:

For melting:  \[ q = \Delta H_{fus} \cdot \text{mass} \]
For boiling:  \[ q = \Delta H_{vap} \cdot \text{mass} \]

Below the boiling point, in the liquid state, any heat energy that is added increases the speed of the water molecules, increasing the kinetic energy. When this happens, **an increase of temperature is observed**. The energy (q) needed to change the temperature (ΔT) depends on the mass (m) and the specific heat capacity (Cp) of the substance which for water is 4.18 J/g°C (1.00 calories/g°C) for water:

\[ q = \Delta T \cdot m \cdot Cp \]
Lesson Lead:
What changes of energy occur as water is heated?

Materials
- hot plate
- 250-mL beaker
- water
- temperature probe
- CBL/CBL2/or LabPro
- TI Graphing Calculator with ChemBio or Datamate
- balance
- hot mitt
- *thermometer (instead of probe/CBL/TI calculator)
- *timer (instead of probe/CBL/TI calculator)
- *graph paper (instead of probe/CBL/TI calculator)

* See Alternative Procedure for lab using thermometers below.

Safety Precautions:
- Always wear safety goggles and a lab apron.
- Use hot mitt to handle hot plate.
- Do not eat or drink anything in a laboratory.

Procedure for CBL/temperature probe:
1. Direction (how to measure temperature by using CBL and graphic calculator)
2. Measure the mass of an empty 250-mL beaker. Add 150mL of water and record the mass of the beaker plus the water.
3. Turn your hot plate to high setting (or as recommended by your teacher). Allow a few minutes for the plate to heat up.
4. Place a temperature probe into the water inside the beaker.
5. Plug the temperature probe cable into Channel 1 of the CBL System. Connect the CBL System to the TI Graphic Calculator with link cable, using the port on the bottom edge of each unit. Firmly press in the cable ends.
6. Turn on the CBL unit and the calculator. Press PRGM and select CHEMBIO or press Apps and choose DATAMATE Press ENTER. Press ENTER again to go to the Main Menu.
7. Set up the calculator and CBL for one temperature probe.
   a. Select SET UP PROBES from the Main Menu.
   b. Enter “1” as the number of probes.
   c. Select TEMPERATURE from SELECT PROBE menu.
   d. Enter “1” as the channel number.
8. Set up the calculator and CBL for data collection.
   a. Select COLLECT DATA from Main Menu.
   b. Select TIME GRAPH from the DATA COLLECTION menu.
   c. Enter “30” as the time between samples, in seconds. (The CBL will collect the data for a total of 15 minutes.)
   d. Enter “-1” as the minimum temperature (Ymin). To enter –1 use (-), not –.
   e. Enter “105” as the maximum temperature (Ymax).
   f. Enter “5” as the temperature increment (Yscl).
9. Place the beaker on the hot plate and begin data collection.
   a. Press ENTER on the calculator to begin the data collection. It will take 30 seconds for the graph to appear with the first data point plotted.
Alternative Procedure for Thermometers:
1. Turn your hot plate to high setting (or as recommended by your teacher) and allow a few minutes for the plate to heat up.
2. Measure the mass of an empty 250-mL beaker. Add 150mL of water and record the mass of the beaker plus the water in Table 1.
3. Measure the initial temperature of the water in °C and record in Table 2.
4. Place the beaker on the hot plate and record the temperature every 30s. Carefully stir the water before taking temperature reading. Do not allow the thermometer to touch the bottom of the beaker when recording the temperature.
5. When the water starts boiling, continue recording the temperature every 30 seconds for an additional three minutes.
6. Carefully remove the beaker from the hot plate, let it cool and then record the volume of the remaining water.

Data Table 1: Mass

<table>
<thead>
<tr>
<th>Measurements</th>
<th>Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Empty beaker</td>
<td></td>
</tr>
<tr>
<td>Beaker + Water</td>
<td></td>
</tr>
<tr>
<td>Mass of water</td>
<td></td>
</tr>
</tbody>
</table>

Data Table 2: Temperature

<table>
<thead>
<tr>
<th>Time (s)</th>
<th>Temperature (°C)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial</td>
<td></td>
</tr>
<tr>
<td>30 s</td>
<td></td>
</tr>
<tr>
<td>60 s</td>
<td></td>
</tr>
<tr>
<td>90 s</td>
<td></td>
</tr>
<tr>
<td>continue..</td>
<td></td>
</tr>
</tbody>
</table>

Analysis and Calculations:

A. **Graph**- Make a Heating Curve graph by plotting temperature (Y-axis) vs. time (X-axis).

B. **Calculations**- (Show all your work.)

1. Find the thermal energy given to the water to raise its temperature to the boiling point (record in Table 3).
   a) Copy the mass (g) of the water from Table 1.
   b) Calculate the change in water temperature:
      \[ \Delta T = (T_{final} - T_{initial}) \]
   c) Calculate the heat energy(q) gained by the water using the equation in both Joules and calories:
      \[ q = \Delta T \cdot m \cdot Cp \]

\[\text{where:} \quad q = \text{heat energy (in Joules or calories)}\]
\[\text{\( \Delta T \) = change in temperature (in °C)}\]
\[m = \text{initial mass of water (g)}\]
\[Cp = \text{specific heat of water} = 4.18 \text{ J/g°C or 1.00 cal/ g °C}\]

Data Table 3

<table>
<thead>
<tr>
<th>Energy to Raise Temperature</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass (g)</td>
</tr>
<tr>
<td>( \Delta T ) (°C)</td>
</tr>
<tr>
<td>( q_1 ) (J)</td>
</tr>
<tr>
<td>( q_1 ) (cal)</td>
</tr>
</tbody>
</table>
2. Find the energy needed for vaporization (record in Table 4).
   a) Determine the mass of the water vapor formed by calculating the change in mass of the water (Initial mass – Final mass)
   b) Use the Heat of Vaporization ($\Delta H_{\text{vap}}$) for water, 2260 J/g, 537 cal/g and calculate the energy ($q$) needed for vaporization in Joules and calories using the equation:

\[
q = \Delta H_{\text{vap}} \cdot \text{mass}
\]

### Conclusion:
1. What is required to bring about a phase change in a substance?
2. What is occurring during the horizontal or flat portion of the Heating Curve graph?
3. At what temperature did the water boil?
4. Describe what is occurring in the sloped area of the Heating Curve graph?
5. Can you add heat energy to an object without increasing its temperature? Explain.
6. What is the total energy in Joules and calories required to convert your water sample into steam? (Hint: add the amounts of energy needed to raise the temperature and the amount of energy for vaporization: $q_1 + q_2$).
7. Compare the amount of energy to raise the temperature and the amount of energy to vaporize your sample.
8. Optional: How much energy is released if the all your water is in the vapor state and condenses back to liquid?

### Extension:
1. Label the heating curve in the spaces provided using the following terms:
2. Ice, liquid water, water vapor, melting, boiling or vaporization, increase in kinetic energy and increase in potential energy, temperature increase and no temperature change.
3. Use arrows to point to the areas in the graph where there is a “temperature increase” and those where there is “no temperature change”.

### Data Table 4

<table>
<thead>
<tr>
<th>Energy for Vaporization</th>
<th>Amount</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of water vapor (g)</td>
<td></td>
</tr>
<tr>
<td>$q_2$ (J)</td>
<td></td>
</tr>
<tr>
<td>$q_2$ (cal)</td>
<td></td>
</tr>
</tbody>
</table>
4. Complete the following chart for all possible phase changes using the first row as an example:

<table>
<thead>
<tr>
<th>Description of Phase Change</th>
<th>Term for Phase Change</th>
<th>Heat Movement (absorbed or released)</th>
<th>Temperature Change?</th>
</tr>
</thead>
<tbody>
<tr>
<td>ex. Solid to liquid</td>
<td>Melting</td>
<td>Heat absorbed</td>
<td>None</td>
</tr>
</tbody>
</table>

<p>| | | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
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<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Change of States
Teacher Notes (Answer Key)

Conclusion:
1. What is required to bring about a phase change in a substance?
   (The addition or removal of energy – primarily in the form of heat)
2. What is occurring during the horizontal or flat portion of the heating curve graph?
   (A change of state takes place during any horizontal flat part of a Heating Curve graph. Vaporization takes place as a liquid changes into a gas. While the change of phase is occurring, temperature stays the same, which means that average kinetic energy remains unchanged. The energy supplied is used to weaken the attraction between the molecules as ice melts or water boils. This means that the increase in heat shows up as an increase in potential energy but not as an increase in kinetic energy.)
3. What are the melting and boiling points for water?
   (0°C and 100°C theoretically although their results may vary due to instrumental variations)
5. Describe what is occurring in the sloped area of the heating curve? (As heat energy is added, heat can be used to speed up the molecules in a substance which would cause an increase in temperature related to an increase in average kinetic energy).
6. Can you add heat energy to an object without increasing its temperature? Explain.
   (Heat energy at a phase temperature does not increase the temperature till the phase change is complete.
7. What is the total energy in Joules and calories required to convert your water sample into steam?
   (The answer varies as students add the amounts of energy needed to raise the temperature and the amount of energy for vaporization: q1 + q2).
8. Compare the amount of energy to raise the temperature and the amount of energy to vaporize your sample.
   (Answers can vary but the energy needed to vaporize the sample should be greater than the energy needed to raise the temperature.
9. Optional: How much energy is released if the all your water is in the vapor state and condenses back to liquid?
   (Answers vary but students should answer the same the amount of energy needed to vaporize that amount of water; energy is absorbed for vaporization but released for condensation.)
Extension:

<table>
<thead>
<tr>
<th>Description of Phase Change</th>
<th>Term for Phase Change</th>
<th>Heat Movement (absorbed or released)</th>
<th>Temperature Change?</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid to liquid</td>
<td>Melting</td>
<td>Heat absorbed</td>
<td>None</td>
</tr>
<tr>
<td>Liquid to solid</td>
<td>Freezing</td>
<td>Heat released</td>
<td>None</td>
</tr>
<tr>
<td>Liquid to gas</td>
<td>Evaporation/boiling</td>
<td>Heat absorbed</td>
<td>None</td>
</tr>
<tr>
<td>Gas to liquid</td>
<td>Condensation</td>
<td>Heat released</td>
<td>None</td>
</tr>
<tr>
<td>Solid to gas</td>
<td>Sublimation</td>
<td>Heat absorbed</td>
<td>None</td>
</tr>
<tr>
<td>Gas to Solid</td>
<td>Deposition</td>
<td>Heat released</td>
<td>None</td>
</tr>
</tbody>
</table>

Temperature increase
Increase in KE

No temperature change
Increase in PE

Ice → liquid melting
Liquid → Vapor (vaporization)
Water vapor

Ice → liquid melting
Liquid water
Temperature increase
Increase in KE
Observing Physical and Chemical Changes
(From MDCPS Physical Science Required Labs)

Benchmarks:
SC.A.1.4.3- The student knows that a change from one phase of matter to another involves a gain or loss of energy. (Also assesses B.1.4.3)

SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.E.2.4.6, SC.E.2.4.7, SC.H.1.2.1, SC.H.1.2.2, and SC.H.2.4.2)

SC.H.2.4.1- The student knows that scientists control conditions in order to obtain evidence, but when that is not possible for practical or ethical reasons, they try to observe a wide range of natural occurrences to discern patterns.

Objective:
• Observe and differentiate between physical and chemical changes.

Background Information:
If you break a piece of chewing gum, you change some of its physical properties, its size and shape. However, you do not change the identity of the materials that make up the gum. When a substance freezes, boils, or condenses, it undergoes physical changes. A change in size, shape, or state of matter is called a physical change.

The smell of rotten eggs, burning of the coal, the foaming of an antacid tablet in a glass of water, or the formation of rust on bike fenders are signs that chemical changes have taken place. A change of one substance to another is called a chemical change.

Lesson Lead:
How would you describe the process of evaporating water from seawater?

Materials:
• safety goggles
• lab apron
• modeling clay
• wax paper
• plastic teaspoon
• 2 plastic cups
• 1 antacid tablet
• tape measure
• cold water
• balance

Safety Precautions:
Always wear safety goggles and a lab apron

Procedure:
1. Read through the procedure to determine what data you need to collect and create a data table to record the data.
2. Put on safety goggles and apron.
3. Take a block of modeling clay. Measure its dimension using a tape measure. Record its mass using the triple beam balance.
4. Shape the clay into a ball and record its mass. Measure the dimensions of the ball with the tape measure.

5. Mold the clay into a shape of your choice. Record the mass and try to measure its dimensions.

6. Take the antacid tablet and place it in a small plastic cup. Measure 2 ml of water into another small plastic cup. Place both cups on the balance and record their combined mass.

7. Crush the antacid tablet and place it back into the plastic cup. Place both the cup with the crushed tablet and the cup with the water on the balance and record their combined mass again.

8. Pour water into the cup with the crushed tablet. Record your observations. Place the cup with the water-tablet mixture and the empty cup on the balance and record their combined mass a third time. (Draw your data table for steps 6-8.)

9. Clean up your area and dispose of the substances as directed by your teacher.

Analysis and Conclusions:
1. Is reshaping the clay a physical or chemical change?
2. Describe the changes that take place with the antacid tablet. Why does the mixture of the antacid tablet and water have a different mass after the reaction?
3. Based on your observation of the antacid tablet. What kind of change takes place when it is placed in water?
4. Does a chemical change take place in this experiment? Explain?
5. How do density, volume, and shape affect this experiment?
Models of Atomic Structure

Benchmarks:
SC.A.2.4.1
The student knows that the number and configuration of electrons will equal the number of protons in an electrically neutral atom and when an atom gains or loses electrons, the charge is unbalanced. CS

SC.C.2.4.2
The student knows that electrical forces exist between any two charged objects.

SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.E.2.4.6, SC.E.2.4.7, SC.H.1.2.1, SC.H.1.2.2, and SC.H.2.4.2)

SC.H.3.4.1- The student knows that performance testing is often conducted using small-scale models, computer simulations, or analogous systems to reduce the chance of system failure. CS

Objective:
• Learn how models facilitate the understanding of scientific phenomena.

Background Information:
Building a model of an atom structure can help you understand how a complex system operates. Something as simple as observing and releasing an inflated balloon can give you information about the forces involved in launching rockets. Building and using a model for an atom can help you understand how the main parts of an atom relate to each other. A model helps us understand something we cannot see directly, usually because it is too large or too small. As scientists have learned more about the atom, they have proposed newer models reflecting the newly gained knowledge.

The atom consists of a positively charged center, or nucleus, surrounded by negatively charged particles called electrons. The two major kinds of particles in the nucleus are protons and neutrons. The number of protons, or atomic number, in a neutral atom is equal to the number of electrons. Electrons are located around the nucleus in a region called electron cloud. The energy differences of the electrons can be described by energy levels.

<table>
<thead>
<tr>
<th>Energy level in atom</th>
<th>Maximum number of electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>2</td>
</tr>
<tr>
<td>2</td>
<td>8</td>
</tr>
<tr>
<td>3</td>
<td>18</td>
</tr>
<tr>
<td>4</td>
<td>32</td>
</tr>
</tbody>
</table>

Lesson Lead:
What is a model? What are some examples of using models in a science class?

Materials:
• white paper
• glue
• colored dots (can be made using a hole-puncher to create dots from peel off labels)
• atom template

Procedure:
1. Make a table like the one shown below. List the elements with atomic numbers of 1 through 20. Determine the number of each kind of particle needed to make up an atom of the element. Write your results in the table.

2. Working in teacher assigned groups, use the atom template pattern provided to create a model of the assigned elements. Use a marker or pen to write the number of protons and neutrons in the center circle. This represents the nucleus of the atom. The number of neutrons (n) is equal to the atomic number (number of protons, p) subtracted from the mass number (M):

\[ n = M - p \]

3. In the outer circles, arrange the colored dots to represent the electrons within the energy levels around the nucleus. Use as many of these electrons as you need for your element. Paste or stick the electrons in their places pairing the electrons to represent filled orbitals.

4. Repeat steps 2 and 3 for the other elements in your assigned group.

Data Table:

<table>
<thead>
<tr>
<th>Element name</th>
<th>Number of protons</th>
<th>Number of neutrons</th>
<th>Number of electrons</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Etc.

Analysis and Conclusions:
1. In a neutral atom, which particles will always be present in equal numbers?

2. Hypothesize what you think would happen to the charge of an atom if one of the electrons were removed from the atom.

3. Except for hydrogen, how many first-level electrons did each atom contain?

4. How is the Bohr model of an atom similar and different to an actual atom?

5. What would you do to make a better model of an atom? Explain using relative sizes of each subatomic particle and location.
IMPORTANT: Cut 1 inch outside of the outer circle. Label with pencil the name of the atom in the back and save these atomic models for Lab Activity entitled Periodic Trends.
Isotopes
(Adapted from “Modeling Isotopes” found in Glencoe Lab Manual p. 197-200)

Benchmarks:
SC.A.2.4.5- The student knows that elements are arranged into groups and families based on similarities in electron structure and that their physical and chemical properties can be predicted.

SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.E.2.4.6, SC.E.2.4.7, SC.H.1.2.1, SC.H.1.2.2, and SC.H.2.4.2)

SC.H.2.4.1- The student knows that scientists control conditions in order to obtain evidence, but when that is not possible for practical or ethical reasons, they try to observe a wide range of natural occurrences to discern patterns.

Objectives
• Determine the isotopic composition of 100 pennies.
• Apply the lessons of the penny-isotope analogy to isotopic data.

Background Information:
The defining characteristic of an atom of a chemical element is the number of protons in its nucleus. A given element may have different isotopes, which are nuclei with the same numbers of protons but different numbers of neutrons. For example, 12C and 14C are two isotopes of carbon. The nuclei of both isotopes contain six protons. However, 12C has six neutrons, whereas 14C has eight neutrons. In general, it is the number of protons and electrons that determines chemical properties of an element. Thus, the different isotopes of an element are usually chemically indistinguishable. These isotopes, however, have different masses.

Between 1962 and 1982, pennies were made of brass, which is an alloy composed of 95% copper and 5% zinc. In 1982, the rising price of copper led to a change in the composition of the penny. Beginning in 1982, pennies have been made of zinc plated with copper. These pennies contain 2.5% copper and 97.5% zinc. In this experiment, the two different types of pennies will represent two isotopes of an element.

Lesson Lead:
What is the isotopic composition of a collection of 100 pennies?

Pre-Lab:
1. What is an isotope?
2. The average atomic mass of the atoms of an element is what is known as a weighted average. In a weighted average, the value of each type of item is multiplied by the number of that type of item. The products are added, and the sum is divided by the total number of items. Use weighted average to solve the following problem: If you have four quarters, five dimes, and nine pennies, what is the average value of the coins? Describe the procedure. Then calculate the answer.
3. Explain how the two different types of pennies are analogous to isotopes of an element.
4. Read the entire laboratory activity. Make a flow chart of the procedure you will follow.
Materials:
- 100 pennies
- balance
- large box top
- 2 plastic cups

Safety Precautions:
- Always wear safety goggles and a lab apron in the lab.

Procedure:
1. Use the plastic cups and large box top to contain pennies so they do not get lost.
2. Measure the mass of ten pre-1982 pennies to the nearest 0.01 g. Record your measurement in Data Table 1. Repeat for post-1982 pennies.
3. Using your data from step 1, calculate the average mass of one pre-1982 penny. Record this average mass in Data Table 1. Repeat for a post-1982 penny.
4. Obtain 100 pennies. Find the mass of the sample to the nearest 0.01 g. Record your measurement in Data Table 2.
5. Divide the sample of 100 pennies into pre-1982 and post-1982 pennies. Record the numbers of each in Data Table 2.
5. Cleanup and Disposal- Follow your teacher's instructions for returning the coins.

Data and Observations:

<table>
<thead>
<tr>
<th>Mass of Pennies (g)</th>
<th>Pennies</th>
<th>Mass (g)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>10 pre-1982</td>
<td></td>
</tr>
<tr>
<td></td>
<td>10 post-1982</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Avg.1 pre-1982 penny</td>
<td></td>
</tr>
<tr>
<td></td>
<td>Avg. 1 post-1982 penny</td>
<td></td>
</tr>
</tbody>
</table>

Data Table 2

<table>
<thead>
<tr>
<th>Data for 100-penny sample</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of 100 pennies (g)</td>
</tr>
<tr>
<td>Number of pre-1982 pennies in 100-penny sample</td>
</tr>
<tr>
<td>Number of post-1982 pennies in 100-penny sample</td>
</tr>
<tr>
<td>Average mass of a penny in 100-penny sample (g)</td>
</tr>
</tbody>
</table>

Analysis and Conclusions:
1. In Procedure step 1, why did you measure the mass often pennies instead of the mass of one penny?
2. Divide the mass of 100 pennies in Data Table 2 by 100 to find the average mass. Record your answer in Data Table 2.
3. Using the mass of pre-1982 and post-1982 pennies from Data Table 1 and the number of each type of penny from Data Table 2, calculate the average mass of a penny in the 100-penny sample. How does your answer compare to the average value calculated in question 2?
4. How is the value you calculated in question 3 analogous to the atomic mass of the atoms in a sample of an element?
5. Calculate the theoretical mass of a pre-1982 penny and a post-1982 penny.
6. The density of copper is 8.96 g/cm³, and that of zinc is 7.13 g/cm³. Using the compositions given in the introduction, the density of a pre-1982 penny is $(0.95)(8.96 \text{ g/cm}³) + (0.05)(7.13 \text{ g/cm}³) = 8.87 \text{ g/cm}³$. Calculate the density of a post-1982 penny.
7. A typical penny has a diameter of 1.905 cm and a thickness of 0.124 cm. What is the volume in cm³ of a typical penny? Hint: $V = (\pi \times r^2)(\text{thickness of penny}).$
8. Using the density and volume values from questions 1 and 2, calculate the theoretical mass of a pre-1982 penny and the mass of a post-1982 penny.
9. Data Table 3 shows the isotopic mass and relative abundance for the most common isotopes of copper and zinc.
   a. How many protons and neutrons are there in a $^{64}\text{Cu}$ nucleus?
   b. How many protons and neutrons are there in a nucleus of $^{64}\text{Zn}$?

Data Table 3

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Atomic number</th>
<th>Mass number</th>
<th>Isotopic mass (amu)</th>
<th>Relative abundance (%)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Copper-63</td>
<td>29</td>
<td>63</td>
<td>62.9298</td>
<td>69.09</td>
</tr>
<tr>
<td>Copper-64</td>
<td>29</td>
<td>64</td>
<td>64.9278</td>
<td>30.91</td>
</tr>
<tr>
<td>Zinc-64</td>
<td>30</td>
<td>64</td>
<td>63.9291</td>
<td>48.89</td>
</tr>
<tr>
<td>Zinc-66</td>
<td>30</td>
<td>66</td>
<td>65.9260</td>
<td>27.81</td>
</tr>
<tr>
<td>Zinc-67</td>
<td>30</td>
<td>67</td>
<td>66.9271</td>
<td>4.73</td>
</tr>
<tr>
<td>Zinc-68</td>
<td>30</td>
<td>68</td>
<td>67.9249</td>
<td>18.57</td>
</tr>
</tbody>
</table>

10. Use the data in Data Table 3 and answer the following questions.
   a. Calculate the atomic mass of copper.
   b. Calculate the atomic mass of zinc.
11. Use the values from Data Table 1 and the answers from question 10 to calculate the following.
   a. How many atoms of copper are in a pre-1982 penny? (Hint: Use Avogadro's number.)
   b. How many atoms of zinc are in a pre-1982 penny?
   c. How many total atoms (copper and zinc) are in a pre-1982 penny?
   d. How many total atoms (copper and zinc) are in a post-1982 penny?

Real-World Chemistry:
1. A nuclear power plant that generates 1000 MW of power uses 3.2 kg per day of 235U.
2. Naturally occurring uranium contains 0.7% 235U and 99.7% 238U. What mass of natural uranium is required to keep the generator running for a day?
Half-Life
(Adapted from Glencoe, Physical Science)

Benchmark:
SC.A.2.4.3 The student knows that a number of elements have heavier, unstable nuclei that
decay, spontaneously giving off smaller particles and waves that result in a small loss of mass
and release a large amount of energy. CS

Objective:
• Make a model that illustrates the half-life of an imaginary isotope.
• Graph and interpret data of the isotope's half-life.
• Compare individual group data with average class data.

Background Information:
Isotopes are atoms of the same element with different atomic masses. These different masses are
a result of having different numbers of neutrons in their nuclei. Isotopes can be stable or
unstable (radioactive). Radioactive isotopes have unstable nuclei that break down in a process
called radioactive decay. During this process, the radioactive isotope is transformed into
another, usually more stable, element. The amount of time it takes half the atoms of a
radioactive isotope in a particular sample to change into another element is its half-life. A half-
life can be a fraction of a second for one isotope or more than a billion years for another isotope,
but it is always the same for any particular isotope.

Lesson Lead:
How can pennies be used to simulate nuclear decay?

Materials:
• 100 pennies
• plastic cup
• container with lid such as a shoebox
• timer or clock with second hand

Procedure:
1. Place 100 pennies, each head-side up, into the container. Each penny represents an atom of
an unstable isotope.
2. Place the lid securely on the container. Holding the container level, shake it vigorously for
20 seconds.
3. Set the container on the table and remove the lid. Remove only pennies that are now in a
tails-up position (decayed nuclei).
4. Count the pennies you removed and record this number in Table 1 under Trial 1. Also record
the number of heads-up pennies that are left (undecayed nuclei).
5. Repeat steps 2 through 4 until there are no pennies left in the container.
6. Repeat steps 1 through 5 and record your data in Table 1 under Trial 2.

Data Analysis
1. Calculate the averages for each time period and record these numbers in Table 1.
2. Graph the average data from Table 1. Graph the number of heads-up pennies remaining on the Y-axis (undecayed atoms or nuclei) against time (on X-axis) using a symbol such as X, O, Δ, etc.
3. Copy the averages from Table 1 into Table 2 under Group I.
4. Record the averages obtained by other groups in your class in Table 2.
5. Determine the totals and then averages for the combined data from all groups and record in Table 2.
6. Graph the average class data for undecayed nuclei on your graph using a different symbol in the same way as you graphed your individual group's data. Be sure to include symbol key for the two sets of data.

