SCIENCE STANDARDS OF LEARNING
ENHANCED SCOPE & SEQUENCE

CHEMISTRY

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Preface

The *Science Standards of Learning Enhanced Scope and Sequence* is a resource intended to help teachers align their classroom instruction with the Science Standards of Learning that were adopted by the Board of Education in January 2003. The *Enhanced Scope and Sequence* contains the following:

- Units organized by topics from the 2003 *Science Standards of Learning Sample Scope and Sequence.* Each topic lists the following:
    - Standards of Learning related to that topic
    - Essential understandings, knowledge, and skills from the *Science Standards of Learning Curriculum Framework* that students should acquire
- Sample lesson plans aligned with the essential understandings, knowledge, and skills from the *Curriculum Framework.* Each lesson contains most or all of the following:
    - An overview
    - Identification of the related Standard(s) of Learning
    - A list of objectives
    - A list of materials needed
    - A description of the instructional activity
    - One or more sample assessments
    - One or more follow-ups/extensions
    - A list of resources
- Sample released SOL test items for each Organizing Topic.

School divisions and teachers can use the *Enhanced Scope and Sequence* as a resource for developing sound curricular and instructional programs. These materials are intended as examples of ways the essential understandings, knowledge, and skills might be presented to students in a sequence of lessons that has been aligned with the Standards of Learning. Teachers who use the *Enhanced Scope and Sequence* should correlate the essential understandings, knowledge, and skills with available instructional resources as noted in the materials and determine the pacing of instruction as appropriate. This resource is not a complete curriculum and is neither required nor prescriptive, but it can be a valuable instructional tool.

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Organizing Topic — Introduction to Chemistry

Standards of Learning

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include

a) designated laboratory techniques;
b) safe use of chemicals and equipment;
c) proper response to emergency situations;
d) manipulation of multiple variables, using repeated trials;
e) accurate recording, organization, and analysis of data through repeated trials;
f) mathematical and procedural error analysis;
g) mathematical manipulations (SI units, scientific notation, linear equations, graphing, ratio and proportion, significant digits, dimensional analysis);
h) use of appropriate technology including computers, graphing calculators, and probeware, for gathering data and communicating results; and
i) construction and defense of a scientific viewpoint (the nature of science).

Essential Understandings, Knowledge, and Skills

Correlation to Textbooks and Other Instructional Materials

The student will use hands-on investigations, problem solving activities, scientific communication, and scientific reasoning to

- apply experimental design used in scientific investigation:
  - Perform and design experiments to test predictions;
  - Predict outcomes when a variable is changed;
- use graphs to show the relationships of the data:
  - Dependent variable (vertical axis)
  - Independent variable (horizontal axis)
  - Scale and units of graph
  - Regression lines;
- identify and properly use the following basic lab equipment: beaker, flask, graduated cylinder, test tube, test tube rack, test tube holder, ring stand, wire gauze, clay triangle, crucible with lid, evaporation dish, watch glass, wash bottle, and dropping pipette;
- identify, locate, and properly utilize MSDS and laboratory safety equipment, including aprons, goggles, gloves, fire extinguishers, fire blanket, safety shower, eye wash, broken glass container, and fume hood;
- express measurements in SI units and know the SI prefixes of *milli-, centi-, deci-,* and *kilo-*;
- read instruments, considering significant figures, and perform mathematical operations using significant figures;
- use appropriate technology, such as graphing calculator and probeware interfaced to a graphing calculator or computer, to collect and analyze data.

Laboratory Safety and Skills

Organizing Topic Introduction to Chemistry

Overview Students focus on laboratory safety and the basic laboratory skills necessary to prevent accidents. With the help of visual aids and technology resources, students locate, identify, and describe the use of lab safety equipment. They undertake guided practice in proper, safe laboratory techniques, using basic lab equipment and innocuous materials such as water, salt, and salt-water solutions.

Related Standards of Learning CH.1a, b, c

Objectives

The students will

- make the following measurements, using the specified equipment:
  - Volume: graduated cylinder, pipette, volumetric flask, burette;
  - Mass: electronic balance or triple-beam balance
  - Temperature: thermometer or temperature probe;
  - Pressure: barometer or pressure probe;
- identify, locate, and know how to use laboratory safety equipment, including aprons, goggles, gloves, fire extinguishers, fire blanket, safety shower, eye wash, broken glass container, and fume hood;
- demonstrate the following basic lab techniques: filtering, decanting, using chromatography, and lighting a gas burner;
- identify the following basic lab equipment: beaker, flask, graduated cylinder, test tube, test tube rack, test tube holder, ring stand, wire gauze, clay triangle, crucible with lid, evaporation dish, watch glass, wash bottle, and dropping pipette;
- understand Material Safety Data Sheet (MSDS) warnings, including handling chemicals, lethal dose (LD) hazards, chemical disposal, and chemical spill cleanup;
- demonstrate safe laboratory practices, procedures, and techniques.

Materials needed

- Safety posters displayed throughout the lab
- *Safety in Science Teaching* manual (see Resources at end of this lesson)
- Lab manual of safety procedures for each student
- Lab safety equipment (e.g., eye wash, safety shower, fire extinguisher, fire blanket) with appropriate signage
- PowerPoint presentation on lab safety (See Resources)
- Copies of the attached activity sheet
- Materials listed on the activity sheet

Instructional activity

Content

This standard provides an introduction to chemistry and safety procedures in the chemistry lab. Students are introduced to scientific vocabulary for chemistry, mathematical manipulations, and techniques for experimentation involving the identification and proper use of chemicals and equipment. They become familiar with the recommended statewide standards for high school laboratory safety. It is intended that students will actively develop scientific investigation, reasoning, and logic skills in the context of the key concepts presented in this standard.

Teacher Notes

The mixture for the "Percent Sand, Salt, Iron Filings, Mystery Substance in a Mixture" lab activity will need to be prepared prior to the beginning of the activity. Because the purpose of the activity is to practice safe techniques in the laboratory, exact measurements for the mixture are not necessary. When creating your mixture, keep in mind that you want the students to be able to separate the mixture and retrieve enough of each substance to be massed. The mystery substance should be an insoluble substance, such as small plastic pellets. Students will be able to separate the various materials by different densities.

Procedure

1. Design a demonstration that provides students with an opportunity to observe and identify laboratory safety concerns. Instruct students not to disclose any of their concerns until the completion of the demonstration. After the demonstration, the students can work in pairs to write their concerns and the possible consequences of not following proper safety procedures.
2. Have the students set up a KWL chart like the one shown below. Have the students first list what they Know about appropriate safety procedure, and then have them list what they Want to know about lab safety. After they have completed their charts, have the students share their "Knows" and "Wants to know," listing them on a KWL chart on the board or overhead.

KWL Chart — Topic: Lab Safety		
What I **Know**	What I **Want** to know	What I **Learned**

3. Show the PowerPoint presentation on lab safety (found at http://www.chem.unl.edu/safety/hslabcon.html). Discuss with the class the most important points.
4. Assist the class in developing a safety guide to be used in the laboratory. Use the Virginia Department of Education safety manual, *Safety in Science Teaching,* (found at http://www.pen.k12.va.us/VDOE/Instruction/safetymanual.pdf) as a reference.
5. Present a set of lab scenarios to the students, and review in relation to each scenario the prevention of accidents in the lab and proper responses to accidents when they happen. These scenarios should include
   - acid splashing into eyes;
   - hair catching on fire; and
   - broken glass cutting the skin and bleeding occurring.
6. At this point, you may wish to have the students develop skits related to various safety rules in the class guide, in which they demonstrate their knowledge and understanding of safety rules through their performances. Also, students not previously familiar with these rules will gain a deeper understanding of them from these skits.
7. Have the students complete their KWL charts by filling in the "What I Know" column. Then have them share their responses to fill in the class KWL chart.
8. The attached "Percent Sand, Salt, Iron Filings, Mystery Substance in a Mixture" activity is designed for students to practice techniques while separating mixtures, transferring solids and liquids quantitatively, filtering and washing solutes, and evaporating salt solutions to dryness. The activity may be adapted, based on the materials available to you.

Sample assessment

- Evaluate each student's laboratory technique during the lab activity.
- Observe students locating, identifying, and using safety equipment in the lab.
- Have students respond to questions such as: "What should you do first if your lab partner spills hydrochloric acid?"

Follow-up/extension

- Have each student make a safety-related poster that focuses on one of the main safety topics, such as the use of goggles during a lab. The poster should include the rule and a visual depiction of the rule, such as a cartoon, sketch, or photograph.

Resources

- PowerPoint presentation on lab safety. http://www.chem.unl.edu/safety/hslabcon.html.
- *Safety in Science Teaching.* Virginia Department of Education. http://www.pen.k12.va.us/VDOE/Instruction/safetymanual.pdf. Safety manual with sample documents.

Percent Sand, Salt, Iron Filings, Mystery Substance in a Mixture

Name: ____________________________________ **Date:** ______________________

Objectives

This activity is designed for you to practice techniques while separating mixtures, transferring solids and liquids quantitatively, filtering and washing solutes, and evaporating salt solutions to dryness.

Safety

1. You will use a variety of equipment and techniques in this activity. Make sketches of the equipment on the back of your activity sheet, and describe precautions you should be aware of before you work with them in the laboratory.
2. Read the procedure carefully. Write safety rules and precautions beside the steps in the procedure to highlight these before beginning the activity.
3. Obtain your teacher's approval on steps 1 and 2 before beginning the activity.

Materials

Electronic balance
Mixture in a cup
Magnet with material to cover
Erlenmeyer flask
Filter paper
Funnel
Pipette
Ring stand
Pipe stem triangle
Hot water
Hot plate and/or burner
Graduate cylinder
Wire gauze
Evaporating dish
Wash bottle

Procedure	Safety Rules and Precautions
1. Mass the cup containing the mixture, two separate sheets of filter paper, and a clean, dry Erlenmeyer flask. Record these masses on your table. Use the wrapped magnet to remove the iron filings. Record the mass of the iron filings.	
2. Prepare a filtering funnel with one sheet of filter paper, properly folding the paper. You may use a few drops of water to help position the paper in the funnel. This will be used to filter a water solution of the mixture. The flask will be used to capture the filtrate. Use a ring stand and pipe stem triangle to hold the funnel. Be certain the ring is cooled before use.	
3. As you rotate the funnel, add the mixture into the dampened funnel. Try to cover the bottom half of the funnel with the mixture. Place the funnel in the ring, and position the flask to capture the filtrate.	
4. Pour about 60 mL of hot water into the graduated cylinder, which has been placed in the sink, adding the hot water to it carefully. Wrap the graduated cylinder with several layers of paper towel to insulate it so you can transport it to your station safely before it cools.	

Procedure

5. Pour 5 to 10 mL of the hot water into the funnel, making sure the flask is underneath the funnel. IMPORTANT! Pour small amounts of the water into the funnel several times because it is more efficient to wash a system several times with small amounts of water than once with a large amount. Do not use more than 40 mL of water, as this will save evaporation time.
6. Devise a method for separating the sand from the mystery substance. You must separate the two substances and remove all the water — both the sand and the mystery substance must be completely dry before massing. Filtration will not work because both substances are insoluble in water. Remember to mass any piece of equipment prior to its use. If you need equipment not at your station, just ask your teacher for needed items.
7. Place the flask on wire gauze on the ring stand. Place the remaining filter paper on top of the flask to prevent splattering (place it on the flask when about 1/2 the liquid has been evaporated). *CAUTION! BE CAREFUL NOT TO ALLOW THE FILTER PAPER TO CATCH ON FIRE.* As you remove most of the liquid, the small amounts of liquid still present may generate steam, which can splatter large amounts of salt out of the dish when applying direct heat. Heat to complete dryness, and then stop heating immediately. Check with your teacher before discontinuing the heating.
8. Allow the dish and salt to cool to room temperature before massing. You should also find the masses of the sand and filter paper when they are completely dry. Complete all calculations, and answer the questions assigned at the end of the sample data sheet. Complete a full lab write-up.

Safety Rules and Precautions

Data Table

Measurements	Mass in grams (g)
Mass of cup & mixture	g
Mass of two sheets of filter paper	g
Mass of Erlenmeyer flask	g
Mass of the recovered iron filings	g
Mass of sample cup	g
Mass of filter paper (evaporation)	g
Mass of filter paper (filtration)	g
Mass of filter paper and sand (dried)	g
Mass of recovered sand	g
Mass of flask, filter paper, salt (dried)	g
Mass of recovered salt	g
Mass of recovered mystery substance	g

Calculations

Percent total mixture recovered:

(total mass of sand, salt, Fe filings recovered ÷ mass of total original mixture) x 100 =

Percent recovery of individual components:

(mass of sand recovered ÷ actual mass of sand) x 100 =

(mass of salt recovered ÷ actual mass of salt) x 100 =

(mass of Fe filings recovered ÷ actual mass of Fe) x 100 =

(mass of mystery substance recovered ÷ actual mass of mystery substance x 100 =

Lab Questions

On a separate sheet of paper, answer the following questions.

1. Define *filtrate, solution, solvent,* and *solute.* Which substances in this lab acted as each of these?
2. Describe the appearance of the filtrate during the evaporation phase. Try to explain what you saw.
3. What effect would using far too much water to dissolve the salt have on the results?
4. What new procedure would you follow if you discovered that the filter paper had torn and some bits of sand and paper were in the evaporating dish along with the filtrate?
5. What are the possible cause(s) for a sand recovery greater than 100% and a related salt recovery of less than 100%?
6. What might be the explanation for a salt recovery greater than 100% with a sand recovery at or very near 100%?
7. Which acts to dissolve the salt more completely: one large rinse of water or several small rinses of water? Why? What are the effects of hot versus cold rinse water?
8. What other errors or poor techniques might result in incorrect results?
9. If you were to redo this experiment, how would you change the procedure?
10. What might the mystery substance be?

Scientific Inquiry: Measurement/Data

(Adapted from *Planning by Design,* a series of lessons prepared by Richmond Public Schools. Used by permission.)

Organizing Topic Introduction to Chemistry

Overview Students focus on the concept of dependent and independent variables in experimental designs by responding to proposed experimental problems with the proper variable combinations. Students also concentrate on the importance of mass and measurement.

Related Standards of Learning CH.1d, e

Objectives

The students will

- record and interpret data from experiments in the form of bar, line, and circle graphs;
- demonstrate research skills, using a variety of resources;
- identify independent and dependent variables, constants, controls, and repeated trials;
- make valid conclusions after analyzing data;
- use research methods to investigate practical problems and questions;
- present experimental results in appropriate written form.

Materials needed

Demonstration materials:

- Two 250-mL beakers
- Warm tap water
- Red food coloring
- Dropper
- Bleach (sodium hypochlorite 5%)

Laboratory materials per group:

- Four antacid tablets
- Two 250-mL beakers
- Forceps
- Stopwatch

Instructional activity

Content/Teacher Notes

Chemistry students should have a complete understanding of experimental design and the terminology involved. Students should know and be able to apply the following terminology:

variable. A factor that is changed in an experiment.

independent variable. The variable that is purposely changed. Each change of a variable is known as a *level* of independent variable.

dependent variable. The variable that changes as a result of changing the independent variable.

hypothesis. A predication about how changing the independent variable will affect the dependent variable. Hypotheses are based on observations, previous experimental results, and information from books and communication with other scientists. A hypothesis is usually an if/then statement in the following form: "If the (independent variable) is (name the change), then the (dependent variable) will (name the effect of the change)." When the hypothesis and the experimental results agree, the hypothesis is supported by the results; when the hypothesis and the experimental results do not agree, the hypothesis is not supported by the results.

constants. The various factors in an experiment that do not change.

control. An unmanipulated group that is the standard for comparison in an experiment.

data. Information collected from the experiment. Data can be a collection of measurements or counts. Measurements are taken using instruments of the metric system. There is no such thing as a perfect measurement or a measurement that is free from error.

repeated trials. The number of times each *level* of the independent variable is tested. Repeated trials are conducted to reduce the effect of errors: repeated trials increase the reliability of the results of an experiment. The greater the number of repeated trials, the more confidence you can place in your data when you say that the hypothesis was or was not supported.

average (mean). When repeated measurements or counts are made, you summarize the data by finding the average or mean. Average (mean) = sum of measurements or sum of counts ÷ number of trials.

Introduction

1. Before class, fill two clear 250 mL beakers with 125 mL of warm water. Add one drop of red food coloring to one beaker. Add 3 or 4 drops of household bleach (sodium hypochlorite) to the other beaker.
2. Show the students the two beakers, and ask them what they think is in each of them. List their answers on the board.
3. After the students have given you various answers, mix the two solutions, and note the reaction. Ask students to make some observations about what they saw happen.
4. Facilitate a class discussion until students arrive at the conclusion that sodium hypochlorite is a chemical that will cause food coloring to “disappear.” Write the chemical formula for sodium hypochlorite, $NaClO$, on the board.
5. Ask students what would have been different about the outcome if you had used a different amount of sodium hypochlorite.
6. Use the Four Question Strategy to set up a procedure for this experiment:
   #1 What materials are readily available for conducting experiments on _________________?
   #2 How do __________________________ act?
   #3 How can you change the set of _________________ materials to affect the actions?
   #4 How can you measure or describe the responses of _________ to the change?
7. To begin an investigation, the teacher may choose one independent variable from question #3 and one dependent variable from question #4 to perform the experiment as a demonstration for students to observe results. All other variables in question #3 must remain the same: they are the constants.

Procedure

1. Give pairs of students the opportunity to design their own experiment regarding the reaction rate of antacid tablets when reacting with water. Give each pair the following materials at their lab station: four antacid tablets, two 250-mL beakers, forceps, and a stopwatch. They may use other materials if they wish, such as an electronic balance, balance paper, watch glass, and graduated cylinder.
2. Have student pairs design their experiment, identifying the independent variable, the dependent variable, and the control. Have them clearly specify the factors that must remain constant throughout their experiment. They must also provide a clear description of the procedure they will use to perform their procedure and clearly express the hypothesis they will be testing.

3. Validate the students' procedures prior to their using any chemicals or equipment in the lab.
4. After the experiments have been conducted and the data have been collected and analyzed, have students prepare to share their results with the class.

Sample assessment

- Assess measurement and data organization skills through lab work and the related data tables and conclusion questions.
- Give a quiz on experimental design to assess the students' abilities to label the dependent and independent variables.
- Give a lab practical covering measurement to help assess skill level and laboratory competency.

Follow-up/extension

- Have students present and conduct their own original research on a practical problem that they face everyday, using independent and dependent variables and a making a hypothesis.

Resources

Suggested Web sites with information on writing lab reports:

- http://www.columbia.edu/cu/biology/faculty/mowshowitz/howto_guide/lab_report.html.
- http://slhs.lafourche.k12.la.us/mstp/Chemistry/How%20to%20write%20a%20lab%20report.pdf.
- http://www.science.smith.edu/departments/Chem/Courses/labreports.htm.
- http://www.mhhe.com/biosci/genbio/maderinquiry/writing.html.

Organizing Topic — Atomic Structure

Standards of Learning

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include

e) accurate recording, organization, and analysis of data through repeated trials;
f) mathematical and procedural error analysis; and
h) use of appropriate technology including computers, graphing calculators, and probeware, for gathering data and communicating results.

CH.2 The student will investigate and understand that the placement of elements on the periodic table is a function of their atomic structure. The periodic table is a tool used for the investigations of

a) average atomic mass, mass number, and atomic number;
b) isotopes, half lives, and radioactive decay;
c) mass and charge characteristics of subatomic particles;
g) electron configurations, valence electrons, and oxidation numbers; and
i) historical and quantum models.

Essential Understandings, Knowledge, and Skills

Correlation to Textbooks and Other Instructional Materials

The student will use hands-on investigations, problem solving activities, scientific communication, and scientific reasoning to

- review location, charge, and relative size of subatomic particles — electron, proton, and neutron;
- examine the periodic table in regard to the following:
  - The atomic number of an element is the same as the number of protons.
  - In a neutral atom, the number of electrons is the same as the number of protons.
  - The average mass for each element is the weighted average of that element's naturally occurring isotopes.
- calculate relative atomic mass;
- explain that an isotope is an atom that has a different number of neutrons than is found in other atoms of the same element and that while some isotopes are radioactive, many are not;
- determine the half life of a radioactive substance;
- describe alpha, beta, and gamma radiation with respect to penetrating power, shielding, and composition;
- recognize that discoveries and insights have changed the model of the atom over time;
- explain the emergence of modern theories based on historical development;
- understand and demonstrate
  - MSDS warnings

  - safety rules for science
  - laboratory safety cautions
  - safe techniques and procedures; ______________________________
- relate the following major insights regarding the atomic model to the principal scientists listed below:
  - Particles: Democritus ______________________________
  - First atomic theory of matter: John Dalton ______________________________
  - Discovery of the electron: J. J. Thomson ______________________________
  - Discovery of the nucleus: Ernest Rutherford ______________________________
  - Discovery of charge of electron: Robert Millikan ______________________________
  - Planetary model of atom: Niels Bohr ______________________________
  - Periodic table: Dmitry Mendeleev, Henry Moseley ______________________________
  - Quantum of energy: Max Planck ______________________________
  - Uncertainty principle: Werner Heisenberg ______________________________
  - Wave theory: Louis de Broglie. ______________________________

Atomic Structure: Elements

Organizing Topic Atomic Structure

Overview Students focus on the structure of the periodic table and the valance electrons in an element. Using this information with the octet rule, students use the power of the table to predict ion formation and the oxidation numbers associated with the ions formed.

Related Standards of Learning CH.2g

Objectives

The students will learn about and gain experiences with the following scientific principles:

- Electrons are added one at a time to the lowest energy levels first (Aufbau Principle).
- An orbital can hold a maximum of two electrons (Pauli Exclusion Principle).
- Electrons occupy equal-energy orbitals so that a maximum number of unpaired electrons results (Hund's Rule).
- Energy levels are designated 1–7. Orbitals are designated s, p, d, and f according to their shapes
- s, p, d, f orbitals relate to the regions of the periodic table.
- Loss of electrons from a neutral atom results in the formation of an ion with a positive charge (cation).
- Gain of electrons by a neutral atom results in the formation of an ion with a negative charge (anion).
- Transition metals can have multiple oxidation states.
- Matter occurs as elements (pure), compounds (pure), and mixtures, which may be homogeneous (solutions) or heterogeneous.
- Important physical properties are density, conductivity, melting point, boiling point, malleability, and ductility.
- Reactivity is the tendency of an element to enter into a chemical reaction.
- Discoveries and insights related to the atom's structure have changed the model of the atom over time.
- The modern atomic theory is called the Quantum Mechanical Model.
- Major insights regarding the atomic model of the atom and the corresponding principal scientists include
    - particles: Democritus
    - first atomic theory of matter: John Dalton
    - discovery of the electron: J. J. Thompson
    - discovery of the nucleus: Ernest Rutherford
    - discovery of charge of electron: Robert Millikan
    - planetary model of atom: Niels Bohr
    - periodic table by atomic mass: Dmitry Mendeleev
    - periodic table by atomic number: Henry Moseley
    - quantum nature of energy: Max Planck
    - uncertainty principle: Werner Heisenberg
    - wave theory: Louis de Broglie.

Materials needed

- Periodic table of elements

- Clay in three different colors
- Toothpicks
- Six 12-inch balloons, two of each of three colors
- One 9-inch balloon of a different color
- String or ribbon
- Fluorescent bulb with some flaking of the white inner coating
- Light fixture for the bulb
- Bar magnet

Instructional activity

Content/Teacher Notes

This lesson requires three 90-minute blocks.

Introduction

1. Ask students to visualize the model of an atom as you draw an electron-shell diagram of a sodium atom. A sodium atom has 11 positively charged protons in its nucleus that will attract 11 electrons having negative electrical charges. The first two electrons attracted to the nucleus will occupy the first electron-shell or energy level as though they were racing around the surface of a Ping-Pong ball with a marble inside it.
2. Now, ask the students to imagine that the marble is the atom's nucleus and the surface of the Ping-Pong ball is the first electron-shell. The nucleus can attract 9 more electrons. However, as those electrons fall toward the nucleus, the electrons that are already there repel them. In order for the forces of attraction (to the nucleus) and repulsion (away from the other electrons) to be balanced, the approaching electrons must reside in a higher electron-shell farther away from the nucleus.
3. Next, ask the students to imagine the Ping-Pong-ball-marble-atom placed inside a hollowed-out orange. According to the mathematical calculations made by the Danish physicist Niels Bohr and his colleagues, the second electron-shell can hold up to eight electrons before other approaching electrons are pushed to even higher energy levels. Therefore, the last electron attracted to the nucleus of the sodium atom will occupy the third energy level as though our hollowed-out orange-Ping-Pong-ball-marble-atom were placed inside a basketball.
4. On the board, draw a model of electrons absorbing energy from one shell and leaping to the next shell, then falling back to the lower shell and giving off energy.
5. Have the students copy this concept into their lab notebooks.
6. Inform the students that Bohr's model of the atom was the first to explain why atoms of different elements give off specific colors of light when heated to very high temperatures: electrons leaping between energy levels emit specific frequencies of electromagnetic radiation.
7. On the board, write the formula $E = hv$, which is the formula for finding the amount of energy emitted by atoms, where E = energy, v = the frequency of the radiation, and h = Planck's constant.
8. Explain that the German physicist Max Planck discovered that the amount of energy in all kinds of electromagnetic energy is always a multiple of Planck's constant. This means that all forms of electromagnetic energy are transmitted in tiny packets, which Planck called "quanta."
9. Assist students in constructing a clay model of atoms, using clay in three different colors and toothpicks. Assign each student a different element to work with. Advise them to space the electrons as far apart as they can within each electron-shell since they are repelling one another at all times.

10. Explain that the maximum number of electrons that can fit in the electrons-shells of atoms in the first three periods of the periodic table is 2, 8, and 8, respectively.
11. Show students how to draw Bohr electron-shell diagrams and write the electron dot structure for each element that they modeled in clay.

Procedure

Part 1: Introduction to Orbitals

1. Explain that each of the energy levels holds sublevels and that each sublevel in turn holds characteristic orbitals. There are four kinds of sublevels — s, p, d, and f. Each of these sublevels can hold a characteristic number of electrons, as follows:
    - s can hold up to 2 electrons.
    - p can hold up to 6 electrons.
    - d can hold up to 10 electrons.
    - f can hold up to 14 electrons.

    Orbitals within these sublevels also have characteristic shapes. The s-orbitals have the shape of a sphere, while p-orbitals have the shape of a dumbbell, and d-orbitals have a clover-like shape, with one type appearing as a dumbbell within a doughnut. The f-orbitals are not well defined.
2. All of these orbitals layer on top of each other to form the electron cloud that is the outer portion of the atom. Each orbital's location is determined by the electrons that are characteristically found there, and each electron's location is determined by the energy it has. It is due to this layering that we describe the structure of the electron cloud as a series of nested spheres. Even though the orbitals take on nonspherical shapes, when the orbitals of a sublevel are combined, they form sphere-like shapes, which nest inside the next larger sphere-like shape. The location of these sphere-like shapes within the nested spheres is determined by their energy levels, that is, the energy levels of the electrons that are in them.
3. In this activity, students will form models of s- and p-orbitals, using balloons. Students should notice that while it takes one balloon to form the spherical s-orbital (holding a maximum of two electrons), each dumbbell-shaped p-orbital requires two balloons to form it. This is due to the limitations of working with balloons, which are quite large, when modeling electrons, which are quite small. To make it easier to see each of the p-orbitals, make sure students pair up two balloons of the same color (x, y, z).
4. Have students blow up the pairs of balloons to form teardrop shapes. Tell them not to blow them to their largest volume, as they will need room for flexibility.
5. Instruct the students to attach the pairs of balloons together with string to form dumbbell shapes. Have them slide the pairs of balloons together at their middles, aligning them to form the x, y, z orbitals of a p-sublevel.
6. Next, have each student blow up the smaller balloon to form a spherical s-orbital and then try to determine the best way to attach it to the p-orbital model in order to illustrate its true location.
7. Once students have figured out how to connect the orbitals, lead them in a discussion, using the following questions.
    - What did you observe about how easily the teardrop/dumbbell shapes fit together? Can you visualize the sphere in the space that the p-orbitals form?
    - Why was it so difficult to place the s-orbital in its proper place? Can it be done using balloons?
    - Where would you find the individual electrons if this were the p-orbital of the element nitrogen?
    - Can you tell which balloon the electron(s) would be in? Why, or why not?

- What do you think the third quantum level orbitals would look like?

Part 2: Orbital Diagrams and Electron Configurations

1. By now the students should be wondering how to put all of the information about orbitals on paper. Using the above activity as a springboard, draw some orbital diagrams on the board. Explain that all orbital diagrams do is use arrows to represent the spin of the electrons. Give them the rules for creating orbital diagrams:
   - The Pauli Exclusion Principle states that an orbital can hold a maximum of two electrons.
   - The Aufbau Principle states that electrons are added one at a time to the lowest energy orbitals available until all the electrons of the atom have been accounted for.
   - Hund's Rule states that electrons occupy equal energy orbitals so that a maximum number of unpaired electrons result.
2. Have the students practice drawing some orbital diagrams.
3. Use the periodic table to help students see the electron configurations. Show them how the periodic table can be divided by the orbitals. Point out that when writing the electron configurations, the sum of the superscripts in an electron configuration represents the total number of electrons in the atom.
4. Have the students practice writing as many electron configurations with their matching orbital diagrams as possible before the class period is over.

Part 3: Ionization, Valence Electrons, and Chemical Properties of Families of Elements

1. After reviewing the structure of atoms according to the Bohr model, explain that electrons in the last electron-shell are "vulnerable," i.e., if an atom has fewer than four electrons in its outer shell, it tends to "lose" them. Losing negatively charged electrons will leave an atom with an excess of positively charged protons in the nucleus, and these protons cannot leave under ordinary circumstances. (Changing the number of protons in the nucleus of an atom changes that atom into an atom of another element; this occurs only in nuclear reactions.) An atom with more protons than electrons becomes a positive (+) ion, called a "cation."
2. Explain that atoms with more than four electrons in their outer shell tend to "gain" electrons. This is because "gaps" remain in unfilled shells far from the nucleus, allowing extra electrons to be attracted to the "exposed" protons of the nucleus. Underscore that outer shells are most stable when they are filled. An atom with more electrons than protons becomes a negative (−) ion, called an "anion." These atoms can also share their electron-shells with other atoms.
3. Hand out periodic tables, and have students label the charges that each family/group of elements carries once the element has lost or gained an electron. Point out some of the physical and chemical properties that these families have because of the activity of their valence electrons.

Part 4: Glow in the Dark

1. In this activity, students will explore some of the properties of a fluorescent light bulb, which usually contains argon and mercury vapor. In a lighted room, have students examine the fluorescent bulb and note where the inner white coating has flaked off. (More information regarding the structure and function of a fluorescent light bulb may be found at http://home.howstuffworks.com/fluorescent-lamp2.htm.)
2. Have the students examine the bulb briefly in a dark room.
3. Return to the lighted room, and connect the bulb to the light fixture. *CAUTION! Unplug the fixture while inserting the bulb.*

4. Turn on the bulb, and have the students carefully observe how it lights up, noting any colors they see. Have them compare the colors they observe at the spots where the white coating has flaked off with those that they see on the rest of the coated bulb. *CAUTION! Avoid staring at the light for long time periods.*
5. Place the bar magnet against the glass bulb, and have students observe and note any changes in the light pattern.
6. Unplug the fixture, and carry it into the dark room.
7. Have students cover one eye while they turn on the bulb for two minutes. Then, have them turn off the bulb, uncover their eye, and observe the bulb closely in the dark.
8. End the lab with a discussion based on the following questions:
    - How does the appearance of the light coming from uncoated spots on the bulb compare to its appearance on the rest of the bulb? *(Light is more white and intense where there is coating on the glass and less white and intense where the coating is missing.)*
    - What evidence do you have that the gas inside the bulb contains charged particles? *(Since the gas appeared to be attracted or repelled by the magnet, one may speculate that there are charged or ionized particles in the bulb. Current flows through the bulb mainly by ionizing the inert gas in the bulb.)*
    - Discuss the properties of the elements in the bulb. Why are they used in a fluorescent bulb?
    - Describe the appearance of the fluorescent bulb after you turned it off in the dark room. What can you say about the coating on the bulb? *(The coating on the bulb continued to emit some light when the electricity was turned off.)*

Sample assessment

- Have students write a two-to-three-paragraph conclusion, summarizing what was illustrated in the labs and what information or insight was gained from it. Remind them not to summarize the procedure except as necessary to explain the conclusion.

Follow-up/extension

- Have students write one paragraph explaining the connection between the balloons and the locations of electrons and telling how this illustrates the structure of the electron cloud.
- Have students create a timeline of historical models of atoms.

Resources

Suggested Web site with information on electrons:

- http://www.iop.org.

Suggested Web site with information on electron configurations:

- http://library.thinkquest.org/3659/structures/electronconfig.html.

Suggested Web site with information on atomic structure:

- http://www.chem4kids.com/files/atom_structure.html.

Suggested Web site with information on ions:

- http://www.chem4kids.com/files/atom_ions.html.

Suggested Web sites with information on elements:

- http://environmentalchemistry.com/yogi/periodic.
- http://www.chemsoc.org/viselements/pages/pertable_j.htm.

Isotope Tic Tac Toe

Organizing Topic Atomic Structure

Overview Student pairs become "certified experts" in modeling the structure of a particular atom. The expert pairs then check other students' work as those students work individually to describe and model the structure of selected atoms while completing a game of Tic Tac Toe. The student Tic-Tac-Toe cards then provide the basis for calculating relative atomic masses as well as for introducing radioactive decay.

Related Standards of Learning CH.2

Objectives

The students will
- review location, charge, and relative size of subatomic particles — electron, proton, and neutron;
- examine the periodic table in regard to the following:
  - The atomic number of an element is the same as the number of protons.
  - In a neutral atom, the number of electrons is the same as the number of protons.
  - The average mass for each element is the weighted average of that element's naturally occurring isotopes.
- calculate relative atomic mass;
- explain that an isotope is an atom that has a different number of neutrons than is found in other atoms of the same element and that while some isotopes are radioactive, many are not;
- determine the half life of a radioactive substance;
- describe alpha, beta, and gamma radiation with respect to penetrating power, shielding, and composition;
- recognize that discoveries and insights have changed the model of the atom over time;
- explain the emergence of modern theories based on historical development;
- relate the following major insights regarding the atomic model to the corresponding principal scientists:
  - particles: Democritus
  - first atomic theory of matter: John Dalton
  - discovery of the electron: J. J. Thompson
  - discovery of the nucleus: Ernest Rutherford
  - discovery of charge of electron: Robert Millikan
  - planetary model of atom: Niels Bohr
  - periodic table by atomic mass: Dmitry Mendeleev
  - periodic table by atomic number: Henry Moseley
  - quantum nature of energy: Max Planck
  - uncertainty principle: Werner Heisenberg
  - wave theory: Louis de Broglie.

Materials needed

- Colored "particle poker chips" with disc magnets attached to them
- 16-cell Tic-Tac-Toe card for each student
- Magnetic white board
- Periodic table of the elements for each student

- Colored pencils or crayons

Instructional activity

Content/Teacher Notes

A simple classroom model for atomic structure can be assembled by using a magnetic white board and poker chips with small disc magnets glued to them to represent protons, neutrons, and electrons. Use the model to review the location, charge, and relative mass of protons, neutrons and electrons. It is not necessary for students to understand electron configurations at this point; simply model an electron cloud around and some distance away from the nucleus.

Each of the 16 cells on the student Tic-Tac-Toe card identifies a specific atom by giving the element's name and the mass number of the isotope and providing space to list the numbers of protons, neutrons, and electrons in the atom. The cells also include information on percent relative abundance for naturally occurring elements or information on half life and decay for radioactive isotopes. See example at right.

Lithium-6
# of protons:__
# of neutrons:__
# of electrons:__
% abundance: 8%

Before undertaking the activity, make "particle poker chips" from construction paper by marking them as follows: + on blue chips (representing protons), − on red chips (representing electrons), and no mark on white chips (representing neutrons).