Table 1: Group Data

<table>
<thead>
<tr>
<th>Shaking Time</th>
<th>Trial 1</th>
<th>Trial 2</th>
<th>Averages</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Number of Heads-up Remaining</td>
<td>Number of Tails-up Removed</td>
<td>Number of Heads-up Remaining</td>
</tr>
<tr>
<td>At 0 s</td>
<td>100</td>
<td>0</td>
<td>100</td>
</tr>
<tr>
<td>After 20 s</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>After 40 s</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>After 60 s</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>After 80 s</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>After 100 s</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>After 120 s</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>After 140 s</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Table 2: Class Data

<table>
<thead>
<tr>
<th>Group</th>
<th>I</th>
<th>II</th>
<th>III</th>
<th>IV</th>
<th>V</th>
<th>VI</th>
<th>VII</th>
<th>VIII</th>
<th>Total</th>
<th>Average</th>
</tr>
</thead>
<tbody>
<tr>
<td>0 s H</td>
<td>100</td>
<td>100</td>
<td>100</td>
<td>100</td>
<td>100</td>
<td>100</td>
<td>100</td>
<td>100</td>
<td>800</td>
<td>100</td>
</tr>
<tr>
<td>H T</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
<td>0</td>
</tr>
<tr>
<td>20 s H</td>
<td></td>
<td></td>
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<td></td>
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<td></td>
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<td></td>
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<tr>
<td>H T</td>
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<td>40 s H</td>
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<td>60 s H</td>
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<tr>
<td>80 s H</td>
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<tr>
<td>100 s H</td>
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<tr>
<td>H T</td>
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<tr>
<td>120 s H</td>
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<td>H T</td>
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<tr>
<td>140 s H</td>
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<td></td>
<td></td>
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</tr>
<tr>
<td>H T</td>
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<td></td>
<td></td>
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<td></td>
<td></td>
</tr>
</tbody>
</table>

H = heads; T = tails
Conclusion:
1. In this model, what represented the process of radioactive decay?
2. Which side of the penny represented the decayed isotope? Which side represented the undecayed atoms?
3. In this model, what was the half-life of the pennies? Explain.
4. What can you conclude about the total number of atoms that decay during any half-life period of the pennies?
5. Why were more accurate results obtained when the data from all groups was combined and graphed?
6. If your half-life model had decayed perfectly, how many atoms of the radioactive isotope should have been left after 80 seconds of shaking?
7. If you started with 256 radioactive pennies, how many would have remained undecayed after 60 seconds of shaking?
Flame Tests

Benchmarks:
SC.A.2.4.6 - The student understands that matter may act as a wave, a particle, or something else entirely different with its own characteristic behavior. CS

SC.B.1.4.4.- The student knows that as electrical charges oscillate they create time-varying electric and magnetic fields that propagate away from the source as an electro-magnetic wave.

Objective:
- Observe the spectra emitted by different ions.
- Identify the metallic ions by the color emitted

Background Information:
Flame tests provide a way to test for the presence of specific metallic ions. The heat of the flame excites the loosely-held electrons in the metal ion prompting the electrons to jump from a ground level to a higher energy level in the atom. As the electrons fall back to their ground state, energy is released from the excited electrons and can be seen as a colored flame. The color is a combination of the wavelengths of each transition and can be used to determine the identity of an unknown ion. Although white light produces a continuous spectrum in which all wavelengths of visible light are present (400-700nm), an excited electron produces one or more specific lines in the spectrum. This unique spectrum corresponds to an element’s distinct electron configuration.

Lesson Lead:
What colors are characteristic of particular metallic ions in a flame test?

Pre-Lab:
1. What precautions should you take when working with an open flame?
2. Is the color of the flame, a chemical or a physical property of these metals? Explain.
3. What is the purpose of the Bunsen burner in this experiment?
4. Which are the electrons that will become excited?
5. List the colors of the visible spectrum in order of increasing wavelength.
6. What is meant by the term frequency of a wave? What are the units of frequency? Describe the relationship between frequency and wavelength

Materials: (per lab group)
- lab goggles
- lab apron
- Bunsen burner
- nichrome wire loop
- wash bottle with distilled water
- well plate
- hydrochloric acid (HCl) 6.0 M in a dropper bottle
- solutions of metallic salts:
  - calcium nitrate (Ca(NO₃)₂)
  - copper nitrate (Cu(NO₃)₂)
  - lithium nitrate (LiNO₃)
  - potassium nitrate (KNO₃)
  - strontium nitrate (Sr(NO₃)₂)
  - sodium chloride (NaCl)
  - unknown solution

Safety Precautions:
- Always wear safety goggles and a lab apron in the lab.
- Tie back hair and secure loose bulky clothing while working near a flame.
- All of the salt solutions (except NaCl) are toxic, if any of the solutions splash on your skin, wash affected area with large amounts of water and notify your teacher.
- Handle carefully the heated nichrome wire. Make sure to hold only from the handle.
- Hydrochloric acid is corrosive to the skin and clothing and the vapors are irritating to the lungs and the eyes. Avoid contact with the solution and inhalation of its vapors. If acid splashes on your skin or clothing wash immediately with water and notify your teacher.

**Note:** This experiment should be done in a Chemistry lab room with appropriate ventilation. If this is not available, the teacher may opt to perform the experiment as a demonstration.

**Procedure:**
1. Put on your goggles and lab apron, and secure your hair.
2. Obtain a well plate and place it on a white sheet of paper. Label each well with the name of the solutions to be tested including a well for the unknown solution. Put a dropperful of each known solution into its corresponding well.
3. Clean the nichrome wire before testing each solution. Rinse the loop with distilled water followed by ringing with the 6.0 M HCl. Place the loop into the flame for about a minute. Observe the color of the clean nichrome wire in the flame; this is the color you should see after you clean the wire for each new test.
4. Dip the nichrome wire into the well with the calcium nitrate (Ca(NO₃)₂) solution and immediately place the loop into the flame. Observe and record the color of the flame in the Data Table. Repeat each test 2 or 3 times before trying a new solution. Clean the nichrome wire between each test (see step 3). Continue testing and recording the flame colors produced by the metallic ions of each solution.
5. Sodium has a very strong color which could affect the color results of your other tests. To prevent this from happening follow proper nichrome wire cleaning procedure and leave the sodium test for last.
6. Obtain an unknown from your teacher and repeat steps 3 and 4.

**Data Table: Flame Test Results**

<table>
<thead>
<tr>
<th>Compound in Solution</th>
<th>Flame Color</th>
</tr>
</thead>
<tbody>
<tr>
<td>calcium nitrate (Ca(NO₃)₂)</td>
<td></td>
</tr>
<tr>
<td>copper nitrate (Cu(NO₃)₂)</td>
<td></td>
</tr>
<tr>
<td>lithium nitrate (LiNO₃)</td>
<td></td>
</tr>
<tr>
<td>potassium nitrate (KNO₃)</td>
<td></td>
</tr>
<tr>
<td>strontium nitrate (Sr(NO₃)₂)</td>
<td></td>
</tr>
<tr>
<td>sodium chloride (NaCl)</td>
<td></td>
</tr>
<tr>
<td>unknown solution</td>
<td></td>
</tr>
</tbody>
</table>

**Analysis and Conclusions:**
1. What particles are found in the chemicals that may be responsible for the production of colored light?
2. Why do different metals have different characteristic flame test colors?
3. The majority of the known compounds tested contain nitrate, yet the colors of the flames were different. What effect would you expect on the flame colors if these compounds were chlorides instead of nitrates?
4. What color did your unknown produce in the flame? What is your unknown?
5. What would be another way of exciting the electrons without using a Bunsen burner?

**Real-World Chemistry:**
1. A firework contains copper chloride and strontium sulfate. What colors would you expect to be produced?
2. When a pan of milk boils over onto the stove the flame turns red-orange. Explain why.

**Extension:**
1. Calculate the wavelength of the yellow light emitted by a sodium lamp, in meters, if the frequency of the radiation is $5.10 \times 10^{14}$/sec.
2. What is the energy of a photon of green light whose frequency is $6.85 \times 10^{14}$/sec?
3. The spectrum of lithium has a red line of 670.8 nanometers. (Hint: 1 meter = $1 \times 10^9$ nm).
   a. Convert the nanometers, using dimensional analysis to meters.
   b. Calculate the frequency of the wave.
   c. Calculate the energy of a photon with this wavelength.
4. How is spectrometry used to determine the composition of stars?
Periodic Trends

**Benchmarks:**

SC.A.2.4.5- The student knows that elements are arranged into groups and families based on similarities in electron structure and that their physical and chemical properties can be predicted.

SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.E.2.4.6, SC.E.2.4.7, SC.H.1.2.1, SC.H.1.2.2, and SC.H.2.4.2)

SC.H.2.4.1- The student knows that scientists assume that the universe is a vast system in which basic rules exist that may range from very simple to extremely complex, but that scientists operate on the belief that the rules can be discovered by careful, systematic study.

SC.H.3.4.1- The student knows that performance testing is often conducted using small-scale models, computer simulations, or analogous systems to reduce the chance of system failure. CS

**Objective:**

- Determine the periodic trends per group of main elements.
- Relate the reactivity of elements to their electron structure.

**Background Information:**

Elements within the same family in the Periodic Table have similar properties and in many cases, one element can be replaced by another element with similar properties. This knowledge is presently being widely used in medicine with radioactive tracers that are being used more frequently to locate tumors in the body and deliver medications for their treatment. This knowledge is also being widely studied in external pollutants and contaminants that replace essential elements within the human body, such as the replacement of calcium in bones by radioactive strontium.

**Lesson Lead:**

How do periodic trends relate to atomic structure?

**Materials:** (8 Periodic Table groups to 8 Lab group)

- Previous models from Models of Atomic Structure Lab
- Colored construction paper (8 different colors for class activity)
- Glue
- Colored markers
- String or wire

**Procedure:**

1. Paste the previous models from the Models of Atomic Structure lab on your assigned colored papers coded for your particular family (1 color of construction paper per Periodic Table group).
2. Using markers write the name, symbol and electron configuration on your construction paper. See sample provided.
3. Research the elements in your assigned Periodic Table group. Find the physical and chemical properties of each element, their practical uses and similarities with any other element in the Periodic Table. If possible obtain pictures of the elements assigned.
4. Add a summary of this information to your construction paper. See sample provided.
5. Create two additional squares for your assigned Periodic Table family. On the first square, label the group number and family name; this will be added as a title to your Periodic Table family.
6. Use the second square to summarize the properties of the elements in your Periodic Table group; this will be attached below your column of elements.
7. Working with other groups, arrange your models according to the order of the Periodic Table. Be sure to include your group title above and group properties below your elements. You can secure each model to the others with string or wire to create a large class Periodic Table.
8. Hang your Periodic Table on the wall and use for future reference.

Analysis and Conclusions:
1. Your lab group will prepare a presentation to the class highlighting your assigned group of elements. As each group presents, take notes on a blank Periodic Table.
2. What can you conclude about the arrangement of valence electrons in each group?
3. How do the chemical and physical properties of an element relate to its electron configuration? Describe any patterns you see.
4. Analyze the information provided in the table showing the energy necessary to remove the first electron (first ionization energy) from several elements.
5. Plot the points of the ionization energy (y axis) vs. the atomic number (x axis) and label each point with the element symbol in the graph provided below.
6. Connect with different color lines the points corresponding to the elements within the same family or group.
7. Based on your graph, which group of elements are the most reactive? Where are they found in the periodic table? How does this relate to their valence electron structure shown in your class created Periodic Table?
8. Which group has the least reactive elements? How does this relate to their valence electron structure?
9. Would you predict the second ionization energy to be higher or lower than the first? Explain.

<table>
<thead>
<tr>
<th>ELEMENT</th>
<th>Z</th>
<th>First Ionization Energy (eV)</th>
</tr>
</thead>
<tbody>
<tr>
<td>He</td>
<td>2</td>
<td>24.6</td>
</tr>
<tr>
<td>Li</td>
<td>3</td>
<td>5.4</td>
</tr>
<tr>
<td>Be</td>
<td>4</td>
<td>9.3</td>
</tr>
<tr>
<td>F</td>
<td>9</td>
<td>17.4</td>
</tr>
<tr>
<td>Ne</td>
<td>10</td>
<td>21.5</td>
</tr>
<tr>
<td>Na</td>
<td>11</td>
<td>5.1</td>
</tr>
<tr>
<td>Mg</td>
<td>12</td>
<td>7.6</td>
</tr>
<tr>
<td>Cl</td>
<td>17</td>
<td>13.0</td>
</tr>
<tr>
<td>Ar</td>
<td>18</td>
<td>15.7</td>
</tr>
<tr>
<td>K</td>
<td>19</td>
<td>4.3</td>
</tr>
<tr>
<td>Ca</td>
<td>20</td>
<td>6.1</td>
</tr>
<tr>
<td>Br</td>
<td>35</td>
<td>11.8</td>
</tr>
<tr>
<td>Kr</td>
<td>36</td>
<td>14.0</td>
</tr>
<tr>
<td>Rb</td>
<td>37</td>
<td>4.2</td>
</tr>
<tr>
<td>Sr</td>
<td>38</td>
<td>5.7</td>
</tr>
</tbody>
</table>

Real-World Chemistry:
1. Your science teacher has limited amounts of bromine for a particular experiment? What is one element that might be used as a replacement?
<table>
<thead>
<tr>
<th>Element Name</th>
<th>Element Symbol</th>
<th>Atomic Number</th>
<th>Atomic Mass</th>
</tr>
</thead>
</table>

**Note:** 1 color of construction paper per Periodic Trend Group

**Noble gas notation**

**Main physical and chemical properties**
Bonding
(Adapted from Living by Chemistry Unit 1 Alchemy Preliminary Edition, General Chemistry)

Benchmarks:
SC.A.2.4.5- The student knows that elements are arranged into groups and families based on similarities in electron structure, and that their physical and chemical properties can be predicted.

SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.H.1.2.1, SC.H.1.2.2, SC.H.2.4.2, SC.E.2.4.6, and SC.E.2.4.7)

Objective:
• Test different substances for their ability to dissolve and conduct an electric current.
• Predict the bonding nature of each substance on the basis of your results.

Background Information:
Bond types can be predicted on the basis of the ability of substances to dissolve and conduct electricity. **Solubility** is a physical property of matter that can be defined as the amount of a substance that can be dissolved in a specified amount of solvent. **Conductivity**, another physical property, describes the ability of a substance to conduct an electric current. An electric current can be produced by the movement of positive or negative particles. In a solid, such as a wire, only the negative charges (free electrons) are able to move. In this lab you will construct a conductivity apparatus which will light up when an electrical circuit is completed.

**Ionic bonds** occur between metals and nonmetals when valence electrons are transferred from the metal to the nonmetal. Most ionic compounds are soluble in water and conduct electricity once dissolved when the ions are free to move in the water. Ionic compounds exhibit high melting and boiling points. **Covalent bonds** occur between two nonmetals by a sharing of valence electrons. Molecular compounds (held together by covalent bonds) do not conduct electricity. Some molecular compounds dissolve in water while others do not. **Metallic bonds** hold metal atoms together by sharing mobile valence electrons that resemble a “sea of electrons.” Metal substances generally do not dissolve in water, but all conduct electricity. Some substances consist of nonmetal atoms joined in a large **covalent network** (essentially gigantic molecules). The majority of such covalent network molecules do not dissolve in water and none conduct electricity.

Lesson Lead:
How can the properties of substances be used to predict the way atoms are held together?

Materials: **Conductivity Apparatus***:
A) tape
B) 9-V battery
C) battery clip
D) bare wire leads
E) resistor
F) LED or buzzer
G) wood backing.(tongue depressor)
Tools to make Conductivity Apparatus*:
• Scissors
• Wire stripper
• Tape solder
• 1 nail
• Pliers
• Matches, Candle or soldering iron


Note: Teacher may opt to use commercially available conductivity testers. Materials for construction of student-made apparatus can be obtained at local hobby/electrical supply stores.

Other:
• small beakers or cups (100 mL or less)
• distilled water (~100 mL)
• wash bottle filled with distilled water
• stirring rod
• 1 cm² of Al foil
• a penny
• rubbing alcohol (isopropyl alcohol) (10mL)

Approximately 1 g of each of the following:
• sucrrose (table sugar)
• NaCl
• SiO₂ (sand)
• paraffin wax shavings
• CaCl₂ (calcium chloride)
• CuSO₄ (copper II sulfate)

Safety Precautions:
• Always wear safety goggles and a lab apron.
• Use conductivity tester only for described activities.

Pre-Lab:
• Predict the conductivity and solubility properties of the substances listed in Table 1, filling in columns 2 and 3 of the table with brief explanations.

Procedure:
1. Make conductivity apparatus: See above diagram for how to construct apparatus.
2. Conductivity- Use the apparatus to test the conductivity of all substances before making any solutions. Be sure to keep the wire test probes apart. Record your results in the second column of Table 2. Rinse the paper clip probes with distilled water between each test.
3. Solubility- Take ~1 g of each solid substance, place it in the small beaker or cup and try to dissolve it with ~10 mL of distilled water. Record your results in the third column of Table 2.
4. **Conducts when dissolved** - If the substance dissolves in water, test the solution with the conductivity apparatus. Make sure to keep the probes apart. Enter your results in the final column of Table 2. If a substance did not dissolve enter NO in this column.

<table>
<thead>
<tr>
<th>Substances</th>
<th>Conducts?</th>
<th>Dissolves?</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Yes/No</td>
<td>Yes/No</td>
</tr>
<tr>
<td><strong>Al (s) Aluminum foil</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>C\textsubscript{12}H\textsubscript{22}O\textsubscript{11}(s) Sucrose (sugar)</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>C\textsubscript{20}H\textsubscript{42} (s) paraffin (wax)</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>C\textsubscript{3}H\textsubscript{8}O (l) isopropyl alcohol</strong></td>
<td></td>
<td>Not Applicable</td>
</tr>
<tr>
<td><strong>CaCl\textsubscript{2} (s) calcium chloride</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Cu (s) Copper</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>CuSO\textsubscript{4} (s) Copper sulfate</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>H\textsubscript{2}O (l) Distilled water</strong></td>
<td></td>
<td>Not Applicable</td>
</tr>
<tr>
<td><strong>NaCl (s) salt, Sodium chloride</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>SiO\textsubscript{2} (s), sand, Silicon dioxide</strong></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Table 2: Test Results

<table>
<thead>
<tr>
<th>Substances</th>
<th>Original substance conducts?</th>
<th>Substance dissolves?</th>
<th>Solution conducts?</th>
<th>Were Table 1 predictions correct?</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Yes/No</td>
<td>Yes/No</td>
<td>Yes/No</td>
<td>Yes/No</td>
</tr>
<tr>
<td>Al (s) Aluminum foil</td>
<td>Yes</td>
<td>Yes</td>
<td>Yes</td>
<td>No</td>
</tr>
<tr>
<td>C₁₂H₂₂O₁₁(s) Sucrose (sugar)</td>
<td>Yes</td>
<td>Yes</td>
<td>No</td>
<td>No</td>
</tr>
<tr>
<td>C₂₀H₄₂ (s) paraffin (wax)</td>
<td>Yes</td>
<td>Yes</td>
<td>Yes</td>
<td>Yes</td>
</tr>
<tr>
<td>C₃H₈O (l) isopropyl alcohol</td>
<td>Not Applicable</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CaCl₂ (s) calcium chloride</td>
<td>Yes</td>
<td>Yes</td>
<td>No</td>
<td>No</td>
</tr>
<tr>
<td>Cu (s) Copper</td>
<td>No</td>
<td>Yes</td>
<td>No</td>
<td>No</td>
</tr>
<tr>
<td>CuSO₄ (s) Copper sulfate</td>
<td>Yes</td>
<td>Yes</td>
<td>Yes</td>
<td>Yes</td>
</tr>
<tr>
<td>H₂O (l) Distilled water</td>
<td>Not Applicable</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaCl (s) salt, Sodium chloride</td>
<td>Yes</td>
<td>Yes</td>
<td>Yes</td>
<td>Yes</td>
</tr>
<tr>
<td>SiO₂ (s), sand, Silicon dioxide</td>
<td>Yes</td>
<td>Yes</td>
<td>Yes</td>
<td>Yes</td>
</tr>
</tbody>
</table>

Analysis

1. Were most of your solubility and conductivity predictions correct? Discuss.
2. Was the intensity of the LED light the same for all conducting substances? Explain.
3. Study your data from Table 2 and group your substances using the diagram on the next page. Write the names of the substances in the first tier of boxes corresponding box to their ability to dissolve. Then classify those substances into the four boxes in the bottom tier (labeled A, B, C, D) to help you group them according to their properties.
**Conclusion**

1. What do the substances that conduct electricity but do not dissolve in water have in common? (Explain).
2. What do the substances that dissolved and conduct electricity have in common? (Explain)
3. What do the substances that dissolved but do not conduct electricity have in common? (Explain)
4. Write a generalization about the substances that did not conduct electricity.
5. Write a generalization about the substances that did conduct electricity.
6. Classify the substances that you tested according to the four ways atoms are held together. What would you call the substances in box A? B? C? D? Hint: Reread the Background Information.
Activities of Metals
(Adapted from Glencoe textbook)

Benchmarks:
SC.A.1.4.1-The student knows that the electron configuration in atoms determines how a substance reacts and how much energy is involved in its reactions. CS

SC.A.1.4.5- The student knows that connections (bonds) form between substances when outer-shell electrons are either transferred or shared between their atoms, changing the properties of substances. CS

Objectives:
• Observe chemical reactions
• Sequence the activities of some metals
• Predict if reactions will occur between certain substances

Background Information:
Some metals are more reactive than others. By comparing how different metals react with the same ions in aqueous solutions, an activity series for the tested metals can be developed. The activity series will reflect the relative reactivity of the tested metals. It can be used to predict whether reactions will occur.

Lesson Lead:
Which is the most reactive metal tested? Which is the least reactive metal tested? Can this information be used to predict whether reactions will occur?

Pre-Lab:
1. Read the entire CHEMLAB.
2. Make notes about procedures and safety precautions to use in the laboratory.
3. Prepare your data table.
4. Form a hypothesis about what reactions will occur.
5. What are the independent and dependent variables?
6. What gas is produced when magnesium and hydrochloric acid react? Write the chemical equation for the reaction.
7. Why is it important to clean the magnesium ribbon? How might not polishing a piece of metal affect the reaction involving that metal?

Materials:
• 1.0M Zn(NO₃)₂
• 1.0M Al(NO₃)₃
• 1.0M Cu(NO₃)₂
• 1.0M Mg(NO₃)₂
• pipettes (4)
• wire cutters

• Cu wire
• Al wire
• Mg ribbon
• Zn metal strips (4)
• emery cloth or fine sandpaper
• 24-well microscale reaction plate

Safety Precautions:
• Always wear safety goggles and a lab apron.
• Use caution when using sharp and coarse equipment
Procedure:
1. Use a pipette to fill each of the four wells in column 1 of the reaction plate with 2 mL of 1.0M Al(NO₃)₃ solution.
2. Repeat the procedure in step 1 to fill the four wells in column 2 with 2 mL of 1.0M Mg(NO₃)₂ solution.
3. Repeat the procedure in step 1 to fill the four wells in column 3 with 2 mL of 1.0M Zn(NO₃)₂ solution.
4. Repeat the procedure in step 1 to fill the four wells in column 4 with 2 mL of 1.0M Cu(NO₃)₂ solution.
5. With the emery paper or sandpaper, polish 10 cm of aluminum wire until it is shiny. Use wire cutters to cut the aluminum wire into four 2.5-cm pieces. Place a piece of the aluminum wire in each row A well that contains solution.
6. Repeat the procedure in step 5 using 10 cm of magnesium ribbon. Place a piece of the Mg ribbon in each row B well that contains solution.
7. Use the emery paper or sandpaper to polish small strips of zinc metal. Place a piece of Zn metal in each row C well that contains solution.
8. Repeat the procedure in step 5 using 10 cm of copper wire. Place a piece of Cu wire in each row D well that contains solution.
9. Observe what happens in each cell. After five minutes, record your observations on the data table you made.

Cleanup and Disposal:
1. Dispose of all chemicals and solutions as directed by your teacher.
2. Clean your equipment and return it to its proper place.
3. Wash your hands thoroughly before you leave the lab.

Analysis and Conclusions:
1. In which wells of the reaction plate did chemical reactions occur? Which metal reacted with the most solutions? Which metal reacted with the fewest solutions? Which metal is the most reactive?
2. The most active metal reacted with the most solutions. The least active metal reacted with the fewest solutions. Order the four metals from the most active to the least active.
3. Compare your activity series with the activity series shown here. How does the order you determined for the four metals you tested compare with the order of these metals?
4. Write a chemical equation for each single-replacement reaction that occurred on your reaction plate.
5. Use the diagram below to predict if a single-replacement reaction will occur between the following reactants. Write a chemical equation for each reaction that will occur.
   a. Ca and Sn(NO₃)₂
   b. Ag and Ni(NO₃)₂
   c. Cu and Pb(NO₃)₂
6. If the activity series you sequenced does not agree with the order in the diagram below, propose a reason for the disagreement.
7. What is the general location of the most active metals (top, bottom, left, or right)? Least active?
Real-World Chemistry:

1. Based on what you learned in this lab, compare copper, silver, platinum and gold as to their value and use in jewelry.
Determining Reaction Rates

Benchmarks:
SC.A.1.4.4- The student experiments and determines that the rates of reaction among atoms and molecules depend on the concentration, pressure, and temperature of the reactants and the presence or absence of catalysts.

SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.E.2.4.6, SC.E.2.4.7, SC.H.1.2.1, SC.H.1.2.2, and SC.H.2.4.2)

SC.H.2.4.1- The student knows that scientists control conditions in order to obtain evidence, but when that is not possible for practical or ethical reasons, they try to observe a wide range of natural occurrences to discern patterns.

Objective:
• Determine the factors that affect the rates of reactions.

Background Information:
Chemical reactions occur when the molecules of two or more elements or compounds, called the reactants, collide and recombine to form a new compound, which is called the product. According to the collision theory, these colliding molecules must first reach the reaction's activation energy for the reaction to occur. Activation energy is the level of energy required for the molecules to collide with enough force to recombine and form a new product.

The rate of reaction describes how fast reactants form products in a chemical reaction. Chemical reactions can be sped up or slowed down by altering the surface area, concentration, and temperature of the reactants.

Lesson Lead:
What effect does a change of concentration have on the collisions of reacting molecules?

Materials:
• Three 250-ml beakers
• Two graduated cylinders
• Stopwatch
• Stirring rods
• Distilled water at room temperature
• 120 ml of Solution A
• 90 ml of Solution B

Safety Precautions:
• Always wear safety goggles and a lab apron
• Wear gloves when handling solutions A & B
Teacher Preparation

Solution A:
Dissolve 4.3 g of potassium iodate (KIO₃) per liter of water.

Solution B:
Make a paste of 4 g soluble starch in a small amount of water. Slowly add paste to 900 mL of boiling water. Boil for several minutes and allow solution to cool. Just before using, add 0.2 g of sodium thiosulfate Na₂S₂O₅ and 5 mL of 1 molar H₂SO₄ and add water to bring the final volume to 1 liter of solution.

Note: Be sure to estimate how many liters you will need according to the total number of students or groups that will be using solutions.