Introduction

1. Distribute "particle poker chips," periodic tables, and Tic-Tac-Toe cards. Select one of the cells on the card to complete as a class. Let students make models at their desks, and then ask them to check for accuracy by comparing their model to the classroom model. Have students draw a representation of the model in their data books.
2. Refer to other cells on the Tic-Tac-Toe card to introduce or review the term *isotope*. Explain that the relative atomic masses shown on the periodic table of the elements are determined from the percent abundances of that element's isotopes as they exist on Earth. Point out that the isotopic ratios can be strongly affected by the source of the element and that the elemental isotopic abundances elsewhere in the universe are different.
3. Ask students to identify the isotopes on the card for which percent abundance information is given. Then point out the radioactive isotopes, and direct students to use a yellow pencil or crayon to shade each of the cells containing a radioactive isotope.

Procedure

1. Certifying "expert" teams: Assign an isotope to each team, and direct students to complete that isotope's information in the cell and make a poker-chip model of it. As the teams work, move from team to team, checking their work for accuracy. Initial the completed cell on each student's card to "certify" him/her as an "expert."
2. Have students draw a representation of their certified accurate model in their data books so they will have it available for checking other students' work during the Tic-Tac-Toe game. For classes of fewer than 32 students, you will be the expert for the remaining isotopes.
3. Playing Tic-Tac-Toe: Have each student work individually to complete a vertical, horizontal, or diagonal Tic-Tac-Toe by filling in information for an isotope on the card, making a model, and getting a signature from one of the isotope's experts to verify the accuracy of the work. Have students make a record of each of their models in their data books.

Observations and Conclusions

1. Have students calculate the relative atomic mass of elements, using the percent abundance data on the cards, and compare the calculated masses to the atomic masses on the periodic table.
2. Explain to students that half life must be determined experimentally by following the decay of a radionuclide over time, and that databases may include more than one value for the half life of a particular isotope based on the reported results of different experiments.
3. Ask students to list the radioactive isotopes on their cards in order of decreasing half life, and discuss the range of half-life periods. Point out that because some of the radioisotopes decay so fast, they are not found in nature although they can be observed as the product of some nuclear reactions. You may wish to have the students construct decay curves for some of the radioisotopes, based on their half lives.
4. Direct students to list the different types of radioactive decay shown on the cards: alpha decay, beta decay, and electron capture. Students can use poker chips to demonstrate radioactive decay if they modify their poker-chip models by using a proton chip and an electron chip stacked on top of each other to represent a neutron. Examples include:
    - Hydrogen-3 decays by the conversion of a neutron to a proton and emission of an electron, forming helium-3.
    - The nucleus of an atom with too few neutrons may gain one more neutron by capturing one of the negatively charged electrons orbiting about the nucleus. This effectively cancels the positive charge on one of the protons, turning it into a neutron. An example of this kind of radioactivity is the decay of beryllium-7 to form lithium-7.
    - Boron-8 decays to form helium-4 by electron capture and alpha emission.
5. Describe penetrating power and shielding of alpha and beta emissions. Explain that after a nuclear decay, the nucleus may still have excess energy to shed. This energy can be given off in the form of a pulse of electromagnetic radiation, called "gamma radiation," with no change in mass or charge. Compare the penetrating power and shielding of gamma radiation to alpha and beta emissions.

Sample assessment

- Give students new isotopes, as in the examples shown below, and assess their ability to do the following:
    - Determine the number of protons, neutrons. and electrons in the atom.
    - Model the structure of the atom.
    - Calculate the relative atomic mass of the atom.
    - Model radioactive decay of the atom.
    - Construct a decay curve for the atom.

Nitrogen-14	**Nitrogen-15**	**Nitrogen-16**	**Nitrogen-17**
# of protons:__	# of protons:__	# of protons:__	# of protons:__
# of neutrons:__	# of neutrons:__	# of neutrons:__	# of neutrons:__
# of electrons:__	# of electrons:__	# of electrons:__	# of electrons:__
% abundance: 99.6%	% abundance: 0.04%	Half life: 7.13 s	Half life: 4.75 s

Follow-up/extension

- The relative masses of atoms are measured using a mass spectrometer. The history of mass spectrometry clearly illustrates the emergence of modern theories based on historical development. The first mass spectrometer was invented in J. J. Thomson's lab at Cambridge at the end of the nineteenth century. Mass spectrometry was used to discover the existence of isotopes of nonradioactive elements. Modern radiometric dating employs accelerator mass spectrometers that can count each particle of a sample and separate all the isotopes, making them useful in radiometric dating. In this connection, you may have students do one or more of the following:

  - Investigate the contributions of Democritus and John Dalton to the atomic model that Thomson used.
  - Relate Rutherford's discovery of the proton, Moseley's arrangement of the elements on Mendeleev's periodic table by atomic number, and Bohr's planetary model to an explanation of the basis for mass spectrometry.
  - Research the relationship between Thomson's discovery of the electron and his contribution to invention of the mass spectrometer as well as the contribution of Millikan to clarifying the relationship.
  - Describe insights regarding electron structure that have emerged as the principles and applications of mass spectrometry have evolved.

- Have students make posters describing the work of one or more of these scientists and showing the model that they used to visualize the atom. Posters can be displayed chronologically to form a timeline on the classroom wall, thus illustrating the historical development of a model for atomic structure from the ancient Greeks to the present.

Resources

- *The Berkeley Laboratory Isotopes Project: Exploring the Table of Isotopes.* http://ie.lbl.gov/education/isotopes.htm.
- *Physics 2000, Science Trek, Isotopes and Radioactivity.* http://www.colorado.edu/physics/2000/isotopes/.
- *WebElements™ Periodic Table.* http://www.webelements.com/.

Radioactive Decay and Half Life

Organizing Topic Atomic Structure

Overview Students model the rate of decay of radioactive isotopes, using a penny model.

Related Standards of Learning CH.2.b

Objectives

The students will
- review location, charge, and relative size of subatomic particles — electron, proton, and neutron;
- examine the periodic table in regard to the following:
  - The atomic number of an element is the same as the number of protons.
  - In a neutral atom, the number of electrons is the same as the number of protons.
  - The average mass for each element is the weighted average of that element's naturally occurring isotopes.
- explain that an isotope is an atom that has a different number of neutrons than other atoms of the same element and that some isotopes are radioactive but many are not;
- determine the half life of a radioactive substance.

Materials needed

For each group of students:
- Container with top
- 100 pennies
- Plastic cup
- Periodic table of the elements
- Attached table of isotopic decay types and half lives

For each student:
- Graph paper
- Ruler
- Attached activity sheet

Instructional activity

Content/Teacher Notes

Before doing this activity, students need to have had instruction in the types of radioactive decay and in the definition of *half life*. Common isotopes to use in this activity are carbon-14, iodine-131, cobalt-60, hydrogen-3, strontium-90, and uranium-238, although any radioactive isotope with a known decay type and half life can be used. It works well to identify the isotope with a sticky note on the top of each group's container.

Procedure

1. Place the students into groups, and have each group perform the activity described on the attached activity sheet. Pennies represent atoms of the given isotope, and any penny that comes up tails upon turning the container has decayed to a new element.
2. Have each student complete the data table, graph, and questions on the activity sheet.

Sample assessment

- Assess the students' completed activity sheets.

Follow-up/extension

- Have students investigate the uses and dangers, if any, of the isotope that they used in the model.

Resources

- *The Berkeley Laboratory Isotopes Project: Exploring the Table of Isotopes.* http://ie.lbl.gov/education/isotopes.htm.
- *Physics 2000, Science Trek, Isotopes and Radioactivity*. http://www.colorado.edu/physics/2000/isotopes/.

Radioactive Decay and Half Life
Activity Sheet

Name: ______________________________________ **Date:** ________________________

Instructions

The 100 pennies in your group's container represent the atoms of a radioactive isotope.

1. Seal the container, and turn it over six times. This represents one half-life period.
2. Remove any pennies that come up tails, and place them in a plastic cup. These pennies represent those atoms that have undergone radioactive decay.
3. Count the heads-up pennies that remain in the original container, and record the number in the data table below.
4. Repeat steps 1–3 with the remaining pennies to represent three additional half-life periods.

Data Table

Half-Life Period	Time (sec.)	Atoms Remaining	Mass of Atoms (amu)
0	0	100	
1			
2			
3			
4			

Data Analysis

On graph paper, graph mass versus time from your data table. Plot all points, and then use the ruler to connect them with a line of best fit. Be sure to label each axis and title your graph.

Conclusions

1. Write the nuclear decay equation for the radioisotope that you were given.

2. For your isotope, find the amount of time that elapses in 3.5 half-life periods. Show your work.

3. Ancient geological formations are often dated by finding the amount of certain uranium isotopes contained in the rock layer.
    - Why are uranium isotopes useful in determining the age of ancient geological formations?
    - How can radioactive dating be useful when the temperatures and pressures to which the geological formation has been exposed have varied so much throughout history?

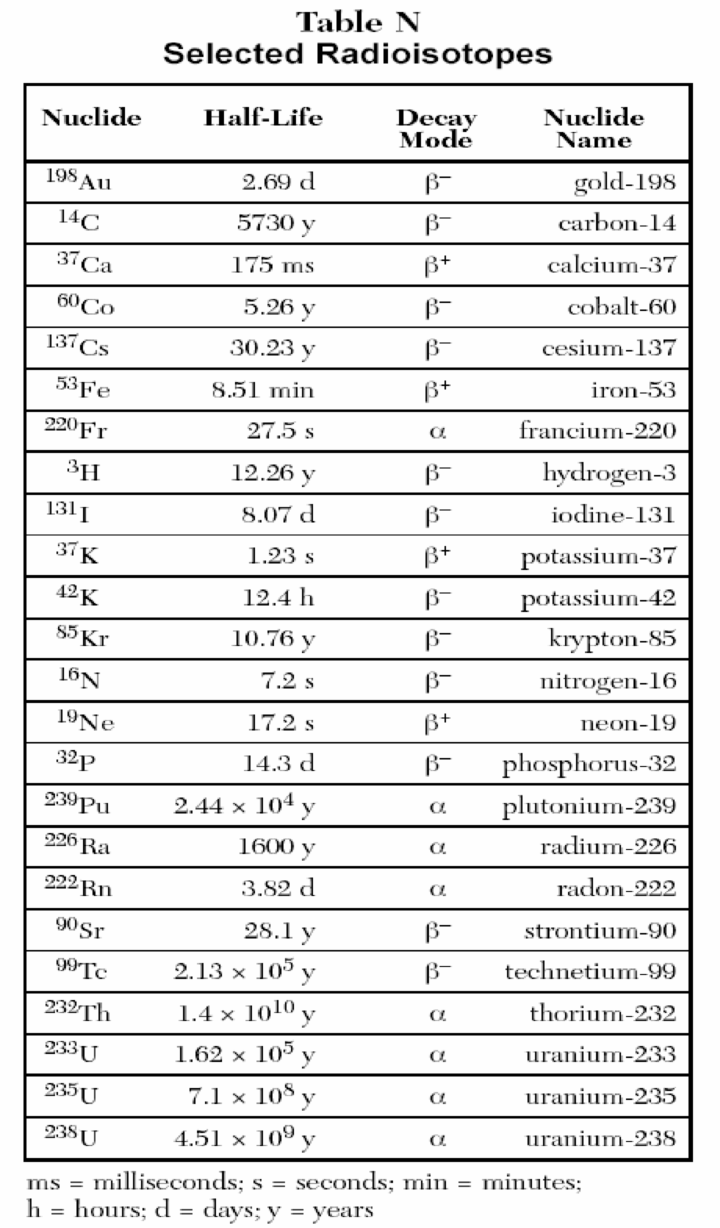


Table N
Selected Radioisotopes

Nuclide	Half-Life	Decay Mode	Nuclide Name
^{198}Au	2.69 d	β^-	gold-198
^{14}C	5730 y	β^-	carbon-14
^{37}Ca	175 ms	β^+	calcium-37
^{60}Co	5.26 y	β^-	cobalt-60
^{137}Cs	30.23 y	β^-	cesium-137
^{53}Fe	8.51 min	β^+	iron-53
^{220}Fr	27.5 s	α	francium-220
^{3}H	12.26 y	β^-	hydrogen-3
^{131}I	8.07 d	β^-	iodine-131
^{37}K	1.23 s	β^+	potassium-37
^{42}K	12.4 h	β^-	potassium-42
^{85}Kr	10.76 y	β^-	krypton-85
^{16}N	7.2 s	β^-	nitrogen-16
^{19}Ne	17.2 s	β^+	neon-19
^{32}P	14.3 d	β^-	phosphorus-32
^{239}Pu	2.44×10^4 y	α	plutonium-239
^{226}Ra	1600 y	α	radium-226
^{222}Rn	3.82 d	α	radon-222
^{90}Sr	28.1 y	β^-	strontium-90
^{99}Tc	2.13×10^5 y	β^-	technetium-99
^{232}Th	1.4×10^{10} y	α	thorium-232
^{233}U	1.62×10^5 y	α	uranium-233
^{235}U	7.1×10^8 y	α	uranium-235
^{238}U	4.51×10^9 y	α	uranium-238

ms = milliseconds; s = seconds; min = minutes;
h = hours; d = days; y = years

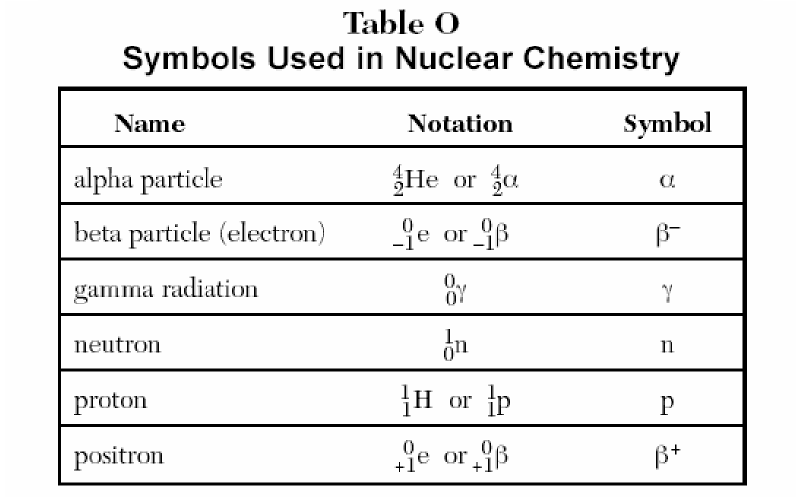


Table O
Symbols Used in Nuclear Chemistry

Name	Notation	Symbol
alpha particle	$^{4}_{2}He$ or $^{4}_{2}\alpha$	α
beta particle (electron)	$^{0}_{-1}e$ or $^{0}_{-1}\beta$	β^-
gamma radiation	$^{0}_{0}\gamma$	γ
neutron	$^{1}_{0}n$	n
proton	$^{1}_{1}H$ or $^{1}_{1}p$	p
positron	$^{0}_{+1}e$ or $^{0}_{+1}\beta$	β^+

Organizing Topic — Properties of Matter

Standards of Learning

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include

a) designated laboratory techniques;
b) safe use of chemicals and equipment;
c) proper response to emergency situations;
d) manipulation of multiple variables, using repeated trials;
e) accurate recording, organization, and analysis of data through repeated trials;
f) mathematical and procedural error analysis;
g) mathematical manipulations (SI units, scientific notation, linear equations, graphing, ratio and proportion, significant digits, dimensional analysis); and
h) use of appropriate technology including computers, graphing calculators, and probeware, for gathering data and communicating results.

CH.2 The student will investigate and understand that the placement of elements on the periodic table is a function of their atomic structure. The periodic table is a tool used for the investigations of

h) chemical and physical properties.

CH.5 The student will investigate and understand that the phases of matter are explained by kinetic theory and forces of attraction between particles. Key concepts include

c) phase changes;
d) molar heats of fusion and vaporization;
e) specific heat capacity; and
f) colligative properties.

Essential Understandings, Knowledge, and Skills

Correlation to Textbooks and Other Instructional Materials

The student will use hands-on investigations, problem solving activities, scientific communication, and scientific reasoning to

- understand that matter is classified by its chemical and physical properties;
- differentiate between physical and chemical properties, using common examples;
- observe and classify matter as elements, compounds, heterogeneous mixtures, or homogeneous mixtures (solutions);
- recognize the following physical properties: density, conductivity, melting point, boiling point, malleability, ductility, and specific heat capacity;
- use probeware to gather data;
- collect volume, mass, and temperature measurements, using appropriate equipment;
- understand and demonstrate
  - MSDS warnings;

  - safety rules for science; ____________________________
  - laboratory safety cautions; ____________________________
  - safe techniques and procedures; ____________________________
- demonstrate the following basic lab techniques: filtering, decanting, using chromatography, lighting a gas burner; ____________________________
- interpret a heating curve graph; ____________________________
- calculate energy change and specific heat; ____________________________
- understand that the solid, liquid, and gas phases of a substance have different energy content; ____________________________
- review location and use of safety equipment; ____________________________
- demonstrate precision in measurement; ____________________________
- understand accuracy in terms of closeness to the true value of a measure. ____________________________

Heat Transfer and Heat Capacity

(From “Heating Curve Lab” and “What Are You Eating?” by Catherine Beck, found on her Web site *Science Education at Virginia Tech* at http://filebox.vt.edu/users/ckeel/home.htm. Used by permission.)

Organizing Topic Properties of Matter

Overview Students collect and analyze data related to heat transfer.

Related Standards of Learning CH.1a, b, g; CH.2i; CH.5c, d, e

Objectives

The students will

- construct a heating curve and explain its components;
- perform calculations involving ΔH_{vap} and ΔH_{fus};
- review phase changes and quantify energy differences;
- review the phases of matter and their energy content;
- understand that specific heat capacity is a property of a substance;
- calculate changes in energy, using heat capacity (C_p) and calorimetry in a lab;
- relate calculations to nutrition and calorie content.

Materials needed

- Two attached lab worksheets
- Materials listed on the lab worksheets

Instructional activity

Part 1: Heating Curve

Introduction

1. Review concepts related to heat transfer.
2. Have students brainstorm phase changes they know.

Procedure

1. Pass out the attached “Heating Curve” lab worksheet, and let students read it over. Go over the procedures and safety issues involved in the lab: Hot plates can become very hot, so be careful not to touch them. Exercise care with hot water and beakers; use wire mesh to set beakers down. Wear goggles and aprons, and tie back long hair. Keep cords away from heat sources.
2. Have students acquire materials and perform the lab. Provide guided instruction and assistance as needed.

Observations and Conclusions

1. Discuss the inquiry questions from the lab and the lab-report format.
2. Using the graph from the lab, explain ΔH_{fus} and ΔH_{vap}. Work some sample problems together as a class.

Part 2: What Are You Eating?

Introduction

1. Review SOL periodic table information.

Procedure

1. Pass out the attached "What Are You Eating?" lab worksheet, and let students read it over. Go over the safety issues involved in the lab. Have groups choose the food with which they will work.
2. Have students acquire materials and perform the lab. Provide guided instruction and assistance as needed. As students complete the lab, have them work in groups on the calculations and inquiry questions.
3. Using lab format, explain the "Specific Heat" section of the worksheet. Provide notes, and work some sample problems together as a class. Have students complete the worksheet.
4. Have students practice C_p problems in groups (10 min.). Answer questions, as needed.

Observations and Conclusions

1. Discuss the inquiry questions from the lab and the lab-report format.
2. Using the graph from the lab, explain ΔH_{fus} and ΔH_{vap}. Work some sample problems as a class.

Sample assessment

- For homework, have students finish their labs and analyses to turn in at the next class period. Have students practice C_p problems to turn in at the next class period.
- Observe students during the lab work, and assess their proficiency.
- Have the students write lab reports according to the format already discussed, and assess the reports according to the lab rubrics shown below.

Follow-up/extension

- Have students practice ΔH_{fus} and ΔH_{vap} problems in groups. Check answers quickly.

Lab Rubrics

Heating Curve Lab

____ 50 points total	Construct a heating curve Two graphs, checked for accuracy (15 pts ea.) Safe and managed completion of the lab during class (20 pts)
____ 20 points total	Phase changes and energy transfer Inquiry questions following lab answered correctly (2 pts ea.)
____ 30 points total	ΔH_{fus} and ΔH_{vap} calculations Problems given on extra sheet completed correctly (5 pts ea.)

What Are You Eating? Lab

____ 40 points	Practice of safe lab procedures and efficient time management Completion of lab and clean up
____ 20 points	Data recorded in lab notebook At least two samples neatly recorded (10 pts ea.)
____ 20 points	Analysis and comparison of own data Correct completion of lab questions (5 pts ea.)
____ 10 points	Comparison among class data Other data present and commented on
____ 10 points	Clear summary Summary present and incorporating ideas addressed in lab

Heating Curve

Name: ______________________________ **Date:** ____________________

Prediction
What will the "heating curve" of water look like as you add constant heat to the water over time? Show your prediction on the graph at right, including appropriate title, labels, units, and scale.

Materials Needed
Graphing calculator or computer
Data-collection device (calculator-based lab [CBL]) with
    temperature probe
250-mL beaker
100-mL graduated cylinder
Ring stand and clamp
Slit stopper
Hot plate or burner
Wire mesh
Ice

Procedure

1. Collect the materials listed above. Put on goggles and aprons, and tie back long hair. *CAUTION! The hot plate will get quite hot, so be careful not to touch it! Exercise care with hot water and beakers. Use wire mesh on which to set beakers down. Keep cords away from the heat source.*
2. Set up the CBL as instructed, and open the program on the calculator as directed.
3. Place the beaker with 35 mL water and ~3 ice cubes onto the hot plate. Suspend the probe in the water, being very careful not to let it touch the sides or bottom of the beaker.
4. Start the CBL data-collection program, and turn the hot plate or burner on the highest setting. Record the temperature every 30 seconds in the following table:

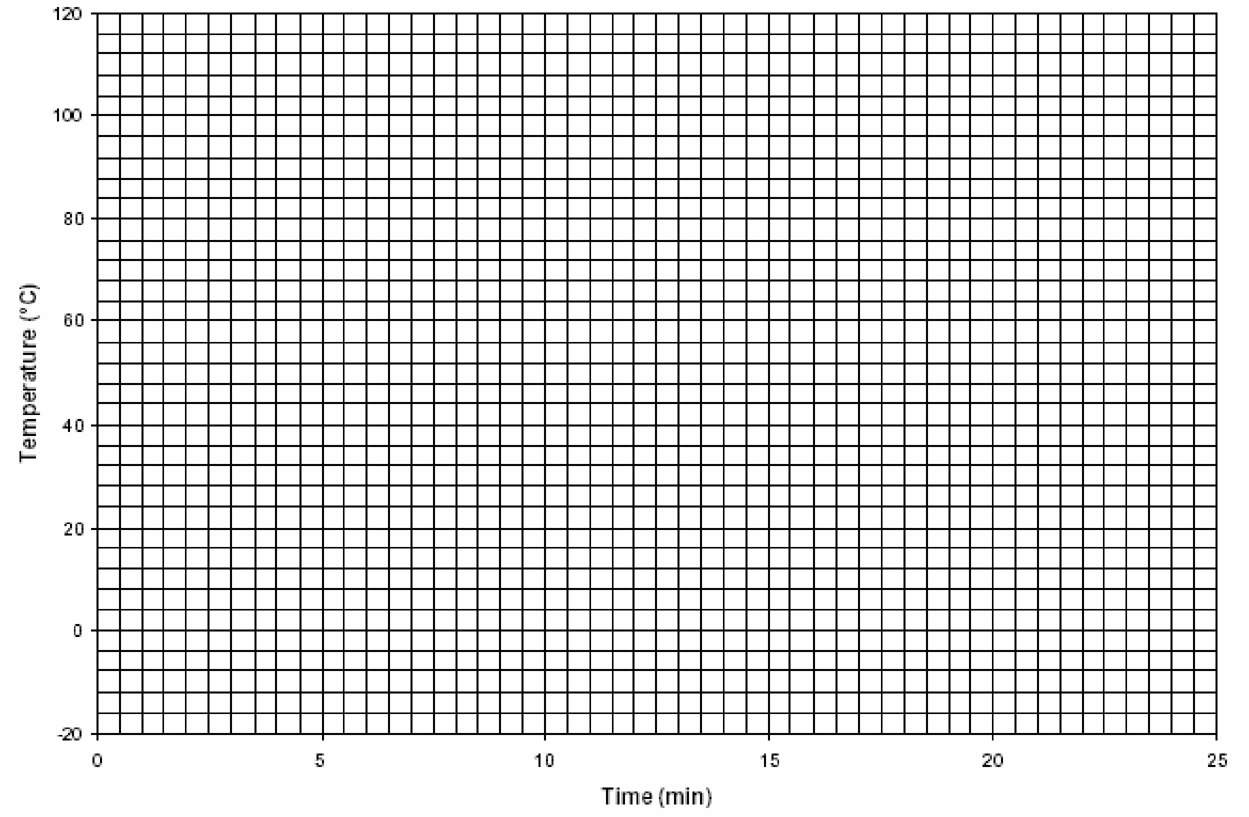
Time (sec.)	**Temperature (°C)**
0	
30	
60	
90	
120	
150	
180	
210	
240	
270	
300	
330	
360	
390	
420	
450	
480	
510	
540	
570	
600	
630	
660	
690	
720	
750	
780	
810	
840	
870	
900	
930	
960	
990	
1,020	
1,050	
1,080	
1,110	
1,140	
1,170	
1,200	
1,230	
1,260	
1,290	
1,300	

5. When the readings level off, take five readings at one temperature.
6. Then, turn off the heat, and carefully remove the beaker from the heat source, setting it on wire mesh. Clean up your area.

Questions

1. Recall the different states of matter (solid, liquid, gas). How do the water molecules differ in the liquid and gas states? Draw and explain.

2. Plot your temperature and time data on the graph below.



3. Label the phase changes on your heating curve above.
4. What happens to the molecules as they begin to boil?

5. Did you stop adding heat at any point during the lab? _______ As heat was added, what happened to the energy of the system?

6. If you were always adding constant heat, why did the temperature trend change?

7. What happens to the energy being absorbed from the heat source? Use the heating curve and your knowledge of atoms to explain.

ΔH_{vap} Practice: $\Delta H = m \Delta H_{vap}$

1. How much heat is needed to vaporize 250 g of water at 100°C and 101.3 kPa pressure?

2. When a quantity of water vapor at 100°C and 101.3 kPa is condensed to the liquid phase, 1.81 × 105 J of heat is released. What mass of water is condensed?

3. What quantity of heat in joules is required to vaporize 600 g of water at 100°C and 101.3 kPa?

What Are You Eating?

Name: ______________________________________ **Date:** _______________________

Instructions

You will be using a metal can filled with water as a "calorimeter." The change in the temperature of the water will tell you how much energy your food contains. *CAUTION! Be careful of the open flame and very hot food!*

Materials Needed

Calculator-based labs (CBL)
Graphing calculator
Matches
Ring stand
Can
Cold water
Marshmallows
Popcorn
Paperclip
Food holder
Temperature probe
Balance

Procedure

1. Obtain and put on goggles and apron.
2. Follow CBL setup instructions found on the lab bench. A couple of steps will be different: Instead of 41 samples every 30 seconds, take 36 samples every 5 seconds. Also, instead of Y values being 10° to 105°C, use Y values of 5° to 50°C.
3. Choose which food your group wants to investigate. Repeat the procedure below three times for each type of food.
4. Carefully measure out 75 mL of cold water, and pour it into the can. Place the can on the ring stand.
5. Measure the initial water temperature with the temperature probe. Wait until the temperature stabilizes, and record it in the chart below.
6. Place the food on the paperclip, place it in the food holder, and mass the food plus holder. Record the initial mass in the chart below.
7. Slide the food holder underneath the can, leaving a 1-inch gap.
8. Have one member of the group light a match and quickly catch the food on fire.
9. Carefully watch the display on the probe as the temperature increases. Record the highest temperature you see. It will take a while to reach this temperature, as the temperature should continue to rise even after the food has stopped burning.
10. Mass of the holder with any remaining food once more, and record the mass below.
11. Empty the water in the can.
12. When the 3 min. is up, the probe will tell you it is "done." Enter past this screen, and a graph will appear. Hit enter again. This takes you to a screen that asks whether you want to repeat. Choose YES, and go back to step 4.

Food: ___________________________			
	Trial 1	**Trial 2**	**Trial 3**
Initial mass: food + holder (g)			
Final mass: food + holder (g)			
Initial water temperature (°C)			
Final water temperature (°C)			
Volume of water			

Specific Heat

Our goal is to calculate the heat (or energy) produced by the food we eat. We know the heat capacity of water (it's constant), and we need to find the mass of the water we used and the change in its temperature. Use the following formula:

$$\mathbf{H = mC_p\Delta \mathit{t}}$$

(heat = mass • heat capacity • change in temperature)

Data Analysis

1. For each sample, 75 mL of water was used. The density of water is roughly 1.0 g/mL. What mass of water did we use?

2. The change in temperature will be different for each trial.
   Δt = final temperature − initial temperature
   Trial 1: Δt = _________ − _________ = _________

   Trial 2: Δt = _________ − _________ = _________

   Trial 3: Δt = _________ − _________ = _________

3. Now we can calculate how much energy is in the food we eat. Use the equation at the top of the page to calculate the heat absorbed by the water (one calculation for each trial).
   First in joules (C_p = 4.18 J/g°C)
   Trial 1 Trial 2 Trial 3

4. Then in Calories (C_p = 0.001 Cal/g°C)
   Trial 1 Trial 2 Trial 3

5. To find out how much energy each gram of food contains, we need to know the mass of the food we used (initial mass − final mass).
   Trial 1 Trial 2 Trial 3

6. Then, divide the Calories you calculated in #3 by the mass from above to give you the energy content of each food sample.
   Trial 1 Trial 2 Trial 3

Questions

1. How accurate do you think this experiment was? What could be improved to make it more accurate?

2. Compare your data to at least one other group’s data. If it’s the same food, how close are the results? What makes them differ? If it’s a different food, what causes the Calories to differ? What’s different about the foods?

3. What foods (either types of food or specific examples — at least two for each) would give you more Calories than the food you chose? Fewer Calories? Why?

4. How do you think the calculations you did today are used in the food industry?

Molar Heat of Fusion for Water

(From “Thermochemistry Lab – The Determination of the Molar Heat of Fusion for Water” by John L. Park, found on his Web site *The Chem Team: A Tutorial for High School Chemistry* at http://dbhs.wvusd.k12.ca.us/webdocs/ChemTeamIndex.html. Copyright 2002. Used by permission.)

Organizing Topic Properties of Matter

Overview Students warm liquid water (the “hot body”) and dump ice (the “cold body”) into it. Careful measurements of mass and temperature changes allow the students to calculate the amount of heat energy required to melt one mole of ice — i.e., the “molar heat of fusion for water.”

Related Standards of Learning CH.5d

Objectives

The students will

- determine the value for the molar heat of fusion for water;
- understand and demonstrate
  - MSDS warnings;
  - safety rules for science;
  - laboratory safety cautions;
  - safe techniques and procedures;
- calculate energy change and specific heat;
- demonstrate precision in measurement;
- understand accuracy in terms of closeness to the true value of a measure.

Materials needed

- Goggles and apron
- Hot plate
- Triple beam balance
- Thermometer
- Styrofoam cup
- Beaker tongs
- Small plastic spoon
- Hot pad for warm beaker
- Two 400-mL beakers
- One 600-mL beaker
- Tap water
- Ice
- Paper towels
- Attached lab worksheet

Instructional activity

Content/Teacher Notes

Heat is the flow of energy due to the temperature difference that exists between a hot body and a cold body. Heat flow will stop when the temperatures of the hot and cold bodies become the same.

CAUTION! Make sure students wear goggles and aprons during the entire course of the lab, which includes all cleanup time. Make sure they handle all glassware with great care, being very careful not to drop or knock over any pieces. Make sure they handle all hot objects with proper care, being very careful to protect themselves from being burned.

Introduction

1. Hand out a “Molar Heat of Fusion for Water” worksheet to each student, and go over the preliminary steps.
2. Have lab partners discuss and decide who will do which steps in the setup procedure.

Procedure

Give students verbal instructions to do the experiment, as follows:

1. Use beaker tongs to grasp the 400-mL beaker containing the warm water. Pour water into the Styrofoam cup until the cup is half full. Immediately mass the cup and warm water, and record this measurement on a data table. Set the hot beaker on the hot pad until the beaker cools. Place the Styrofoam cup inside the 600-mL beaker to stabilize it.
2. Measure the temperature of the warm water in the Styrofoam cup, and record this value in the data table. Leave the thermometer in the water as you go on to the next step.
3. *Immediately* after recording the temperature, add the equivalent of a handful of ice cubes to the warm water. Be very careful not to add any cold water (melted ice) in the 400-mL beaker to the warm water. Do not allow any splash to occur.
4. Using the thermometer, gently stir the ice in the water. Your goal is to lower the temperature of the warm water to a single digit value *and have no ice remaining*. If you decide you need to add more ice to do this, add it one piece at a time, and keep stirring gently without stopping. Once the temperature has stopped going down, record it on the data table.
5. Remove the thermometer from the water. Remove the cup from the 600-mL beaker. Mass the cup and cold water, and record this value in the data table. You are now finished with the experimental portion of run #1 of this lab.
6. Make a second run of the experiment by drying the Styrofoam cup, reheating the water in the 400-mL beaker to about 55°C, and repeating steps 1–5 above.
7. Pour all water into the sink, and return any unused ice to where you got it.
8. Dry all equipment that is wet, including the table top. Put all materials onto a dry paper towel near the back of the table.
9. Make sure you have a copy of *all* data; do not depend on your partner being present the next day.
10. Wait for the teacher's signal before removing goggles and aprons.

Observations and Conclusions

Have the students do the following:

1. Calculate the mass of the warm water in the Styrofoam cup.
2. Calculate the temperature change that the warm water underwent as it melted the ice.
3. Calculate the amount of heat lost, in joules, by the warm water as it melted the ice.
4. Calculate the mass of the ice that melted.
5. Calculate the amount of heat, in joules, that heated the melted ice from 0°C to the final temperature.
6. Calculate the amount of heat, in joules, that actually did melt ice.
7. Calculate, to the 0.01 place, a) the heat of fusion for water in joules per gram, and b) the molar heat of fusion for water in kilojoules per mole.
8. Calculate the percent error for the value in 7b above. The true value is 6.02 kJ/mol.

Resources

- Park, John L. "The Determination of the Molar Heat of Fusion for Water." *The Chem Team: A Tutorial for High School Chemistry.* http://dbhs.wvusd.k12.ca.us/webdocs/Thermochem/Lab-MolarHeatFusionWater.html and http://dbhs.wvusd.k12.ca.us/webdocs/Thermochem/Lab-MHFW-Comments.html#disc.

Molar Heat of Fusion for Water
Lab Worksheet

Directions

After discussing the following steps with your lab partner, set up the lab by doing the following preliminary steps:

1. Prepare for the lab by putting on goggles and aprons.
2. Zero the triple beam balance.
3. Make sure the Styrofoam cup is clean, dry, and empty.
4. Mass the Styrofoam cup, create a data table, and record the measurement in your table.
5. Put 150 mL of tap water into a 400-mL beaker, and put it on the hot plate. Turn the hot plate to full power, and put a thermometer in the water. Stir the water gently with the thermometer. You want the water to be between 60° and 70°C. Monitor the temperature of the water, and when it gets to about 55°C, turn the hot plate off.
6. After completing step 5, fill the second 400-mL beaker to the top with ice.
7. Continue the lab by following your teacher's verbal instructions for each step.

The Colligative Properties of Solutions

Organizing Topic Properties of Matter

Overview Students determine the densities of water, antifreeze, and a variety of water-antifreeze solutions. They graph density versus percent solution antifreeze and then determine and plot the boiling points and freezing points of the various water-antifreeze solutions.

Related Standards of Learning CH.1; CH.5f

Objectives

The students will

- collect volume, mass, and temperature measurements, using appropriate equipment;
- relate concentration to colligative properties of solutions;
- demonstrate precision in measurement.

Materials needed

Skills-development activity:

- Graduated cylinder
- Centigram balance
- Ten 15-mL test tubes with stoppers
- Test tube rack
- Commercial antifreeze (environmentally safe antifreeze is recommended)
- 2.5 mL corn syrup (optional)
- 7.5 mL propylene glycol (optional)

(Note: Check to be sure that it is permissible for students to use these materials at your school.)