Procedure:
1. Measure 60 mL of Solution A and pour into a 250 mL beaker. Add 10 mL of distilled water and stir.
2. Measure 30 mL of Solution B and pour into another 250 mL beaker. Place this beaker on top of a white sheet of paper in order to see changes.
3. Add the 70 mL of Solution A to the Solution B. Stir rapidly. Record the time it takes for the reaction to occur.
4. Rinse the reaction beaker and repeat the steps listed above for the other amounts shown in the data table. Complete the data table.

Data Table:

<table>
<thead>
<tr>
<th>Solution A (mL)</th>
<th>Water (mL)</th>
<th>Solution B (mL)</th>
<th>Reaction Time (s)</th>
</tr>
</thead>
<tbody>
<tr>
<td>60</td>
<td>10</td>
<td>30</td>
<td></td>
</tr>
<tr>
<td>40</td>
<td>30</td>
<td>30</td>
<td></td>
</tr>
<tr>
<td>20</td>
<td>50</td>
<td>30</td>
<td></td>
</tr>
</tbody>
</table>

Analysis and Conclusions:
1. What visible indication can you describe that suggests that a chemical reaction is taking place?
2. Plot a graph of Time in seconds (y-axis) vs. Volume of Solution A (x-axis).
3. Describe the trend shown in your graph?
4. What happens when you add more water to Solution A?
5. What is kept constant during this experiment?
6. How does concentration affect the rate of reaction? Explain in terms of collisions.
Hydrated Crystals
(Adapted from Glencoe textbook p.342-343)

Benchmark:
SC.B.1.4.2-The student understands that there is conservation of mass and energy when matter is transformed. CS

Objectives
• Heat a known mass of hydrated compound until the water is removed.
• Calculate the experimental and theoretical percentages of water in the hydrate.
• Predict the empirical formula for magnesium sulfate hydrate.

Background Information:
Hydrates are crystalline compounds with water molecules incorporated in their structure. The ratio of moles of water to one mole of the compound is a small whole number which can be determined experimentally by heating the hydrate to remove the water. For example, in the hydrated compound copper(II) sulfate pentahydrate (CuSO_4·5H_2O), the ratio is 5:1. Once the water has been removed from a hydrated compound it has become anhydrous.

The percentage error between the experimental and theoretical percentages of water in a hydrate can be calculated using the following equation:

\[
\% \text{ error} = \frac{|\text{experimental value} - \text{theoretical value}|}{\text{theoretical value}}
\]

Lesson Lead: Can water be present in a solid compound?

Pre-Lab:
1. Read the entire laboratory directions.
2. Prepare all written materials that you will take into the laboratory. Be sure to include safety precautions, procedure notes, and a data table.
3. Explain how you will obtain the mass of water and the mass of anhydrous MgSO_4 contained in the hydrate from your data.
4. How will you convert the masses of anhydrous MgSO_4 and water to moles?
5. How can you obtain the formula for the hydrate from the moles of anhydrous MgSO_4 and the moles of water?

Materials:
• hotplate
• balance (preferably 2 decimal place)
• hydrated MgSO_4 (Epsom salts)
• 2 beakers (50 or 100 mL and 400 or 600 mL)
• hot mitts or beaker tongs

Safety Precautions:
• Always wear safety goggles and a lab apron.
• Hot objects will not appear to be hot.
Procedure:
1. Allow hot plate to warm up on the high setting.
2. Measure to the nearest 0.01 g the mass of a clean, dry 50 or 100 ml beaker and record.
3. Add about 3 g hydrated MgSO₄ to the beaker and measure the mass of the beaker plus hydrate to the nearest 0.01 g. Record the mass.
4. Record your observations of the solid hydrate.
5. Place the small beaker on the hot plate and cover with the large beaker.
6. After a few minutes write down what you observe happening inside the large beaker, and then carefully remove the large beaker while wearing a hot mitt.
7. Continue heating the small beaker and contents for a total of about 10 minutes.
8. Remove the small beaker from the hot plate, allow it to cool, and determine the mass of beaker plus contents. Record this mass.
9. Reheat the beaker and contents for 5 minutes, cool, and measure the mass again.
10. If these last two masses do not agree within 0.02 g, you should reheat the beaker and contents a third time.
11. Dispose of the MgSO₄ according to your teacher’s directions, and then clean the beakers.

<table>
<thead>
<tr>
<th>Table 1</th>
<th>Observations</th>
</tr>
</thead>
<tbody>
<tr>
<td>Solid Hydrated MgSO₄</td>
<td></td>
</tr>
<tr>
<td>Large beaker during initial heating</td>
<td></td>
</tr>
<tr>
<td>Solid Anhydrous MgSO₄</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Table 2</th>
<th>Mass Data</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass of small beaker</td>
<td></td>
</tr>
<tr>
<td>Mass of beaker + MgSO₄ hydrate</td>
<td></td>
</tr>
<tr>
<td>Mass MgSO₄ hydrate</td>
<td></td>
</tr>
<tr>
<td>Mass of beaker + anhydrous MgSO₄</td>
<td></td>
</tr>
<tr>
<td>Mass anhydrous MgSO₄</td>
<td></td>
</tr>
<tr>
<td>Mass of water in MgSO₄ hydrate</td>
<td></td>
</tr>
</tbody>
</table>

Analysis and Conclusions:
Percentage of Water:
1. Calculate the percentage of water in the hydrated crystals of MgSO₄ using your experimental data.
2. Assuming that the correct formula for the hydrate is MgSO₄ · 7H₂O, calculate the theoretical percentage of water in the hydrated crystals.
3. Calculate your percentage of error by comparing your experimental and theoretical percentages of water in the hydrate, using the formula in the Background section.
**Hydrate Formula**
4. Calculate the moles of anhydrous MgSO₄.
5. Calculate the moles of water removed from the hydrate by heating.
6. Determine the ratio of moles of water to moles of anhydrous MgSO₄.
7. Using this ratio, give your experimentally determined, predicted formula for hydrated MgSO₄.

**Further Questions**
8. Compare your observations of the hydrated and anhydrous crystals.
9. The method used in this experiment is not suitable for determining the percentage of water in all hydrates. Explain why this may be so.
10. What might you observe if the anhydrous crystals were left uncovered overnight?
11. If you were to repeat this experiment, what would you change in order to decrease your % error?

**Real-World Chemistry:**
1. Packets of the anhydrous form of a hydrate are sometimes used to keep cellars from being damp. Is there a limit to how long a packet could be used?
2. Gypsum (CaSO₄·2H₂O) is a mineral used for making wallboard for construction. The mineral is stripped of three-quarters of its water of hydration in a process called calcinning. Then, after mixing with water, it hardens to a white substance called plaster of Paris. Infer what happens as calcinated gypsum becomes plaster of Paris.
A Bagged Chemical Reaction

**Benchmarks:**

**SC.A.1.4.4**- The student experiments and determines that the rates of reaction among atoms and molecules depend on the concentration, pressure, and temperature of the reactants and the presence or absence of catalysts.

**SC.H.1.4.1**- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.E.2.4.6, SC.E.2.4.7, SC.H.1.2.1, SC.H.1.2.2, and SC.H.2.4.2.)

**SC.H.2.4.1**- The student knows that scientists control conditions in order to obtain evidence, but when that is not possible for practical or ethical reasons, they try to observe a wide range of natural occurrences to discern patterns.

**Objectives:**

- Observe the changes associated with a chemical reaction including heat changes, changes in an indicator, bubble of gas released.
- Determine which compounds are reactants and which are products through the use of a chemical equation.
- Use indicators in order to identify whether a solution is acidic or basic.

**Background Information:**

Chemical reactions occur when molecules come together to form new products. In our bodies and environments, chemical reactions occur continuously to help run our lives. During a reaction, chemical bonds are broken and remade. Usually, color changes, gases being released, changes in temperature, and/or formation of solute characterize chemical reactions. These events describe changes in energy or solubility of the compound, meaning new products were produced.

Chemical reactions can be defined by a chemical equation in which reactants and products are characterized by chemical symbols. All chemical reactions are accompanied by a change in energy. Some reactions release energy to their surroundings (usually in the form of heat) and are called **exothermic**. For example, sodium and chlorine react so violently that flames can be seen as the exothermic reaction gives off heat. On the other hand, some reactions need to absorb heat from their surroundings to proceed. These reactions are called **endothermic**. A good example of an endothermic reaction is that which takes place inside of an instant "cold pack." Commercial cold packs usually consist of two compounds - urea and ammonium chloride in separate containers within a plastic bag. When the bag is bent and the inside containers are broken, the two compounds mix together and begin to react. Because the reaction is endothermic, it absorbs heat from the surrounding environment and the bag gets cold.

**Lesson Lead:**

What in our world produces carbon dioxide? And what makes oxygen using carbon dioxide?

**Materials:**

- Safety goggles & Lab apron
- Calcium chloride pellets (CaCl₂)
- Baking soda, sodium bicarbonate (NaHCO₃)
• Phenol red solution
• Measuring cup or graduated cylinder
• 2 plastic teaspoons
• Plastic cup
• 1-gallon Ziploc-type bag
• 2 twist ties or rubber bands
• Water

Safety Precautions:
1. Remind students that there is NO eating or drinking during the lab.
2. CO2 is produced in the bags. Make sure the area is well-ventilated before releasing all the gas.
3. Do not let students ingest baking soda or calcium chloride. Avoid contact with eyes or mouth. If ingested in small amounts neither are toxic, but if ingested in larger amounts give student a full glass of water and contact a medical facility.
4. Watch for bags exploding. Ensure the students shake the bag away from their faces and clothes.
5. Once bags get tightly filled with gas, release the CO2. If the bag does explode, all the products are non-toxic and can be washed off.
6. The bags can be disposed of in the trash because all products are non-toxic.

Procedure:
1. Place 2 teaspoons calcium chloride pellets into one corner of the bag.
2. Twist off the corner to separate the calcium chloride from the rest of the bag. Follow teacher instructions and use either a rubber band or twist tie to secure.
3. Place two teaspoons of baking soda into the opposite corner of the bag.
4. Twist off the corner to separate the baking soda from the rest of the bag. Follow teacher instructions and use either a rubber band or twist tie to secure.
5. Fill a cup with about 10 ml of water and phenol red solution.
6. Pour the phenol red solution into the bag.
7. Carefully remove as much air as possible and close the bag.
8. Carefully untwist the two corners (if you used a rubber band, use scissors to cut the rubber band, but be careful not to cut the bag).
9. Hold the two corners of the bag apart. One partner may hold both corners, or each partner may hold one side.
10. Quickly observe any immediate changes in the corners.
11. Release the calcium chloride and baking soda allowing the liquid to mix with the two substances.
12. Observe the reaction until it comes to a complete stop. Record all your observations.

Analysis and Conclusions:
1. What signs of chemical reaction did you observe?
2. Would you characterize either reaction as exothermic? Which reaction? Why?
3. Would you characterize either reaction as endothermic? Which reaction? Why?
4. What was the function of the phenol red in the experiment?
5. During this reaction, baking soda (NaHCO₃) is combined with calcium chloride (CaCl₂) in water. Write the balanced chemical equation for the reaction that occurred.
6. What type of gas is in your bags?
Teacher Notes:

The chemical equation for the reaction is shown here:
NaHCO₃(s) + CaCl₂(s) → CaCO₃(s) + NaCl(aq) + HCl(aq)
NaHCO₃(s) + HCl(aq) → H₂O(l) + CO₂(g) + NaCl(aq)

Therefore, CaCO₃ or calcium carbonate, which is the main component of chalk, is produced during the reaction. Also, NaCl and H₂O make a mixture of salt water in the bag. Lastly, the gas in the bag is carbon dioxide, which is exhaled from our bodies and then taken in by plants, so they can generate oxygen. The reaction bubbles to release the gas and gets hot because it is exothermic and releases energy from chemical bonds that were broken. Exothermic means energy out and endothermic is energy into reaction.

Watch for bags exploding. Ensure the students shake the bag away from their faces and clothes. Once bags get tightly filled with gas, release the CO₂. If the bag does explode, all the products are non-toxic and can be washed off.
A Mole Ratio
(Adapted from Glencoe ChemLab and MiniLab)

Benchmarks:
SC.A.1.4.4- The student experiments and determines that the rates of reaction among atoms and molecules depend on the concentration, pressure, and temperature of the reactants and the presence or absence of catalysts.

SC.H.1.4.1- The student knows that investigations are conducted to explore new phenomena, to check on previous results, to test how well a theory predicts, and to compare different theories. (Also assesses SC.E.2.4.6, SC.E.2.4.7, SC.H.1.2.1, SC.H.1.2.2, and SC.H.2.4.2,)

SC.H.2.4.1- The student knows that scientists control conditions in order to obtain evidence, but when that is not possible for practical or ethical reasons, they try to observe a wide range of natural occurrences to discern patterns.

Objectives:
• Observe a single replacement reaction.
• Measure the masses of iron and copper.
• Calculate the moles of each metal and the mole ratio.
• Calculate the percent yield of copper.
• Determine the limiting reactant of the reaction.

Background Information:
Iron reacts with copper (II) sulfate in a single replacement reaction. By measuring the mass of iron that reacts and the mass of copper metal produced, you can calculate the ratio of moles of reactant to moles of product. This mole ratio can be compared to the ratio found in the balanced chemical equation.

Lesson Lead:
• Which reactant is the limiting reactant?
• How does the experimental mole ratio of Fe to Cu compare with the mole ratio in the balanced chemical equation?
• What is the percent yield?