Inquiry-lab activities:

- Solutions of various percent antifreeze prepared in the skills-development lab
- Thermometer or temperature probe
- Boiling chips
- Hot plate or burner
- Styrofoam cup

Instructional activity

Content/Teacher Notes

Colligative properties are properties of solutions that depend on the amount of solute particles in the solution (concentration) and are independent of the nature of the solute. Freezing point depression, boiling point elevation, vapor pressure lowering, and osmotic pressure are all *colligative properties.*

When water is the solvent, the boiling point of water will increase 0.512°C for each 76 grams of propylene glycol (antifreeze) added to 1,000 grams of water. The freezing point of water will decrease 1.86°C for each 76 grams of propylene glycol added to 1,000 grams of water.

In this activity, students determine the boiling and freezing points of various solutions. They notice that density is a good indicator of boiling point. When students determine the freezing point of the solutions, they notice that as the density increases, the freezing point steadily decreases but then begins to increase. This discrepant event helps students realize that density is correlated to, but is not the cause of, the changes in the boiling and freezing points.

This laboratory exercise gives students an opportunity to notice that as antifreeze is added to a solution, the density of the solution increases. As the density of the solution increases, the boiling point of the solution also increases and the freezing point decreases initially but then begins to increase.

CAUTION! Ethylene-glycol-based antifreeze is highly toxic. Biodegradable antifreeze is recommended in order to eliminate many disposal problems. For another safe alternative substance with the same density as toxic antifreeze, use a mixture of 2.5 mL corn syrup added to 7.5 mL propylene glycol. Additionally, the use of boiling chips will help prevent super-heating.

Introduction

1. Have a class discussion about the calculation of density and the proper use of a balance. Students may need instruction about boiling point and freezing point determination, depending on their previous lab experience.
2. Explain that boiling points of water-antifreeze solutions are easy to determine in a laboratory and require minimal equipment, although students will need to be extremely careful. Determining the freezing points of water-antifreeze solutions is much more difficult due to the extreme cold required. However, since substances melt and freeze at the same temperature, the freezing points of water-antifreeze solutions are equal to the melting points of the equivalent ice-antifreeze solutions, which are easily determined.

Skills-Development Activity

Introduction

1. Begin by explaining that density is a characteristic property of matter: different substances exhibit different densities. *Density* is defined as the mass of a substance divided by its volume. Tell the students that in the lab, they will determine the densities of water, antifreeze, and various solutions of antifreeze and water.

Procedure

1. Have the students determine and record on a data chart the mass of a clean, dry 10-mL graduated cylinder.
2. Have students fill the graduated cylinder with 10.0 mL of tap water and determine and record on their data chart the mass of the graduated cylinder plus the water.
3. Instruct students to empty the water and dry the graduated cylinder.
4. Next, have students carefully fill the graduated cylinder with 10 mL of antifreeze and then determine and record the mass of the graduated cylinder plus the antifreeze.
5. Have students pour the antifreeze into a labeled test tube and save it for the next experiment.
6. Tell students to pour 1.0 mL of antifreeze into the dry graduated cylinder and then fill it to 10.0 mL with water. They should record the mass and then pour this solution into a test tube labeled "10% solution." Have them save the solution for the next experiment.
7. Tell students to pour 2.0 mL of antifreeze into another dry graduated cylinder and then fill it to 10.0 mL with water. They should record the mass and then pour this solution into a test tube labeled "20% solution." Have them save this solution for the next experiment.
8. Have students make 10.0 mL each of 30%, 40%, 50%, 60%, 70%, 80%, and 90% antifreeze solutions, recording the mass of each solution, labeling according to its percentage, and saving for the next experiment.
9. Instruct students to clean up appropriately.

10. Finally, have the students calculate the density of each of the solutions that were prepared and then graph density versus percent solution of antifreeze. Let pure water be 0 percent antifreeze and pure antifreeze be 100 percent antifreeze. Ask: “Is density a good indicator of concentration?” Have them explain their reasoning. Have them describe the general shape and trends in the graph they constructed.

Inquiry-Lab Activities

1. The boiling points of the various water-antifreeze solutions: Have students pour 2 to 3 mL of one of the solutions to be tested in a clean test tube and add a boiling chip to prevent super-heating. Instruct them in the safe way to expose the test tube to the heat source. Have them measure the boiling point by recording the temperature at which the liquid *first* starts to boil. Remind them that the boiling point changes upon prolonged boiling of a solution; therefore, they should record the boiling point as early as possible. Have the students repeat this process for each of the nine solutions, recording the boiling points in their data charts.
2. The freezing points of the various water-antifreeze solutions: Have the students measure 90 grams of ice and place it in a Styrofoam cup. Next, have them place a thermometer in the ice and stir carefully, reading the temperature every 30 seconds until the temperature remains constant. Inform them that this is the melting point of ice and that *the freezing point of water equals the melting point of ice.* Have students then add 10 grams of antifreeze to the 90 grams of ice in the cup and measure the freezing point of the 10% antifreeze solution by recording the lowest temperature reached. Record the freezing point in their data chart. You may wish to assign various lab stations the following ice-antifreeze mixtures: 80 grams ice to 20 grams antifreeze, 70 grams ice to 30 grams antifreeze, 60 grams ice to 40 grams antifreeze, 50 grams ice to 50 grams antifreeze, 40 grams ice to 60 grams antifreeze, 30 grams ice to 70 grams antifreeze, 20 grams ice to 80 grams antifreeze. Record the results on the chalkboard or the overhead.

Observations and Conclusions

- Have students explain what happened to the density of the water-antifreeze solution as the percentage of antifreeze changed.
- Have students identify what happened to the freezing point when they had at least 50% antifreeze. Have them explain the reasons this happened.
- Challenge students to predict how density is related to boiling point and freezing point.
- Discuss the fact that antifreeze is used in a car’s radiator to help prevent it from freezing or boiling over. Ask: “Now that you have experimented with different percent concentrations of water-antifreeze and know how changing percent concentration of antifreeze affects the boiling and freezing points of the solutions, what percent solution would be most effective in preventing freezing and boiling in your car’s radiator?”

Sample assessment

- Use the formal written lab report as an evaluation tool.

Resources

- *Colligative Properties: Introduction.* Purdue University Department of Chemistry. http://www.chem.purdue.edu/gchelp/solutions/colligi2.html.

Thermochemistry: Heat and Chemical Changes

Organizing Topic Properties of Matter

Overview Students focus on describing some physical properties of solutions. They learn why some solutions conduct electricity while others do not. They also learn about three colligative properties of solutions.

Related Standards of Learning CH.5

Objectives

The students will
- recognize that solid, liquid, and gas phases of a substance have different energy content;
- understand that specific amounts of energy are absorbed or released during phase changes;
- explain that specific heat capacity is a property of a substance;
- know that the number of solute particles changes the freezing point and boiling point of a pure substance;
- recognize that a liquid's boiling point and freezing point are affected by changes in atmospheric pressure;
- understand that a liquid's boiling point and freezing point are affected by the presence of certain solutes;
- calculate energy changes, using specific heat capacity;
- calculate energy changes, using molar heat of fusion and molar heat of vaporization;
- perform calorimetry calculations.

Materials needed

For Introduction:
- Large rubber bands
- 13 x 18 cm piece of sheet metal
- 13 x 18 cm piece of Styrofoam
- Approximately 40 x 20 cm piece of wood

For Lab Activities:
- Thermometers
- Ring stands and clamps
- Metal rods
- Bunsen burners
- Heat-resistant gloves
- Erlenmeyer flasks
- Water
- Food coloring
- One-hole rubber stopper
- Glass tubing
- Ice
- Two-hole rubber stopper
- Hot plate
- Large beaker or bucket
- Goggles
- Tongs
- Snack-food packages
- Soda crackers and box
- Empty soda can
- Dissecting needles
- Pictures of frogs, fish, or reindeer
- Large picture of a car
- Flat-bottomed Florence flasks
- Bag of raisins
- Bag of dried cranberries
- Pint-size, unopened glass bottle of soda water
- Rock salt
- Styrofoam cup

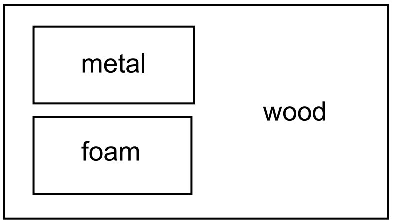
Instructional activity

Content/Teacher Notes

The activities in this lesson require six 90-minute class periods.

CAUTION! Use great care with laboratory glassware; although it is tempered to withstand drastic temperature changes, it is possible that it will shatter. Follow the MSDS General Safety Precautions.

Introduction



1. Before the introductory activity, make a “temperature-comparison board” by gluing a 13 x 18 cm piece of sheet metal and a 13 x 18 cm piece of Styrofoam to a piece of wood that is about 40 x 20 cm.
2. After cautioning students about misuse, give each student a clean, medium-sized rubber band.
3. Have students do this procedure, following your verbal instructions:
   a. Hook your index fingers through the ends of the rubber band. Without stretching the rubber band, place it against your upper lip or forehead, and note its temperature.
   b. Move the rubber band away from your skin. Quickly stretch and hold it, and again place it against your skin. Note any temperature change.
   c. Fully stretch the rubber band, and then allow it to return to its original state. Once more, place it against your skin, and note any temperature change.
   d. Repeat steps 3.b and 3.c until you are certain of the temperature change in each step.
4. Ask the students the following questions:
   - Did the rubber band feel cool or warm after it was stretched in step 3.b?
   - Did the rubber band feel cool or warm after it returned to its original shape in step 3.c.
   - What is heat?
   - In what direction does heat flow?

   Students will discover that the rubber band feels warmer after it is stretched and cooler after it is allowed to relax. Students may infer that heat is related to the observed temperature changes. From their experience, students may realize that heat flows from a warm object to a cool object. The stretching of rubber is *exothermic*; the reverse process — relaxing the stretched rubber — is *endothermic*. Define *exothermic reactions* and *endothermic reactions.* Stress that in an exothermic reaction, heat is released, and the energy of the products is *less* than that of the reactants. Draw a graph on the board to depict this. Emphasize that in an endothermic reaction, the energy of the products is *greater* than that of the reactants. Graph this on the board.
5. Review *thermochemistry* as the study of heat changes occurring during chemical reactions.
6. Pass around the piece of wood with the sheet metal and Styrofoam attached, and ask students to put their hand on each of the three surfaces and describe the temperature of each. Usually students respond that the metal feels the coolest and the Styrofoam feels the same temperature as their hands. Explain that all three surfaces are actually the same temperature — the temperature of the room, which is usually considerably cooler than their hand — and that when they touch those surfaces, heat is being transferred from their hand to them. Metals conduct heat away from the hand more rapidly than Styrofoam does, so the metal “feels” cooler than the Styrofoam.

Procedure

Heat versus Temperature; Conducting Heat Energy through a Solid

1. Begin with a review of the arrangement of atoms in a crystal, emphasizing that the atoms are held together in an orderly arrangement that gives the crystal a definite shape.
2. Ask students to consider what would happen to the atoms if they began to absorb energy from, for example, an open flame. Point out that the atoms would move around more vigorously, slamming into one another with greater momentum. This could change the arrangement of atoms in the crystal and alter its shape. The crystal could melt and eventually vaporize. Explain that this change would be the result of an increase in the momentum of the atoms in the system. The transfer of energy from atom to atom in a solid is called *conduction.*

4. Ask students to consider what would happen if a thermometer were to be placed against the crystal while it is being warmed. Have student analyze the system at the atomic level and higher and write a brief statement that explains why the liquid inside the thermometer reads higher and higher as the crystal is warmed. Guide students to the conclusion that the thermometer does not measure heat directly; instead, it reflects the average kinetic energy of the atoms in the system. *Heat* is a measure of the total energy of a system. The heat energy released during a chemical change in a substance can be measured using a *calorimeter.* The unit of heat energy is the *calorie*: one *calorie* is the amount of energy needed to raise the temperature of 1 gram of pure water 1 degree Celsius.
5. Distribute thermometers that read both degrees Fahrenheit and degrees Celsius. Remind students that each thermometer uses several scales, Fahrenheit and Celsius, to read the same amount of average kinetic energy; therefore, for example, 32°F = 0°C = 273K. Show students how to use the temperature conversions to change temperatures from Fahrenheit to Celsius to Kelvin.
6. <u>Conducting Heat Energy through a Solid</u>: In the lab, have students track the change of temperature of a metal being heated, as follows: Have them clamp a thermometer to a ring stand and also clamp a metal rod to the same ring stand. Direct them to use heat-resistant gloves while heating the metal rod with a Bunsen burner, reading and recording the temperature every 15 seconds for 2 minutes. Then, ask them to use this data to make a line graph of the change of temperature over time.

Transferring Heat though a Liquid or Gas

1. Pour a small amount of water and several drops of food coloring into a flask, and stopper the flask with a one-hole rubber stopper. Place a piece of glass tubing through the stopper and into the colored water. Rub your hands together, place them over the bottom of the flask, and have the students watch closely and record their observations. Next, rub ice over the bottom of the flask in the same place you positioned your hands, and have students watch and record their observations.
2. Fill a small Erlenmeyer flask with water and several drops of food coloring, and stopper the flask with a two-hole rubber stopper. Warm the flask on a hot plate on a low setting for two minutes. While the flask is warming, fill a large beaker or bucket with cold water. When the water in the flask is hot, don goggles and use tongs or heat-resistant gloves to transfer the flask to the bottom of the large beaker or bucket. *CAUTION! Use great care with laboratory glassware; although it is tempered to withstand drastic temperature changes, it is possible that it will shatter. Follow the MSDS General Safety Precautions.* Have students record their observations.
3. Discuss the students' recorded observations from both labs. In the first, they should have observed the colored water rising in the tube as the air in the flask expands due to heat convection and then the colored water sinking in the tube as the air contracted. In the second lab, they should have observed the colored, hot water escaping from the flask when it was placed in the large beaker or bucket of cold water.

3\. Review the results of the previous activity, repeating the distinction between heat energy and temperature. Explain that the metal bar used in the previous activity did not hold heat very well but transferred the heat very quickly; in other words, the metal was an excellent conductor of heat but did not have the capacity to hold or store heat. Scientists can measure the capacity of a substance to hold or store heat. The capacity of a substance to store chemical energy is called *specific heat.* Water has a specific heat equal to 1 because it takes one calorie of energy to raise the temperature of 1 gram of water 1°C. The specific heat of iron, on the other hand, is only 0.11; that is, it takes 0.11 calories to raise the temperature of 1 gram of iron 1°C — only about one-tenth the amount of energy needed to raise the temperature of an equal amount of water.

4. Have students use reference materials to create their own Specific Heat of Substances Chart. Assign some substances that they should know about, and have them choose some that they want to learn about. Many periodic tables give the specific heat of each element. You may wish to have them graph each substance listed on the chart.
5. Ask students also to determine the factors that would make a substance a good insulator. Ask: "Would such a substance have a low or high specific heat?" *(Insulators have a high specific heat. They store heat energy and prevent its transfer.)* Have students use the reference materials to make a list of good insulators.
6. Have students present their Specific Heat of Substances Charts and insulator lists.

Measuring Calories

1. Before this lab, have students bring in examples of the outside packaging from their favorite snack foods.
2. Remind students that *conduction* is the transfer of energy through a solid. Now, point out that energy can also be transferred from one particle to another in a liquid or a gas, and define *convection* as the transfer of energy through a fluid (i.e., a liquid or gas).
3. Review the definition of a *calorie* — the amount of energy needed to raise the temperature of 1 gram of pure water 1 degree Celsius. Therefore, raising the temperature of 100 grams (100 mL) of water 1 degree Celsius would require 100 calories.
4. Explain the difference between a food Calorie and an energy calorie: 1 food Calorie is equal to 1,000 energy calories, as measured using a thermometer or calorimeter. Have students read the ingredients label on the packaging of their favorite snacks, paying particular attention to the calories-per-serving information. Then, have them read the ingredients label on a soda cracker box. Example: "Calories per serving = 140; serving size 10 crackers." Ask: "How many food Calories are there in one soda cracker?" *(14)* "How many energy calories are in one cracker?" *(14,000)* Burning one cracker and using *all* of the released energy to heat 100 mL of water would heat the water to about 140°C — quite a lot of heat! Explain that in the lab that the students are about to perform, they will see that most of the heat from the burning cracker will be lost.
5. Have the students conduct the following lab:
   a. Pour 100 mL of water from a beaker into an empty soda can.
   b. Secure the soda can with two ring clamps, and lower a thermometer into the can just below the surface of the water.
   c. Skewer a soda cracker onto the end of a dissecting needle, and clamp it to the ring stand below the soda can. Record the temperature of the water. Hold a lighted match to the soda cracker until the cracker burns on its own. Record the temperature of the water when the cracker has completely burned. Have students determine the difference between the theoretical rise in temperature (about 140°C) and what they actually observed. Have them explain why this happened.
6. At the board, perform some calorimetry calculations. Have students practice more calorimetry calculations for homework.

The Colligative Properties of Solutions

1. Before the students enter the room, write on the board or make a banner that reads as follows:

   Colligative property
   depends on the *concentration* of particles,
   not their identity!

   Also, place a large picture of a car at the front of the room.
2. As students arrive, give each of them a picture of a frog, fish, or reindeer, and point out the large picture of a car at the front of the room. Have the students try to figure out the following riddle: What do these three animals have in common with a car? When they have guessed long enough, give them a clue: antifreeze. If no one "gets it," tell them these animals, like a car, can survive being frozen. Scientist believe that a substance in the cell of these animals act as a natural antifreeze, which prevents their cells from freezing. Although fluids surrounding their cells may freeze, the cells themselves do not.
3. Ask students, "What happens to the molecules in water when water freezes?" *(The molecules form a crystalline lattice.)* "How does a solute change the freezing point of a solvent?" *(It might slow down the formation of the crystal lattice.)* "Why do we throw salt down on ice in the winter?" *It lowers the freezing point, causing the ice to melt.*
4. Point out that there are special properties of solutions. The physical properties of solutions differ from those of the pure solvent used to make the solution. These properties depend only on the number of particles dissolved in a given mass of solvent. The properties, called *colligative properties,* include vapor pressure lowering, boiling-point elevation, and freezing-point depression. Remind students that *vapor pressure* is the pressure exerted by vapor that is in dynamic equilibrium with its liquid in a closed system. Point out that a solution with particles that are not easily vaporized always has a lower vapor pressure than the pure solvent.
5. Place three beakers of water in front of the students, and inform them that at this point, all three beakers have the same vapor pressure. The vapor pressure of a nonvolatile solution (a solution filled with particles that are not easily vaporized) is less than the vapor pressure of the pure solvent. Make sure students are writing the key points down in their lab books. Equilibrium is established between the liquid and the vapor in the pure solvent.
7. Add three raisins to the first beaker to represent the particles of the solute glucose. Now, point out that solvent particles form shells around the solute particles, thus reducing the number of free solvent particles able to escape the liquid. Equilibrium is eventually re-established at a lower vapor pressure. Use more raisins and some dried cranberries to show additional examples: in the second beaker, place three raisins and three cranberries to represent the particles in sodium chloride; and in the third beaker, place three raisins and six cranberries to represent the particles in calcium chloride. As a review, have the students determine the number of particles that belong in each solvent, based on the molecular formulas of sodium chloride and calcium chloride.
8. Ask the students, "How many particles in solution are produced by each formula unit of aluminum bromide?" *(4)* "How many moles of particles would 3 mol Na_3PO_4 give in a solution?" *(12 mol of particles)* Stress that adding a nonvolatile solute to a solvent deceases the vapor pressure.
9. Define *boiling point* of a substance as the temperature at which the vapor pressure of the liquid phase equals the atmospheric pressure. Ask the students, "If the vapor pressure is now lower because of the addition of the solute, then what happens to the boiling point of the solution?" *(It rises.)* Students will probably understand more readily if this is explained in terms of particles. Attractive forces exist between the solvent and solute particles. It takes additional kinetic energy for the solvent particles to overcome the attractive forces that keep them in the liquid. Thus, the

presence of a solute elevates the boiling point of the solution. The magnitude of the boiling point elevation is proportional to the number of solute particles dissolved in the solvent. For example, the boiling of water increases by 0.51°C for every mole of particles that the solute forms when dissolved in 1,000g of water. This is why boiling point is a colligative property.)

Freezing Point

1. When water freezes, the particles of the solid take an orderly pattern. The presence of a solute in the water disrupts the formation of this pattern. Point out that the solution will still freeze, but at a lower freezing point.
2. Have students observe the freezing point depression of water by the addition of a substance such as rock salt (NaCl) to ice-water mixture. You might have each group of students use a different salt. Provide each group with a thermometer, Styrofoam cup, and a mixture of ice and water, and ask them to measure the initial temperature of the ice-water mixture. Then have them measure the lowest temperature reached after the addition of the salt. Also, have them investigate the rate at which the frozen salt water melts to discover whether the salt water melts faster or slower than plain water.

Observations and Conclusions

- Ask students, “Will 1 mole of sugar have the same effect as 1 mol of table salt, NaCl, in lowering the freezing point of water? Explain your answer.” *(1 mol of table salt will have approximately twice the effect in lowering the freezing point of water as 1 mol of sugar, because 1 mol of NaCl produces 2 moles of solute particles in solution, while 1 mole of sugar produces only 1 mole of solute particles.)*
- Have the students
  - define *thermochemistry*
  - define *specific heat capacity*
  - identify colligative properties of solutions
  - calculate energy changes, using specific heat capacity
  - perform calorimetry calculations.

Sample assessment

- Have the students perform the following lab and complete the lab write up in one class period. This lab is a good example to link the ideas of freezing point depression and gas solubility. Each student will need a pint-size, unopened glass bottle of soda water, ice, a large beaker, and rock salt.
- Have the students remove the label on their bottle, pack the bottle in a large beaker of ice, and sprinkle rock salt on top of the ice.
- Approximately 10 minutes later, have students remove their bottle from the ice. (The soda water should still be liquid.) Next, have students open their bottle. As the gas effervesces from the water, the water in the bottle will instantly freeze.
- Ask student the following questions:
  - How did the temperature of the ice bath compare to the normal freezing point of water? *(The temperature was lower than 0°C because of freezing point depression caused by the salt.)*
  - Why did the soda water remain a liquid while in the ice bath? *(The CO_2 gas dissolved in the water lowered its freezing point.)*
  - What happened to the CO_2 when the bottle was opened, and why? *(The CO_2 came out of solution because the pressure was lowered when the bottle was opened.)*

  - Why did the water freeze when the bottle was opened? *(When the CO_2 came out of solution, the freezing point of the water was raised to 0°C, but the temperature of the water was already below 0°C.)*

Follow-up/extension

- Have students research how solubility and colligative properties play a role in human physiology, environmental science, plant growth, and industry. Have them visit the Web site *Kitchen Chemistry* at http://personal.cfw.com/~rollinso/SciFood.html.

Resources

- *Newton North High School Chemistry home page.* http://www.newton.mec.edu/tigerscience/chem/Chemhomepage.htm. Contains chemistry lesson plans with worksheets.

Organizing Topic — Electron Configuration and the Periodic Table

Standards of Learning

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include

- g) mathematical manipulations (SI units, scientific notation, linear equations, graphing, ratio and proportion, significant digits, dimensional analysis).

CH.2 The student will investigate and understand that the placement of elements on the periodic table is a function of their atomic structure. The periodic table is a tool used for the investigations of

- a) average atomic mass, mass number, and atomic number;
- b) isotopes, half lives, and radioactive decay;
- c) mass and charge characteristics of subatomic particles;
- d) families or groups;
- e) series and periods;
- f) trends including atomic radii, electronegativity, shielding effect, and ionization energy; and
- g) electron configurations, valence electrons, and oxidation numbers.

Essential Understandings, Knowledge, and Skills

Correlation to Textbooks and Other Instructional Materials

The student will use hands-on investigations, problem solving activities, scientific communication, and scientific reasoning to

- use, for any neutral atom of a particular element, the periodic table to determine atomic number, atomic mass, the number of protons, the number of electrons, and the number of neutrons; ____________________
- point out that
  - the Periodic Law states that when elements are arranged in order of increasing atomic number, their physical and chemical properties show a periodic pattern, which is periodicity; ____________________
  - the periodic table is arranged by increasing atomic numbers; ____________________
  - periods and groups are named by numbering column and rows; ____________________
- understand that
  - electron configuration is the arrangement of electrons around the nucleus of an atom, based on their energy level; ____________________
  - atoms can gain or lose electrons within the outer energy level; ____________________
- use an element’s electron configuration to determine the number of valence electrons and possible oxidation numbers; ____________________
- apply the following principles of electron configuration:
  - Aufbau Principle; ____________________
  - Pauli Exclusion Principle; ____________________

  - Hund's Rule;
  - Energy levels are designated 1–7. Orbitals are designated s, p, d, and f according to their shapes. These orbitals relate to regions of the periodic table.
  - Loss of electrons from neutral atoms results in the formation of an ion with a positive charge (cation).
  - Gain of electrons by a neutral atom results in the formation of an ion with a negative charge (anion).
- identify the location of the following on the periodic table: alkali metals, alkaline earth metals, transition metals, halogens, noble gases, and metalloids;
- determine that
  - vertical columns, called "groups," have similar properties because of similar valence electron configurations;
  - horizontal rows, called "periods," have somewhat predictable properties based on an increasing number of outer orbital electrons;
- graph data to determine relationships and trends;
- identify the following trends in the periodic table:
  - Shielding effect is constant across the period and increases within given groups from top to bottom.
  - Atomic radius decreases from left to right and increases from top to bottom within given groups.
  - Ionization energies generally increase from left to right and decrease from top to bottom of a given group.
  - Electronegativity increases from left to right and decreases from top to bottom.

Element Family Reunion

Organizing Topic Electron Configuration and the Periodic Table

Overview Students construct knowledge about families of elements. Each student becomes an "expert" in the physical and chemical properties of one element and then brings that expertise to "cooperative family-of-elements teams." Each team constructs a description of the trends and properties exhibited by their elements and uses this information to predict properties for missing elements. As teams present the "family reunion scrapbooks" to the class, students work individually with a template of the periodic table to summarize trends in families and identify trends in periods.

Related Standards of Learning CH.2

Objectives

The students will

- use, for any neutral atom of a particular element, the periodic table to determine atomic number, atomic mass, the number of protons, the number of electrons, and the number of neutrons;
- point out that
  - the Periodic Law states that when elements are arranged in order of increasing atomic number, their physical and chemical properties show a periodic pattern, which is periodicity
  - the periodic table is arranged by increasing atomic numbers
  - periods and groups are named by numbering column and rows;
- understand that
  - electron configuration is the arrangement of electrons around the nucleus of an atom, based on their energy level
  - atoms can gain or lose electrons within the outer energy level;
- use an element's electron configuration to determine the number of valence electrons and possible oxidation numbers;
- apply the following principles of electron configuration:
  - Aufbau Principle
  - Pauli Exclusion Principle
  - Hund's Rule
  - Energy levels are designated 1–7. Orbitals are designated s, p, d, and f according to their shapes. These orbitals relate to regions of the periodic table.
  - Loss of electrons from neutral atoms results in the formation of an ion with a positive charge (cation).
  - Gain of electrons by a neutral atom results in the formation of an ion with a negative charge (anion).
- identify the location of the following on the periodic table: alkali metals, alkaline earth metals, transition metals, halogens, noble gases, and metalloids;
- determine that
  - vertical columns, called "groups," have similar properties because of similar valence electron configurations;
  - horizontal rows, called "periods," have somewhat predictable properties based on an increasing number of outer orbital electrons;
- graph data to determine relationships and trends;

- identify the following trends in the periodic table:
  - ° Shielding effect is constant across the period and increases within given groups from top to bottom.
  - ° Atomic radius decreases from left to right and increases from top to bottom within given groups.
  - ° Ionization energies generally increase from left to right and decrease from top to bottom of a given group.
  - ° Electronegativity increases from left to right and decreases from top to bottom.

Materials needed

- Access to Internet databases and library references about the chemical elements
- Computer spreadsheet and presentation software
- Attached blank template for the periodic table of the elements

Instructional activity

Content/Teacher Notes

During the initial phase of this activity, students will work independently to gather information. They will need directions to both electronic and printed sources of information and help in interpreting that information. When students meet in cooperative teams to analyze and synthesize information about the team's family/group and period of elements and to develop a presentation for the rest of the class, they will need help in using spreadsheet and presentation software.

This project can be completed within five or six class periods. Students may work individually in the school media center and/or classroom, as well as at home, but they must use class time for working in cooperative teams and for completing the project. Approximately three hours of class time should be allocated for this teamwork.

Introduction

1. Before beginning the activity, give students the opportunity to observe, in the laboratory and on the Internet, and describe several common elements and some of the elements' reactions.
2. Have students practice using a spreadsheet by selecting a periodic property, such as density, and graphing it as a function of increasing atomic number for the first 20 elements.

Procedure

Individual assignments:

1. Using the online *Periodic Table of the Elements: A Resource for Elementary, Middle School, and High School Students* from the Los Alamos National Laboratory's Chemistry Division (see Resources at the end of this lesson), assign each student an element from a list of elements that includes four from each of the following families:
   - Group 1 — **alkali metals:** lithium, potassium, rubidium
   - Group 2 — **alkaline earth metals:** beryllium, magnesium, strontium, barium
   - Group 15 — nitrogen, phosphorus, antimony, bismuth
   - Group 16 — oxygen, selenium, tellurium, polonium
   - Group 17 — **halogens:** fluorine, chlorine, bromine, iodine
   - Group 18 — **noble gases:** helium, neon, argon, krypton

   Groups 13 and 14 may be added if class size necessitates it. Transition elements may be added for capable classes. It is important to have at least four elements in a group represented so that trends can be seen when students meet in their cooperative teams. It is also important that at least one element in a family/group be missing so that teams can predict properties. Finally, it is ideal to have as many groups from the periodic table represented as possible.

2. Give students a list of information to find about their assigned elements. The list could include
   - important isotopes
   - electron configuration and orbital filling diagram
   - valence electrons
   - electron dot diagram
   - common ions, including electron configurations for ions
   - physical properties, including boiling point, melting point, electrical conductivity, density, atomic radius, shielding effect and ionization potentials, and electronegativity
   - chemical properties, including reactivity in oxygen, water, and acids.

   The list could also include general information, such as abundance on earth, sources, uses, and historical information.

Cooperative family-of-elements teams:

1. After the individual assignments are complete, have students meet in family-of-elements teams (based on valence electron configurations) to compile and examine properties.
2. Have each team make a spreadsheet to compile quantitative data (boiling point, melting point, density, radius, conductivity, and ionization potential) and use the spreadsheet to make a graph(s) showing changes in properties as a function of atomic number in their family.
3. Ask teams to determine which element(s) are missing from their families and to use their graphs, as well as other information collected, to predict properties of the missing element(s). Allow students to refer back to the computer databases to verify predictions.
4. Have each team prepare and submit a summary of the family properties that includes their graphs, the predictions for the missing elements, and an evaluation of the accuracy of the predictions. Evaluate these summaries for accuracy and completeness.
5. Allow time for each team to prepare a presentation for the “family reunion scrapbook.” Presentations may be in the form of a poster, PowerPoint presentation, or Web pages that share the team’s family-of-elements information with the rest of the class.

Observations and Conclusions

1. Give each student a copy of the attached template for the periodic table. As students view the family reunion scrapbook, direct them to summarize family descriptions and trends on the template. Then ask students to use the family information to identify trends across periods and to annotate their periodic table. Focus the students’ attention on answers to the following questions:
   - How do valence electron configurations change going across periods?
   - Where are the families that commonly form anions? That form cations?
   - How do boiling point, melting point, and density change?
   - What trends in atomic radius, shielding effect, first ionization energy, and electronegativity can you identify?

Sample assessments

- Students’ work in the classroom provides an opportunity for authentic assessment of group interactions: using time and resources effectively, working cooperatively, and enlisting the suggestions and contributions of each team member. Develop rubrics for evaluating the students’ completion of independent element assignments, participation in the teamwork, and the family reunion scrapbooks. Make the rubrics available to the students before they begin work on the project. The scrapbooks can be evaluated according to completeness and accuracy of information presented, evidence of contributions by all team members, evidence that the standards listed for this project have been met, and the effectiveness of the graphical and visual representations of family properties.

- Assess students' understanding of periodic table relationships in an essay format by asking each student to relate trends in atomic radius to trends in one other variable, assigning different variables to individual students or to pairs of students working together.

Follow-up/extension

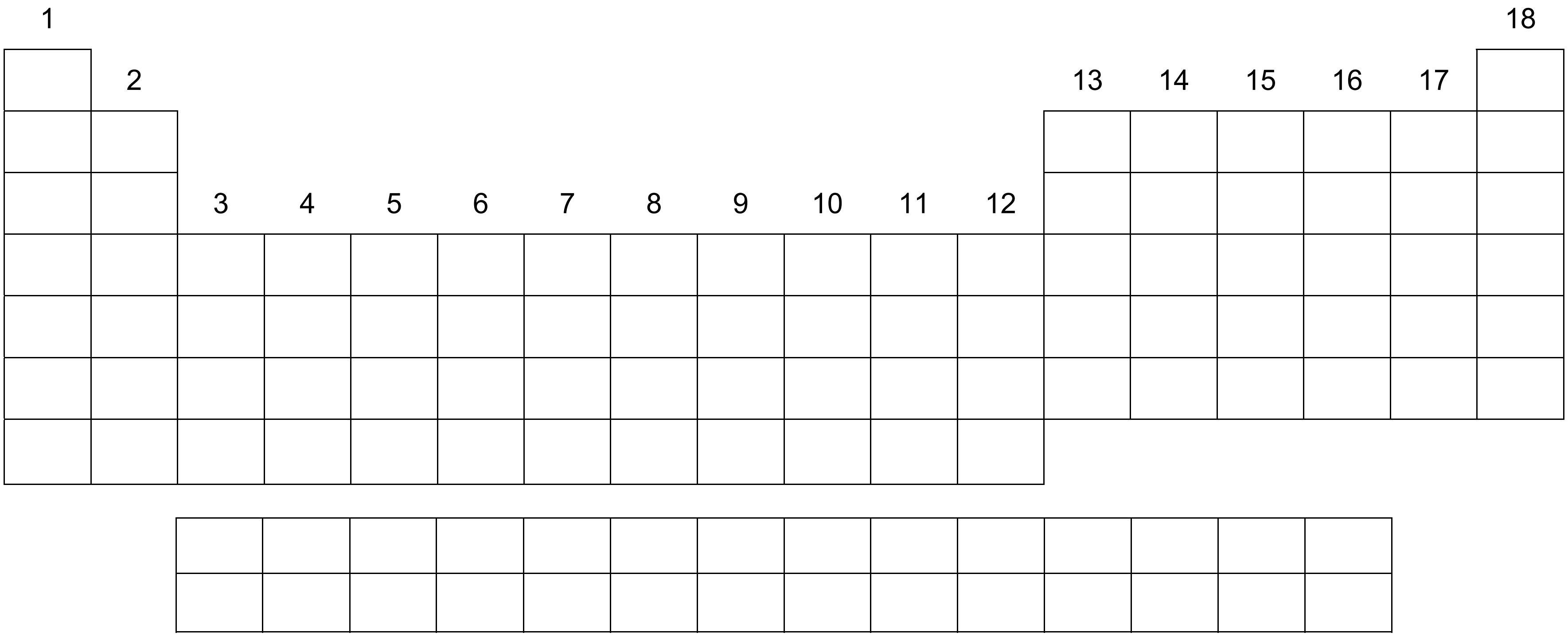
- Have students model electron dot diagrams for elements, using index cards with the element symbol and atomic number and small candies or dried beans to represent valence electrons.
- Have students use these models to demonstrate the formation of cations and anions. These models can be used in subsequent activities to model oxidation–reduction and the formation of ionic and covalent bonds, as well as Lewis dot diagrams for covalent molecules.

Resources

- *Periodic Table of the Elements: A Resource for Elementary, Middle School, and High School Students.* Los Alamos National Laboratory's Chemistry Division. http://pearl1.lanl.gov/periodic/default.htm.
- *WebElements™ Periodic Table.* http://www.webelements.com/.

Periodic Table of the Elements

Name: ____ **Date:** ____________________



Directions

1. Add group and period numbers to this periodic table.
2. Write the symbol, and fill in the Lewis Dot Diagrams for the 1st four rows and all elements in groups (columns) 1, 2, and 13–18.
3. Use a colored pencil to shade in the two columns where electron configurations end with s^1 and s^2.
4. Find the six columns that end with a “p” sublevel, and shade these in with a different color.
5. Use a third color to shade in the columns that are filling “d” sublevels.
6. Annotate the table with some sample data to show how boiling point, melting point, and density change across periods.
7. Use colored arrows to show trends in atomic radius, shielding effect, first ionization energy, and electronegativity, going down groups and going across periods. Be sure to provide a color code for the arrows.

Atomic Structure: Periodic Table

Organizing Topic Electron Configuration and the Periodic Table

Overview Students focus on the structure of the periodic table of the elements and the names of the portions — metals, nonmetals, and metalloids — as well as the group names and special series names.

Related Standards of Learning CH.2d, e, f

Objectives

The students will

- understand that the Periodic Law states that when elements are arranged in order of increasing atomic numbers, their physical and chemical properties show a periodic pattern. *Periodicity* is regularly repeating patterns or trends in the chemical and physical properties of the elements arranged in the periodic table;
- understand that the names of groups and periods on the periodic chart are “alkali metals,” “alkaline earth metals,” “transition metals,” “halogens,” “noble gases,” and “metalloids”;
- understand that periods and groups are named by numbering columns and rows;
- understand that some elements, such as oxygen, hydrogen, fluorine, chlorine, bromine, and nitrogen, naturally occur as diatomic molecules;
- understand that electronegativity increases from left to right within a period and decreases from top to bottom within a group;
- understand that shielding effect is constant within a given period and increases within given groups from top to bottom;
- understand that atomic radius decreases from left to right and increases from top to bottom within given groups;
- understand that ionization energies generally increase from left to right and decrease from top to bottom of a given group;
- use an element’s electron configuration to determine the number of valence electrons and possible oxidation numbers.

Materials needed

- Periodic table of the elements
- Colored pencils
- Rulers
- Construction paper

Instructional activity

Introduction

1. Lead the class in completing a KWL chart about periodic families.
2. Review the following background information with the class:
    - The periodic table has seven periods (horizontal rows) and 18 groups or families (vertical columns).
    - If you analyze the periods, you see that all of the elements in a period have the same number of electron energy levels: the first period has 1 electron energy level; the second period has 2 electron energy levels; and so forth.

- There is a pattern in the groups as well: the elements in group 1 have 1 electron in their outer energy level; the elements in group 2 have 2 electrons in their outer energy level; the elements in group 13 have 3 electrons in their outer energy level; the elements in group 14 have 4 electrons in their outer energy level; and so forth.
- Groups 3 through 12 are not included in the trend. That is because they are **transition elements,** which follow a slightly different trend due to the overlap of energy levels and the way those energy levels fill.
- The number of electrons in an element's outer energy level determines the element's chemical properties. Therefore, since all of the elements of a group have the same number of electrons, they will react similarly to each other. The elements in the first group have a special name — the **alkali metals**. The elements in group 2 also have a special name, the **alkaline earth metals**. The alkali metals are the most reactive group of metals on the periodic table. Francium is the most reactive of the group; reactivity increases as you go from the top to the bottom of this column.
- Group 17 is the **halogen** group of elements. The halogen group is the most reactive nonmetal group on the periodic table. Fluorine is the most reactive nonmetal in this group, and reactivity decreases as you go from the top of the column to the bottom.
- Group 18 is known as the **noble gas** or inert gas group. It is named this because all of its elements are stable and unlikely to react or bond with other elements. Notice that all of these elements have 8 electrons in their outer shell, with the exception of Helium, which has only 2 electrons in its outer energy level. This, however, is not really an exception because Helium's single energy level (unexcited) can hold only a maximum of 2 electrons.

Procedure

1. Give students guided practice in identifying the group, series, and section in which a given element is located. Further guided practice should be used to introduce group and period trends of atomic radii, ionic radii, and electronegativity across periods and down groups.
2. Give students some sample elements from each period, and have them identify the period and group within which each element is located. Repeat this exercise until 10 students in a row can answer correctly.
3. Give students some sample pairs of elements across periods and down groups, and have them predict changes in atomic radii, ionic radii, and electronegativity. Repeated this exercise until 8 out of 10 students answer correctly.
4. Have students do a "Photographic Periodic Table" project in which each student is assigned to bring in a photograph showing a particular element's use in today's world. Help students recognize uses of pure elements as opposed to compounds. The use depicted in the photograph should be specific to the element form and not a compound, if possible.

Sample assessment

- Quiz students on the periodic table, focusing on group and series names by general recognition and by locating elements within a group and/or period.
- Have students visit the interactive *The Periodic Table Challenge* at http://www.ilpi.com/genchem/periodicquiz.html to test whether they know the elements by position in the periodic table. Here they see a blank table and must type in the element symbols in the correct places. This site gives hints to help students remember/guess each element.

Follow-up/extension

- Have students create an enlarged two-dimensional model of the periodic table, color-coded with the following periodic trends: atomic numbers, masses (rounded), symbols, alkali metals, alkaline earth metals, transition metals, lanthanide series, actinide series, other metals, nonmetals, noble gases, metalloids, and halogens.

Resources

- "Atom Basics: Ions." *Rader's Chem4Kids.com*. http://www.chem4kids.com/files/atom_ions.html.
- "Atom Basics: Structure." *Rader's Chem4Kids.com*. http://www.chem4kids.com/files/atom_structure.html.
- "Electron Configurations." *CHEMystery.* http://library.thinkquest.org/3659/structures/electronconfig.html.
- *The Periodic Table Challenge.* http://www.ilpi.com/genchem/periodicquiz.html.
- *A Visual Interpretation of the Table of Elements.* http://www.chemsoc.org/viselements/.
- *WebElements™ Periodic Table.* http://www.webelements.com/.

Organizing Topic — Bonding, Nomenclature, and Formula Writing

Standards of Learning

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include

- g) mathematical manipulations (SI units, scientific notation, linear equations, graphing, ratio and proportion, significant digits, dimensional analysis); and
- h) use of appropriate technology including computers, graphing calculators, and probeware, for gathering data and communicating results.

CH.2 The student will investigate and understand that the placement of elements on the periodic table is a function of their atomic structure. The periodic table is a tool used for the investigations of

- a) average atomic mass, mass number, and atomic number;
- b) isotopes, half lives, and radioactive decay;
- c) mass and charge characteristics of subatomic particles;
- g) electron configurations, valence electrons, and oxidation numbers; and
- h) chemical and physical properties.

CH.3 The student will investigate and understand how conservation of energy and matter is expressed in chemical formulas and balanced equations. Key concepts include

- a) nomenclature;
- b) balancing chemical equations;
- c) writing chemical formulas (molecular, structural, and empirical; and Lewis diagrams); and
- d) bonding types (ionic and covalent).

CH.5 The student will investigate and understand that the phases of matter are explained by kinetic theory and forces of attraction between particles. Key concepts include

- f) colligative properties.

Essential Understandings, Knowledge, and Skills

Correlation to Textbooks and Other Instructional Materials

The student will use hands-on investigations, problem solving activities, scientific communication, and scientific reasoning to

- recognize that bonds form to achieve stability;
- explain the Law of Multiple Proportions and the Law of Definite Composition;
- differentiate between empirical, molecular, and structural formulas;
- identify and use
  - chemical formulas;
  - coefficients, chemical symbols, and subscripts;
- illustrate how negative and positive ions are formed and how to represent them;
- summarize the following concepts about ionic bonding:
  - Ionic bonds involve the transfer of electrons.

  - Ionization energy is the amount of energy needed to remove an electron from an atom in the gas phase. ____________________________
  - Elements with low ionization energy form ions easily. ____________________________
- recognize that transition metals can have multiple oxidation states; ____________________________
- summarize the following concepts about covalent bonding:
  - Covalent bonds involve sharing of electrons. ____________________________
  - Polar molecules result when a molecule behaves as if one end were positive and the other negative. ____________________________
  - Electronegativity is the measure of attraction of an atom for electrons in a covalent bond. ____________________________
- name and write formulas for certain elements that naturally occur as diatomic molecules, including oxygen, hydrogen, and nitrogen; ____________________________
- name binary and ionic compounds, using the roman numeral system where appropriate; ____________________________
- name compounds, using the IUPAC system; ____________________________
- recognize the formulas and names of the polyatomic ions carbonate, sulfate, nitrate, hydroxide, phosphate, and ammonium; ____________________________
- know chemical formulas for certain common substances, including water, carbon monoxide, carbon dioxide, sulfur dioxide, and carbon tetrafluoride; ____________________________
- draw Lewis dot diagrams to show covalent bonding; ____________________________
- predict, draw, and name linear, bent, trigonal planar, tetrahedral, and trigonal pyramidal molecular shapes; ____________________________
- recognize polar and nonpolar molecules; ____________________________
- calculate percent composition; ____________________________
- understand and demonstrate
  - MSDS warnings;
  - safety rules for science;
  - laboratory safety cautions;
  - safe techniques and procedures. ____________________________

A Crystal Lab

Organizing Topic Bonding, Nomenclature, and Formula Writing

Overview Students explore the properties of ionic and metallic crystals.

Related Standards of Learning CH.1a, b, c; CH.3d

Objectives

The students will
- recognize that bonds form to achieve stability;
- illustrate how ionic compounds are formed from positive and negative ions;
- understand that ionic bonds involve the transfer of electrons;
- recognize that metals have crystalline structures;
- understand and demonstrate
  - MSDS warnings
  - safety rules for science
  - laboratory safety cautions
  - safe techniques and procedures.

Materials needed

Ionic crystal formation:
- Distilled water
- Beaker
- Burner
- Filter paper
- Watch glasses
- Stereoscope
- Sodium chloride ($NaCl$)
- Copper(II) sulfate pentahydrate ($CuSO_4 \bullet 5\ H_2O$)
- Potassium aluminum sulfate [$KAl(SO_4)_2$]
- Copper wire
- Aluminum wire
- 0.1 *M* silver nitrate ($AgNO_3$) solution
- 0.1 *M* copper(II) nitrate [$Cu(NO_3)_2$] solution
- Sodium thiosulfate ($Na_2S_2O_3 \bullet 5\ H_2O$) crystals
- Test tube
- Utility clamp
- Cold running water
- Bright window or other good light source

Instructional activity

Content/Teacher Notes

Students will find this lab a fun way to learn the nomenclature for ionic compounds. They will take real ownership of their crystals and will want to check on them daily. They may even want to continue growing larger crystals after school, an activity that can yield some fantastic results.

Copper sulfate pentahydrate will grow beautiful blue rhombic crystals. The quality of the crystals greatly depends on the quality of the copper sulfate compound used. If you have crystalline reagent grade cupric sulfate, you will get first-rate crystals. Try to avoid using powered copper sulfate, as it does not produce very good results. Students can get great results using copper sulfate purchased from hardware or lawn care stores.

Students can change the rate at which the solution evaporates by covering it or leaving it uncovered in order to see how this affects the size of the crystals.

Potassium aluminum sulfate (alum) will grow clear hexagonal crystals, and sodium chloride will grow clear cubic crystals. Potassium aluminum sulfate can be purchased at the grocery store. The silver crystals are really beautiful and are amazing to watch under the stereoscope because they look like shiny silver trees.

The following background information is essential to the understanding of this activity:

- The behavior of substances in the solid state is governed mainly by the way in which the atoms, ions, or molecules of the substance are arranged and the forces that hold them together.
- These experiments deal only with ionic and metallic crystalline solids. Each of these is distinguished by the kinds of particles that make up the three-dimensional structure of the solid. If the structure has a regular pattern, then the solid is described as "crystalline." A familiar example would be sodium chloride (NaCl), in which sodium ions and chloride ions alternate in an extended structure known as a "crystal lattice." In this structure, each ion occupies a point in space called a "lattice point." This is an example of an ionic crystalline solid since ions are the structural unit. Any crystalline solid would exhibit a similar orderly arrangement of particles. Some solids, however, do not have such organization on the molecular level. Substances such as glass are really super-cooled liquids or amorphous solids because they lack a definite crystal-lattice organization. In this experiment, we are concerned only with crystalline solids.
- **Ionic crystals** are, of course, composed of ions. These are held tightly in place in the crystal lattice by strong electrostatic forces because of the opposite charges of the ions. These very strong forces contribute to high melting points. Although the ions are charged, they cannot move in the solid; hence, these crystals in their solid form do not conduct electricity. However, when melted, most of the attractions in the lattice structure have been overcome, and molten ionic compounds conduct electricity very well since the charged ions are now free to move. Many ionic compounds also dissolve in water. The amount that dissolves is completely separated into ions and thus makes a solution which is a good conductor of electricity.
- Ionic crystals get larger by adding layers of ions on the outside of the base structure. Sometimes a "seed crystal" is added to form this base. Crystals can be grown from a solution of a substance dissolved in a liquid; when the liquid evaporates, the dissolved substance joins together in a crystal structure. The more slowly the liquid cools and evaporates, the larger and more pure the crystal structure will be.
- **Metallic crystals** consist of metal cations positioned at the lattice points of the crystal. The valence electrons are given up to a common "sea" of electrons that flows freely throughout the metal cations. This property in metallic bonding makes most metals malleable. Metals are excellent conductors of electricity because of all the mobile electrons. This conductivity exists in both the solid and liquid state.
- **Supersaturated solutions** are highly unstable solutions in which a saturated solution has been forced to hold more of a dissolved substance than it should be able to hold at that temperature. Supersaturated solutions are formed by quickly and carefully cooling saturated solutions. The

crystals will not form unless the solution is jarred or a "seed crystal" is introduced. Once the crystallization process starts, it continues rapidly until all of the solution is crystallized.

- The test tube should feel warm during the crystallization process. Crystallization is an exothermic process that gives off heat energy. When bonds form, energy is released in the form of heat. The crystal-lattice structure is more stable and lower in energy than the solution, so by forming the crystal lattice, the compound increases its stability and decreases its potential energy.
- Sodium thiosulfate is a hydrated compound that contains water molecules bonded to the lattice structure. When heated, the water molecules are removed, and the remaining substance dissolves in the water. This makes it easy to make a supersaturated solution of sodium thiosulfate.

Introduction

1. Explain to the students that they will be dealing with only ionic and metallic crystalline substances in this experiment. Each of these is distinguished by the kinds of particles that make up the three-dimensional structure of the solid. If the structure has a regular pattern, then the solid is described as "crystalline." Ionic crystals are, of course, composed of both cations and anions. Metallic crystals consist of only metal cations positioned at the lattice points of the crystal.

Procedure

Experiment 1: Ionic Crystal Formation

1. Place approximately 50 mL of distilled water in a clean beaker, and begin heating. Be sure to keep the water below boiling temperature.
2. Slowly add solid sodium chloride, and allow it to dissolve while stirring.
3. Continue to add the salt until no more will dissolve and a small amount of undissolved solid remains in the bottom of the beaker. This is now a saturated solution.
4. Filter the hot solution into a clean, dry beaker, and cover with a watch glass, leaving a small opening for evaporation to take place.
5. Set the beaker aside, and do not disturb it until the crystals have completely formed.
6. When the crystals have completely formed, observe their structure under the stereoscope.
7. Repeat the above procedure with copper(II) sulfate pentahydrate and again with potassium aluminum sulfate.

Experiment 2: Metallic Crystal Formation

1. Place a small piece of clean copper wire on a watch glass, and place under the stereoscope.
2. Use a plastic pipette to add a small amount of silver nitrate solution ($AgNO_3$) to the copper metal, and observe the growth of silver (Ag) crystals. The reaction is expressed as $2\ AgNO_3 + Cu \rightarrow Cu(NO_3)_2 + 2\ Ag$. *CAUTION! $AgNO_3$ causes skin discoloration.*
3. Observe the growth of the metal crystals under the stereoscope.
4. Repeat the above procedure with aluminum wire and a copper(II) nitrate [$Cu(NO_3)_2$] solution. The reaction is expressed as $3\ Cu(NO_3)_2 + 2\ Al \rightarrow 2\ Al(NO_3)_3 + 3\ Cu$.

Experiment 3: Supersaturated Solution Crystal Formation

1. Fill a large, clean test tube slightly more than half full with sodium thiosulfate ($Na_2S_2O_3 \cdot 5\ H_2O$) crystals. Sodium thiosulfate is a hydrated crystal: it contains water molecules chemically bonded to the crystal-lattice structure.

2. Using a utility clamp, heat the test tube containing the sodium thiosulfate. As you heat the substance, the water will come out of the crystals, and then the rest of the substance will dissolve in the water. Heat the test tube until all of the crystals dissolve.
3. When the solution begins to boil, heat it *gently* so it will not shoot out of the test tube. *CAUTION! Do not point the test tube at anyone, including yourself.* Boil the solution for a few seconds to rinse any remaining crystals from the upper part of the test tube; then turn off the burner.
4. At this point, you will have a saturated solution. Hold the test tube of hot liquid very still under cold running water, or place it in a beaker of cold water. Do not shake or stir the solution. Do not let any water run into the solution, as this would dilute it. Cool the liquid for several minutes, and then slowly remove the test tube from the running water or water bath. Wait about 10 seconds. If the test tube still feels warm, cool it a little longer; if it feels cool, stop cooling it.
5. By cooling the saturated solution rapidly and carefully, you have created a supersaturated solution. The cooled solution actually contains more solid than it should contain at that temperature. This is a very unstable solution, and it will not stay supersaturated very long. It can be easily changed by dropping in a seed crystal of the solid.
6. Place the test tube of cool liquid in a bright window or near a good light. Be careful not to shake or disturb the solution at this point. Drop a single crystal of sodium thiosulfate into the liquid in the test tube, and observe what happens. Feel the sides of the test tube for the release of heat energy as the crystal-lattice structure forms.
7. Repeat this process several times with the same test tube and contents. Each time, you create a more supersaturated solution, and crystallization will occur more rapidly.
8. Clean up your test tube as directed by your instructor. (You may have to throw it away.)

Observations and Conclusions

1. Observe your crystals under the stereoscope, and draw the crystal structures you see.
2. Describe the differences and similarities between the ionic and metallic crystals you have made.

Sample assessment

- Organize and hold a crystal growing contest with carefully specified rules. Judge the students' crystals on the basis of shape and color, and grade their lab reports on their crystals.
- Have students research the actual crystalline geometry for the crystals they grow, and then discuss any differences between the lab-grown crystals and the expected geometry.

Follow-up/extension

- Challenge students to grow larger and more complex crystals. There is a lot of information on the Internet about crystal-growing projects. Have them research a crystal they would like to work on further, and encourage them to experiment with different crystal-growing techniques.

Molecular Model Building

Organizing Topic Bonding, Nomenclature, and Formula Writing

Overview Students use molecular model kits and draw Lewis dot diagrams to explore molecular geometry and molecular polarity.

Related Standards of Learning CH.1a, b, c; CH.3c, d

Objectives

The students will

- summarize the following concepts about covalent bonding:
    - Covalent bonds involve sharing of electrons.
    - Polar molecules result when a molecule behaves as if one end were positive and the other negative.
    - Electronegativity is the measure of attraction of an atom for electrons in a covalent bond.
- draw Lewis dot diagrams to show covalent bonding;
- predict, draw, and name linear, bent, trigonal planar, tetrahedral, and trigonal pyramidal molecular shapes;
- recognize polar and nonpolar molecules;
- understand and demonstrate
    - MSDS warnings
    - safety rules for science
    - laboratory safety cautions
    - safe techniques and procedures.

Materials needed

- Molecular model kits (or toothpicks and small Styrofoam balls or gum drops)

Instructional activity

Content/Teacher Notes

In this activity, the student will gain valuable experience working with some hands-on molecular models as they practice drawing Lewis dot structures and predict shapes according to the **valence shell electron pair repulsion rule** (VSEPR).

You will need a classroom set of basic molecular model kits. Number each kit, and sign them out by number; this will greatly reduce the loss of parts from one lab section to the next. If you do not have access to molecular model kits, you can use toothpicks and gumdrops or small Styrofoam balls, but the bond angles will be difficult to achieve.

Students who find the molecules in this lab easy to build can be assigned additional, advanced molecules. Suggested advanced molecules and ions are listed below, grouped according to their shape:

Linear

CO_2
CS_2
OCS
BeF_2
NO_2^+
NCF
SCN^-
OCN^-
N_2O

Tetrahedral arrangement

V-shaped (bent)
- SCl_2
- $SeBr_2$
- BrO_2^-
- ClO_2^-
- OF_2

Trigonal pyramidal
- NF_3
- XeO_3
- SO_3^{2-}
- PO_3^{2-}
- $SOCl_2$
- PCl_3
- ClO_3^-
- IO_3^-

Tetrahedral
- SO_4^{2-}
- XeO_4
- PO_4^{3-}
- BF_4^-
- CF_4
- BrO_4^-
- SiF_4
- SiO_4^{4-}
- NF_4^+
- $AsCl_4^+$
- PF_4^+
- $SiCl_4$
- ClO_4
- CCl_4
- NOF_3
- SO_2Cl_2

Trigonal bipyramidal arrangement

Linear
- XeF_2
- KrF_2
- ClF_2^-
- BrF_2^-
- IF_2^-

T-shaped
- ClF_3
- BrF_3
- IF_3
- $XeOF_2$
- XeF_3^+

See-saw
- SCl_4
- IO_4^{3-}
- XeO_2F_2
- PF_4^-
- $AsCl_4^-$

Trigonal bipyramidal
- SbF_5

Octahedral arrangement

Square planar
- IF_4^-
- BrF_4^-

Square pyramidal
- $XeOF_4$
- ClF_5
- SeF_5^-
- TeF_5^-
- $ClOF_4^-$
- IF_5
- BrF_5
- XeF_5^+
- $SbCl_5^{2-}$

Octahedral
- AlF_6^{3-}
- SiF_6^{2-}
- PF_6^-
- IO_6^{5-}
- IF_6^+

Introduction

1. Review with students the following basic points:
    - The properties of molecules depend not only on their molecular composition and structure, but also on their shape. Molecular shape determines a compound’s state of matter, boiling point, freezing point, viscosity, volatility, surface tension, and the nature of its reactions.
    - The geometry of a small molecule can be predicted by examining the central atom and identifying the number of atoms bonded to it and the number of unshared electron pairs surrounding it.
    - The shapes of molecules may be predicted by using the VSEPR rule, which states that electron pairs around a central atom will position themselves to allow for the maximum amount of space between them.

  - Covalent bonds can be classified as either polar covalent or nonpolar covalent. In a **polar covalent bond,** electrons are more attracted to the atom with the greater electronegativity, resulting in a partial negative charge on that atom. The atom with the smaller electronegativity value acquires a partial positive charge. In a **nonpolar covalent bond,** the electrons are shared equally, and there is no unequal distribution of charge.
  - Molecules made up of covalently bonded atoms can be classified as either polar or nonpolar. The geometry of the molecule determines whether it is polar or not. For example, if polar bonds are symmetrically arranged around a central atom, their charges may cancel each other out and the molecule is nonpolar. If, on the other hand, the arrangement of the polar bonds is asymmetrical, the electrons will be attracted more to one end of the molecule and a polar molecule or dipole results.

2. Tell students that in this activity, they will construct ball-and-stick models of covalent molecules and name the geometry and the polarity of each molecule.

Procedure

1. Distribute molecular model kits (or toothpicks and small Styrofoam balls or gum drops) to each team of students. Provide each group with enough parts to build several models before having to take them apart to build new ones.
2. Have students build models of assigned molecules by following these steps:
   - Sum all of the molecule's valence electrons.
   - Place the central atom, which is usually the first atom except for H.
   - Arrange the remaining atoms around the central atom.
   - Arrange the electrons around the atoms to obey the octet rule for all elements except H. (C, N, O, and F always obey the octet rule.)
   - Obey the duet rule for H.
   - Form single bonds if possible. (Form multiple bonds only if absolutely necessary.)
   - Elements in the third row or below on the periodic table can hold extra electrons if they are the central atom. (They have d-orbitals that can hold the extra electrons).
3. Have each student draw Lewis dot diagrams for the following atoms or molecules:

N
S
CH_4
H_2
N_2
O_2
Cl_2
H_2O
HCN
NH_3
CO_2
H_2O_2
O_3
SO_4^{2-}
BrF_3
XeO_4
H_2Se
CO_3^{2-}
ICl_3
CH_2Cl_2
C_2H_2
C_6H_6

Follow-up/extension

- Reinforce the concept of conservation of mass in a chemical reaction by doing the following:
  - Have students build models of the compounds CH_4 and O_2.
  - Have students write and balance the equation for the reaction that occurs between CH_4 and O_2.
  - Then, have students build the number of CH_4 and O_2 molecules that are present in the balanced equation.
  - When these models are complete, have students "react" the methane and oxygen models by breaking the bonds in the reactants and reforming the product molecules. They should see that they still have the same number of atoms in the product molecules that they had in the reactants.

This process can be done for a variety of reactions.

Sample assessment

Use the following questions and/or assignments for assessment:

- Draw the Lewis diagram for oxygen.
- Draw the Lewis diagram for SF_4.
- How many lone pairs of electrons are on the sulfur atom in the sulfite ion, SO_3^{2-} ?
- Use the VSEPR theory to predict the molecular geometry of ICl_3.
- Use the VSEPR theory to predict the molecular geometry of SO_2.
- Use the VSEPR theory to predict the molecular geometry of H_2Se.
- Use the VSEPR theory to predict the molecular geometry of NH_4^+.
- What is the polarity of the water molecule? Explain your answer.

Structure and Polarity of Molecules Lab
Molecular Geometry Charts

Basic Structures

Total # of e^- pairs	# of bonding pairs	# of lone e^- pairs	Molecular geometry	Bond angles
2	2	0	Linear	180°
3	3	0	Trigonal planar	120°
3	2	1	Bent	<120°
4	4	0	Tetrahedral	109.5°
4	3	1	Trigonal pyramidal	<109.5°
4	2	2	Bent	<109.5°
4	1	3	Linear	180°

More Advanced Structures

Total # of e^- pairs	# of bonding pairs	# of lone e^- pairs	Molecular geometry	Bond angles
5	5	0	Trigonal bipyramidal	90°, 120°
5	4	1	See-saw	180°, 90°, <120°
5	3	2	T-shaped	90°
5	2	3	Linear	180°
6	6	0	Octahedral	90°
6	5	1	Square pyramidal	<90°
6	4	2	Square planar	90°
6	3	3	T-shaped	90°
6	2	4	Linear	180°

For each molecule listed in the box below, use a separate sheet of paper to

1. determine the total number of e^- pairs
2. draw the Lewis structure
3. determine the number of bonding e^- pairs
4. determine the number of lone e^- pairs
5. state the shape of the molecule
6. state whether the molecule is polar or nonpolar.

H_2
N_2
CO_2
O_2
CH_4
HBr
H_2O
NH_3
C_2H_5OH
HCN
H_2O_2
SO_3
$SeBr_2$
PCl_3
CF_4
C_2H_2
CH_3NH_2
BCl_3
$XeOF_4$
Br_3^-
SeF_4
IF_5
AsF_5
XeF_4
IF_3
BH_2^-
$COCl_2$
N_2F_2
C_6H_6
$CH_3CH_2CH_2OCH_2CH_2CH_3$
$CH_3CHCHCH_2CH_2CH_3$
C_6H_{12}
acetone
isopropyl alcohol
acetic acid
boric acid

Mystery Anions

Organizing Topic Bonding, Nomenclature, and Formula Writing

Overview Students use some simple tests to identify anions in a solution.

Related Standards of Learning CH.1a, b, c; CH.3a

Objectives

The students will

- recognize the formulas and names of the polyatomic ions carbonate, sulfate, nitrate, and hydroxide;
- identify and use
  - chemical formulas
  - coefficients, chemical symbols, and subscripts;
- understand and demonstrate
  - MSDS warnings
  - safety rules for science
  - laboratory safety cautions
  - safe techniques and procedures.

Materials needed

- Distilled water
- Sodium carbonate
- 0.2 *M* $Ba(OH)_2$ [or $Ca(OH)_2$] solution
- 6.0 *M* HCl solution
- Sodium chloride
- 0.1 *M* silver nitrate solution
- Magnesium sulfate solid
- 0.2 *M* $BaCl_2$ solution
- Sodium iodide
- Bleach
- Dropper
- Starch solution
- Test tubes

Instructional activity

Content/Teacher Notes

This simple lab will give students a chance to see that unknown compounds can be identified according to the chemical behaviors of their anions. The lab can be used at the same time the students are memorizing the polyatomic ions.

Introduction

1. Explain to the students that this lab will introduce them to qualitative analysis, the area of chemistry concerned with the identification of substances by their chemical reactions. Students will observe the reactions of some simple salts and will identify unknown samples by testing their reactivity.

2. Review all safety rules and precautions needed in the following experiments.
3. Review all waste-disposal procedures needed in the following experiments:
   - Place all solutions and precipitates containing silver in labeled waste jars.
   - Dispose of all solutions not containing silver by pouring them down the sink.
   - Do not leave insoluble solids in the sink, but dispose of them in the trash can.

Procedure

Experiment 1: The Carbonate Ion, CO_3^{2-}

1. Place a pea-sized amount of sodium carbonate solid in a small, clean test tube that has been rinsed with distilled water.
2. Add to it 1 or 2 drops of 6.0 *M* HCl solution *(CAUTION! Strong acid)*, and record your observations.
3. Confirm the presence of CO_2 by dripping a drop of 0.2 *M* $Ba(OH)_2$ solution down the side of the test tube as the gas bubbles are forming. Watch carefully for the precipitate of barium carbonate in the drop of solution running down the side of the test tube. [You can substitute $Ca(OH)_2$ solution for the $Ba(OH)_2$ solution, in which case you would have a precipitate of calcium carbonate.]

Experiment 2: The Chloride Ion, Cl^-

1. Place a pea-sized amount of sodium chloride solid in a small, clean test tube that has been rinsed with distilled water.
2. Add 15 drops distilled water.
3. Then, add 3 or 4 drops of 0.1 *M* silver nitrate solution, and mix the contents. Observe the formation of a white precipitate of silver chloride. Record your observations.
4. Now, test your tap water for the presence of the chloride ion. Place about 2 mL of *tap* water (without NaCl) into a clean, rinsed test tube, and add the silver nitrate solution. Look carefully for precipitate. You should look through the test tube the long way — through the opening — and compare what you see to a test tube containing only distilled water. What can you conclude about the presence of chloride ions in the tap water?

Experiment 3: The Sulfate Ion, SO_4^{2-}

1. Place a pea-sized amount of magnesium sulfate solid in a small, clean test tube that has been rinsed with distilled water.
2. Add about 1 mL of distilled water.
3. Then, add a few drops of 0.2 *M* $BaCl_2$ solution, and mix the contents. Observe the formation of a precipitate of barium sulfate. Record your observations.

Experiment 4: The Iodide Ion, I^-

1. Place a pea-sized amount of sodium iodide solid in a small, clean test tube that has been rinsed with distilled water.
2. Add 1 mL of distilled water.
3. Then, add 5 drops of bleach *(CAUTION! Avoid getting bleach on your skin.)*, and mix the contents.
4. Add a dropper-full of starch solution to the mixture, and record your observations. Confirm the presence of I^- by observing the reaction between the starch and the reaction mixture. Record your observations.

Experiment 5: An Unidentified Ion

1. Give students an unidentified sample of one of the ions studied in the previous experiments, along with the necessary materials to carry out tests to identify the ion.
2. Allow students to carry out the experiments to identify the ion, making sure they use all appropriate safety precautions.
3. Have students record their observations in their notebooks. Emphasize that they must write down exactly what they do in the lab and all their observations.
4. Have students turn in their lab reports with the mystery ion identified and their supporting evidence cited.

Sample assessment

- Assess students' lab reports identifying the mystery ion.

Follow-up/extension

- Demonstrate the testing of several other mystery anions, such as those detailed below, and have the students record their observations. Alternatively, have students research other possible tests for these and/or other anions. Hold a class discussion on the results and conclusions.

  The Carbonate Ion, CO_3^{2-} and Hydrogen Carbonate Ion, HCO_3^-

  1. Place a pea-sized amount of baking soda in a small, clean test tube that has been rinsed with distilled water.
  2. Add to it 1 or 2 drops of sulfuric acid (H_2SO_4) *(CAUTION! Strong acid)*, and record your observations.
  3. Confirm the presence of CO_2 by dripping a drop of $Ba(OH)_2$ solution down the side of the test tube as the gas bubbles are forming. Watch carefully for precipitate of barium carbonate in the drop of solution running down the side of the test tube.

  The Chloride Ion, Cl^-

  1. Place a pea-sized amount of sodium chloride in a small, clean test tube that has been rinsed with distilled water.
  2. Take the tube to the fume hood, and carefully add 1 or 2 drops of sulfuric acid. *CAUTION! Avoid breathing the gas formed.*
  3. Carefully lower a piece of moist pH indicator paper into the tube as the gas evolves. Indicator paper is impregnated with a colored compound which is sensitive to acids and bases. It is used to categorize solutions and gases as strongly acidic (red), neutral (no significant color change), or basic (blue). Record your observations.
  4. Write a balanced chemical equation to describe the reaction taking place.

  The Chloride Ion, Cl^-

  1. Place a pea-sized amount of NaCl in a small, clean test tube that has been rinsed with distilled water.
  2. Add 15 drops of distilled water and one drop of 3.0 *M* nitric acid (HNO_3) solution. *(CAUTION! Strong acid)*
  3. Then, add 3 or 4 drops of 0.1 *M* silver nitrate solution, and mix the contents. What happens? Record your observations.

4. Write a balanced chemical equation for the observed reaction. The purpose of the nitric acid in this reaction is to prevent the precipitation of undesired silver salts, like AgOH, which can occur with non-acidic conditions. It is not a reactant in the balanced net ionic equation for this reaction of chloride ion and silver nitrate.

The Chloride Ion, Cl^-

1. To test your tap water for the presence of chloride ion, place about 2 mL of *tap* water (without NaCl) into a clean, rinsed test tube
2. Add 1 drop of 3.0 *M* HNO_3 solution.
3. Then, add the silver nitrate. Look carefully for precipitate. You should look through the test tube the long way — through the opening — and compare what you see to a test tube containing only distilled water. What can you conclude about the presence of chloride ions in the tap water?
4. Write a balanced chemical equation for the observed reaction.

The Sulfate Ion, SO_4^{2-}

1. Place a pea-sized amount of Epsom salts (magnesium sulfate) in a small, clean test tube that has been rinsed with distilled water.
2. Add 1 or 2 drops of sulfuric acid. Compare your observations with those made above for the carbonate ion, and record them in your notebook.
3. Write a balanced chemical equation for the observed reaction.

The Sulfate Ion, SO_4^{2-}

1. Place a pea-sized amount of Epsom salts (magnesium sulfate) in a small, clean test tube that has been rinsed with distilled water.
2. Add about 1 mL of distilled water, and mix the contents.
3. Add 1 drop of 3.0 *M* nitric acid solution and then 1 or 2 drops of 0.2 *M* $BaCl_2$ solution. Compare the result of this experiment to those seen above for the carbonate ion.
4. Write a balanced chemical equation for the observed reaction.

The Iodide Ion I^-

1. Place a pea-sized amount of sodium iodide (NaI) in a small, clean test tube that has been rinsed with distilled water.
2. Add 1 mL of distilled water, and mix the contents.
3. Then, add 5 drops of bleach *(CAUTION! Avoid getting bleach on your skin.),* and record your observations.
4. Write a balanced chemical equation for the observed reaction.
5. Confirm the presence of I_2 and I^- by observing the reaction between a cornstarch packing peanut and the reaction mixture in your test tube. Tear off a piece of a packing peanut, and push it into the solution with a stirring rod. Record your observations.

The Iodide Ion I^-

1. Place a pea-sized amount of sodium iodide (NaI) in a small, clean test tube that has been rinsed with distilled water.
2. Add 1 mL of distilled water, and mix the contents.

3. Add 1 drop of 3.0 *M* HNO_3 solution and then 3 or 4 drops of 0.1 *M* silver nitrate solution, and record your observations.
4. Write a balanced chemical equation for the observed reaction.