Pre-Lab:
1. Is it important that you know you are using the hydrated form of copper (II) sulfate?
2. Would it be possible to use the anhydrous form? Why or why not?

Materials:
• iron metal filings, 20 mesh
• copper(II) sulfate pentahydrate (CuSO₄ · 5H₂O)
• 400-mL and 150-mL beakers
• 100-mL graduated cylinder
• weighing paper (filter paper can be used)
• balance
• hot plate
• beaker tongs
• distilled water
• stirring rod
Safety Precautions:
1. Remind students that there is NO eating or drinking during the lab.
2. Always wear safety glasses and a lab apron.
3. Hot objects will not appear to be hot.
4. Do not heat broken, chipped, or cracked glassware.
5. Turn off the hot plate when not in use.

Procedure:
1. Measure and record the mass of a clean, dry 150-mL beaker.
2. Place approximately 12 g of copper (II) sulfate pentahydrate into the 150-mL beaker and measure and record the combined mass.
3. Add 50 mL of distilled water to the copper (II) sulfate pentahydrate and heat the mixture on the hot plate at a medium setting. Stir until the solid is dissolved, but do not boil. Using tongs, remove the beaker from the hot plate.
4. Measure approximately 2 g of iron metal filings onto a piece of weighing paper. Measure and record the exact mass of the filings.
5. While stirring, slowly add the iron filings to the hot copper (II) sulfate solution.
6. Allow the reaction mixture to stand, without stirring, for 5 minutes to ensure complete reaction. The solid copper metal will settle to the bottom of the beaker.
7. Use the stirring rod to decant (pour off) the liquid into a 400-mL beaker. Be careful to decant only the liquid.
8. Add 15 mL of distilled water to the copper solid and carefully swirl the beaker to wash the copper. Decant the liquid into the 400-mL beaker.
9. Repeat step 8 two more times.
10. Place the 150-mL beaker containing the wet copper on the hot plate. Use low heat to dry the copper.
11. Remove the beaker from the hot plate and allow it to cool.
12. Measure and record the mass of the cooled 150-mL beaker and the copper.

Cleanup and Disposal
1. Make sure the hot plate is off.
2. The dry copper can be placed in a waste container. Wet any residue that sticks to the beaker and wipe it out using a paper towel. Pour the unreacted copper (II) sulfate and iron(II) sulfate solutions into a large beaker in the fume hood.
3. Return all lab equipment to its proper place.
4. Wash your hands thoroughly after all lab work and cleanup is complete.

Data Table:

<table>
<thead>
<tr>
<th>Mass of empty 150-mL beaker</th>
<th>Mass of 150-mL beaker + CuSO₄ · 5H₂O</th>
<th>Mass of CuSO₄ · 5H₂O</th>
<th>Mass of iron filings</th>
<th>Mass of 150-mL beaker and dried copper</th>
<th>Mass of dried copper</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Observations:</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Analysis and Conclusions:
1. What evidence did you observe that confirms that a chemical reaction occurred?
2. Write a balanced chemical equation for the single-replacement reaction that occurred.
3. From your data, determine the mass of copper produced.
4. Use the mass of copper to calculate the moles of copper produced.
5. Calculate the moles of iron used in the reaction.
6. Determine the whole number ratio of moles of iron to moles of copper from your data in steps 4 and 5.
7. What is the expected mole ratio of iron to copper from the balanced equation? Compare this ratio to the experimental mole ratio calculated using your data (#6).
8. Use the balanced chemical equation to calculate the mass of copper that should have been produced from the sample of iron you used.
9. Use this number and the mass of copper you actually obtained to calculate the percent yield.
10. What was the source of any deviation from the mole ratio calculated? from the chemical equation? How could you improve your results?

Extensions:
1. Determine the expected yield of copper metal (g) from the CuSO₄ · 5H₂O used?
2. Compare this answer to number 8 and determine the limiting reactant?

Real-World Chemistry:
1. A furnace that provides heat by burning methane gas (CH₄) must have the correct mixture of air and fuel to operate efficiently. What is the mole ratio of oxygen to methane gas in the combustion of methane?
2. Automobile air bags inflate on impact because a series of gas-producing chemical reactions are triggered. To be effective in saving lives, the bags must not overinflated or underinflated. What factors must automotive engineers take into account in the design of air bags?
Energy Content of Foods and Fuels
(From Vernier Chemistry Lab Manual)

Benchmark:
SC.B.1.4.1 - The students understands how knowledge of energy is fundamental to all the scientific disciplines (e.g., the energy required for biological processes in living organisms and the energy required for the building, erosion, and rebuilding of the Earth.

SC.B.1.4.3 - The student knows that temperature is a measure of the average translational kinetic energy of motion of the molecules in an object. CS

SC.B.1.4.6 - The student knows that the first law of thermodynamics relates the transfer of energy to the work done and the heat transferred. CS

SC.B.1.4.7 - The student knows that the total amount of usable energy always decreases, even though the total amount of energy is conserved in any transfer. CS

Objective:
• Determine the energy content of foods and fuels and compare to standard values.
• Understand how energy is produced from chemical reactions that take place all around us.

Background Information:
All human activity requires “burning” food and fuel for energy. In this experiment, you will determine the energy released (in kJ/g) as various foods and fuels burn such as cashews, marshmallows, peanuts, popcorn, paraffin wax, and ethanol. You will look for patterns in the amounts of energy released during burning of the different foods and fuels.

Lesson Lead:
What is the energy efficiency of different fuels?

Materials:
• Lab Pro or CBL 2 system*
• TI graphing calculator*
• Temperature probe*
• DataMate program
• 2 Food samples
• food holder (paper clip + foil)
• wooden splint
• candle (large diameter)
• aluminum foil squares
• *thermometer (instead of probe/CBL/TI calculator)
• *graph paper (instead of probe/CBL/TI calculator)
• utility clamp and slit stopper
• 2 Stirring rods
• ring stand and 4-inch ring
• 100-mL Graduated cylinder
• soda can
• cold water
• matches
• goggles

Safety Precautions:
• Always wear safety goggles and a lab apron.
• Do not eat or drink anything in a laboratory.
• Tie back hair and secure loose bulky clothing while working near a flame.

NOTE: If anyone is allergic to nuts, you can substitute with potato chips or other snack foods.
Part A: Energy Content of Foods

Procedure:

1. Obtain and wear goggles and a lab apron. Tie back long hair and secure loose fitting clothes.

2. Plug the Temperature Probe into Channel 1 of the LabPro or CBL 2 interface. Use the link cable to connect the TI Graphing Calculator to the interface. Firmly press in the cable ends.

3. Turn on the calculator and start the DATAMATE program. Press CLEAR to reset the program.

4. Set up the calculator and interface for the Temperature Probe.
   a. Select SETUP from the main screen.
   b. If the calculator displays a Temperature Probe in CH 1, proceed directly to Step 5. If it does not, continue with this step to set up your sensor manually.
   c. Press ENTER to select CH 1.
   d. Select TEMPERATURE from the SELECT SENSOR menu.
   e. Select the Temperature Probe you are using (in °C) from the TEMPERATURE menu.

5. Set up the data-collection mode.
   a. To select MODE, press once and press ENTER.
   b. Select TIME GRAPH from the SELECT MODE menu.
   c. Select CHANGE TIME SETTINGS from the TIME GRAPH SETTINGS menu.
   d. Enter “6” as the time between samples in seconds.
   e. Enter “100” as the number of samples. The length of the data collection will be 10 minutes.
   f. Select OK to return to the setup screen.
   g. Select OK again to return to the main screen.

6. Obtain a piece of one of the two foods assigned to you and a food holder. To make the food holder shape the paper clips into a small tripod and place on a piece of aluminum foil. Find and record in Data Table 1 the initial mass of the food sample and food holder.

7. Determine and record the mass of an empty can. Add 50 mL of cold water to the can. Determine and record in the mass of the can and water.

8. Set up the apparatus as shown in Figure 1. Use a ring and stirring rod to suspend the can about 2.5 cm (1 inch) above the food sample. Use a utility clamp to suspend the Temperature Probe in the water. The probe should not touch the bottom of the can. Remember: The Temperature Probe must be in the water for at least 30 seconds before you do Step 9.

9. Select START to begin collecting data. Record the initial temperature of the water, t1, in your data table (round to the nearest 0.1°C). Note: You can monitor temperature in the upper-right corner of the real-time graph displayed on the calculator screen. Remove the food sample from under the can and use a wooden splint to light it. Quickly place the
burning food sample directly under the center of the can. Allow the water to be heated until the food sample stops burning. CAUTION: Keep hair and clothing away from open flames.

10. Continue stirring the water until the temperature stops rising. Record this maximum temperature, t2. Data collection will stop after 10 minutes (or press the STO key to stop before 10 minutes has elapsed).

11. Determine and record the final mass of the food sample and food holder.
12. To confirm the initial (t1) and final (t2) values you recorded earlier, examine the data points along the curve on the displayed graph. As you move the cursor right or left, the time (X) and temperature (Y) values of each data point are displayed below the graph.

13. Press ENTER to return to the main screen. Select START to repeat the data collection for the second food sample. Use a new 50-mL portion of cold water. Repeat Steps 6-12.

14. When you are done, place burned food, used matches, and partially-burned wooden splints in the container provided by the teacher.

Analysis

1. Find the mass of water heated for each sample and record in Data Table 1.
2. Find the change in temperature of the water, Δt, for each sample and record in Data Table 1.
3. Calculate the heat absorbed by the water, q, using the equation \( q = C_p \cdot m \cdot \Delta t \) where q is heat, \( C_p \) is the specific heat capacity, m is the mass of water, and \( \Delta t \) is the change in temperature. For water, \( C_p \) is 4.18 J/g°C. Change your final answer to kJ.
4. Find the mass (in g) of each food sample burned.
5. Use the results of Steps 3 and 4 to calculate the energy content (in kJ/g) of each food sample.
6. Record your results and the results of other groups in Data Table 2: Class Results Table.
## Data Table 1: Group Results

<table>
<thead>
<tr>
<th>Food type</th>
<th>Amounts</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial mass of food and holder</td>
<td>g</td>
<td>g</td>
</tr>
<tr>
<td>Final mass of food and holder</td>
<td>g</td>
<td>g</td>
</tr>
<tr>
<td>Mass of food burned</td>
<td>g</td>
<td>g</td>
</tr>
<tr>
<td>Mass of can and water</td>
<td>g</td>
<td>g</td>
</tr>
<tr>
<td>Mass of empty can</td>
<td>g</td>
<td>g</td>
</tr>
<tr>
<td>Mass of water heated</td>
<td>g</td>
<td>g</td>
</tr>
<tr>
<td>Final temperature, $t_2$</td>
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<td>°C</td>
</tr>
<tr>
<td>Initial temperature, $t_1$</td>
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<td>°C</td>
</tr>
<tr>
<td>Temperature change, $\Delta t$</td>
<td>°C</td>
<td>°C</td>
</tr>
<tr>
<td>Heat, $q$</td>
<td>kJ</td>
<td>kJ</td>
</tr>
<tr>
<td>Energy content in kJ/g</td>
<td>kJ/g</td>
<td>kJ/g</td>
</tr>
</tbody>
</table>

## Data Table 2: Class Results

<table>
<thead>
<tr>
<th></th>
<th>Marshmallows</th>
<th>Peanuts</th>
<th>Cashews</th>
<th>Popcorn</th>
</tr>
</thead>
<tbody>
<tr>
<td>Average for each food type:</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Marshmallows</td>
<td>kJ/g</td>
<td>kJ/g</td>
<td>kJ/g</td>
<td>kJ/g</td>
</tr>
<tr>
<td>Peanuts</td>
<td>kJ/g</td>
<td>kJ/g</td>
<td>kJ/g</td>
<td>kJ/g</td>
</tr>
<tr>
<td>Cashews</td>
<td>kJ/g</td>
<td>kJ/g</td>
<td>kJ/g</td>
<td>kJ/g</td>
</tr>
<tr>
<td>Popcorn</td>
<td>kJ/g</td>
<td>kJ/g</td>
<td>kJ/g</td>
<td>kJ/g</td>
</tr>
</tbody>
</table>
Conclusion:
1. Which food had the highest energy content? The lowest energy content?
2. Food energy is often expressed in a unit called a Calorie. There are 4.18 kJ in one Calorie. Based on the class average for peanuts, calculate the number of Calories in a 50-g package of peanuts.
3. Two of the foods in the experiment have a high fat content (peanuts and cashews) and two have high carbohydrate content (marshmallows and popcorn). From your results, what generalization can you make about the relative energy content of fats and carbohydrates?