The Iodide Ion I^-

1. Place a very small amount of NaI in a small, clean test tube.
2. In the fume hood, carefully add 1 or 2 drops of sulfuric acid. *CAUTION! Avoid breathing the gases formed.* Record your observations.
3. Write a balanced chemical equation for the observed reaction.
4. Confirm the presence of I_2 and I^- by observing the reaction between a cornstarch packing peanut and the reaction mixture in your test tube. First, add about 1 mL of distilled water, and mix the contents. Then, tear off a piece of a packing peanut, and push it into the solution with a stirring rod. Record your observations.

Mystery Iron Ions

Organizing Topic Bonding, Nomenclature, and Formula Writing

Overview Students use mole ratios to determine the formula for an iron compound.

Related Standards of Learning CH.2

Objectives

The students will
- recognize that transition metals can have multiple oxidation states;
- name binary and ionic compounds, using the roman numeral system where appropriate;
- investigate and understand that quantities in a chemical reaction are based on molar relationships;
- identify and use
  - chemical formulas
  - coefficients, chemical symbols, and subscripts.

Materials needed

- 1 mL of phenolphthalein indicator (0.13 g of solid indicator dissolved in a 50% ethanol/water mixture) or universal indicator
- 2 mL of 0.1 *M* $FeCl_3$ solution (0.82 g of solid iron(III) chloride dissolved in 30 mL of solution)
- 10 mL of 0.1 *M* NaOH solution (0.60 g of NaOH dissolved in 150 mL of solution)
- 24-well plates
- Toothpicks
- Thin stem dropper pipettes

Instructional activity

Content/Teacher Notes

Students will use a double replacement reaction to determine the charge on the iron ion in an iron(III) chloride solution. The reaction is complete when the NaOH solution has no more metal ions with which to react. The indicator phenolphthalein will be used to show when excess hydroxide ions have been added.

Introduction

1. Explain to the students that in this experiment, they will determine the formula for a transition metal salt and find out whether the compound is iron(II) chloride or iron(III) chloride.

Procedure

Have the students do the experiment, as follows:

1. Prepare a data table like the one shown at right.
2. Fill a clean, dry pipette with the mystery iron chloride solution (0.1 *M* $FeCl_3$), and add 5 drops of the solution to each of four wells in the well plate. Make all the drops the same size by holding the pipette at the same angle for all drops.

Trial	# Drops of NaOH
1	
2	
3	
4	

3. Place one drop of the phenolphthalein indicator into the wells, and stir to mix well.

4. Fill another clean, dry pipette with the sodium hydroxide solution (0.1 *M* NaOH), and add it drop by drop to the first well. Stir the solution with a clean toothpick after adding each drop. Keep count of the number of drops you are adding to the well.
5. Continue adding the NaOH solution until the pink color of the indicator phenolphthalein just begins to show up clearly. The pink color change indicates that all of the iron chloride has reacted.
6. Record in your data table the number of drops of NaOH that you added to the iron chloride solution.
7. Repeat the procedure in steps 4–6 for the solution in the other three wells. Be sure to use a clean toothpick for each trial. Be as consistent as possible when judging the appearance of the pink color.
8. Dispose of the chemicals as directed by your teacher.

Observations and Conclusions

Prompt conclusions by having the students do or answer the following:

1. Write a balanced equation for a double replacement reaction between sodium hydroxide and
   - iron(II) chloride *($FeCl_2 + 2\ NaOH \rightarrow Fe(OH)_2 + 2\ NaCl$)*
   - iron(III) chloride. *($FeCl_3 + 3\ NaOH \rightarrow Fe(OH)_3 + 3\ NaCl$)*
2. Assuming that all of the drops were the same size, how do the numbers of moles of compound in each drop of the solution compare? *(Drops of equal size will contain an equal number of moles of the compound.)*
3. Calculate the average number of drops of NaOH that were needed to react fully with the iron chloride solution. *(Answers will vary but should average about 15 drops. This gives a ratio of 5 drops iron solution to 15 drops of the NaOH solution.)*
4. Using your answers to the questions above, determine how many moles of NaOH are needed to react with each mole of iron chloride.
5. Compare your answers to the previous question with your balanced equations in question #1. Was your iron chloride compound iron(II) chloride or iron(III) chloride? *(From the 5:15, or 1:3, ratio, students should interpret that they have an iron(III) chloride solution.)*
6. Share your data with other lab groups. Calculate a class average for the ratio of moles of NaOH to moles of iron chloride.
7. Compare the class average with your results: using the class average as the accepted value, calculate your percent error.
8. What are some possible sources of error in this experiment?

Follow-up/extension

- Using the same procedures, identify an unidentified solution of copper chloride as either copper(I) chloride or copper(II) chloride.

Properties of Compounds and Chemical Formulas

Organizing Topic Bonding, Nomenclature, and Formula Writing

Overview Students group unknown compounds according to their chemical behaviors and chemical formulas.

Related Standards of Learning CH.1a, b, c; CH.2g, h

Objectives

The students will
- recognize that reactivity is the tendency of an element to enter into a chemical reaction;
- recognize that bonds form to achieve stability;
- understand and demonstrate
  - MSDS warnings;
  - safety rules for science;
  - laboratory safety cautions;
  - safe techniques and procedures.

Materials needed

- 24-well plates
- Unknown solution 1: 5.0 grams of $NaBr$ dissolved in 100 mL of distilled water
- Unknown solution 2: 5.0 grams of $Ca(NO_3)_2$ dissolved in 100 mL of distilled water
- Unknown solution 3: 1.25 grams of Li_2CO_3 dissolved in 100 mL of distilled water
- Unknown solution 4: 4.0 grams of $Ba(NO_3)_2$ dissolved in 100 mL of distilled water
- Unknown solution 5: 4.5 grams of NaI dissolved in 100 mL of distilled water
- Unknown solution 6: 7.5 grams of K_2CO_3 dissolved in 100 mL of distilled water
- Unknown solution 7: 5.8 grams of $NaCl$ dissolved in 100 mL of distilled water
- Unknown solution 8: 3.8 grams of Na_2CO_3 dissolved in 100 mL of distilled water
- Unknown solution 9: 16.0 grams of $Sr(NO_3)_2$ dissolved in 100 mL of distilled water
- Test solution A: 7.5 grams of $Pb(NO_3)_2$ dissolved in 100 mL of distilled water
- Test solution B: 15.0 grams of Na_2SO_4 dissolved in 100 mL of distilled water
- Test solution C: 4 drops of 1.0 *M* HCl added to 1 mL of 1% phenolphthalein indicator and then diluted to 100 mL with distilled water
- Test solution D: 19 mL of concentrated HNO_3 added to 100 mL of distilled water *(CAUTION! Concentrated HNO_3 is very corrosive.)*
- Toothpicks
- Goggles
- Flasks or beakers
- Dropper pipettes

Instructional activity

Content/Teacher Notes

Before undertaking this experiment, be sure your school's chemical hygiene plan allows for the use of all the chemicals used. Be sure to have a container available to collect the waste since some of the compounds contain heavy metals and should not be poured down the sink drain.

In this experiment, students will group unknown compounds according to their chemical behaviors by

- mixing two solution together to see if a precipitate forms
- observing which precipitates are dissolved by another solution
- recording any color changes that occur when two solutions are mixed together.

After analyzing the data, students will place each unknown substance into a group with other substances that show similar behavior and characteristics.

Unknown and test solutions can be placed in labeled flasks or beakers in a central location. The students can then fill labeled dropper pipettes with the unknown and test solutions and return to their lab stations to fill their well plates. If you have enough dropper bottles available, these could be filled with the test solutions, and a set could be provided to each lab table or group.

Although students may not have worked with chemical equations yet, you may wish to share the following reactions with them to reinforce similarities in chemical behaviors. They can then see that the compounds with similar formulas undergo very similar reactions.

$2\ NaBr + Pb(NO_3)_2 \rightarrow 2\ NaNO_3 + PbBr_2$
$2\ NaI + Pb(NO_3)_2 \rightarrow 2\ NaNO_3 + PbI_2$
$2\ NaCl + Pb(NO_3)_2 \rightarrow 2\ NaNO_3 + PbCl_2$

$Ca(NO_3)_2 + Na_2SO_4 \rightarrow 2\ NaNO_3 + CaSO_4$
$Ba(NO_3)_2 + Na_2SO_4 \rightarrow 2\ NaNO_3 + BaSO_4$
$Sr(NO_3)_2 + Na_2SO_4 \rightarrow 2\ NaNO_3 + SrSO_4$

$Li_2CO_3 + Pb(NO_3)_2 \rightarrow PbCO_3 + 2\ LiNO_3$
$K_2CO_3 + Pb(NO_3)_2 \rightarrow PbCO_3 + 2\ KNO_3$
$Na_2CO_3 + Pb(NO_3)_2 \rightarrow PbCO_3 + 2\ NaNO_3$

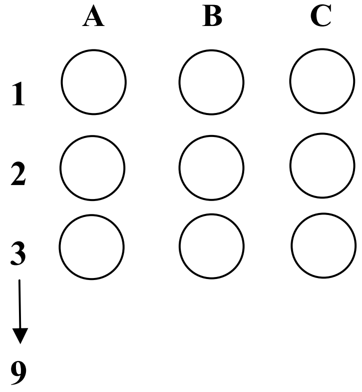
Introduction

1. Announce to the class that this laboratory experiment will enable them to group unknown compounds according to their chemical behaviors. Explain that they will mix two solutions together to see if a precipitate forms. They will observe which precipitates are dissolved by another solution. Finally, they will record any color changes that occur when two solutions are mixed together. After analyzing their data, they will place each unknown substance into a group with other substances of similar behavior and characteristics.

Procedure

Have students do the experiment, as follows:

1. Label the well plate vertically 1–9 for the nine unknown solutions, and horizontally A–C for test solutions A, B, and C, as shown below:



2. Place about 10 drops of unknown solution 1 into each of the three wells A, B, C in the first row. *CAUTION! Some of the solutions are corrosive. Avoid direct contact.*
3. To well A, add 5 drops of test solution A, and observe the results.

4. To well B, add 5 drops of test solution B, and observe the results.
5. To well C, add 5 drops of test solution C, and observe the results.
6. Use a clean toothpick to stir for 10 seconds any solution that does not show an immediate reaction, and then observe again. Record all results on a blank data chart similar to the filled-in one below.

Unknown solution	**A Precipitate? (yes/no)**	**B Precipitate? (yes/no)**	**C Color (describe color)**	**D Precipitate dissolved? (yes/no)**
1	*yes*	*no*	*none*	*no*
2	*no*	*yes*	*none*	*yes*
3	*yes*	*no*	*pink*	*yes*
4	*no*	*yes*	*none*	*yes*
5	*yes (yellow)*	*no*	*none*	*no*
6	*yes*	*no*	*pink*	*yes*
7	*yes*	*no*	*none*	*no*
8	*yes*	*no*	*pink*	*yes*
9	*no*	*yes*	*no*	*yes*

7. To any mixture that produces a precipitate, add 15 drops of test solution D, stir with a clean toothpick for 20 seconds, and then observe again. Solution D is to be added only to those mixtures that produce a precipitate.
8. Repeat steps 2–7 for each of the other unknown solutions 2–9. Record all results.
9. Pour waste material from the wells into the waste container. Rinse out the well plates thoroughly for use in the next laboratory. Clean up your work area, and wash your hands.
10. Analyze your observations by arranging the nine unknown solutions into groups according to similarities in their chemical behaviors. Justify your groups with data from the laboratory results.
11. After being provided with the chemical formula for each of the nine unknown substances, examine the similarities in the compounds, based on their formulas. Relate the similar chemical formulas to the similar chemical behaviors.

Observations and Conclusions

Possible Groups:

1. Solutions 1, 5, and 7 had no reaction with B or C and formed a precipitate with A that did not dissolve with D. Solutions 1 (NaBr), 5 (NaI), and 7 (NaCl) are sodium salts of halogens (sodium halides). (Students may place #5 in a group by itself due to the yellow color of the precipitate.)
2. Solutions 2, 4, and 9 had no reaction with A or C and formed a precipitate with B that dissolved with D. Solutions 2 [$Ca(NO_3)_2$], 4 [$Ba(NO_3)_2$], and 9 [$Sr(NO_3)_2$] are alkaline earth nitrates.
3. Solutions 3, 6, and 8 had no reaction with B, turned pink with C, formed a precipitate with A that dissolved with D. Solutions 3 [Li_2CO_3], 6 [K_2CO_3], and 8 [Na_2CO_3] are alkali metal carbonates.

Sample assessment

- Have students place the nine unknown solutions into groups according to similar chemical behaviors. Have them present their groupings to the class, and then hold a class discussion of any differences in groupings. Have students examine the similarities in the compounds based on their chemical formulas and relate similar chemical formulas to similar chemical behaviors.

Follow-up/extension

- Have students research the compounds used in this lab to find additional similarities in both physical and chemical behaviors.

Matter and Energy: Equations and Formulas

Organizing Topic Bonding, Nomenclature, and Formula Writing

Overview Students focus on the concepts of naming elements, molecules, and compounds. They also examine the Law of Conservation of Matter through balancing chemical equations and writing chemical formulas. Students make 3-D models and infer conclusions from them.

Related Standards of Learning CH.3a, b, c

Objectives

The students will
- recognize that when atoms form compounds, they undergo a chemical change, such as losing or gaining electrons;
- understand ions, cations, and anion;
- understand that ions may be positive or negative and are formed when an atom or a group of atoms either gains electrons or loses electrons;
- recognize that cations are positive and are identified as metals. They lose electrons when chemically combined. In naming compounds, the cations are named first, using their names as they appear on the periodic table;
- recognize that anions are negative and are identified as nonmetals. They gain electrons when chemically combined. In naming compounds, the anions are named second or after cations. The ending of their elemental name is changed to *-ide*;
- understand that a binary compound consists of a compound that contains two different elements;
- understand that a ternary compound contains the atoms from three different elements;
- recognize that Roman numerals are used to name cations, which have more than one charge;
- learn how to write compounds by using the "crisscross method";
- learn that in a balanced chemical equation, the same number of atoms of an element or compound is found on both the reactant side and the product side.

Materials needed

Demonstration materials:
- Bare cooper wire
- Zinc chips
- Charcoal
- Empty covered jar
- Beaker of the compound water
- 250-mL beaker
- 250-mL flask
- White vinegar
- Steel wool
- Balloon
- Balance
- Periodic table of the elements

Lab materials per group:
- Solution of 4 g of NaOH dissolved in 1 L
- Solution of 22 g of $CuSO_4$ dissolved in 1 L

- Solution of NH_4OH
- Solution of 24 g of $Zn(NO_3)_2 \times 3\ H_2O$ dissolved in 1 L [or 30 g of $Zn(NO_3)_2 \times 6\ H_2O$ in 1 L]
- Four 3-oz. plastic cups
- Three 5-oz. plastic cups
- Balance
- Graduated cylinder
- Four pipe cleaners
- Large Styrofoam ball
- 4 small Styrofoam balls
- Metric ruler
- Protractor

Instructional activity

Content/Teacher Notes

For complete understanding of modern chemical nomenclature, the students must be familiar with the following terms: *ion, cation, anion, binary, ternary, molecule, molecular compound, ionic compound, chemical formula, molecular formula, formula unit, polyatomic ions.*

The concept of empirical formulas can be introduced on day four, but it is best taught after the students have been introduced to the mole.

Note that the iron in the steel wool used in the demonstration rapidly oxidizes (rusts) to form iron(III) oxide. The result is a rusty piece of steel wool and a balloon that has been completely sucked into the flask. However, the mass both before and after the reaction is the same. This should provide you with a nice hook to begin a discussion on the Law of Conservation of Matter.

Introduction

1. Display samples of common elements (e.g., bare copper wire, zinc chips, charcoal), an empty covered jar to represent diatomic molecules of oxygen gas and nitrogen gas (which comprise 99% of the air we breathe), and a beaker of the compound water.
2. Using the sample displayed elements, help students come up with working definitions of the following terms:
    - **element:** any substance that cannot be split into simpler substances by ordinary chemical means
    - **molecule:** a particle that has two or more atoms bonded together
    - **compound:** a substance with different kinds of atoms
3. Distribute copies of the periodic table of the elements, or have students use the one in their textbook. Point out the seven diatomic elements: H, N, C, O, F, Cl, and Br.
4. Explain the rules for writing a basic ionic compound formula:
    - Ionic compounds usually contain a metal and a nonmetal.
    - Use charges to write the formula.
    - If the ionic compound is a transition metal, a Roman numeral gives the charge.
    - The *-ide* ending indicates a binary compound without a polyatomic ion.
5. Explain the rules for using polyatomic ions in ionic compounds:
    - Each polyatomic ion is a complete unit; *never* break it up or change the numbers.
    - Use charges to write the formula, just as you would with regular ionic compounds.
    - Most names end in *-ate* and *-ite*; only a few (e.g., cyanide, hydroxide) have an *-ide* ending.

6. Explain the rules for writing a formula for a molecular compound:
   - Molecular compounds are nonmetals.
   - Do *not* use charges to write the formula.
   - The prefix gives the number of each element.
7. Explain the rules for writing the name of a compound:
   - Check if it is ionic or molecular.
   - If it is molecular, the prefix gives the number of each element.
   - If it is ionic, name the first element, then name the second ion. If it is a single element, then just change the ending to *-ide*. If it is more than one element, it is a polyatomic ion (which you have memorized). Use the name you memorized.
   - Check if the metal is a transition metal. If it is, use a Roman numeral.
   - Figure out the charge by working backwards. Examples:
     - ° $FeCl_2$ would be iron chloride, until you noticed that iron is a transition metal. To figure the charge, first find the charge of the negative ion (chloride). It is Cl^{-1}. Since there are two of these chlorides given in the formula $FeCl_2$, the total negative charge is −2. Therefore the iron must be a +2 to offset it. The name is "iron(II) chloride."
     - ° PtO_2: O has a −2 charge, and there are two of them, so the total negative charge is −4. Therefore, the platinum (Pt) must have a charge of +4. The name is "platinum(IV) oxide."
     - ° Fe_2O_3: O has a −2 charge, and there are three of them, so the total negative charge is −6. Therefore, the iron (Fe) must have a total positive charge of +6. Because there are 2 Fe, each of them must have a charge of +3. The name is "iron(III) oxide."
8. Assist students in writing the chemical symbols of elements and the chemical formulas for molecules and compounds.

Procedure

Part 1: Demonstration of the Law of Conservation of Matter

1. Tear off an egg sized piece of steel wool. Be careful not to ball it up too tightly.
2. Place the steel wool into the 250-mL beaker, and add white vinegar until the entire piece of steel wool is immersed. Soak the steel wool for 4–7 minutes.
3. Remove the steel wool from the vinegar, and wring out any excess liquid.
4. Place the steel wool into the 250-mL flask, and cover the opening of the flask with a balloon.
5. Mass the entire steel wool-balloon-flask system, and record the results.
6. Allow the system to sit for 30–45 minutes.
7. Then, observe the results. Again, mass the steel wool-balloon-flask system, and record.
8. Provide students with an unbalanced version of the following chemical equation, and have them practice balancing it prior to proceeding to the main portion of the activity:

$$4\,Fe(s) + 3\,O_2(g) \rightarrow 2\,Fe_2O_3(s)$$

Part 2: Lab for Balancing Equations

CAUTION! Sodium hydroxide is caustic and must be handled carefully. Students should wear gloves. If they do get it on their skin, it will feel slippery, just like liquid soap, and should be washed off immediately. Copper solutions can cause eye infections, so students should wash their hands after handling such solutions.

Have the students do the experiment, as follows:

1. Measure 60 mL of NaOH solution in a graduated cylinder, and then pour into a small (3-oz.), clean plastic cup.
2. Rinse the graduated cylinder completely before making the next measurement.
3. Measure 60 mL of $CuSO_4$ solution in the graduated cylinder, and then pour it into a clean 5-oz. cup.
4. Carefully place the two solutions together on the balance, making sure the solutions do not mix. Mass the solutions and the cups together, and record the combined mass.
5. Pour the NaOH into the 5-oz. cup with the $CuSO_4$ solution. Allow the solutions to mix, and record your observations.
6. Mass both cups and the mixture again. Record the new mass. By how much did the mass change?
7. Repeat the process in steps 1–4 above for the combinations listed in the data section below. Do not allow the solutions to mix before measuring the initial mass.

Reaction	**Mass (g) before**	**Mass (g) after**	**Observations**
NaOH and $CuSO_4$			
NH_4OH and $CuSO_4$			
NH_4OH and $Zn(NO_3)_2$			

8. Have students complete the following equations and balance them:
   - ___ NaOH + ___ $CuSO_4 \rightarrow$ ______________________________
     [$2\ NaOH + CuSO_4 \rightarrow Na_2SO_4 + Cu(OH)_2$]
   - ___ NH_4OH + ___ $CuSO_4 \rightarrow$ ______________________________
     [$2\ NH_4OH + CuSO_4 \rightarrow (NH_4)_2SO_4 + Cu(OH)_2$]
   - ___ NH_4OH + ___ $Zn(NO_3)_2 \rightarrow$ ______________________________
     [$NH_4OH + Zn(NO_3)_2 \rightarrow NH_4NO_3 + Zn(OH)_2$]

Part 3: Chemistry Geometrics Lab

In this lab, students will create models of molecules that have different molecular shapes. They will write molecular formulas and create structural formulas of several compounds. Have the students do the experiment, as follows:

1. Stick one pipe cleaner into each of the four small Styrofoam balls.
2. Stick the other end of two of the pipe cleaners into the large Styrofoam ball so that the two small balls are as far apart from each other as possible.
3. Note the shape of this model of a molecule, and use the protractor to measure the angle between the two small balls. Record this measurement in the data table.
4. Detach the small balls from the model, and then attach three small balls to the large ball so that the three balls are as far apart from each other as possible.
5. Note the shape of this molecule model, and measure the angle between two of the small balls. Since the balls should be as far apart from each other as possible, the angles between all pairs of small balls should be the same. Record this measurement in the data table.
6. Repeat steps 4 and 5, using all four small balls.
7. Record the following data in another table: Independent Variable = Number of Balls; Dependent Variable = Angle between the Small Balls
8. Sketch the three molecule models you created, and name each. *(Linear, Trigonal Planar, Tetrahedral)*

9. Under each sketch, describe the relationship between the number of atoms in a molecule and the size of the angles between the bonds in the molecule. *(The greater the number of atoms, the smaller the angle between atoms.)*
10. Explain why atoms arrange themselves in their molecule so that they are as far apart from each other as possible. *(Students should suggest that the electron clouds around atoms repel each other, thus forcing the atoms apart to the greatest possible degree. The fact that the valence electrons determine reactivity for forming compounds should be stressed.)*

Part 4: Introduction to Empirical Formulas

1. Before the end of this lesson, introduce the concept of empirical formulas by giving the definition of the term, showing some examples of empirical formulas, and comparing the difference between empirical formulas and molecular formulas. Tell students that in the weeks to come, they will be working with “moles” calculating empirical formulas. This should pique their interest and keep them wondering.

Observations and Conclusions

Use the following questions to prompt conclusions:

1. Why is it necessary to balance the equation of a chemical reaction? *(Balanced chemical equations obey the Law of Conservation of Matter and thus are true representations of what actually occurs in nature.)*
2. Why do atoms arrange themselves in their molecule so that they are as far apart from each other as possible? *(Electron clouds around atoms repel each other.)*
3. What are four different “shorthand” systems chemists use to write chemical formulas? *(Molecular, structural, empirical, and Lewis dot diagrams. Students may forget Lewis dot because it was introduced in a previous lesson, but use this as reinforcement.)*

Follow-up/extension

- Have students invent their own board game on the concepts of nomenclature, balancing equations, and/or writing chemical formulas.

Resources

- *Science Education Tutorials and Practice Exercises: General Chemistry :: Balancing Equations :: Tutorial.* http://www.wfu.edu/~ylwong/balanceeq/balanceq.html.
- *Science Help Online Chemistry: Writing Chemical Formulas.* http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson53.htm. This Web site provide instructions for and exercises in writing ionic and molecular formulas.

Organizing Topic — Chemical Reactions and Equations

Standards of Learning

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include
- a) designated laboratory techniques;
- b) safe use of chemicals and equipment; and
- c) proper response to emergency situations.

CH.3 The student will investigate and understand how conservation of energy and matter is expressed in chemical formulas and balanced equations. Key concepts include
- d) bonding types (ionic and covalent);
- e) reaction types (synthesis, decomposition, single and double replacement, oxidation–reduction, neutralization, exothermic, and endothermic); and
- f) reaction rates and kinetics (activation energy, catalysis, and degree of randomness).

CH.4 The student will investigate and understand that quantities in a chemical reaction are based on molar relationships. Key concepts include
- b) stoichiometric relationships; and
- f) chemical equilibrium

Essential Understandings, Knowledge, and Skills	Correlation to Textbooks and Other Instructional Materials
The student will use hands-on investigations, problem solving activities, scientific communication, and scientific reasoning to	
• recognize that reactivity is the tendency of an element to enter into a chemical reaction;	____________________
• understand that elements and compounds react in different ways;	____________________
• infer that the conservation of matter is represented by balanced equations;	____________________
• recognize and write equations for the major types of chemical reactions — synthesis, decomposition, single replacement, double replacement, and redox reactions;	____________________
• evaluate a chemical reaction and write equations, determine formulas, and balance chemical equations, using coefficients;	____________________
• understand the concept of limiting reactants in a reaction;	____________________
• identify the following principles relative to chemical reactions:	
° Spontaneous reactions may be fast or slow.	____________________
° Randomness (entropy), heat content (enthalpy), and temperature affect spontaneity.	____________________
° Chemical reactions based on the net heat energy are exothermic (heat producing) and endothermic (heat absorbing).	____________________
° Reaction rates/kinetics are affected by activation energy, catalysis, and the degree of randomness (entropy).	____________________

  - Catalysts decrease the amount of activation energy needed.
  - Reactions can occur in two directions simultaneously.
  - Le Chatelier's principle indicates the qualitative prediction of direction of change with temperature, pressure, and concentration.
  - A reaction is said to reach equilibrium when the forward reaction rate equals the reverse reaction rate.
- interpret reaction-rate diagrams;
- identify the limiting reactant in a reaction;
- calculate percent yield of a reaction;
- understand and demonstrate
  - MSDS warnings;
  - safety rules for science;
  - laboratory safety cautions;
  - safe techniques and procedures.

Predicting Products and Writing Equations

Organizing Topic Chemical Reactions and Equations

Overview Students conduct a series of chemical reactions, predict the products, and write balanced chemical equations that describe what occurs in the reactions.

Related Standards of Learning CH.1a, b, c; CH.3e

Objectives

The students will
- recognize and write equations for the major types of chemical reactions — synthesis, decomposition, single replacement, double replacement, and redox reactions;
- evaluate a chemical reaction and write equations, determine formulas, and balance chemical equations, using coefficients;
- chemical reactions based on the net heat energy are exothermic (heat producing) and endothermic (heat absorbing);
- understand and demonstrate
    - MSDS warnings
    - safety rules for science
    - laboratory safety cautions
    - safe techniques and procedures.

Materials needed

- Wooden splints
- pH paper or strips
- Plastic pipettes
- Test tubes in a variety of sizes
- Beakers in a variety of sizes
- Test tube rack
- Tongs
- Candle
- Scoop
- Burner
- Watch glass
- Ethanol
- Magnesium metal (ribbon form)
- Manganese dioxide (catalyst)
- 3% hydrogen peroxide solution
- Iron filings
- Zinc metal (Mossy zinc is fine.)
- 3.0 *M* HCl and 1.0 *M* HCl (Exact concentration is not crucial.)
- Solid $CuCO_3$
- Solid $(NH_4)CO_3$
- 0.1 *M* KI (Exact concentration is not crucial.)
- 0.1 *M* $Pb(NO_3)_2$ (Exact concentration is not crucial.)
- Solid $NaHCO_3$
- Copper metal pieces

- 0.1 *M* $AgNO_3$ solution (Exact concentration is not crucial.)
- 0.1 *M* NaOH and 0.1 *M* HCl (These concentrations need to be equal to be accurate.)
- 0.1 *M* Na_2CO_3 solution (Exact concentration is not crucial.)
- 0.1 *M* $CuSO_4$ (Exact concentration is not crucial.)
- Small pieces of Ca metal
- Phenolphthalein indicator

Instructional activity

Content/Teacher Notes

Before undertaking these lab experiments, be sure your school's chemical hygiene plan allows for the use of the chemicals required.

These lab experiments can be conducted by students or done by the teacher as demonstrations. If you do not feel confident that your students have sufficient lab skills to safely do some of the reactions involving flammable materials, do those as demonstrations. These experiments can be conducted over several days as you introduce the types of reactions, or you can do them on one day as a review of the different types of reactions. Partners can work in larger groups to do certain assigned experiments and then present their results to the rest of the class either as a demonstration or simply as a discussion.

Depending on the level of the students' knowledge at this point, they may write a word equation first, then translate that to a formula equation, and finally balance the equation; or they may write the balanced equation directly.

Students will need to know the positive results of the splint tests for the gases hydrogen, oxygen, and carbon dioxide. If they are not familiar with the splint tests for these gases, you can demonstrate them prior to the lab. (Hydrogen gas will "pop" when tested with a burning splint. Oxygen gas will relight a glowing splint. Carbon dioxide will extinguish a burning splint.)

Solution concentrations are not really crucial for most of the lab except of the neutralization reaction. For that, about equal drops of acid and base are needed to reach the neutralization point; hence, the molarity of the NaOH and HCl for this part are important.

Introduction

1. Tell the students that chemists observe what is happening in a chemical reaction and try to describe what is happening in a clear and concise way. A chemical equation uses formulas and symbols to describe the substances that are involved in a reaction, the physical state of the substances, the use of a catalyst, and the changes in energy that occur. In this experiment, you will perform a series of reactions and make careful observations of the changes that occur. Using simple tests and your knowledge of chemistry, you will identify the products of the reactions. With this information, you will then write and balance chemical equations to describe the reactions. Finally, you will classify the reactions as to type: combustion, oxidation, synthesis, decomposition, single replacement, double replacement, or acid-base.

Procedure

Either demonstrate or have students conduct the following reactions. Have the students write a balanced chemical equation for each reaction, showing the product(s), and classify the reaction as to type: combustion, oxidation, synthesis, decomposition, single replacement, double replacement, or acid-base. Some reactions may have more than one classification. Direct students to indicate in their balanced equations all phases of matter (*s, l, g*) for reactants and products. Also, have them indicate whether the reaction is exothermic or endothermic by adding energy as either a reactant or a product.

1. Place a very small amount (no more than 4 drops) of ethyl alcohol (ethanol, C_2H_5OH) on a watch glass on the lab table, and ignite it with a match. Use crucible tongs to hold the match. *CAUTION! Ethyl alcohol is very flammable. Remove all flammable materials from the lab area.*

   ($2\ C_2H_5OH + 6\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$*; combustion; exothermic. The alcohol burns with a nonluminous blue flame. If you turn off the lights, the flame will be more visible.)*

2. Ignite a very small piece (less than 0.5 in.) of Mg metal in the burner flame. *CAUTION! Magnesium metal burns with a very bright light. Do not stare directly at the flame. Remove all flammable materials from the lab area.*

   ($2\ Mg + O_2 \rightarrow 2\ MgO$*; oxidation; synthesis; exothermic.)*

3. Place a very small amount (fill only the very tip of your scoop) of solid catalyst, MnO_2, in a test tube that is half filled with 3 percent hydrogen peroxide. Test with a glowing splint the gas that is given off.

   ($2\ H_2O_2 \rightarrow 2\ H_2O + CO_2$*; decomposition; endothermic. The MnO2 acts a catalyst for the reaction.)*

4. Sprinkle a very small amount of Fe filings directly into a candle flame. Do not use your burner flame for this. *CAUTION! Metal powders are very flammable and reactive. Use only a small amount of Fe, and remove all flammable materials from the lab area.*

   ($4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$*; oxidation; synthesis; exothermic. The iron fillings spark in the flame: Fe is used in fireworks to make gold sparks.)*

5. Place a test tube in the test tube rack, and fill it half way with 3.0 *M* HCl. Place a small piece of zinc metal in the tube, and collect the gas that is released by using a test tube holder to hold another test tube upside directly above and touching the reaction tube. Test the gas in the collection test tube by inserting a burning splint. *CAUTION! 3.0 M HCl is corrosive. The gas that is released is very reactive and explosive. DO NOT ignite the gas that is produced.*

   ($Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$*; single replacement; exothermic. The zinc and HCl make hydrogen gas that gives a sharp "pop" when ignited.)*

6. Heat a small amount (about ¼ of a test tube) of solid $CuCO_3$ in a test tube and test the gas that is produces with a burning splint.

   ($CuCO_3 \rightarrow CuO + CO_2$*; decomposition; endothermic. The* CO_2 *gas puts out the burning splint, and the blue-green* $CuCO_3$ *turns black as the CuO forms.)*

7. Heat a very small amount (just enough to fill the bowl of the test tube) of $(NH_4)_2CO_3$ in a test tube. Place a damp piece of pH paper in the vapors released during the reaction. *CAUTION! The vapors being released are irritating to the throat and eyes. Conduct this procedure in a fume hood or in a room with good ventilation.*

   [$(NH_4)_2CO_3 \rightarrow 2\ NH_3 + H_2O + CO_2$*; decomposition; endothermic. The ammonia gas that is produced reacts with the damp pH paper to form a base.]*

8. Add a few drops of 0.1 *M* KI solution to a few drops of 0.1 *M* $Pb(NO_3)_2$ solution in a small test tube. *CAUTION! Do not dispose of these chemicals in the sink. Place them in the labeled waste container.*

   [$2\ KI + Pb(NO_3)_2 \rightarrow 2\ KNO_3 + PbI_2$*; double replacement. A yellow precipitate of* PbI_2 *forms.]*

9. Add several drops of 1.0 *M* HCl solution to a small amount of solid $NaHCO_3$ in a test tube. Test with a burning splint the gas that is released. *CAUTION! HCl is corrosive.*

   ($NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$ *acid-base. The carbonated base produces carbon dioxide gas, which puts out the burning splint.)*

10. Place a small piece of Cu metal in a clean dry test tube, and add several drops of 0.1 *M* $AgNO_3$ solution. Add just enough $AgNO_3$ solution to cover the metal. *CAUTION! $AgNO_3$ causes skin discoloration.*

    [$Cu + AgNO_3 \rightarrow Ag + Cu(NO_3)_2$; single replacement. (The $AgNO_3$ solution reacts with the Cu metal to produce a silver coating on the copper surface.]

11. Put 20 drops of 0.1 *M* NaOH solution in a test tube, and add 5 drops of universal indicator. Then, add 20 drops, one at a time, of 0.1 *M* HCl solution to the test tube, tapping the test tube gently as you add it to allow it to mix. *CAUTION! HCl and NaOH are corrosive.*

    ($NaOH + HCl \rightarrow NaCl + H_2O$; acid-base; double replacement. The indicator color changes from purple to green as the base is neutralized.)

12. Add about 1 mL of a solution of 1.0 *M* Na_2CO_3 to about 1 mL of a solution of 1.0 *M* $CuSO_4$ in a test tube.

    ($Na_2CO_3 + CuSO_4 \rightarrow CuCO_3 + Na_2SO_4$; double replacement. A cloudy, blue precipitate of $CuCO_3$ forms.)

13. Add a very small piece of Ca metal to a 50-mL beaker that is about half filled with water. When the reaction is finished, add a few drops of phenolphthalein indicator. *CAUTION! Calcium metal is very reactive and produces a flammable gas. Do not pick the calcium metal up with your hands, but use forceps or tongs. DO NOT have open flames during this reaction.*

    [$Ca + 2\ H_2O \rightarrow Ca(OH)_2 + H_2$; single replacement; exothermic. The $Ca(OH)_2$ forms a cloudy, white precipitate. This turns pink upon the addition of the indicator phenolphthalein, indicating a base. Hydrogen gas is released.]

14. Heat a small piece of copper metal directly in a burner flame.

    ($2\ Cu + O_2 \rightarrow 2\ CuO$; oxidation; synthesis; endothermic. A black coating of the oxide will form on the surface of the copper metal.)

Sample assessment

- Have students type up a formal lab report or simply complete a lab report sheet.
- If the lab is done as a demonstration, ask the students to write the balanced equations and classify the reactions as each reaction is completed. Have the students compare and discuss their answers.
- Have students work in groups to do assigned parts of the lab and then present their results to the rest of the class as a demonstration.

Follow-up/extension

- Manganese dioxide acts as a catalyst for the decomposition of hydrogen peroxide. Have students research the mechanism by which catalysts act — i.e., find out how catalysts make chemical reactions go faster.

What Affects the Rate of a Chemical Reaction?

Organizing Topic Chemical Reactions and Equations

Overview Rates of speed of chemical reactions depend on the ability of molecules to make consistent and effective collisions with each other. Students will study various conditions and the effects they have on the collision frequency of substances.

Related Standards of Learning CH.1a, b, c; CH.3f

Objectives

The students will
- identify the following relative to chemical reactions:
  - Reaction rates/kinetics are affected by activation energy and catalysis.
  - Catalysts decrease the amount of activation energy needed.
- Understand and demonstrate
  - MSDS warnings
  - safety rules for science
  - laboratory safety cautions
  - safe techniques and procedures.

Materials needed

- Test tubes
- Ice water bath
- Hot water bath
- Distilled water
- 250-mL plastic bottle
- Steel wool
- Test tube rack
- Five 250-mL beakers
- 0.1 *M*, 1.0 *M*, 3.0 *M*, and 6.0 *M* HCl solutions
- Small pieces of zinc
- Powdered zinc
- 3% H_2O_2 solution
- 0.1 *M* $CaCl_2$ solution
- 0.1 *M* NaCl solution
- 0.1 *M* $FeCl_3$ solution
- 0.1 *M* KNO_3 solution
- 0.1 *M* $Fe(NO_3)_3$ solution
- Alka-Seltzer tablets

Instructional activity

Content/Teacher Notes

When and how fast will a chemical reaction occur? Chemical kinetics is the study of rates of chemical reactions and reaction mechanisms. Many factors influence reaction rate. Two of the most important are the nature and properties of the reactants themselves, including particle size (surface area). Other factors, such as concentration, temperature, pressure, or the presence of a catalyst, will all also affect the rate of a reaction. Collision theory can explain how these factors affect reaction rate.

Collision-theory concepts:

- For a chemical change to occur, old bonds must be broken (an endothermic process) and new bonds must be formed (an exothermic process).
- The reactants must collide with each other to form products.
- The reactants must collide with each other at the correct angle and the correct molecular orientation.
- The collisions between reactants must be effective — i.e., they must have enough energy (called "activation energy").
- Activation energy is the minimum energy needed for a reaction to occur.
- If the rate of collisions increases, the reaction rate increases.
- Anything that increases the chances of collisions between reactants also increases the number of collisions that are effective.
- If the percentage of effective collisions increases, the reaction rate increases.

Effect of concentration on reaction rate:

- An increase in the concentration of reactants results in an increase in the reaction rate. At higher concentrations, the molecules of the reactants are closer to each other; therefore, collisions occur more frequently, a higher percentage of collisions are effective, and the reaction rate increases.

Effect of surface area on reaction rate:

- In a solid, only surface particles can interact with the other reactants. If the solid is divided into smaller pieces, then there is greater surface area; therefore, more particles are able to react, and the reaction rate increases.

Effect of temperature on reaction rate:

- An increase in the temperature of reactants usually results in an increase in the reaction rate. At higher temperatures, the molecules of the reactants move faster; therefore, collisions occur more frequently, a higher percentage of collisions are effective, and the reaction rate increases. A general rule of thumb is that on average, the reaction rate doubles for every 10ºC rise in temperature.

Effect of a catalyst on reaction rate:

- A catalyst is a substance that changes the reaction rate without being consumed by the reaction.
- A catalyst acts by lowering the activation energy required for a reaction to take place, thus allowing the reaction to occur more rapidly.

Introduction

1. Tell the students that the rate of a chemical reaction depends on the ability of molecules or ions to make consistent and effective collisions with each other. In this experiment, you will study various conditions and the effect they have on the collision of molecules. You will observe the effects of temperature, concentration, particle size (surface area), and catalysts on the rates of chemical reactions and then explain these effects in terms of the collision theory.

Procedure

Experiment 1: Effect of Temperature on Reaction Rate

1. Pour about 5 mL of 6.0 *M* HCl into each of three clean test tubes. Place one of the tubes in an ice water bath maintained at 0°C. Place another in a hot water bath maintained at 60°C. Maintain the

third test tube at room temperature. Allow about 10 minutes for the tubes to reach equilibrium temperature.

2. Cut three small pieces of zinc to the same size (0.5 x 2.0 cm). Clean the pieces with steel wool, if necessary.
3. Note the time, and drop one piece of zinc into each of the three test tubes.
4. Watch the reactions, and note the time at which each reaction ceases. Which reaction was fastest? Which was slowest?

Experiment 2: Effect of a Catalyst on Reaction Rate

1. Measure 90 mL of distilled water into a clean 250-mL plastic bottle, and add 10 mL of 3% H_2O_2. Label the solution as 0.3% H_2O_2. This will be your test solution.
2. Rinse six clean test tubes and a 10-mL graduated cylinder with the 0.3% H_2O_2. Discard the rinsing.
3. Measure 5 mL of the H_2O_2 solution into each of the six test tubes, and place them in a test tube rack.
4. Add 5 drops of each of the following solutions to separate test tubes of H_2O_2:
    - 0.1 *M* $CaCl_2$
    - 6.0 *M* HCl
    - 0.1 *M* NaCl
    - 0.1 *M* $FeCl_3$
    - 0.1 *M* KNO_3
    - 0.1 *M* $Fe(NO_3)_3$
5. Observe each solution, and report the rate of gas evolution from each, using the terms *fast, slow, very slow,* or *none* to describe the rate. Describe the catalytic effect as *high, low,* or *none*.

Experiment 3: Effect of Concentration on Reaction Rate (at Constant Temperature)

1. Pour 5 mL of each of the following HCl solutions into separate clean test tubes:
    - 0.1 *M*
    - 1.0 *M*
    - 3.0 *M*
    - 6.0 *M*.
2. Cut four small pieces of zinc (1 x 1 cm). Clean with steel wool if necessary. Drop a piece of zinc into each of the acid solutions, and record the starting time and the ending time of each reaction. Compare the reaction times. Which was fastest? Which was slowest?

Experiment 4: Effect of Particle Size or Surface Area on Reaction Rate

1. Cut a piece of zinc (0.5 x 2.0 cm). Clean with steel wool if necessary.
2. Determine the mass of the zinc to the nearest 0.01 g and record.
3. Place the piece of zinc in a clean, dry test tube.
4. Measure an equal mass of powdered zinc into a second clean, dry test tube.
5. Place both test tubes in a rack, and add 5 mL of 1.0 *M* HCl to each.
6. Observe the reactions for several minutes, and record your observations. Compare the reaction times. Which was fastest? Which was slowest?

Observations and Conclusions

Use the following questions to prompt conclusions:

1. Which ionic compounds used in Experiment 2 were effective catalysts? Using your data, identify the ion responsible for the catalytic activity.
2. Many reaction rates approximately double for every 10°C increase in temperature. Are the results obtained in Experiment 1 consistent with this general rule?
3. Describe the effect of temperature on the reaction rate. Explain this effect in terms of the collision theory of reactions.
4. Describe the effect of concentration on the reaction rate. Explain this effect in terms of the collision theory of reactions.
5. Describe the effect of particle size or surface area on the reaction rate. Explain this effect in terms of the collision theory of reactions.
6. Describe the effect of a catalyst on the reaction rate. Explain this effect in terms of the collision theory of reactions.

Sample assessment

- Have students type up a formal lab report or simply complete a lab report sheet and answer the questions.
- Have students work as groups to do assigned parts of the lab and then present their results to the rest of the class as a demonstration.

Follow-up/extension

- Have groups of students explore the following reaction to study further the effect of temperature on reaction rate:
  1. Fill each of five 250-mL beakers with 100 mL of water adjusted to the following temperatures: 0°C, 10°C, 20°C, 30°C, and 40°C. Adjust the temperatures by using ice water, tap water, and hot water.
  2. Add one Alka-Seltzer tablet to each beaker, and time the reaction until every last bit of the tablet stops fizzing. Record the results in a data table.
  3. Using the data, construct a graph of the reaction time versus temperature.
  4. How did the reaction rate vary with the temperature? Explain. A general rule of thumb is that on average, the reaction rate doubles for every 10ºC rise in temperature. Does this rule hold true for this reaction?

Which Way Will It Go? Equilibrium and Le Chatelier's Principle

Organizing Topic Chemical Reactions and Equations

Overview Students observe chemical and physical reactions that undergo equilibrium changes.

Related Standards of Learning CH.4f

Objectives

The students will
- identify the following relative to chemical reactions:
  - Reactions can occur in two directions simultaneously.
  - Le Chatelier's principle indicates the qualitative prediction of direction of change with temperature, pressure, and concentration.
  - A reaction is said to reach equilibrium when the forward reaction rate equals the reverse reaction rate.

Materials needed

- Distilled water
- Test tubes
- Stirrer
- Pipette
- Ice water bath
- Hot water bath
- Ethanol
- Solid NaCl
- Concentrated HCl
- Bromothymol blue indicator
- 0.1 *M* HCl
- 0.1 *M* NaOH
- Solid $CoCl_2 \cdot 6\,H_2O$
- 0.1 *M* silver nitrate solution

Instructional activity

Content/Teacher Notes

In many chemical reactions, the reactants are not completely converted into products. When all apparent chemical change has ceased, often significant amounts of the reactant and product species remain in the reaction mixture. Just as the reactants can interact to form products at a particular speed, called the "forward rate," the products can interact to re-form the reactants at a speed called the "reverse rate."

In a typical chemical reaction, initially the reactant concentrations are high, and the product concentrations are low. Thus the initial forward rate is much faster than the initial reverse rate. However, the forward reaction slows down as the reactants are depleted, and at the same time, the reverse reaction speeds up due to the increasing amounts of products. The two rates become exactly equal at a particular combination of reactant and product concentrations. The reactants are then replenished by the reverse reaction just as quickly as they are consumed in the forward reaction. As there is no net change in the amounts of reactants and products present in the reaction mixture, the reaction is said to be in a state of dynamic equilibrium.

Henri Le Chatelier's principle states that if a stress is applied to a system in equilibrium, the system will react in such a way as to relieve that stress and restore equilibrium under a new set of conditions. To relieve stress, a system can do only one of two things: (1) form more products, using up reactants, or (2) reverse the reaction and form more reactants, using up products. Le Chatelier proposed that the effect of changes in pressure, temperature, and reactant and product concentrations upon a chemical equilibrium will always be such as to oppose the applied change.

Introduction

1. Put the students into groups of 3 or 4. Tell the groups that as you do the following demonstrations, they are to watch the demonstrations closely, ask necessary questions, and write down their observations. The students will then discuss within their groups the results of the demonstrations and come to conclusions about what is happening to the equilibrium systems. Each group will then present their results to the class.

Procedure

Demonstration 1: Equilibrium in a Saturated Sodium Chloride Solution

$$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$$

1. Pour some solid NaCl into a test tube, and fill the tube ¾ full of distilled water.
2. Cork and shake the tube to form a saturated solution.
3. If all the NaCl dissolves, add some more until excess solid is obtained.
4. Decant the solution into a second test tube.
5. To this saturated solution of NaCl, add some Cl^{-} ions in the form of *concentrated* HCl.
6. Have the students make observations and explain the results. *(Results: The addition of the Cl^{-} ion from the HCl shifts the equilibrium to the left, resulting in solid NaCl forming in the solution and settling to the bottom of the test tube.)*

Demonstration 2: Equilibrium with an Acid-Base Indicator

Acid-base indicators like bromothymol blue have large organic molecules that can gain and lose hydrogen ions to form substances that have different colors. The reaction of the indictor bromothymol blue can be illustrated as follows:

$$\underset{\textit{yellow}}{HIn(aq)} \rightarrow H^{+}(aq) + \underset{\textit{blue}}{In^{-}(aq)}$$

In this reaction, HIn is the neutral indicator molecule, and In^{-} is the indicator ion after the molecule has lost a hydrogen ion. This demonstration also shows that equilibrium reactions can easily be caused to go in either direction.

1. Fill a test tube about half full of distilled water.
2. Add several drops of bromothymol blue indicator solution. Then, add drops of 0.1 *M* HCl while stirring. This will cause the H^{+} ion concentration to *increase*.
3. Have students note the color of the indicator.
4. Next, stir the solution constantly while adding 0.1 *M* NaOH drop by drop until no further color change occurs. Adding OH^{-} ions causes the H^{+} ion concentration to *decrease*.

5. Now, try to add just the right amount of acid back to this solution to cause it to be neutral or green (half of the indicator molecules are blue and the other half are yellow) in color after it is stirred.
6. Have the students make observations and explain the results. *(Results: The addition of the* H^+ *ion from the HCl shifts the equilibrium to the left, resulting in the HIn form of the indicator molecule. The addition of the* OH^- *ion from the NaOH shifts the equilibrium to the right as the* H^+ *ion is removed during its reaction with the added* OH^- *ion to form water. The shift to the right results in the* In^- *form of the indicator molecule.)*

Demonstration 3: Equilibrium with Cobalt Complex Ions

$$\underset{pink}{[Co(H_2O)_6]^{2+}}(aq) + 4\ Cl^-(aq) \rightarrow \underset{blue}{[CoCl_4]^{2-}} + 6\ H_2O\ \Delta H = +\ 50\ \text{kJ/mol}$$

Changing the Concentration of Other Ions

1. Dissolve a small amount of solid $CoCl_2 \bullet 6\ H_2O$ in a beaker containing ethanol. The solution should be purple; if it is pink, add a little concentrated HCl, drop by drop while stirring, until it is purple.
2. Put some of the solution into each of two test tubes.
3. To one of the test tubes, add drops of concentrated HCl, one drop at a time while stirring. Have students note the result.
4. To the second tube, add drops of 0.1 *M* silver nitrate solution, one drop at a time while stirring.
5. Have the students make observations and explain the results. *(Results: The addition of the concentrated HCl shifts the equilibrium to the right, resulting in a deep blue color. This is due to the addition of the* Cl^- *ion. The addition of the* $AgNO_3$ *shifts the equilibrium to the left, resulting in a pink color. The added* Ag^+ *ion reacts with the* Cl^- *ion to remove it by forming the insoluble compound AgCl. The equilibrium shifts left, producing a cloudy pink color. The cloudy appearance is due to the precipitate of AgCl.)*

The Effect of Changing the Temperature

1. Take some of the cobalt chloride and HCl solution from step 3 above. The solution should be violet — between the original pink-red and the bright blue.
2. If it is not violet, adjust the color by carefully adding distilled water, one drop at a time while stirring, until a violet color is obtained. Alternatively, add concentrated HCl, depending on the color of the solution and the direction you need to go. You now have a solution that contains both $[Co(H_2O)_6]^{2+}$ and $[CoCl_4]^{2-}$ ions.
3. Divide this violet solution into three medium-size test tubes. Place one tube in an ice bath. Place the second tube in a hot water bath maintained at between 80ºC and 90ºC, and maintain the third tube at room temperature.
4. After the tubes have reached the desired temperatures, have the students note and record the colors of each of the three solutions. *(Results: The solution in the hot water bath is deep blue. The reaction is endothermic as noted by the +ΔH value. Heat is a reactant in this reaction. Addition of heat shifts the reaction equilibrium to the right, forming the blue complex. The solution in the ice water bath is pink. Removal of heat shifts the equilibrium to the left, forming the pink complex. The solution at room temperature has an equilibrium position between the pink and the blue; it remains purple.*
5. Determine whether the color changes are reversible by switching the test tubes in the hot and cold water baths and allowing them to come to room temperature.

Sample assessment

Use the following questions to evaluate the students understanding of the concepts of Le Chatelier's principle.

- Novelty devices for prediction of rain contain cobalt(II) chloride and are based on the following equilibrium: $\underset{\textit{purple}}{\mathrm{CoCl_2}(s)} + 6\ \mathrm{H_2O}(g) \longleftrightarrow \underset{\textit{pink}}{\mathrm{CoCl_2} \bullet 6\ \mathrm{H_2O}(s)}$
  What color will such an indicator be if it is about to rain?
- If you wanted to use the reaction
  $\mathrm{CO}(g) + 2\ \mathrm{H_2}(g) \longleftrightarrow \mathrm{CH_3OH}(g) + \text{heat}$
  for producing methanol ($\mathrm{CH_3OH}$) commercially, would high temperature or low temperature favor a maximum yield?
- Arsenic can be extracted from its ores by first reacting the ore with oxygen (called "roasting") to form solid $\mathrm{As_4O_6}$, which is then reduced by using carbon:
  $\mathrm{As_4O_6}(s) + 6\ \mathrm{C}(s) \longleftrightarrow \mathrm{As_4}(g) + 6\ \mathrm{CO}(g)$
  Predict the direction of the equilibrium shift in response to each of the following changes in conditions:
  - Addition of carbon monoxide
  - Addition of carbon
  - Removal of $\mathrm{As_4}$
- For each of the following reactions, predict how the equilibrium concentration of the products will be affected if the temperature is increased:
  - $\mathrm{N_2}(g) + \mathrm{O_2}(g) \longleftrightarrow 2\ \mathrm{NO}(g) \qquad \Delta H = +181\ \mathrm{kJ}$
  - $2\ \mathrm{SO_2}(g) + \mathrm{O_2}(g) \longleftrightarrow 2\ \mathrm{SO_3}(g) \qquad \Delta H = -198\ \mathrm{kJ}$
- The following chemical system
  $2\ \mathrm{W}(aq) + \mathrm{X}(g) \longleftrightarrow 3\ \mathrm{Y}(g) + 2\ \mathrm{Z}(g) + \text{heat}$
  is at equilibrium.
  - How would the equilibrium change if the concentration of W were increased?
  - How would the equilibrium change if the temperature were increased?
  - In what direction would the equilibrium shift if a catalyst were added to the system?
  - Predict what would happen to the concentration of W if the temperature were decreased.
  - If the temperature of the system were decreased, what would happen to Z?
  - If the concentration of W were decreased, how would the equilibrium change?
  - In what direction would the equilibrium shift if the concentration of Y were increased?
  - What would happen to the concentration of Y if the concentration of W were increased?
  - What are four specific things that could be applied to this system that would favor the reverse reaction?

Organizing Topic — Stoichiometry

Standards of Learning

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include

a) designated laboratory techniques;
e) accurate recording, organization, and analysis of data through repeated trials; and
g) mathematical and procedural error analysis.

CH.4 The student will investigate and understand that quantities in a chemical reaction are based on molar relationships. Key concepts include

a) Avogadro's principle and molar volume; and
b) stoichiometric relationships.

Essential Understandings, Knowledge, and Skills	Correlation to Textbooks and Other Instructional Materials
The student will use hands-on investigations, problem solving activities, scientific communication, and scientific reasoning to	
• summarize the following basic concepts of stoichiometry:	
° Atoms and molecules are too small to count by usual means.	____________________
° A mole is a way of counting any type of particle, such as atoms, molecules, and formula units.	____________________
° Stoichiometry involves quantitative relationships.	____________________
° Stoichiometric relationships are based on mole quantities in a balanced equation.	____________________
• know and use the following:	
° Avogadro's number;	____________________
° Molar volume;	____________________
° Molar mass;	____________________
° Total grams of reactant(s) equal total grams of product(s).	____________________
° Equal volumes of gases at the same temperature and pressure have the same number of particles.	____________________
• make calculations involving the following relationships:	
° Mole-mole;	____________________
° Mass-mass;	____________________
° Mole-mass;	____________________
° Mass-volume;	____________________
° Mole-volume;	____________________
° Volume-volume;	____________________
• illustrate that	
° scientific notation is used to write very small and very large numbers;	____________________
° ratios and proportions are used in calculations;	____________________
° the last digit of any valid measurement must be estimated and is therefore uncertain;	____________________

  - ° dimensional analysis is a way of translating a measurement from one unit to another unit; ______________________________
- make the following measurements, using the specified equipment:
  - Volume: graduated cylinder and pipette; ______________________________
  - Mass: electronic balance; ______________________________
  - Pressure: barometer or pressure probe; ______________________________
  - Temperature: thermometer or temperature probe; ______________________________
- identify the limiting reactant in a chemical reaction; ______________________________
- calculate percent yield of a reaction. ______________________________

Moles Lab Activities

Organizing Topic Stoichiometry

Overview In a variety of activities, students practice mole conversions at varying levels of difficulty. These conversions have applications to various content areas in chemistry and are therefore useful for practice throughout their Chemistry course, not just when studying stoichiometry.

Related Standards of Learning CH.1a, b, c, g; CH.4

Objectives

The students will

- perform various measurements of mass and volume;
- perform various calculations/conversions, using molar mass, Avogadro's number, and molar volume;
- practice rules of scientific notation and significant figures;
- analyze experimental data.

Materials needed

(The materials needed for each mole lab activity are listed on the attached activity sheets with the activity.)

Instructional activity

Content/Teacher Notes

The mole is the basic counting unit used in chemistry and is used to keep track of the amount of matter being measured or transferred. Performing calculations using molar relationships is essential to understanding chemistry. Because of its importance, the concept of the mole, its value, and basic conversions should be introduced very early in the course and revisited often, ideally in every unit, in order for students to understand the concept completely and become proficient with its use.

This lesson offers a wide variety of lab activities in which students practice mole conversions at varying levels of difficulty. These conversions have applications to various content areas so that students can practice these concepts throughout a chemistry course, not just when studying stoichiometry. The common theme of these activities is to have students do simple lab measurements of various substances encountered in everyday life and then perform simple mole calculations and respond to conceptual questions related to the concept. If students approach the idea of the mole both conceptually and mathematically, then they will be able to handle the wide variety of problems that rely on the mole. The idea is to provide a solid conceptual and analytical understanding of the mole concept.

Although the first activity is designed to give students a solid understanding of a counting unit and relative masses as a foundation for understanding the mole, students should be introduced to the quantity of the mole and its role as a counting unit before starting these activities. A review of basic scientific notation and rules for significant figures is also recommended. The activities start with conversions involving elements followed by compounds and then by simple reactions.

Doing one or more moles lab activities in each unit you teach will give students plenty of practice and time to become proficient with all of the basic conversions and calculations. While these activities could be done together in one discrete unit, it will prove more effective to students' long-term retention to have them do these activities throughout the course.

Teacher notes for each activity are found in the list below. Among other things, these notes indicate the recommended unit in which the activity should take place and an estimated time for students to complete the work.

Procedure

(The procedures for the moles lab activities are found on the attached lab activity sheets.)

Observations and Conclusions

(Specific information will be found in the teacher notes for each activity. Keep in mind that if mole calculations are applied in context throughout the course, students will become proficient in the calculations and will acquire an excellent conceptual understanding of the stoichiometric principles of chemistry.)

Sample assessment

- All of these concepts can be tested with traditional assessments, of course. It is best, however, if occasional authentic assessments are done as well. Many of the extensions included with the activities can be used as assessment tools.
- Ask students to design their own investigation related to a moles lab activity. You will find examples of these in some of the extensions as well.
- Another option is to have students design their own Moles Lab Activity Sheet for other students to complete. This can be done experimentally, if time permits, or with mock data provided with the sheets. Also require students to submit a grading rubric for the activity they design. Student-generated assessments can be useful for quizzes and tests.

Follow-up/extension

(These are listed with the various moles lab activities.)

Teacher Notes for Moles Lab Activities

Moles Lab Activity 1: PCU (Popcorn Counting Units)

Time: Students will need 20–30 minutes to do initial calculations and collect data. Part 3 could be completed outside of class.

Application: This activity should be used when introducing isotopes and relative atomic masses, which requires the simultaneous introduction of the concept of the mole as a counting unit. The activity also provides reinforcement of scientific notation.

Helpful Hints/Suggestions: Samples can be placed in small plastic containers or baggies. Students may have difficulty understanding that for the data table in Part 2, the PCU is the *number* determined in Part 1 and should be the same for each type of bean. Students may also have difficulty in completing the extension table unless you either explain in advance what they are to do or have them practice such calculations previous to the activity. Some of the confusion comes from not understanding that the tables in Parts 1 and 3 are actually worksheets, not data tables. Not everyone will get the exact same value for a PCU, but this is correct. For honors classes, this fact may be an interesting discussion topic involving lab errors and differences among the kernels and balances.

Answers to Selected Questions: The answers to question #9 need to be recorded in a class data table so that question #10 can be answered. The answer to each part of question #13 should be the same and equal to answer #2. Question #19 should relate to the small size of the atom and the need for a large number of them in order for them to be seen and measured. The answer to question #19 is C-12, the reference isotope for atomic masses.

Moles Lab Activity 2: Elements

Time: Students will need about 5–10 minutes at each lab station to do initial calculations and measurements. They will need a total of 15–20 minutes to complete all questions, some of which could be completed at home.

Application: This activity should be used when introducing elements, chemical symbols, and moles-to-atoms calculations.

Helpful Hints/Suggestions: This activity is actually a set of six activities dealing with six different elements — one for each lab station. It is easy to make additional activities, as the format of each is consistent. Also, if you do not have one or more of these elements on hand, it is easy to substitute others, using the same format. Setting up each sample at a lab station and having students rotate between them works best. Make sure each sample is clearly labeled and stays that way.

These activities go quickly once students get through the first two. It helps if the first group at each station has to get your signature, confirming that their calculations are correct, before they move on. They then become that element's "experts," a role they will like. This will help insure they really do know what they are doing and will allow for other students to have more than one person to verify their work and/or answer questions. It also allows you to determine quickly who is having difficulty with the concept and to whom to offer additional help.

The extension for the aluminum activity requires students to weigh out one mole of aluminum foil and make a creative sculpture. Students don't always understand this from the directions, so it may need some further explanation. They can make a mole, but they should realize that it can be formed into anything.

Answers to Selected Questions: None of the questions should give students any great difficulty. It may help for you to do the extension questions ahead of time so you have an answer key.

Moles Lab Activity 3: Compounds

Time: Students will need about 5–10 minutes at each lab station to do initial calculations and measurements. They will need a total of 15–20 minutes to complete all questions, some of which could be completed at home.

Application: This activity should be used when introducing formulas of compounds, molecular masses, and percent composition.

Helpful Hints/Suggestions: This activity is actually a set of four activities. Directions on the handouts have been intentionally kept short so that the students will focus on the concept rather than the directions. Samples can be placed in small plastic containers or baggies.

Water: When doing water, remind them about keeping the balance pans dry. You need to tell them ahead of time that the water is going to be weighed directly in the graduated cylinder, so if they are not using an electronic balance that has a tare function, they will need to find the mass of the cylinder as well. You will need to alter the directions on the handout accordingly.

Chalk: If it is difficult to go outside for use of the sidewalk chalk, then use the blackboard or large pieces of paper on the floor. Students really seem to enjoy and remember this activity. Encourage them to be stylistic and artistic in drawing their names. If you do go outside, clearly state limits as to where and what they write. Depending on the nature of your group, you many need to remind them that this is very public so they need to keep it “clean.” Be certain that students copy what they draw into their data books.

“Candium”: This lab needs to be done with great cleanliness. If you use store-bought cups for samples and measuring, then a reward to students when they finish the activity is that they may eat their “compound” or “element” as long as it has not been contaminated (touched anything in the lab). If you do permit this, you should make it very clear that you are using nonstandard lab materials and that this is *not* a standard practice in a chemistry lab!

Answers to Selected Questions: The answers to most questions require basic conversions. The extension for NaCl deals with colligative properties that will probably not have been covered at this point. You can either require them (honors students) to look it up on their own, or talk about deicing and antifreeze as the questions come up. Introducing in extensions topics that will be covered later in the course helps students put things in context and drives home the point about interconnectedness and cumulative knowledge. The extension for sidewalk chalk requires students to differentiate between ionic and covalent bonds. This topic should be covered if you are working on compounds and their formulas.

Moles Lab Activity 4: Solutions

Time: Students will need about 5–10 minutes at each lab station to do initial calculations and measurements. They will need a total of 15–20 minutes to complete all questions, some of which could be completed at home.

Application: This activity should be used when introducing chemical formulas, dissociation, mass-to-molecules/ions calculations, and solutions.

Helpful Hints/Suggestions: This activity is actually a set of two activities. Some students will need help with adding molar masses of these hydrated salts. Alum can be purchased at the store. One solution is supposed to be saturated, and the other is not — a concept that is also being introduced at this time. You should also reinforce the concept of the smallness of ions/particles that pass through the filter paper. It is helpful to use the copper solution to reinforce the concept of dissociation and that the particles do not disappear but are simply too small to be seen.

Answers to Selected Questions: The questions in the extensions concerning what is left on the filter paper and what is in the filtrate aim at having the students identify the species present and the form that they are in — i.e., atoms, ions, or molecules.

Moles Lab Activity 5: Synthesis of a Compound

Time: Students will need about 5–10 minutes at each lab station to do initial calculations and measurements. They will need a total of 15–20 minutes to complete all questions, some of which could be completed at home.

Application: This activity should be used when introducing chemical formulas, percent composition, and empirical formulas.

Helpful Hints/Suggestions: The copper has an interesting series of color changes before it turns black, but students may miss it if they are not paying attention. This reaction is not efficient, and errors will be high. You should get enough conversion that the percent oxygen will be less than the first oxide so that that should be their choice for their product. The high error in this lab is actually a good teaching opportunity, not only for sources of error (incomplete reactions), but also for ways reactions really occur. The oxidation is not instantaneous and is greatly dependent on surface area and amount of heat. This leads into a discussion of collision theory. Because students' data will vary, this lab is good for having a discussion of class-data accuracy and precision. You get a little better data with copper powder than granules, but not enough to offset the safety issue, if that is a concern. Either way, a lot of the copper will not react, which can be seen if the product is stirred after it cools. If you have metal rods for stirring, students can stir while the copper is heating. However, this can lead to problems of sample loss and possible burns if the students are not careful, and it decreases error only a little bit. Another way to decrease error is to use Bunsen burners/clay triangles and heat for a longer period of time. However, the reduction of error may not be worth the added time when the focus should be on the reaction and calculations. Remind students of the safety rules for working with very hot materials.

Answers to Selected Questions: Weaker students might have a little trouble with predicting the possible oxides and with the Roman numeral naming system: they often remain confused about the Roman numeral indicating the number of metal atoms. This is a good reaction to try to clear up this confusion. Some students may have trouble figuring out the equation for Pre-Lab #4; they need to start with the oxide produced in #3 (Cu_2O) and react it with O_2 again to form CuO. For question #3 after the analysis, some students have a hard time with the concept that Cu_2O is less fully oxidized than CuO; therefore, focus on the theoretical percent O to help them understand this.

Moles Lab Activity 6: Single Replacement and Percent Yield

Time: Students will need about 30 minutes at each lab station to do initial calculations and measurements. They will need a total of 15–20 minutes to complete all questions, some of which could be completed at home.

Application: This activity should be used when introducing chemical equations/types, mass-to-mass conversions, and percent yield.

Helpful Hints/Suggestions: Copper(II) chloride is expensive, and shipping the solid can be hazardous. You can either buy the solution premixed, or you can substitute an acidified copper(II) sulfate solution of the same molarity (the solution needs to be made in 0.5–1.0 *M* HCl). All copper solutions should be handled carefully by students, and you must be sure that the necessary safety procedures are followed. The students should get a measurable temperature change during this lab. If you do the extension to recover the copper, there are problems that occur. The copper on the filter paper needs to be washed thoroughly with water, especially if you are using acidified copper(II) sulfate. As the copper dries, it tends to form hydroxides from the exposure to moisture and the air. This could mean you will end up with more than a 100% yield. For an honors class, this is a good teaching point about experimental errors and unwanted secondary reactions. In a regular class, however, it may distract them from focusing on the mass-to-mass conversions that they learning. It may be possible to reduce this error by washing with acetone before drying, but this might not be worth the effort. In any case, it is important to demonstrate to students that science is not perfect!

Answers to Selected Questions: Weaker students will have some difficulty justifying the reaction from experimental evidence. It is surprising that they may not realize that the solution still having some blue color is proof that the copper solution is not the limiting reactant and that the color of the product matches the color of copper. Students also tend to confuse theoretical and experimental yields.

Moles Lab Activity 7: Alka-Seltzer

Time: Students will need about 15–20 minutes to do initial calculations and measurements. They will need a total of 15–20 minutes to complete all questions, some of which could be completed at home.

Application: This activity should be used when introducing Gas Laws.

Helpful Hints/Suggestions: In this lab, you are ignoring the role of the citric acid in Alka-Seltzer to simplify the reaction so that students can focus on the gas conversions. It is a good idea to acknowledge this up front to avoid questions later that will distract weaker students. Students may forget to put the weighing dish back on the balance after adding the tablet to the water. If you do not have a barometer, then access the weather information online, and use the reported pressure; it will be close enough. (Students will have errors anyway.) The weather service gives pressure in inches of Hg, so students may need help with the conversion to atm. The extension questions in this lab could actually be turned into a student investigation. Students could vary amounts of Mg or acid and see the effect on yield and/or limiting reactant.

Answers to Selected Questions: The conversion described in the last bullet under "Analysis" can be done in several ways. You can leave it up to the students, or for weaker students, you may want to specify which way to do it. Another possibility is to have one partner do it one way and the other partner do it the other way to prove that both ways yield the same answer.

Moles Lab Activity 8: Conservation of Mass

Time: Students will need about 15–20 minutes to do initial calculations and measurements. They will need a total of 15–20 minutes to complete all questions, some of which could be completed at home.

Application: This activity should be used when introducing balancing equations and conservation of mass.

Helpful Hints/Suggestions: This lab is very straightforward. Some students will have trouble with loss of mass because they do not seal the bag correctly or they spill some vinegar before they get the bag sealed. You may want to model sealing up the bag while keeping the beaker upright. If students do not measure out the reactants correctly, they may generate enough gas to pop the bag open; deal with this as a lab-safety and carelessness issue.

Answers to Selected Questions: If this is the first time students have been asked to calculate percent error, they will need help. They will confuse theoretical and experimental values. They may also have some calculator-entry issues with order of operations.

Moles Lab Activity 9: Percent Water in a Hydrate

Time: Students will need about 30 minutes to do initial calculations and measurements. They will need a total of 15–20 minutes to complete all questions, some of which could be completed at home.

Application: This activity should be used when introducing chemical formulas, percent composition, and hydrated salts.

Helpful Hints/Suggestions: Weaker students will need help with adding molar masses of these hydrated salts. A desiccator will reduce errors, but it is not absolutely required. Be sure to include this in your discussion of errors, if you do not use one. Other hydrated salts can be used, but copper(II) sulfate has a good color change. [Cobalt(II) chloride also has a good color change and is a good substitute, if you are allowed to order it.] It is good for students to get visual confirmation of the dehydration. You will need to demonstrate how to transfer the hot evaporating dish. This lab usually calls for heating in a crucible, but we have found that the increase in errors when using this method is offset by the students' being able to see the process. It also gives you more to talk about when analyzing errors and suggestions for improvement when students are asked to design their own procedures. Be sure to go over the necessary lab safety procedures. The extension in this activity can be used as an assessment as well.

Answers to Selected Questions: Analysis bullet 4: Students often have difficulty with precision versus accuracy in class data. You might want to review this in the pre- or post-lab discussion. Analysis bullet 6: Students often have misconceptions about whether this is a chemical or physical change. Reviewing this would be helpful. Extension bullet 4: If students do not realize that heating the salt to dehydrate is endothermic, they may not realize that the reverse process must be exothermic.

Moles Lab Activity 1: PCU (Popcorn Counting Units)

Materials

A container of each of the following: popcorn kernels, kidney beans, pinto beans, peas, lima beans, and navy beans; a large, unopened bag of popcorn kernels; balance

Objective

To devise a new counting unit, use it in calculations, and compare it to the use of a **mole**.