Part B: Energy Content of Fuels
Procedure:
1. Obtain and wear goggles.
2. Plug the Temperature Probe into Channel 1 of the LabPro or CBL 2 interface. Use the link cable to connect the TI Graphing Calculator to the interface. Firmly press in the cable ends.
3. Turn on the calculator and start the DATAMATE program. Press CLEAR to reset the program.
4. Set up the calculator and interface for the Temperature Probe.
   a. Select SETUP from the main screen.
   b. If the calculator displays a Temperature Probe in CH 1, proceed directly to Step 5. If it does not, continue with this step to set up your sensor manually.
   c. Press ENTER to select CH 1.
   d. Select TEMPERATURE from the SELECT SENSOR menu.
   e. Select the Temperature Probe you are using (in °C) from the TEMPERATURE menu.
5. Set up the data-collection mode.
   a. To select MODE, press ENTER once and press ENTER.
   b. Select TIME GRAPH from the SELECT MODE menu.
   c. Select CHANGE TIME SETTINGS from the TIME GRAPH SETTINGS menu.
   d. Enter “6” as the time between samples in seconds.
   e. Enter “100” as the number of samples. The length of the data collection will be 10 minutes.
   f. Select OK to return to the setup screen.
   g. Select OK again to return to the main screen.
6. Set the candle on a piece of aluminum foil or any type of candle holder. Wider based candles (ex. 1 inch diameter) are preferable because they do not tip over.
7. Find and record in Table 1 Part 1 the combined mass of the candle and aluminum foil.
8. Determine and record the mass of an empty can. Add 100 mL of chilled water to the can. Determine and record the mass of the can and water.
9. Set up the apparatus as shown in Figure 1. Use a ring and stirring rod to suspend the can about 5 cm above the wick. Use a utility clamp to suspend the Temperature Probe in the water. The probe should not touch the bottom of the can. Remember: The Temperature Probe must be in the water for at least 30 seconds before you do Step 10.
10. Select START on the calculator to begin collecting data. Monitor temperature (in °C) on the calculator screen for about 30 seconds and record the initial temperature of the water, \( t_1 \), in your data table. Light the candle and heat the water until its temperature reaches 40°C and then extinguish the flame. CAUTION: Keep hair and clothing away from an open flame.

11. Continue stirring the water until the temperature stops rising. Record this maximum temperature, \( t_2 \). Data collection will stop after 10 minutes (or press the \( \text{STO} \) key to stop before 10 minutes has elapsed).

12. Determine and record in Table 1 Part 1, the final mass of the cooled candle and foil, including all drippings.

13. To confirm the initial \( (t_1) \) and final \( (t_2) \) values you recorded earlier, examine the data points along the curve on the displayed graph. As you move the cursor right or left, the time \( (X) \) and temperature \( (Y) \) values of each data point are displayed below the graph.

14. Press \( \text{ENTER} \) to return to the main screen. Select START to repeat the data collection using ethanol in an alcohol burner. Repeat Steps 7-13. Be sure to use 200 mL of chilled water in Step 8.

Analysis:
1. Find the mass of water heated.

2. Find the change in temperature of the water, \( \Delta t \).

3. Calculate the heat absorbed by the water, \( q \), using the formula in the introduction of this experiment. For water, \( C_p \) is 4.18 J/g°C. Change your final answer to kJ.

4. Calculate the heat absorbed by the water, \( q \), using the equation

\[
q = C_p \cdot m \cdot \Delta t
\]

where \( q \) is heat, \( C_p \) is the specific heat capacity, \( m \) is the mass of water, and \( \Delta t \) is the change in temperature. For water, \( C_p \) is 4.18 J/g°C. Change your final answer to kJ.

5. Find the mass of paraffin burned.

6. Calculate the heat of combustion for paraffin in kJ/g. Use your Step 3 and Step 4 answers.

7. Calculate the % efficiency in both trials of the experiment. Divide your experimental value (in kJ/g) by the accepted value, and multiply the answer by 100. The accepted heat of combustion of paraffin is 41.5 kJ/g, and for ethanol the value is 30.0 kJ/g. Record in Data Table 2.
Data Table 3: Paraffin Fuel

<table>
<thead>
<tr>
<th>Amount</th>
</tr>
</thead>
<tbody>
<tr>
<td>g</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Initial mass of fuel + container</th>
<th>g</th>
</tr>
</thead>
<tbody>
<tr>
<td>Final mass of fuel + container</td>
<td>g</td>
</tr>
<tr>
<td>Mass of fuel burned</td>
<td>g</td>
</tr>
<tr>
<td>Mass of can and water</td>
<td>g</td>
</tr>
<tr>
<td>Mass of empty can</td>
<td></td>
</tr>
<tr>
<td>Mass of water heated</td>
<td>g</td>
</tr>
<tr>
<td>Final temperature, $t_2$</td>
<td>°C</td>
</tr>
<tr>
<td>Initial temperature, $t_1$</td>
<td>°C</td>
</tr>
<tr>
<td>Temperature change, $\Delta t$</td>
<td>°C</td>
</tr>
</tbody>
</table>

Data Table 4 Heat

<table>
<thead>
<tr>
<th>Heat, $q$</th>
</tr>
</thead>
<tbody>
<tr>
<td>kJ</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Heat of combustion, in kJ/g</th>
</tr>
</thead>
<tbody>
<tr>
<td>kJ/g paraffin</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>% efficiency</th>
</tr>
</thead>
<tbody>
<tr>
<td>%</td>
</tr>
</tbody>
</table>

Conclusion:

1. Based on your results, which fuel produces more energy per gram burned? Give an explanation for the difference. (Hint: Ethanol, C$_2$H$_5$OH, is an oxygenated molecule; paraffin, C$_{25}$H$_{52}$, does not contain oxygen.)

2. Suggest some advantages of using ethanol (or paraffin) as a fuel.

3. Discuss heat loss factors that contribute to the inefficiency of the experiment.

Extensions: (Student Independent Research)

1. How does the energy from the sun become stored in foods and fuels?

2. How does this energy flow through biological systems and through the processes that shape the Earth?
Comparing Rates of Evaporation
(Adapted from Glencoe Textbook p 410-411)

Benchmark:
SC.A.1.4.2 - The student knows that the vast diversity of the properties of materials is primarily due to variations in the forces that hold molecules together. CS

SC.C.2.4.5 - The student knows that most observable forces can be traced to electric forces acting between atoms or molecules. CS

Objectives:
• Measure and compare the rates of evaporation for different liquids.
• Classify liquids based on their rates of evaporation.
• Predict which intermolecular forces exist between the particles of each liquid.

Background Information:
Several factors determine how fast a sample of liquid will evaporate. The volume of the sample is a key factor. A drop of water takes less time to evaporate than a liter of water. The amount of energy supplied to the sample is another factor. In this lab, you will investigate how the type of liquid and temperature affect the rate of evaporation.

Lesson Lead:
How do intermolecular forces affect the evaporation rates of liquids?

Pre-Lab:
1. Read the entire lab directions. Use the data table on the next page.
2. What is evaporation? Describe what happens at the molecular level during evaporation.
3. List the three possible intermolecular forces. Which force is the weakest? Which force is the strongest?
4. Look at the materials list for this lab. Consider the five liquids you will test. Predict which liquids will evaporate quickly and which will take longer to evaporate. Give reasons for your predictions.
5. To calculate an evaporation rate, you would divide the evaporation time by the quantity of liquid used. Explain why it is possible to use the evaporation times from this lab as evaporation rates.
6. Make sure you know how to use the stopwatch provided. Will you need to convert the reading on the stopwatch to seconds?

Materials:
- distilled water
- ethanol
- isopropyl alcohol
- acetone
- household ammonia
- droppers (5)
- small plastic cups (5)
- grease pencil or marking pen
- masking tape
- paper towel
- square of waxed paper
- stopwatch

Safety Precautions:
• Always wear safety goggles and a lab apron.
• Wear gloves because some of the liquids can dry out your skin.
• Avoid inhaling any of the vapors, especially ammonia.
• There should be no open flames in the lab; some of the liquids are flammable.

Procedure:
1. Use a grease pencil or masking tape to label each of five small plastic cups. Use A for distilled water, B for ethanol, C for isopropyl alcohol, D for acetone, and E for household ammonia.
2. Place the plastic cups on a paper towel.
3. Use a dropper to collect about 1 mL of distilled water and place the water in the cup A.
4. Place the dropper on the paper towel directly in front of the cup. Repeat with the other liquids.
5. Place a square of waxed paper on your lab surface.
6. Plan where on the waxed paper you will place each of the 5 drops that you will test. The drops must be as far apart as possible to avoid mixing.
7. Have your stopwatch ready. Collect some water in your water dropper and place a single drop on the waxed paper. Begin timing. Time how long it takes for the drop to completely evaporate. While you wait, make two drawings of the drop. One drawing should show the shape of the drop as viewed from above. The other drawing should be a side view at eye level. If the drop takes longer than 5 minutes to evaporate, record _ 300 in your data table.
8. Repeat step 5 with the four other liquids.
9. Use the above procedure to design an experiment in which you can observe the effect of temperature on the rate of evaporation of ethanol. Your teacher will provide a sample of warm ethanol. Record your observations.

Cleanup and Disposal:
1. Crumple up the waxed paper and place it in the container assigned by your teacher.
2. Place unused liquids in the containers specified by your teacher.
3. Wash out all droppers and test tubes except those used for distilled water.

Data Table 1: Evaporation Data

<table>
<thead>
<tr>
<th>Liquid</th>
<th>Evaporation time (s)</th>
<th>Shape of liquid drop</th>
</tr>
</thead>
<tbody>
<tr>
<td>Distilled water</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ethanol</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Isopropyl alcohol</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Acetone</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Household ammonia</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Analysis and Conclusions:
1. Which liquids evaporated quickly? Which liquids were slow to evaporate?
2. Based on your data, in which liquid(s) are the attractive forces between molecules most likely to be dispersion forces?
3. Make a generalization about the shape of a liquid drop and the evaporation rate of the liquid.
4. What is the relationship between surface tension and the shape of a liquid drop? What are the attractive forces that increase surface tension?
5. The isopropyl alcohol you used is a mixture of isopropyl alcohol and water. Would pure isopropyl alcohol evaporate more quickly or more slowly compared to the alcohol and water mixture? Give a reason for your answer.
6. Household ammonia is a mixture of ammonia and water. Based on the data you collected, is there more ammonia or more water in the mixture? Use what you learned about the relative strengths of the attractive forces in ammonia and water to support your conclusion.
7. How does the rate of evaporation of warm ethanol compare to ethanol at room temperature? Use kinetic-molecular theory to explain your observations.
8. How could you change the procedure to make it more precise?

Real-World Chemistry:
1. The vapor phases of liquids such as acetone and alcohol are more flammable than their liquid phases. For flammable liquids, what is the relationship between evaporation rate and the likelihood that the liquid will burn?
2. Suggest why a person who has a higher than normal temperature might be given a rubdown with rubbing alcohol (70% isopropyl alcohol).
3. Table salt can be collected from salt water by evaporation. The water is placed in large, shallow containers. What advantage do these shallow containers have over deep containers with the same overall volume?
Boyle’s Law Experiment
(Adapted from Prentice Hall Chemistry Connections Lab Manual)

Benchmark:
SC.B.1.4.1 - The student understands how knowledge of energy is fundamental to all the scientific disciplines (e.g., the energy required for biological processes in living organisms and the energy required for the building, erosion, and rebuilding of the Earth.

SC.B.1.4.3 - The student knows that temperature is a measure of the average translational kinetic energy of motion of the molecules in an object.

Objective:
- Determine the relationship between pressure and volume.

Background Information:
In this investigation, you will observe the behavior of a gas, using a device called a Boyle’s law apparatus. The apparatus consists of a graduated syringe with a movable piston. Initially, the syringe is adjusted to trap a volume of gas at the same pressure as its surroundings. The piston then does not move because the pressure exerted by the gas in the syringe equals the pressure of the atmosphere pushing on the piston. If the piston is pushed downward, it compresses the gas trapped in the syringe. If the pressure on the piston is then decreased, the pressure of the trapped gas will push the piston up.

In order to read the volume of trapped gas correctly, you must always read the measurement on the side of the piston that is in contact with the gas. Because air is a mixture of gases—mostly nitrogen and oxygen—that behaves physically as a single gas, the data from this lab can be treated as data for a single gas. As the pressure of the air changes, you will monitor and collect data on the resulting changes in volume. You can then use your data to find the atmospheric pressure and determine how closely your results agree with Boyle’s law.

Lesson Lead:
How does the volume of an enclosed sample of gas change as the pressure of the gas is changed?

Pre-Lab:
1. State Boyle’s law in your own words. Then write the mathematical equation for Boyle’s law.
2. When the piston in the Boyle’s law apparatus is at rest, what is the relationship between the pressure of the trapped gas and the pressure on the outside of the piston?
3. What are the possible sources of external pressure on the piston during this investigation?
4. What is the benefit of collecting three sets of data in the investigation?
5. In what ways can you minimize the risk of injury or damage to equipment from falling books?

Materials:
- safety goggles
- Boyle’s law apparatus
- ring stand clamp
- 5 chemistry textbooks
- 2 pens or pencils of different colors
Safety Precautions:
- Wear appropriate safety goggles.
- Don’t let the students aim the pressurized syringe at anyone as the syringe tip cap could shoot off.

Note: The increasing load of books on the piston may become unsteady. Falling books can injure the person measuring the gas volumes and damage the Boyle’s law apparatus. Steady the books by resting them slightly against the ring or by nudging them into balance as you would with wooden building blocks.