Procedure

Part 1

1. Weigh out 5.0 grams of popcorn kernels on the balance, and count the number of kernels there are in 5.0 grams. This *number* will be called **"1 PCU"** (1 popcorn counting unit).
2. Complete the following equation in your data record: **1 PCU = _____ kernels = 5.0 g of kernels**
3. Show how you would calculate the number of kernels in 3 PCUs:
4. Show how you would calculate the number of kernels in 20.0 grams of popcorn:
5. Show how you would calculate the mass, in grams, of 100 popcorn kernels:
6. Complete the following worksheet:

Number of popcorn kernels	Number of PCUs	Mass of popcorn kernels (g)
	2	
	10	
	500	
		10.0
		650.0
		5.0×10^5
1		
498		
7,000		
5.0×10^8		

7. Use the balance to find the mass, in grams, of the unopened bag of popcorn kernels. Mass: _____
8. Use the mass of the popcorn bag and your PCU to determine how many kernels are in the bag. Show your work here, and record your answer on the class data table.
9. Based on the class data table, what is the average number of kernels in the popcorn bag? ______
10. How close to the class average number is the number you found? _________________________
11. Explain what could account for the different numbers of kernels calculated by each student.
12. We can use a PCU just like a dozen is used. When we count out a dozen eggs, bagels, and marbles, we know that the mass of each dozen will not be the same. Would you expect 1 PCU of lima beans to weigh the same as 1 PCU of popcorn kernels? _________ Explain you answer.

13. How many popcorn kernels have you determined to be equal to 1 PCU? ________ If you were counting out 1 PCU of marbles, how many marbles would you count out? _________ If you were counting out 1 PCU of each type of bean, how many of each would you count out? ________
14. Count out 1 PCU of pinto beans. This will be the number of pinto beans equal to the number of kernels in one PCU. Use a balance to determine the mass of 1 PCU of pinto beans and record in the table below.

Part 2

15. Complete the data table at right, keeping in mind that *the number of particles in a PCU is always the same*.
16. Is the *number* of kidney beans in 1 PCU more than, less than, or equal to the *number* of navy beans in 1 PCU? _____________________________
17. How does the *mass* of 1 PCU of kidney beans compare to the mass of 1 PCU of navy beans? ____________________________________ How can you account for the differences in mass that you observed? ____

Type of particle	**Number of particles in 1 PCU**	**Mass of 1 PCU**
Pinto bean		
Kidney bean		
Lima bean		
Pea		
Navy bean		

18. Would 5.0 grams of kidney beans be more than, less than, or equal to the mass of 1 PCU of kidney beans? ________________________ Would 10.5 grams of peas be more than, less than, or equal to the mass of 1 PCU of peas? ____
19. Why is a **mole** a better unit than a PCU for counting atoms? ______________________________
    How many particles are in a mole? ____

Part 3: Extension

20. Fill in the following worksheet:

Element	**Symbol**	**Mass of 1 mole**	**Number of particles**	**Number of moles**	**Mass of sample (g)**
Carbon			6.02×10^{23}	1	12
Carbon			1.2×10^{24}	2	
Carbon				3	36
Carbon				0.5	
Carbon					3
Magnesium					24.3
Silicon					14
Neon			6.02×10^{22}		
Iron				3	

21. Explain how the mass of one mole of magnesium atoms compares to the mass of one mole of iron atoms. ____
22. Just as our masses in this lab can be based on popcorn kernels, the atomic masses of each element on the periodic table can be (and are) based on one element. What element is it? ______

Moles Lab Activity 2: Elements — Aluminum

Materials

Empty aluminum can; balance; aluminum foil

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Find and record the mass of one mole of aluminum.
2. Find and record the mass of the aluminum can.
3. Answer the following questions:
    - Does the aluminum sample contain more than, less than, or exactly one mole of aluminum?
    - How many moles of aluminum atoms are in one aluminum can?
    - How many individual atoms of aluminum are in one aluminum can?
4. Check your answers with the student aluminum experts, and ask them to initial your original data to certify that they are correct.

Extension

1. How many cans would you need to have one mole of aluminum?
2. Make a “sculpture” out of aluminum foil, using exactly one mole of the foil.

Moles Lab Activity 2: Elements — Carbon

Materials

Sample of carbon; balance

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Find and record the molar mass of carbon.
2. Find and record the mass of the carbon sample.
3. Answer the following questions:
    - Does the carbon sample contain more than, less than, or exactly one mole of carbon?
    - How many moles of carbon atoms are in the carbon sample?
    - How many individual atoms of carbon are in the carbon sample?
4. Check your answers with the student carbon experts, and ask them to initial your original data to certify that they are correct.

Extension

1. A person weighing about 78 kg is about 18% carbon by mass. What is the mass (in grams) of carbon in this person? How many moles of carbon are in this person?
2. Graphite is one allotropic form of the element carbon. Do some research to define *allotrope,* and describe the structure and properties of graphite and of other carbon allotropes.

Moles Lab Activity 2: Elements — Copper

Materials

Sample of copper; balance; pre-1982 penny

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Find and record the molar mass of copper.
2. Find and record the mass of the copper sample.
3. Answer the following questions:
    - Does the copper sample contain more than, less than, or exactly one mole of copper?
    - How many moles of copper atoms are in the copper sample?
    - How many individual atoms of copper are in the copper sample?
4. Check your answers with the student copper experts, and ask them to initial your original data to certify that they are correct.

Extension

1. Determine the mass of a pre-1982 penny.
2. Answer the following questions:
    - How many moles of copper are in the penny?
    - How many atoms of copper are in the penny?
    - How many pennies are needed to make a mole of copper?
    - Why would these calculations be invalid for post-1982 pennies?

Moles Lab Activity 2: Elements — Iron

Materials

10 iron nails; balance; iron filings; small cup; magnetic retriever

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Find and record the molar mass of iron.
2. Find and record the mass of 10 iron nails.
3. Answer the following questions:
    - Do 10 nails contain more than, less than, or exactly one mole of iron?
    - How many moles of iron atoms are in the 10 nails?
    - How many individual atoms of iron are in the 10 nails?
4. Check your answers with the student iron experts, and ask them to initial your original data to certify that they are correct.

Extension

1. How many iron nails are needed to get one mole of iron atoms?
2. Pick up some iron filings with the magnet. Scrape these off into a weighing cup, and determine the mass of the collected iron filings. Calculate how many moles of iron filings were collected.
3. Estimate the volume of one mole of iron filings, and describe how you determined this.

Moles Lab Activity 2: Elements — Silicon

Materials

Sample of silicon; balance

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Find and record the molar mass of silicon.
2. Find and record the mass of the silicon sample.
3. Answer the following questions:
    - Does the silicon sample contain more than, less than, or exactly one mole of silicon?
    - How many moles of silicon atoms are in the silicon sample?
    - How many individual atoms of silicon are in the silicon sample?
4. Check your answers with the student silicon experts, and ask them to initial your original data to certify that they are correct.

Extension

Silicon is to geologists what carbon is to biologists. It makes up 28% of the earth’s crust and is found in many minerals. Sand, quartz, and glass are all made up of silicon dioxide (SiO_2).

1. Answer the following questions:
    - Is silicon dioxide an element or a compound?
    - How many moles of silicon are in one mole of silicon dioxide?
    - How many moles of oxygen are in one mole of silicon dioxide?
    - How many atoms of silicon are in one mole of silicon dioxide?
    - How many atoms of oxygen are in one mole of silicon dioxide?
2. Silicon is a *metalloid*. Look up the properties of a metalloid, and explain why metalloids are so useful in making semiconductors for computers and other electronics.

Moles Lab Activity 2: Elements — Sodium

Materials

A small bag of snack crackers or other small snack food; balance

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Find and record the molar mass of sodium.
2. Answer the following questions:
    - How many mg of sodium are in one serving of snack crackers?
    - How many g of sodium are in one serving of snack crackers?
    - How many moles of sodium are in one serving of snack crackers?
    - How many individual atoms of sodium are in one serving of snack crackers?
3. Check your answers with the student sodium experts, and ask them to initial your original data to certify that they are correct.

Extension

Healthy American adults should restrict their sodium intake to no more than 2,400 milligrams per day. This is about 1¼ teaspoons of table salt (sodium chloride [NaCl]).

1. Answer the following questions:
    - What is the maximum number of moles of sodium recommended in your diet? How many sodium atoms would this be?
    - If 1 teaspoon salt = 2,000 mg sodium, how many tablespoons of salt would you need to get a mole of sodium? (3 tsp = 1 tbsp)

Moles Lab Activity 3: Compounds — Water

Materials

Water; graduated cylinder; balance

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Mass 50.0 mL of water, using the graduated cylinder. Be sure to subtract out the mass of the cylinder. Record the mass of the water.
2. Find and record the molar mass of water (H_2O).
3. Answer the following questions:
    - Is 50.0 mL of water less than, equal to, or more than one mole of water?
    - How many moles of water are in 50.0 mL of water?
    - How many molecules of water are in 50.0 mL of water?
    - How many individual atoms of hydrogen are in 50.0 mL of water?
    - What is the density of your water sample?
4. Check your answers with the student water experts, and ask them to initial your original data to certify that they are correct.

Extension

1. Answer the following questions:
    - Calculate the percent by mass of each element (H and O) in water (H_2O).
    - Calculate the mass of hydrogen in your 50.0 mL of water.
    - Calculate the mass of oxygen in your 50.0 mL of water.
    - Use the density of water to calculate the volume of one mole of water, in mL.
    - One swallow of water is about 20 mL of water. How many moles of water are in one swallow? How many molecules of water are in one swallow?

Moles Lab Activity 3: Compounds — Sodium Chloride

Materials

Sample of sodium chloride; test tube; weighing dish; balance; sample of calcium chloride

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Use a weighing dish to determine the mass of the sodium chloride sample in the test tube.
2. Find and record the molar mass of sodium chloride (NaCl).
3. Answer the following questions:
    - Is the amount in the sample less than, equal to, or more than one mole of sodium chloride?
    - How many moles of sodium chloride are in the sample?
    - How many molecules of sodium chloride are in the sample?
    - How many individual ions (both anions and cations) are in the sample?
4. Check your answers with the student sodium chloride experts, and ask them to initial your original data to certify that they are correct.

Extension

1. Calculate the percent by mass of each element (Na and Cl) in the salt sodium chloride (NaCl).
2. Calculate the percent by mass of each element (Ca and Cl_2) in the salt calcium chloride ($CaCl_2$).
3. Pure water freezes at 0°C. When substances are dissolved in water, the solute particles affect the intermolecular attractions of the water, decreasing the freezing point and elevating the boiling point. This is called a “colligative property.” The magnitude of the change is determined by the number (moles) of solute particles dissolved in solution. Which will lower the freezing point of water more — one mole NaCl or one mole of $CaCl_2$? Why?

Moles Lab Activity 3: Compounds — Chalk

Materials

Sidewalk chalk; balance

Procedure

1. Find the mass of a piece of sidewalk chalk.
2. Write your name on the sidewalk or parking lot the piece of chalk, but do not use it up completely.
2. Find the mass of your piece of chalk following your writing.
3. Answer the following questions, taking the necessary measurements and recording them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.
    - What is the molar mass of chalk (calcium carbonate [$CaCO_3$])?
    - What is the mass of your piece of chalk before writing?
    - What is the mass of your remaining chalk after writing?
    - How many grams of chalk did you leave on the ground?
    - Is the amount left on the ground more than, equal to, or less than one mole of chalk?
    - How many moles of chalk did you leave on the ground?
    - How many molecules of chalk did you leave on the ground?
    - How many individual atoms of oxygen did you leave on the ground?
4. Check your answers with the student calcium carbonate experts, and ask them to initial your original data to certify that they are correct.

Extension

1. Calculate the percent oxygen by mass of chalk (calcium carbonate [$CaCO_3$]).
2. Answer the following questions:
    - How many calcium (Ca) atoms did you leave on the ground?
    - How many atoms (total of all types) did you leave on the ground?
    - Is chalk an ionic or covalent compound? Why?

Moles Lab Activity 3: Compounds — The Fictitious Compound “Candium”

Materials

A sample of “candium” (a mixture of 12 M&M’s and 8 Skittles); 2 small paper cups; balance

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Assume that each kind of candy (“Mm” and “Sk”) in the sample of “candium” represents a different type of atom and that the sample is the compound. Write the formula in the format Mm_xSk_y for the compound “candium,” using the smallest whole-number ratio of Mm to Sk.
2. Find and record the molar mass of candium. (Assume your sample is one mole.)
3. Find and record the mass of the Mm in your sample.
4. Find and record the mass of the Sk in your sample.
5. Answer the following questions:
   - What is the average mass of an Mm (the molar mass of Mm)?
   - What is the average mass of an Sk (the molar mass of Sk)?
   - What is the percent by mass of Mm in candium? Use this equation:

     $$\textbf{\% by mass Mm} = \frac{\textbf{molar mass Mm} \times \textbf{number of Mm atoms in candium formula} \times \textbf{100}}{\textbf{molar mass of candium}}$$

   - What is the percent by mass of Sk in candium? Use a similar equation.
6. Check your answers with your teacher, and ask him/her to initial your original data to certify that they are correct.

Extension

“Candium Isotope” Activity

1. Assume your candium sample is a sample of an *element*, not a compound. In this element, Mm and Sk are the two naturally occurring isotopes of the element. What is the percent abundance (by number) of each isotope in your sample?
2. What is the average atomic mass of the element candium? Use the average mass of each isotope as its mass number and its percent abundance to calculate, using the following formula:

   $$\textbf{[Mass \#}_{\textbf{(Mm)}}\ \textbf{(\%}_{\textbf{Mm}}\textbf{) + Mass \#}_{\textbf{(Sk)}}\ \textbf{(\%}_{\textbf{Sk}}\textbf{)]} \div \textbf{100 = average atomic mass}$$

3. Compare the average atomic mass to the molar mass (total mass of all candies in the sample). What is your percent error?

Moles Lab Activity 4: Solutions — Aqueous Copper(II) Sulfate Pentahydrate

Materials

Beaker or plastic cup; copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5\ H_2O$) crystals; graduated cylinder; water; filter paper; funnel; ring stand and ring; watch glass or Petri dish

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Find and record the molar mass of *hydrated* copper(II) sulfate (copper(II) sulfate pentahydrate [$CuSO_4 \cdot 5\ H_2O$]).
2. Find and record the mass of 0.0125 moles of copper(II) sulfate pentahydrate.
3. Weigh out 0.0125 moles of copper(II) sulfate pentahydrate. Using a graduated cylinder to measure, place 25.0 mL of water into a plastic cup or beaker.
4. Record the color changes you observe as you add the 0.0125 moles of copper(II) sulfate pentahydrate to the water and stir to dissolve completely.
5. Answer the following questions:
   - Is the mixture heterogeneous or homogeneous?
   - What is the solute in this solution?
   - What is the solvent?
   - How many copper atoms are in one mole of copper(II) sulfate?
   - How many copper atoms are in your 25.0 mL of copper(II) sulfate solution?
   - What is the percent water by mass in copper(II) sulfate?
   - What is the percent copper in the 0.0125 moles of copper(II) sulfate?

Extension

1. Predict the appearance of the filtrate (the substances that pass through the filter) if you were to filter the copper(II) sulfate solution.
2. Assemble a filtration apparatus, using a ring stand and ring, a plastic funnel, a beaker, and a piece of filter paper. Sketch and label the apparatus in your data book.
3. Prime the filter with tap water, and filter the copper(II) sulfate solution.
4. Describe the filtrate, and compare it to your prediction. Account for similarities and/or differences. What substances are found in the filtrate?
5. How many oxygen atoms are in one mole of copper(II) sulfate ($CuSO_4$)? Calculate the number of oxygen atoms in your 25.0 mL of saturated copper(II) sulfate solution.
6. Draw a model of an aqueous copper(II) ion by drawing Cu^{2+} surrounded by water molecules. Which end of the water molecules is attracted to the Cu^{2+} ion?
7. Based on the copper(II) ion (Cu^{2+}) and the formula for copper(II) sulfate, what is the charge of the sulfate ion (SO_4)?
8. Place 1 mL of your filtrate on a watch glass, and allow the water to evaporate (this may take a day or two). Draw the shape of the crystals that form.

Moles Lab Activity 4: Solutions — Alum

Materials

Beaker or plastic cup; alum (hydrated aluminum potassium sulfate [$AlK(SO_4)_2 \bullet 12\ H_2O$]) crystals; water; filter paper; funnel; ring stand and ring

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Find and record the molar mass of alum (hydrated aluminum potassium sulfate [$AlK(SO_4)_2 \bullet 12\ H_2O$]).
2. Find and record the mass of 0.050 moles of alum.
3. Weigh out 0.050 moles of alum. Using a graduated cylinder to measure, place 25.0 mL of water into a plastic cup or beaker. Add the 0.050 moles of alum to the 25.0 mL of water, and mix thoroughly. Record you observations.
3. Answer the following questions:
    - Is the mixture saturated or unsaturated?
    - What is the solute in this solution?
    - What is the solvent?
    - How many aluminum atoms are in one mole of aluminum potassium sulfate?
    - How many aluminum atoms are in your 25.0 mL of aluminum potassium sulfate solution?
    - What is the percent water by mass in aluminum potassium sulfate?
    - What is the mass of sulfur in the 0.050 moles of aluminum potassium sulfate?
4. Check your answers with your teacher, and ask her/him to initial your original data to certify that they are correct.

Extension

1. Predict the appearance of the filtrate (the substances that pass through the filter) if you were to filter the aluminum potassium sulfate solution.
2. Assemble a filtration apparatus, using a ring stand and ring, a plastic funnel, a beaker, and a piece of filter paper. Sketch and label the apparatus in your data book.
3. Prime the filter with tap water, and filter the $AlK(SO_4)_2 \bullet 12\ H_2O$ solution.
4. Describe the filtrate, and compare it to your prediction. Account for similarities and/or differences. What substances are found in the filtrate?
5. How many sulfur atoms are in one mole of alum? Calculate the number of sulfur atoms in your 25.0 mL of saturated alum solution.
6. How many oxygen atoms are in one mole of alum? Calculate the number of oxygen atoms in your 25.0 mL of saturated alum solution.
7. Place 1 mL of this solution on a watch glass, and allow the water to evaporate (this may take a day or two). Draw the alum crystals remaining after the water evaporates. Show clearly the shape of one crystal.

Moles Lab Activity 5:
Synthesis of an Oxide of Copper

Materials

Hot plate; evaporating dish; balance; granular or powdered copper

Pre-Lab

Copper has two common ions: copper(I) (Cu^+) and copper(II) (Cu^{2+}). When heated in air, the copper will react with oxygen gas (O_2) to form an oxide. The Roman numeral is used in the name to differentiate between the two possible oxides of copper. Remember that the Roman numeral indicates the *charge* on the metal, not the *quantity.*

1. Determine the formula for each oxide of copper, and indicate its name. Check with the teacher or the student copper experts for accuracy of these answers before proceeding.
2. Determine the percent oxygen in each oxide.
3. Based on your answers from the previous step, which oxide should form first in this experiment? Why? Write the balanced equation for its formation.
4. The second oxide cannot be formed directly, but is formed from the further oxidation of the first oxide. Write a balanced equation for the formation (synthesis) of the second oxide, using the first oxide and oxygen gas as the reactants.
5. Check that the answers to the pre-lab are correct before performing the experiment below.

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly. You must wear goggles when heating.

1. Find and record the molar mass of copper.
2. Find and record the mass of a clean, dry evaporating dish.
3. Determine the mass of 0.30 moles of copper.
4. Place 0.30 moles of copper into the evaporating dish, and record the total mass before heating.
5. Place the evaporating dish on the hot plate, and turn it on to the highest setting. Heat strongly for at least 15 minutes, and record at least four qualitative observations while heating.
6. Use crucible tongs to carefully remove the evaporating dish from the hot plate, and allow the dish to cool for 5 to 10 minutes
7. Find the mass of the dish and oxide of copper after heating, and record.
8. Determine the mass of the copper(x) oxide formed.
9. Determine the mass of oxygen reacted (the difference in mass before and after heating).
10. After the product has cooled, observe it carefully for any signs of an incomplete reaction. Record observations.

Analysis

1. Answer the following questions:
    - What is the experimental percent oxygen in your product? Report this in the class data table.
    - Based on your percent oxygen, which oxide most likely was formed?
    - Based on this result, what is your percent error?

- Look up the color of copper(I) oxide and copper(II) oxide. Does this information support your data? Why, or why not?
- Is the class data accurate and/or precise? Why, or why not?
- What experimental errors would account for any differences you observed?

Extension

1. Iron can also form two oxides. Look up the possible oxidation states (charges) for iron, write the formulas for both oxides, and indicate their names.
2. Determine the theoretical percent oxygen of each oxide.
3. In an experiment similar to the one described above, a student reacted 1.20 g of iron and determined that it gained 0.5 grams upon heating. Determine the percent oxygen in the product.
4. Determine the empirical formula for the oxide formed.
5. Identify which oxide was formed in the experiment. What evidence do you have to support this conclusion?
6. Look up the color of each possible oxide of iron. What would be the color of the product in the experiment? Why?
7. In forming the oxides of transition metals with multiple oxidation states by heating them in air, identify the two things that most likely determine which oxide is formed.

Moles Lab Activity 6: Single Replacement and Percent Yield

Materials

Aluminum foil; beaker or plastic cup with 100 mL of 0.5 *M* aqueous copper(II) chloride [$CuCl_2(aq)$] (or an acidified copper(II) sulfate solution); 100-mL graduated cylinder; balance; thermometer; filter paper; funnel; funnel stand; 250-mL or larger beaker

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Take a piece of aluminum foil with a mass of approximately 0.5 grams. Measure and record the exact mass (to 2 decimal places) of the piece of foil.
2 Use the exact mass of your foil and the molar mass of aluminum (to 2 decimal places) to calculate the number of moles in your piece of aluminum. Record.
3. Gently tear the piece of aluminum into small pieces, and put them into a beaker with 100 mL of 0.5 *M* aqueous copper(II) chloride [$CuCl_2(aq)$]. Observe carefully, and record at least 10 observations during the reaction, numbering each observation. One aspect of these observations should be temperature: determine whether there is a change in the temperature of the solution by recording the temperature of the solution as the reaction proceeds. The initial temperature and final temperature should count as 2 of your 10 observations.
4. Save your solution for the extension activity, and clean up.

Analysis

This is a single replacement reaction.

1. Predict and write the formulas for the two products of the reaction of aluminum metal $Al(s)$ with the copper(II) chloride.
2. Write and balance the equation for the reaction, and place the heat term on the appropriate side.
3. Explain why the reaction is a single replacement reaction.
4. What physical evidence indicates that your reaction actually occurred as predicted by the balanced equation?
5. Which reactant is the limiting reactant? What observations support this choice?
6. Draw a representative model for the *reactants,* including water molecules, aqueous ions, and aluminum atoms, and label each correctly.
7. Draw a representative model of the *products,* including water molecules, aqueous ions, and copper atoms, and label each correctly.
8. Calculate your *theoretical yield,* using the number of moles in your piece of aluminum, and the mole ratio from the balanced equation, calculate the number of grams of copper that would be produced from the reaction.
9. Draw and label an energy diagram for this reaction. Label both axes — the reactants and the products. Use arrows to indicate the heat of reaction (ΔH), and indicate the appropriate sign for ΔH.

Extension

1. Measure and record the mass of a piece of filter paper.

2. Filter the products of the reaction in the experiment above. Wash the filtrate with distilled water, and place the filter paper in the drying oven over night.
3. Remove the filter paper from the drying oven, let cool, and then find the mass of the filter paper and copper product. Record observations.
4. Determine your experimental yield of copper.
5. Compare your experimental yield to the theoretical yield to calculate your percent yield.
6. Indicate two major sources of error in this lab, and predict their effect on the percent yield.
7. Using the volume and molarity of the $CuCl_2$ in the third step of the Procedure above, calculate the number of *moles* of copper ions present in the initial solution.
8. Calculate the number of *grams* of copper ions that were present in the original solution.
9. Compare the grams of Cu ions present in the original solution with the number of grams produced from the reaction (the theoretical yield). Which is greater?
10. Did all of the copper ions in solution become copper atoms? Why, or why not?
11. Prove mathematically which reactant was limiting. Does this support your predictions above?
12. How many total *atoms* of copper were produced (based on your theoretical yield)?
13. Use the balanced equation to prove that this reaction follows the Law of Conservation of Mass.

Moles Lab Activity 7: Alka-Seltzer

Materials

250-mL beaker; 100-mL graduated cylinder; water; thermometer; barometer; Alka-Seltzer tablet; balance

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Measure and record the mass of an Alka-Seltzer tablet.
2. Place 100 mL of water in a 250-mL beaker, and record the combined mass of the water and beaker.
3. Combine and record the masses from steps 1 and 2.
4. Drop the tablet into the beaker, and record three observations during the reaction.
5. Once the reaction has stopped, record the mass again.
6. Clean up, and answer the following questions.

Analysis

1. The antacid in Alka-Seltzer is sodium bicarbonate ($NaHCO_3$). It decomposes to form CO_2 and NaOH. Write a balanced equation for the decomposition of the sodium bicarbonate. (The sodium hydroxide [NaOH] reacts with the citric acid in the Alka-Seltzer to form a neutralized solution.)
2. Answer the following questions, taking the necessary measurements and recording all measurements clearly, with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.
   - What was the mass of the CO_2 gas that was released?
   - How many moles of CO_2 were released?
   - How many liters of the CO_2 were released? Assume standard temperature and pressure (STP).
   - What amount of sodium bicarbonate is in each tablet? (Use the information on the Alka-Seltzer package to determine this.) Convert this mass to grams.
   - What is the number of moles of CO_2 that theoretically should have been produced?
   - How many liters would this be at STP?
3. Calculate the percent yield of the amount of gas actually produced by the tablet.
4. Explain two possible sources of error.
5. The room conditions in this experiment are not really STP (it would be awfully cold). Measure the room temperature and barometric pressure. Convert the temperature to degrees Kelvin and the pressure to atmospheres. What is the volume of CO_2 produced, experimentally, at room conditions? (You many use either the Ideal Gas Law or the Combined Gas Law.)

Extensions

Consider the results if you were to do a similar lab by reacting magnesium metal (Mg) with hydrochloric acid (HCl). Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Write a balanced chemical equation to represent this reaction, and indicate which product would be a gas.
2. Answer the following questions:
   - How many moles of hydrogen gas would be needed to produce the same volume as the CO_2 produced in your experiment under the same conditions of temperature and pressure? Why?
   - What mass of hydrogen gas (H_2) would this be?
   - What mass of magnesium would be needed to produce this amount of hydrogen gas?
   - Would you need an excess of acid to produce this amount of hydrogen? Why, or why not?
   - What would be he limiting reactant in this case? Why?

Moles Lab Activity 8: Conservation of Mass — Reaction of Vinegar and Baking Soda

Materials

Large plastic zip bag; baking soda (sodium bicarbonate [$NaHCO_3$]); two 50-mL beakers or small cups; vinegar (acetic acid [$HC_2H_3O_2$]); thermometer or temperature probe

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Make a data table in which to record class data.
2. Place a plastic zip bag on the balance, and record its mass.
3. Add about 4 grams of baking soda (sodium bicarbonate [$NaHCO_3$]) to the bag, and record the mass again. Determine how many moles of $NaHCO_3$ are in the bag.
4. Measure 30.0 mL of vinegar (acetic acid [$HC_2H_3O_2$]) into a graduated cylinder, and then pour the vinegar into a 50-mL beaker or cup.
5. Be certain the outside of the beaker is completely dry. Carefully set the plastic zip bag containing the baking soda on the lab bench. Set the beaker inside the bag, making sure none of the vinegar spills out. Seal the bag.
6. Zero the balance. Carefully set the bag on the balance, making sure none of the vinegar spills out of the beaker. Determine the total mass of the filled bag, and record on the class data table the mass of the bag + baking soda + vinegar *before* the reaction begins.
7. Remove the bag from the balance. Keeping the bag sealed, invert the beaker, and record at least three qualitative observations in your data book. Be sure to note if there is any temperature change.
8. When the reaction is complete, determine the total mass of the filled bag, and record the mass of the bag + baking soda + vinegar *after* the reaction is complete.
9. Determine the change in mass, and record.
10. Empty the bag into the sink (be careful with the beaker), and rinse out the bag with tap water.

Analysis

1. Use the following formula to calculate the percent error:

$$\textbf{percent error} = \frac{\textbf{mass after} - \textbf{mass before}}{\textbf{mass before}} \times 100$$

   (The theoretical value for this lab is the mass with which you began, and the experimental value is the mass that you measured after the reaction was complete.)
2. Answer the following questions:
   - Based on the Law of Conservation of Mass, what should have been the percent error?
   - Do your results support the Law of Conservation of Mass? Why, or why not?
   - Is the class data accurate and/or precise? Why, or why not?
   - What experimental errors would account for any differences you observed?

Extension

1. The products of this reaction are carbon dioxide, water, and sodium acetate ($NaC_2H_3O_2$). Write the balanced equation for the reaction, and identify two ways in which this reaction can be

classified. Calculate the theoretical yield (mass) of carbon dioxide, based on the moles of baking soda reacted.

2. Use the materials above to design and describe and/or carry out an experiment to support your calculation. How will the procedure need to change in order to determine the amount of carbon dioxide produced?
3. Calculate your percent error.
4. Carry out the following experiment to determine if the amount of baking soda ($NaHCO_3$) affects the mass of CO_2 produced.
   a. Identify the independent variable and the dependent variable for this experiment.
   b. Identify the constants.
   c. Write an appropriate hypothesis, and support your reasoning.
   d. Carry out this experiment with at least four different quantities of baking soda. Make sure some are lower than and some are higher than the quantity used in the original lab.
   e. Graph the results. Make sure your graph has
      - the IV and DV on the correct axis
      - scaling that is appropriate for the data collected
      - both axes labeled with the quantity and appropriate units
      - a best fit line or curve drawn.
   f. Identify the general relationship (trend) between your IV and DV, as shown by the graph. Does this match your hypothesis? Why, or why not?
   g. What might account for any discrepancies seen at larger quantities of baking soda?

Moles Lab Activity 9: Percent Water in a Hydrate

Materials

Copper sulfate pentahydrate ($CuSO_4 \cdot 5\ H_2O$) crystals; hot plate; evaporating dish; crucible tongs; eye dropper; desiccator (if available)

Pre-Lab

In a *hydrated* compound, water molecules are actually absorbed into the ionic crystal lattice. The formula for a *hydrate* tells the number of water molecules (water of hydration) attached to each formula unit of the ionic compound.

1. Identify the ratio of Cu^{2+} ions to SO_4^{2-} ions in copper(II) sulfate pentahydrate.
2. Identify the number of water molecules per formula unit of copper(II) sulfate.
3. Explain why heating would cause this substance to dehydrate (become *anhydrous*).
4. Calculate the molar mass of copper sulfate pentahydrate ($CuSO_4 \cdot 5\ H_2O$). Show all steps in your calculation.
5. Calculate the percent by mass of water in this hydrate (theoretical value).

Procedure

Take the necessary measurements, and record them with units. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. Measure and record the mass of a clean, dry evaporating dish.
2. Weigh out approximately 5.0 grams of $CuSO_4 \cdot 5\ H_2O$ into your evaporating dish. Record the exact mass, correct to 2 decimal places, of the evaporating dish + hydrated copper(II) pentahydrate sulfate crystals.
3. Place the evaporating dish on the hot plate. In complete sentences, write at least two qualitative observations of the copper(II) sulfate pentahydrate crystals.
4. Put on goggles, and turn on the hot plate to heat the copper(II) sulfate pentahydrate in the evaporating dish for 7 to10 minutes or until the color change is complete. Observe carefully, and record at least three qualitative observations in your data log.
5. Remove the evaporating dish from the hot plate, using the crucible tongs. Let the dish cool for 10 minutes. Do this in a desiccator, if one is available. Use the tongs to transfer the dish to the balance. Record the mass of evaporating dish and *anhydrous* copper(II) sulfate.
6. Repeat steps 4 and 5 until a *constant* mass is obtained.
7. Using a dropper, add 2 or 3 drops of water to the anhydrous copper(II) sulfate. Record your observations.

Analysis

Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. What is the mass of water lost in the reaction?
2. Calculate the experimental percent of water in the hydrated copper(II) sulfate pentahydrate, using the mass of water lost and the original mass of the hydrate. Record this in the class data table.
3. Using the above calculation and the theoretical percent of water in hydrated copper(II) sulfate pentahydrate, based on its formula ($CuSO_4 \cdot 5\ H_2O$), calculate the percent error in your experiment.
4. Based on the class data, are the results of this experiment *accurate* and/or *precise*? Explain your reasoning.

5. What would happen if you left the anhydrous copper(II) sulfate sitting out overnight? Why?
6. Is this a chemical change or a physical change? Why?

Extension

1. Calculate the formula for hydrated lithium nitrate ($LiNO_3 \bullet xH_2O$), based on mole ratios determined from the following laboratory data collected in an experiment similar to the one you just completed:
    - mass (g) of hydrated lithium nitrate: 17.00 g
    - mass (g) of anhydrous lithium nitrate: 9.53 g
2. Find the formula, and indicate the name of a second hydrate containing 76.9% $CaSO_3$ and 23.1% H_2O by mass.
3. What is the function of a desiccator? What experimental error will you incur if you do not use one in this lab? Why?
4. Would rehydration of the anhydrous salt be endothermic or exothermic? Why?

Finding the Formula and Percent Composition of an Ionic Compound

Organizing Topic Stoichiometry

Overview Students use mole ratios to determine the formula and percent composition of an ionic compound.

Related Standards of Learning CH.3c; CH.4b

Objectives

The students will
- calculate percent yield of a reaction;
- investigate and understand that quantities in a chemical reaction are based on molar relationships;
- identify and use
  - chemical formulas
  - coefficients, chemical symbols, and subscripts.

Materials needed

- Standard test tubes (18 x 150 mm)
- Wooden splints
- 3.0 *M* HCl
- Beral pipettes
- Magnesium ribbon

Instructional activity

Content/Teacher Notes

Percent Composition (Percent by Mass)

To determine the percent composition (percent by mass) of an ionic compound,
- calculate the molar mass of the compound by writing its formula and adding up the masses of each element in it
- calculate the percent of each element in the compound (the percent composition) by dividing the mass of each element by the molar mass of the entire compound and then multiplying by 100.

As an example, for ammonium sulfate, the steps are as follows:
- For ammonium sulfate, $(NH_4)_2SO_4$, the molar mass is 132.154 g/mol
- percent H = 8(1.008 g) ÷ 132.145 g = 0.061 × 100 = 6.10% H
  percent N = 2(14 g) ÷ 132.145 = 0.212 × 100 = 21.2% N
  percent S = 1(32 g) ÷ 132.145 = 0.242 × 100 = 24.2% S
  percent O = 4(16 g) ÷ 132.145 = 0.484 × 100 = 48.5% O [or percent O = 100 − (6.10 + 21.2 + 24.2) = 48.5% O]

Empirical Formula Determination

One of the most important things we can learn about a compound is its chemical formula. The formula for a compound that expresses the smallest whole-number ratio of the atoms present is called an “empirical formula.” For example, a compound with the formula $C_4H_8O_4$ has the same empirical formula as a compound with formula $C_6H_{12}O_6$: the empirical formula for both of these compounds is CH_2O.

The formula for a molecular compound that gives the actual number of atoms per molecule is called a "molecular formula." The molecular formula $C_6H_{12}O_6$ can be represented as a whole-number multiple of the empirical formula: $(CH_2O)_6$.

To calculate the empirical formula for a compound, we first determine the relative masses of the various elements that are present. One way to do this is to measure the masses of elements that react to form the compound. For example, suppose we weigh out 0.2636 g of pure nickel metal into a crucible and heat the metal in air so that the nickel can react with oxygen to form a nickel oxide compound. After the sample has cooled, we weigh it again and find its mass to be 0.3354 g. The gain in mass is due to the oxygen that reacts with the nickel to form the oxide. Therefore, the mass of oxygen present in the compound is the total mass of the product minus the mass of the nickel: 0.3354 g − 0.2636 g = 0.0718 g. Note that the mass of nickel present in the compound remains the same as the nickel metal originally weighed out. So we know that the nickel compound contains 0.2636 g nickel and 0.0718 g oxygen.

To determine the empirical formula for a compound,

- convert the grams of each element in the compound to moles of atoms of each element
- determine the whole-number mole ratio of each element to the other(s)
- write the formula.

For the above example involving nickel, the steps are as follows:

- 0.2636 g Ni ÷ 58.71 g per mol Ni atoms = 0.00449 mol Ni atoms
  0.0718 g O ÷ 16.00 g per mol O atoms = 0.00449 mol O atoms
- These mole quantities contain an equal number of atoms. It is clear from the moles of atoms of each substance that the mole ratio is 1:1.
- Therefore, the formula is NiO. This is the empirical formula, expressing the smallest whole-number ratio of atoms. All ionic compounds are expressed as empirical formulas — that is, we would never write, for example, Ni_2O_2.