Procedure:
1. Work with a partner so that one person operates the apparatus (see figure below) while the other steadies the books and keeps track of the procedural steps.
2. Put on your safety goggles. Secure the Boyle’s law apparatus with a ring stand and clamp. Adjust the initial volume (about 30 mL) to atmospheric pressure as directed by your teacher.
3. Test the apparatus by pushing down on the piston with your hand slowly and steadily until the volume of the trapped gas is reduced to 15 mL. Release the piston and note whether it returns to initial volume. If not, check and adjust the seal at the syringe opening. (Note: if red tip seals are too loose try replacing it with a small rubber stop with a small indentation).
4. Place the apparatus on a flat, steady surface, such as a sturdy table or the floor. Record the initial volume at 0 books of pressure in the data table.
5. Place one book on the piston and record the resulting volume of trapped gas in the data table. Add a second book and record the gas volume. Continue adding books and recording the resulting volumes until all 5 books are resting on the piston. Remember to steady the books, especially when the apparatus is being read.
6. Remove all the books from the piston and reset the apparatus to the initial volume recorded in Step 2.
7. Repeat Step 5 and 6 two more times, remembering to reset the apparatus between sets of trials.
8. Clean up your work area.
Data Table:

<table>
<thead>
<tr>
<th>Pressure (#books)</th>
<th>$V_1$(mL)</th>
<th>$V_2$(mL)</th>
<th>$V_3$(mL)</th>
<th>$V_{\text{avg}}$(mL)</th>
<th>$1/V_{\text{avg}}$(mL$^{-1}$)</th>
<th>$P_{\text{atm}}$</th>
<th>$P_{\text{total}}$</th>
<th>$P_{\text{total}} \times V_{\text{avg}}$</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
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</tr>
</tbody>
</table>

Analysis and Calculations:

1. Find the average of each set of three volumes and record these averages in the data table.

2. For Graph #1 plot the pressure (in books) on the horizontal axis, and the average volume, $V_{\text{avg}}$ on the vertical axis. Draw a smooth line through the points.

3. According to Boyle’s law, pressure and volume have an inversely proportional relationship. If this idea is correct, you should obtain a linear relationship (straight line) when you plot pressure versus the corresponding inverse of the average volume. Calculate the inverse, $1/V_{\text{avg}}$, of each volume and record these values in the appropriate column of the data table.

4. For Graph #2, plot the pressures in units of “books” on the horizontal axis versus their corresponding $1/V_{\text{avg}}$ values (See sample Graph 2 below).

5. The line obtained for the second plot crosses the vertical axis of the graph above the origin, which tells you that there is pressure on the gas even when there are no books on the piston. Consider that $1/V = 0$ only when the total pressure on the gas is zero (and the volume is infinitely large). The additional pressure is the atmospheric pressure. To find this pressure in units of books, extend the plot of $P_{\text{books}}$ versus $1/V_{\text{avg}}$ on Graph #2 to the point where it intersects the horizontal axis. At this point, $1/V_{\text{avg}} = 0$. The scale distance from this point to the origin is the atmospheric pressure measured in books. Using your graph, determine this value (See Sample Graph 2 below).

6. Record the value of $P_{\text{atm}}$ in each row of the table in the proper column.

7. Add the value you found for atmospheric pressure ($P_{\text{atm}}$) to pressure in books ($P_{\text{books}}$) for each trial and record these values of $P_{\text{total}}$ in the table.

$$P_{\text{total}} = P_{\text{books}} + P_{\text{atm}}$$

7. Calculate the product of $P_{\text{total}} \times V_{\text{avg}}$ for each trial and record these values in the data table.
Analysis and Conclusions:
1. What is the benefit of repeating the measurements 3 times?

2. Explain the relationship between pressure and volume shown in Graph#1.

3. Describe the results shown in Graph #2.

4. Look at the values you calculated in the last column of the data table. How do they compare?

5. State Boyle’s Law.

6. Was your hypothesis proven? Explain why or why not.

7. What do these values mean in terms of Boyle’s law?

Real-World Chemistry:

1. When you use a bicycle pump to inflate a tire you push on the pump and air moves into the tire. In order for the pump to work, air pressure must be greater in the pump than in the bicycle tire. The air will move from the pump to the tire, causing inflation. How does the principle of Boyle’s law come into effect in the operation of bicycle pump?

Sample Graph
Benchmark:
SC.A.1.4.2-The student knows that the vast diversity of the properties of materials is primarily due to variations in the forces that hold molecules together. CS

Objectives:
- Prepare a saturated solution of KCl in water at an assigned temperature.
- Measure the mass of KCl dissolved in a certain mass of water at the assigned temperature.
- Calculate the solubility of KCl at your assigned temperature.
- Graph class solubility data as a function of temperature for KCl in water.

Background Information:
A homogeneous mixture of a solute in a solvent is called a solution. An unsaturated solution is capable of dissolving additional solute for a given amount of solvent. When the particular amount of solvent can dissolve no additional solute, the solution is called saturated. Any additional solute added to a saturated solution will collect on the bottom of the container and remain undissolved. The amount of solute that can be dissolved in a given amount of solvent at a specific temperature and pressure is defined as the solubility of the solute. Solubility is dependent upon temperature. In this activity, you will determine the solubility of a salt at different temperatures and will plot a solubility curve for the solute.

Lesson Lead: How do you determine the solubility curve for a given salt?

Pre-Lab:
1. How will you know when the solution is saturated?
2. Why must a saturated solution be obtained in order to make a solubility curve?
3. Read over the entire laboratory activity. Hypothesize what will happen to the solubility of KCl as the temperature is increased. Record your hypothesis.

Hypothesis: [Student writes]

Materials:
- potassium chloride (KCl)
- distilled water
- balance
- evaporating dish (or 100-mL beaker)
- 25-mL graduated cylinder
- watch glass
- 250 or 400mL beaker
- hot plate or burner with ring stand, 2 rings & wire gauze
- test tube (18X150 mm)
- utility clamp
- glass stirring rod
- thermometer
- funnel
- cotton wadding
- tongs &/or hot mitts
- graph paper

Safety Precautions:
- Always wear safety goggles and a lab apron.
- Never taste any substance used in the lab.
- Test tube and evaporating dish may be cause burns
- Use caution around hot items.
Procedure:
1. Your teacher will assign your group a temperature between 20ºC and 90ºC.
2. Determine the mass of a clean, dry evaporating dish (or 100 mL beaker) with watch glass cover. Set aside.
3. Put 15 mL of distilled water in the test tube and add about 10 g of KCl. Immerse the test tube in the large (250 mL) beaker.
4. Heat the water until it reaches your assigned temperature. Maintain this temperature for 10 minutes, stirring every few minutes and rechecking the temperature.
5. Set up a funnel for filtering the solution, using a small wad of cotton instead of filter paper, placing the previously weighed evaporating dish (or 100 mL beaker) below the funnel.
6. Remove the test tube from the water bath, being careful that any solid at the bottom of the test tube is undisturbed. Decant about half of the solution into the funnel.
7. Weigh the evaporating dish, watch glass, cover and contents when filtering is complete.
8. Evaporate the water from the solution by heating rapidly at first and then more slowly. Start the evaporation without the watch glass and then cover the evaporating dish with the watch glass to prevent loss of KCl by spattering.
9. After the evaporating dish has cooled, measure the mass again. Reheat until the mass changes by less than 0.02 g.

Cleanup and Disposal:
1. Turn off the hot plate and allow it to cool.
2. Make sure all glassware is cool before emptying the contents.
3. Dispose of all chemical as directed by your teacher.
4. Return all lab equipment to its proper place.
5. Clean up your work area.

Analysis:
1. Record your data calculation answers in Table 1.
2. Record class data in Table 2
3. Make a line graph of the solubility of KCl in grams KCl/100 g water (Y-axis) versus the temperature in ºC (X-axis).

Table 1. Individual Group Data

<table>
<thead>
<tr>
<th>Assigned Temperature</th>
<th>Mass of evaporating dish and cover</th>
<th>Mass of evaporating dish and cover plus KCl solution</th>
<th>Mass of KCl solution</th>
<th>Mass of evaporating dish and cover plus dry KCl</th>
<th>Mass of dry KCl</th>
<th>Mass of water that was in solution</th>
<th>Grams KCl in one g water</th>
<th>Grams KCl in 100 g water</th>
</tr>
</thead>
</table>
Table 2. Class Data

<table>
<thead>
<tr>
<th>Temperature (°C)</th>
<th>Mass of KCl per 100 g water</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
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<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Conclusion:
1. Describe your findings on the effect of temperature on the solubility of KCl.
2. Predict the solubility of KCl in water at 15°C and 95°C
3. Compare the results of this lab with the predictions of your hypothesis. Explain possible reasons for any disagreement.

Real World Chemistry:
1. In a dishwasher, the temperature of the water is very hot. Explain why it is better to use hot water in a dishwasher rather than cold water.
2. Unlike solids for which solubility in a liquid generally increases with increasing temperature, the solubility of a gas in a liquid usually decreases as the temperature increases. Knowing this, explain why you should never heat a can containing a carbonated soft drink.
Percentage of Acetic Acid in Vinegar
(Adapted from Holt ChemFile Microscale Experiments)

Benchmark:  
SC.A.1.4.4 - The student experiments and determines that the rates of reaction among atoms and molecules depend on the concentration, pressure, and temperature of the reactants and the presence or absence of catalysts. AA

Objective:  
• Determine the end point of an acid-base titration  
• Calculate the molarity of acetic acid in vinegar  
• Calculate the percentage of acetic acid in vinegar.

Background Information:  
Vinegar, which is naturally produced from the fermentation of apple cider in the absence of oxygen, contains acetic acid (CH₃COOH) at a concentration 4.0 to 5.5 %. The exact concentration of acetic acid can be determined by titration with a standard base such as sodium hydroxide. This can be done on a microscale as described below by counting drops of base needed to neutralize the acid in vinegar, or on a larger scale with a standard 50 mL buret setup.

Materials:  
• apron  
• goggles  
• 10 mL graduated cylinder  
• 24-well plate or 3 small beakers  
• 2 thin-stemmed pipets or droppers  
• phenolphthalein indicator  
• 2.0 mL standardized 0.6 M NaOH  
• 2.0 mL white vinegar  
• stirrer

Safety Precautions:  
• Wear safety goggles and a lab apron.  
• Do not touch chemicals

Procedure:  
Calibrate pipet droppers:  
1. Put about 5 mL of water in the 10 mL graduated cylinder and read the exact volume. Record this reading in the Calibration Data Table.  
2. Fill the pipet with water. Holding the pipet in a vertical position, transfer exactly 20 drops of water to the graduated cylinder. Record the new volume in the graduated cylinder as the Final volume for Trial 1.  
3. The Final volume for Trial 1 will then be the Initial volume for Trial 2 as 20 more drops are added to the graduated cylinder.  
4. Record the new Final volume and repeat for a third trial.  
5. Mark this pipet as Acid.  
6. Repeat steps 1 to 4 and mark the second pipet as Base.  
7. Fill in the Calibration Data Table with the Average Volume of 20 Drops and the Average Volume Per Drop for both pipets.
Calibration Data Table

<table>
<thead>
<tr>
<th>Trial</th>
<th>Acid Pipet</th>
<th>Base Pipet</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Initial</td>
<td>Final</td>
</tr>
<tr>
<td>1</td>
<td></td>
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<td>2</td>
<td></td>
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<tr>
<td>3</td>
<td></td>
<td></td>
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<tr>
<td>Average Volume of 20 Drops</td>
<td>--</td>
<td>--</td>
</tr>
<tr>
<td>Average Volume Per Drop</td>
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</tbody>
</table>

Titration:
1. Drain the water from the Acid pipet, rinse the Acid pipet with vinegar, and discard this vinegar. Then fill the pipet with fresh vinegar.
2. Hold the pipet vertically and add 20 drops of vinegar and 1 drop of phenolphthalein to 3 wells of the well plate or to the 3 small beakers.
3. Drain the water from the Base pipet, rinse the Base pipet with NaOH solution, and discard this rinse solution. Then fill the pipet with fresh NaOH solution.
4. Hold the Base pipet vertically and add NaOH solution drop by drop to one well or beaker with gentle swirling after each drop is added. Continue adding drops of NaOH solution until the pink phenolphthalein color remains for 30 s. Record the number of drops of NaOH added in the Titration Data Table.
5. Repeat step 4 with the 2 other 20-drop vinegar samples refilling the NaOH pipet if necessary. Record the observed number of drops in the Titration Data Table.

6. Cleanup: Clean all equipment used and dispose of the chemicals as directed by your teacher. Wash your hands.

Titration Data Table

<table>
<thead>
<tr>
<th>Trial</th>
<th>Number of Drops Added</th>
<th>Volume Added (mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Vinegar</td>
<td>NaOH</td>
</tr>
<tr>
<td>1</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
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<tr>
<td>3</td>
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</tbody>
</table>

Analysis:
1. Calculate the volumes of vinegar and NaOH for each trial and record in the Titration Data Table. Show your calculations.
2. From the molarity of the standardized NaOH solution (provided by your teacher) calculate the number of moles of NaOH used to neutralize the acetic acid in each trial. Show work.
3. Write the balanced equation for the neutralization of acetic acid by sodium hydroxide.

4. From the calculations in steps 2 and 3, and the mole ratio from the balanced equation of #3 above, calculate the number of moles of acetic acid neutralized by NaOH each trial. Show work.

5. From the moles of acetic calculated in #4 and the volumes of vinegar used in each trial, calculate the molarities of acetic acid in each trial. Then average your results.

6. Calculate the molar mass of acetic acid, CH₃COOH.

7. Using your average molarity of acetic acid in vinegar (#5), calculate the mass of acetic acid in 100 mL (0.100 L) of vinegar. Hint: Find the mass of acetic acid in 1L and then the mass in 100 mL.

8. Assume that the density of vinegar is close to 1.00 g/mL, so that the mass of 100 mL of vinegar will be 100.g. Determine the percentage of acetic acid in your sample of vinegar.

Conclusion:

1. Why was phenolphthalein used in each titration? Could you have done your titrations without phenolphthalein?

2. Why were you instructed to hold the pipets in a vertical position when you used them?

3. What are some possible sources of error in your procedure?

Real World Chemistry:

1. Why must a company producing vinegar check the molarity or percentage of acetic acid in their product?
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