For a compound containing 1.3813 g of Pb, 0.00672 g H, 0.4995 g As, and 0.4267 g O, the steps to find the empirical formula for the compound are as follows:

- 1.3813 g Pb ÷ 207.2 g per mol Pb = 0.006667 mol Pb
  0.00672 g H ÷ 1.008 g per mol H = 0.006667 mol H
  0.4995 g As ÷ 74.92 g per mol As = 0.006667 mol As
  0.4267 g O ÷ 16.00 g per mol O = 0.026669 mol O
- Divide these mole values by the smallest number of moles to get the whole-number mole ratio:
  0.006667 mol Pb ÷ 0.006667 = 1 mol Pb
  0.006667 mol H ÷ 0.006667 = 1 mol H
  0.006667 mol As ÷ 0.006667 = 1 mol As
  0.026669 mol O ÷ 0.006667 = 4 mol O
- Therefore, the empirical formula for the compound is $PbHAsO_4$.

When a 0.3546 g sample of vanadium metal is heated in air, it reacts with oxygen to form an oxide compound with a mass of 0.6330 g. To calculate the empirical formula for this sample of vanadium oxide, the following steps are used:

- 0.6330 g − 0.3546 g = 0.2784 g (mass of oxygen that reacted)
- 0.3546 g V ÷ 50.94 g per mol V = 0.006961 mol V
  0.2784 g O ÷ 16.00 g per mol O = 0.01740 mol O
- Divide these mole values by the smallest number of moles to get the whole-number mole ratio:

0.006961 mol V ÷ 0.006961 = 1.000 mol V
0.01740 mol O ÷ 0.006961 = 2.500 mol O

- Since one of these numbers (2.500) is not a whole number, multiply both numbers by the lowest possible number (2) that will result in two whole numbers and therefore form a whole-number ratio:
  1.000 × 2 = 2 V
  2.500 × 2 = 5 O
  Hence, the whole-number ratio is 2:5. (Note that this step is necessary only if the mole numbers are not already whole numbers.)
- Therefore, the empirical formula for this sample of vanadium oxide is V_2O_5.

Introduction

1. Introduce the activity by telling the students that they will do a lab in which they will form an ionic compound from the elements magnesium and chlorine. They will determine the empirical formula for this compound by comparing the moles of each element that react and determining the mole ratio of one element to the other. They will also determine the percent composition of the compound.

Procedure

Have the students do the experiment, as follows:

1. Make a data chart with spaces for the following data:
   - Mass of clean, dry test tube
   - Mass of magnesium ribbon
   - Mass of test tube and product
   - Mass of product
   - Observation of gas splint test
   - Identity of gas
2. Select a standard 18 x 150 mm test tube, and make sure it is clean and dry. Find the mass of the test tube, and record it in the data chart.
3. Find the mass of approximately 0.10 grams of magnesium ribbon. It is not necessary to have exactly 0.10 grams, but do not be too far off. Record in the data chart the exact mass of the magnesium ribbon used.
4. Fill a beral pipette with 3.0 *M* HCl. *(CAUTION! Strong acid)* Add the acid to the test tube, a few drops at a time. Run the drops down the inside wall of the test tube, holding the test tube at a gentle angle. *CAUTION! Do not add the acid all at once.*
5. While the reaction is proceeding, test the gas with a burning splint. What is the identity of the gas?
6. Continue to add the acid until the reaction is complete — i.e., when there is no solid magnesium visible. *CAUTION! Try to avoid adding excess acid.*
7. Set up a Bunsen burner. Using a *small* flame, start to evaporate off the liquid in the test tube. Be sure to hold the test tube at a 45-degree angle. Keep the test tube moving in the flame, and do not heat just the bottom of the test tube. *CAUTION! Do not point the test tube at anyone, including yourself.*
8. When evaporation appears complete, test by removing the test tube from the flame and inverting a *clean, dry test tube* over the mouth of your test tube. If you observe condensation forming in the top test tube, then continue heating the original test tube a few minutes longer and test again with another clean, dry test tube. Be patient; the heating process will probably take about 10 to 15 minutes.

9. When you are sure your product is dry, let your test tube cool completely. Find the mass of the test tube and product, and record this mass in the data chart.
10. When you are finished, clean up. You may wash your product down the sink.

Observations and Conclusions

Prompt conclusions by having students do the following:

1. Calculate the number of moles of magnesium in the compound.
2. Calculate the mass of chlorine that reacted with the magnesium.
3. Calculate the number of moles of chlorine in the compound.
4. Find the whole-number mole ratio of Mg to Cl in the compound. (Dividing the number of moles of each element by the smaller number of moles gives the whole-number mole ratio.)
5. Write the formula for the compound, and compare it to the known formula $MgCl_2$.
6. Calculate the percent Mg and the percent Cl in the compound from both your lab data and the known chemical formula $MgCl_2$.
7. How do the percent compositions compare with each other?
8. What might be some causes of error in your experiment?

Sample assessment

- Use the students’ lab observations and calculations for assessment.
- Use the questions on the attached handout for assessment. Have the students work individually or with their lab partners.

Follow-up/extension

- Have students find the empirical formulas for additional ionic compounds, such as $ZnCl_2$ and $AlCl_3$, using the same basic procedures.

Finding the Formula and Percent Composition of an Ionic Compound

Directions: Answer the following questions. Show all your calculations, rounding your answers to the teacher-specified number of significant digits and labeling units clearly.

1. What is the percent by mass of each element in the compound $C_{14}H_{18}N_2O_5$?
2. What is the percent composition of sucrose ($C_{12}H_{22}O_{11}$)?
3. What is the percent of phosphorous in $Sn_3(PO_4)_2$?
4. What is the percent composition of aluminum oxide?
5. What is the percent composition of heptaphosphorous decoxide?
6. What is the percent by mass of iron in iron(III) chloride?
7. An oxide of aluminum metal is formed by the reaction of 4.151 grams of Al with 3.692 grams of O. What is the empirical formula for this compound?
8. A chemist analyzed an unknown compound and collected the following data: 0.8007 g C, 0.9333 g N, 0.2016 g H, 2.133 g O. What is the empirical formula for the compound? What is the percent composition of the compound?
9. In a lab experiment, it was observed that 0.6884 grams of lead combines with 0.2356 grams of oxygen. What is the empirical formula for this compound?
10. Cisplatin is used to treat cancerous tumors. It has the following composition: 65.02 g Pt, 9.34 g N, 2.02 g H, and 23.63 g Cl. What is the empirical formula for cisplatin?
11. The most common form of nylon is 63.68 g C, 12.38 g N, 9.80 g H, and 14.14 g O. What is the empirical formula for this common form of nylon?
12. If 1.000 g of barium metal is heated in a stream of pure oxygen gas, 1.117 grams of an oxide is produced. What is the empirical formula for this oxide?
13. A compound has the following composition: 58.84 g Ba, 13.74 g S, and 27.43 g O. What is the empirical formula for this compound?
14. If cobalt metal is mixed with sulfur and heated strongly, 100 grams of a compound is produced that contains 55.06 grams of cobalt . What is the empirical formula for the compound?
15. When 3.269 grams of zinc metal is heated in pure oxygen, the sample gains 0.800 grams of oxygen in forming the oxide. What is the empirical formula for the oxide?
16. A compound consists of 65.45 g C and 5.492 g H and 29.06 g O. What is the empirical formula for the compound?
17. Phosphorous forms two oxides. One has 56.34 g P and 43.66 g O. The other has 43.64 g P and 56.36 g O. What are the empirical formulas for these two oxides?
18. Magnetite is an iron ore. 100 g contains 72.4 g Fe and 27.6 g O. What is its empirical formula?
19. What is the percentage of water by mass in the compound $CoCl_3 \bullet 6\ H_2O$?
20. A compound with 26.2 g N, 7.5 g H, and 66.3 g Cl is sometimes used to salt icy roads. What is the empirical formula for this compound?

Aspirin Analysis

Organizing Topic Stoichiometry

Overview Students use mole ratios and stoichiometry to determine the amount of acetylsalicylic acid in an aspirin tablet.

Related Standards of Learning CH.1a, b, c; CH.3d; CH.4b

Objectives

The students will
- investigate and understand that quantities in a chemical reaction are based on molar relationships;
- identify and use
    - chemical formulas
    - coefficients, chemical symbols, and subscripts.

Materials needed

- 50-mL burettes
- Beaker
- 0.2 *M* NaOH
- Ethanol
- Variety of brands of aspirin tablets
- Phenolphthalein indicator
- 125-mL Erlenmeyer flasks

Instructional activity

Content/Teacher Notes

In this lab, students will analyze aspirin (acetylsalicylic acid) tablets for the acetylsalicylic acid content, using a volumetric analysis technique called "titration." Not all of the mass of an aspirin tablet is acetylsalicylic acid, as there are starch binders added to hold the tablet together. Sodium hydroxide reacts with the acetylsalicylic acid but not with the starch binders, as shown in the following reaction:

$C_9H_8O_4$	+	NaOH	→	$NaC_9H_7O_4$	+	H_2O
acetylsalicylic acid		sodium hydroxide		sodium acetylsalicylate		water

This is an acid-base reaction in which the acetylsalicylic acid reacts with the base sodium hydroxide to produce the salt sodium cetylsalicylate and water (acid + base → salt + water).

To determine when the reaction is complete, students will use a small amount of the indicator phenolphthalein. Phenolphthalein is an indicator because it changes color when all the aspirin has been reacted. This is called the "end point" of the reaction. If one continues adding NaOH after the end point, the solution will turn dark red. You want to stop the titration when the solution is a pale pink color.

Students will use a burette to measure the volume of sodium hydroxide solution that reacts with the aspirin. With a burette, you can add small amounts of NaOH and measure the volume more accurately. This method of analysis is called "titration."

Introduction

1. Introduce the lab by telling students that in this method of analysis, they will use a substance, sodium hydroxide (NaOH), whose concentration they know to analyze a substance, aspirin, whose concentration they do not know.
2. Introduce or review the terms *burette, indicator, phenolphthalein, end point,* and *titration,* using the information under “Content/Teacher Notes” above. Make certain that the students understand these terms and how they are used.

Procedure

Aspirin Brand A Lab Data	
Initial burette reading	
Mass of aspirin tablet	
Total volume of aspirin solution	
Final burette reading	
Volume of NaOH reacted	
Aspirin Brand B Lab Data	
Initial burette reading	
Mass of aspirin tablet	
Total volume of aspirin solution	
Final burette reading	
Volume of NaOH reacted	

1. Make a data table like the one at right.
2. Fill a 50-mL burette almost to the top with the 0.2 *M* NaOH solution. Fill the tip of the burette by draining some the NaOH into a beaker. Drain the NaOH until the level is at the zero mark. Record this initial burette reading in your data table.
3. Find and record the mass of an aspirin tablet of brand A.
4. Place the aspirin tablet into a 125-mL Erlenmeyer flask, and add 25 mL of warm water and 15 mL of ethanol to the flask. Swirl the mixture until the aspirin tablet dissolves.
5. Add two or three drops of the phenolphthalein indicator to the solution in the flask and swirl. (Without the indicator, you would not know when the reaction is complete.)
6. Place the flask on a piece of white paper, and lower the burette so that the tip is inside the neck of the flask. Slowly begin adding the NaOH solution from the burette while gently swirling the contents of the flask. With each addition of NaOH, you will see a pink color appear and then quickly fade away. When the pink color lasts for a few seconds before fading away, begin adding the NaOH drop by drop.
7. Stop adding NaOH when you see a faint pink color that remains for at least 30 seconds before fading; this is the end point of the reaction. If you add too much NaOH, the solution will turn red, and you will have gone past the end point. You will have to start over again.
8. Find and record the total volume of the aspirin solution (to be used in step 2 of the “Observations and Conclusions” below).
9. Record the level of the NaOH in the burette tube at the end point as the final burette reading.
10. Subtract the final burette reading from the initial burette reading to find the volume of NaOH that was added. This amount is the volume of NaOH that was reacted by aspirin tablet A. Record this amount.
11. Rinse out the flask, and repeat the experiment with a different brand of aspirin.

Observations and Conclusions

1. Have students compare the amounts of NaOH reacted in the two experiments. These amounts relate directly to the amounts of aspirin in the two tablets. Was there much difference in the amounts of aspirin in the two brands?
2. The molarity of the NaOH is known to be 0.2 *M*, and the mole ratio of the acetylsalicylic acid to the NaOH is 1:1. Have students use this known molarity of the NaOH, the volume of the NaOH reacted, and the total volume of the aspirin solution to calculate the molarity of the aspirin in the solution.

3. Knowing the mass of the aspirin tablet and given the molar mass of acetylsalicylic acid as 180 g/mol, calculate the percent acetylsalicylic acid and the percent binder in the aspirin tablet.

Sample assessment

- Have students analyze the contents of various brands of aspirin and compare the results.

Follow-up/extension

- Have students research the quality control tests that are done in industry to monitor the amount of aspirin contained in tablets.

Organizing Topic — Kinetic Theory

Standards of Learning

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include
  f) mathematical and procedural error analysis;
  g) mathematical manipulations (SI units, scientific notation, linear equations, graphing, ratio and proportion, significant digits, dimensional analysis); and
  h) use of appropriate technology including computers, graphing calculators, and probeware, for gathering data and communicating results.

CH.4 The student will investigate and understand that quantities in a chemical reaction are based on molar relationships. Key concepts include
  c) partial pressure;
  d) gas laws; and
  e) solution concentrations.

CH.5 The student will investigate and understand that the phases of matter are explained by kinetic theory and forces of attraction between particles. Key concepts include
  a) pressure, temperature, and volume;
  b) vapor pressure;
  c) phase changes;
  d) molar heats of fusion and vaporization;
  e) specific heat capacity; and
  f) colligative properties.

Essential Understandings, Knowledge, and Skills

Correlation to Textbooks and Other Instructional Materials

The student will use hands-on investigations, problem solving activities, scientific communication, and scientific reasoning to

- recognize the following principles relative to the Kinetic-Molecular Theory:
  - Atoms and molecules are in constant motion. ______________________
  - The theory is a model for predicting and explaining gas behavior. ______________________
  - Forces of attraction between molecules determine the physical changes of state. ______________________
  - Pressure, temperature, and volume changes can cause a change in physical state. ______________________
  - Solid, liquid, and gas phases of a substance have different energy content. ______________________
- recognize the following properties of gases:
  - Gases have mass and occupy space, and relatively large distances separate gas particles from each other. ______________________
  - Gas particles are in constant, rapid, random motion and exert pressure as they collide with the walls of their containers. ______________________

  - An Ideal Gas does not exist, but this concept is used to model gas behavior. ______________________
  - A Real Gas exists, has intermolecular forces and particle volume, and can change states. ______________________
  - Gas molecules with the lightest mass travel fastest. ______________________
- state the following laws:
  - Boyle's Law; ______________________
  - Dalton's Law of Partial Pressures; ______________________
  - Charles' Law; ______________________
  - The Ideal Gas Law (PV = nRT);
- solve problems and interpret graphs involving all gas laws; ______________________
- use pressure units such as kPa and mm of Hg; ______________________
- identify the forces of attraction as hydrogen bonding, dipole-dipole attraction, and van der Waals forces; ______________________
- perform investigations in which polar substances dissolve ionic or polar substances; nonpolar substances dissolve nonpolar substances; ______________________
- define *vapor pressure* as a property of a substance determined by intermolecular forces; ______________________
- understand that
  - specific amounts of energy are absorbed or released during phase changes; ______________________
  - boiling point of liquids is affected by changes in atmospheric pressure; ______________________
- define *specific heat capacity*; ______________________
- calculate energy changes using molar heat of fusion and molar heat of vaporization; ______________________
- perform calorimetry calculations; ______________________
- calculate energy changes using specific heat capacity; ______________________
- interpret a phase diagram of water; ______________________
- graph and interpret a heating curve; ______________________
- review solute, solvent, and solution types, and calculate solution concentration; ______________________
- recognize that
  - the number of solute particles changes the freezing point and boiling point of a pure substance; ______________________
  - the freezing point and boiling point of a liquid are affected by the presence of certain solutes; ______________________
  - polar substances dissolve ionic or polar substances; nonpolar substances dissolve nonpolar substances. ______________________

States of Matter

Organizing Topic Kinetic Theory

Overview Students explore the Kinetic-Molecular Theory (KMT) as it relates to the states of matter and phase changes. They apply this knowledge to determining energy relationships and calculations involving calorimetry.

Related Standards of Learning CH.1; CH.4; CH.5

Objectives

The students will

- explain that all matter is in constant motion and that the speed and proximity of particles differentiates the three states of matter;
- recognize that changes of state have corresponding energy changes and are affected by the attractive forces between particles;
- recognize that changes of state are affected by pressure, volume, and temperature;
- define *heat capacity* and explain that heat capacity is affected by physical state and substance composition;
- perform energy calculations involving heat capacity, heat of fusion, and heat of vaporization;
- perform and calculate experimental values, using a calorimeter;
- interpret graphs, including phase diagrams and heating/cooling curves.

Materials needed

- Large, empty pretzel jar, one-fourth filled with 1-inch white Styrofoam balls and two or three colored Styrofoam balls
- Ice cube with temperature probe frozen inside
- Colored pencils or markers
- Temperature probes
- 250-mL beakers
- Hot plates
- Balances
- Calorimeters (a double Styrofoam cup will work)
- Ice cubes
- Picture-hanging wire or other strong, thin wire
- Vacuum pump
- Attached worksheets

Instructional activity

Content/Teacher Notes

From their previous science courses, students should already be familiar with the three basic states of matter and their physical descriptions. The goal of this activity is to connect this knowledge with a deeper understanding of why these physical descriptions are valid, based on the orientation and movement of particles in each state.

The activity will focus on providing a visual model of particle motion and the role that energy plays. Students should gain the understanding that there is a direct relationship between temperature and molecular motion and that relative molecular motions determine physical state.

Students should also be able to explain the energy changes that cause changes of state. If it has not already been covered, these changes should be discussed in the context of endothermic and exothermic changes. Basic calorimetry calculations will be used in determining heat content and change.

Introduction

1. Explain the Kinetic-Molecular Theory (KMT). Ask students to define *temperature* and *pressure,* using the KMT. Explain that *hot* and *cold* are relative terms and have very little meaning in chemistry. Model this by asking the students to touch the metal on their chair or desk with one hand and touch the wood with the other hand. Ask them which is colder. Point out that both substances have been in the same room together for a long time, and ask if they would have the same temperature if measured with a thermometer.
2. Explain why one substance feels cold and the other feels warmer even though they are actually at the same temperature. Bring in the idea of conductivity and metal versus nonmetal.
3. Remind students that all objects in the room and the air in the room are at the same temperature, unless there is an obvious reason for them to be at a different temperature — e.g., in direct sun, by a heat or AC register. Ask students whether the air particles are moving at the same speed as the desk molecules.
4. Use their answers to explain what Kinetic Energy is and the differences in particle motion of solids, liquids, and gases at the same temperature. Ask a student to take the pretzel jar with Styrofoam balls and model the behavior of particles in solids, showing the balls vibrate but not move place to place. Have another student model the movement of particles in liquids (a little place-to-place movement of the balls), and another, the behavior of particles in gases (great place-to-place movement). During this process, reinforce how the motion and orientation of particles relates to the physical description of each state of matter.
5. Ask students the relationship between kinetic energy and changes of state. Have a student use the pretzel jar to model ice heating up to become steam and then cooling back down to become ice. Make sure he/she narrates whether energy is being put in or taken out as the phase changes are simulated. Make sure the student shows all phase changes.
6. Explain that the energy needed to change ice to steam can be measured and calculated. Demonstrate this by using an ice cube with a temperature probe frozen inside.
    - Connect the temperature probe.
    - Place the ice cube in a beaker on a hot plate, and heat.
    - Record temperature-and-time readings until the water boils.
7. While this is demonstration is running, hand out the attached “Heating Curve” worksheet. Have students complete Part 1.
8. Explain the difference between heat capacity, heat of fusion, and heat of vaporization. Ask students why the temperature levels off during a phase change and where the heat that is still being added is going.
9. Show students the equation for calculating heat in a calorimeter: $H = m(\Delta T)C_p$. Make sure they understand each symbol and why it is in the equation. This is a good time to stress that heat and temperature are not the same but are directly related. Ask them why this equation cannot be used during a phase change (no ΔT). Then explain how heat is calculated during a phase change, using heat of fusion and heat of vaporization.
10. Have students complete Part 2 of the “Heating Curve” worksheet.

Part 1: Determining the Heat of Fusion of Ice

Procedure

1. Find the mass of the empty calorimeter, and record.
2. Add about 100 mL of warm water (about 50–60°C) to the calorimeter, and record the combined mass.
3. Find and record the initial temperature of your water.
4. Obtain an ice cube from the cooler, quickly find its mass, and record
5. Immediately add the ice to the water, and record the water temperature every 30 seconds, if using a temperature probe, or when it levels off, if using a thermometer. Be sure to note the final temperature once the ice has completely melted.
6. Clean up materials.
7. Calculate the change in temperature of the water.
8. Calculate the heat lost by the water, in joules.
9. Calculate the amount of heat the ice cube gained, and explain your answer.
10. Calculate the molar heat of fusion for the ice. (Find the moles of ice first.) Record the heat of fusion in the class data table.
11. Look up the actual molar heat of fusion, and calculate your percent error.
12. Are the class results accurate and/or precise? Why, or why not?
13. Identify two major sources of error and their effect on your outcome.

Part 2: Determining the Heat Needed to Boil Water

Procedure

1. Find the mass of a 250-mL beaker.
2. Pour 100 mL of ice water from an ice water bath (be sure there is no ice in your sample) into your beaker, and quickly find its mass.
3. Find and record the initial temperature of your ice water, and place the beaker on the hot plate. Heat to boiling.
4. Record the temperature when the water reaches the boiling point.
5. Turn off the hot plate, and clean up.
6. Determine the temperature change.
7. Calculate the heat gained by the water.
8. Compare this to the heat calculated in Part 2 of the "Heating Curve" worksheet for heating liquid water to boiling. Calculate your percent error.
9. Identify two major sources of error and their effect on your outcome.

Observations and Conclusions

1. There will be fairly large errors in determining the heat of fusion. Students should be bothered that they got more heat than they should have. Be sure to have a discussion about whether the ice started at 0° C or not. Ask the students how they could figure this out and what effect it might have had on the experimentally determined value.
2. Ask the students if there is another way to melt ice other than adding heat. Discuss the role of volume and pressure on physical states and phase changes. As a part of this discussion, ask

students how ice skating works — i.e., how it is possible to glide so easily across the surface of a solid.

3. Demonstrate the phenomenon of melting caused by pressure by passing a thin, strong wire through an ice cube. This is easy to do if the ice cube is supported on top of a narrow bottle so that you can pull straight down. Wear gloves or wrap your hands with towels to keep the wire from cutting into them. You need to exert a good amount of pressure, so you might want a stronger student to do this. An alternative is to put lead or balance weights on each end of the wire and let it happen over time during the class period. Ask students to explain why the ice cube is still in one piece after the wire has passed through completely.
4. Hand out the attached "Phase Diagram" worksheet, and explain that Ideal Gas behavior is favored at high temperatures and low pressures. Ask the students to explain why. Based on their answers, ask them to identify which region on the graph is most likely to be a gas. Have them label that region. Ask them what conditions of temperature and pressure favor the solid state. Have them label this region. Then, have them label the liquid region. Ask them to predict the significance of the places all three curves intersect. Ask them to identify which curve represents all of the boiling points for this substance and which curve represents all of the melting points for this substance. Then have them name the process which occurs when changing from conditions of X to conditions of Z. Finally, have them complete the worksheet questions.

Sample assessment

- Ask students to write a journal of a molecule of a specified substance as it travels a heating and/or cooling curve. Have them narrate what happens to the molecule's behavior (movement and orientation) and energy. The phase changes should be accurate to the substance specified.

Follow-up/extension

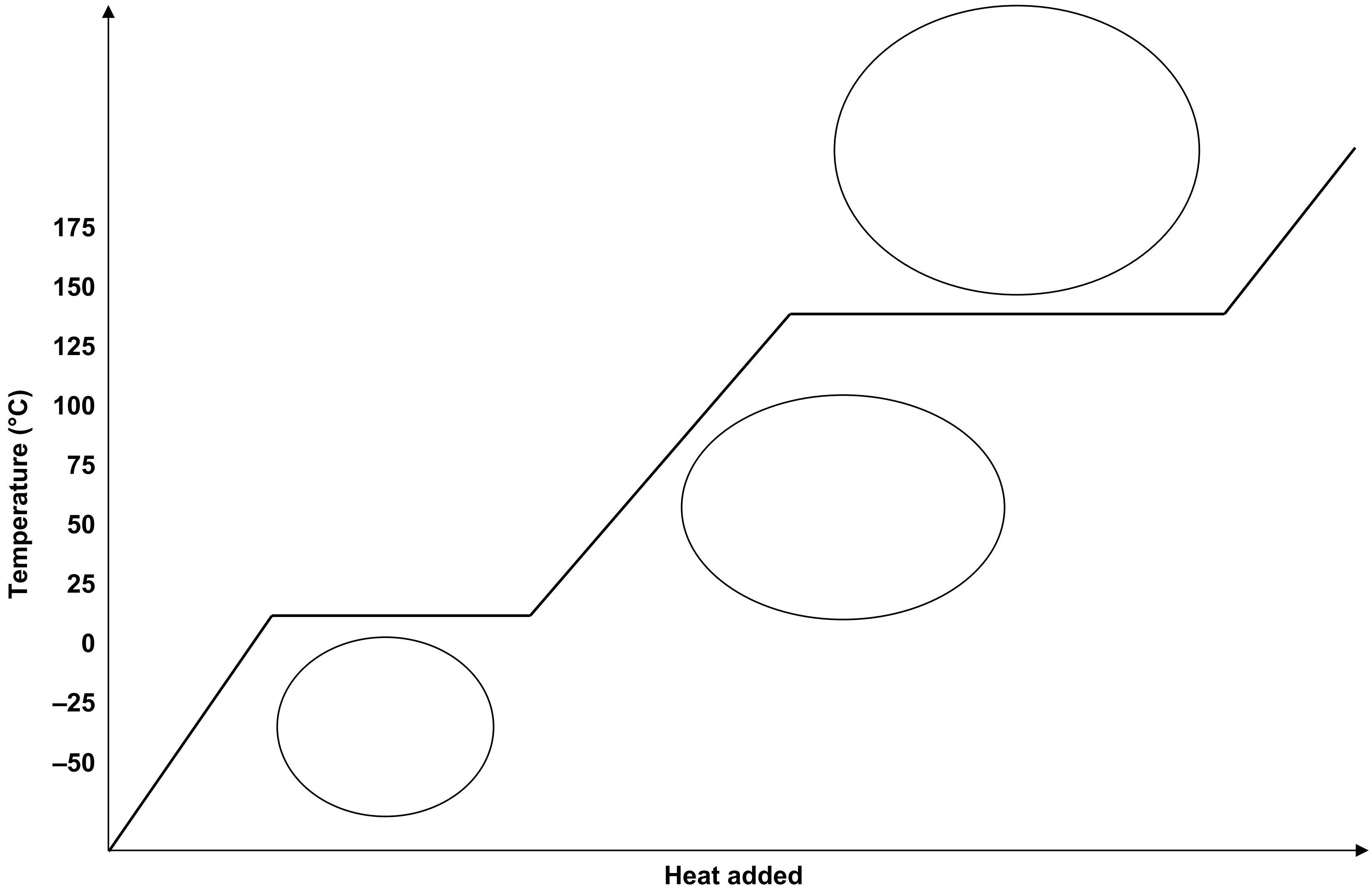
- Have students design an investigation into a phase change, including, perhaps, colligative properties. Encourage them to keep it simple and small in focus. The idea is for students to identify variables and constants and to test out a hypothesis. Some examples are listed below. Remember, it is not that important that they get an answer to their question or that the question is based in valid science (it can be useful to let them test some common myths). This should be about the process and being able to explain why things work or do not work. Possible investigations might include the following;
  - Does salt water freeze and/or boil faster than pure water?
  - Does cold water freeze faster than hot water?
  - Does the amount/type of solute affect the boiling point?
  - Does alcohol boil at the same temperature as water?
  - Does the rate of melting depend more on mass or more on surface area?

Resources

- *Molecular Level Motion of Different Phases.* http://www.chamotlabs.com/Phases.html. This Web site simulates phase transitions.
- *Phase Transitions.* http://intro.chem.okstate.edu/1515SP01/Lecture/Chapter12/PLMPhase.html. This Web site has a good particle-motion simulation.

Heating Curve for Water

Name: ______________________________ **Date:** ____________________



Heating Curve for Water

Name: ______________________________________ Date: _______________________

Instructions

Using the “Heating Curve for Water” graph, complete the following:

1. In what temperature range, in °C, is water a solid? _____________ On the graph, find the circle in this temperature range, and label it “Solid.” Inside this circle, draw a picture of how water *molecules* look in the solid phase.
2. In what temperature range is water a liquid? _____________ Find the circle in this temperature range, and label it “Liquid.” Inside this circle, draw a picture of water *molecules* in the liquid phase.
3. In what temperature range is water a gas? _____________ Find the circle in this temperature range, and label it “Gas.” Inside this circle, draw a picture of water *molecules* in the gas phase.
4. At what temperature does water melt? ________ Find the line that corresponds to this temperature, and *above the line,* label it “Melting.” Indicate whether this change is endothermic or exothermic.
5. At what temperature does water freeze? ________ Find the line that corresponds to this temperature, and *below the line,* label it “Freezing.” Indicate whether this change is endothermic or exothermic.
6. At what temperature does water boil? ________ Find the line that corresponds to this temperature, and *above the line,* label it “Boiling.” Indicate whether this change is endothermic or exothermic.
7. At what temperature does water condense (change from gas to liquid)? ________ Find the line that corresponds to this temperature, and *below the line,* label it “Condensing.” Indicate whether this change is endothermic or exothermic.
8. Use a chemistry book to look up the definition of *temperature*. What does temperature measure?

   ____

9. Explain the relationship between temperature, motion, and position of particles inside a container.

   ____

10 At what temperature is both the liquid and gas phase of water present inside a container? ______

11. At what temperature is both the solid and liquid phase of water present?________

12. Label the following equations as melting, condensing, boiling, or freezing:
    - $H_2O(s) \rightarrow H_2O(l)$ _______________________
    - $H_2O(l) \rightarrow H_2O(g)$ _______________________
    - $H_2O(l) \rightarrow H_2O(s)$ _______________________
    - $H_2O(g) \rightarrow H_2O(l)$ _______________________
    - Do these equations represent a chemical or physical change? ________________

13. Calculate the heat needed to raise 30.0 g of ice from −10°C to 0°C (heat capacity of ice = 2.03 J/g°C).

    ____

14. Calculate the heat needed to melt 30.0 g of ice (heat of fusion = 6.02 kJ/mol).

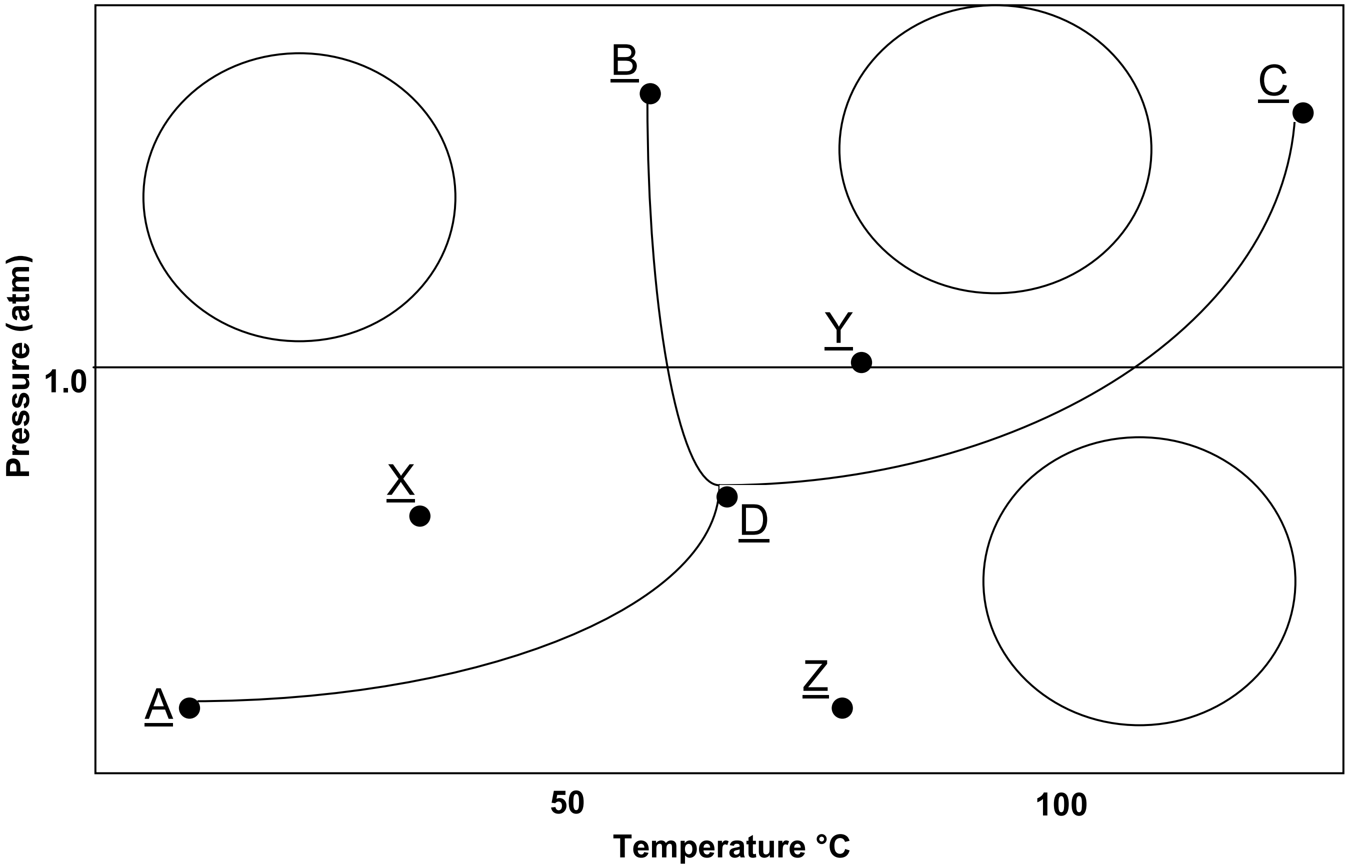
    ____

15. Calculate the heat needed to raise 30.0 g of water from 0°C to 100°C (heat capacity of water = 4.18 J/g°C).

    ____

Phase Diagram

Name: ______________________________________ **Date:** _______________________



1. Inside of each circle in the diagram above, label the physical state present in that region, and draw the orientation of the particles in that state.
2. Identify the phase change represented by each of the following curves:
   - DC ____________________________
   - BD ____________________________
   - AD ____________________________
3. What states of matter are present at point D? ______________________________ What is this point called? ___________________________________
4. Estimate the normal melting point for this substance: ____________. Estimate the normal boiling point for this substance: ____________.
5. What physical state is present at point Y? _______________________________
6. What process occurs when you move from X to Y? _________________________
7. What process occurs when you move from Z to Y? _________________________
8. What process occurs when you move from Y to X? _________________________
9. What process occurs when you move from X to Z? ________________________
10. What is the significance of point C? What is this called? __________________________________
11. Would this substance be more dense as a solid or as a liquid? (Think about whether increasing pressure of the solid would make it melt or freeze.) ________________________ Why?

Vapor Pressure and Colligative Properties

Organizing Topic Kinetic Theory

Overview Students explore the relationship between intermolecular forces and vapor pressure. They apply this knowledge in an investigation of colligative properties of solutions.

Related Standards of Learning CH.1; CH.4; CH.5

Objectives

The students will

- use pressure units, such as kPa, atm, and mmHg;
- identify forces of attraction between molecules, and evaluate the relative magnitude of the various intermolecular forces;
- investigate relative attractions of different intermolecular forces by simple solubility tests;
- investigate the relationship between vapor pressure and intermolecular forces;
- investigate the relationship between vapor pressure, air pressure, and temperature;
- investigate the relationship between particle concentration and vapor pressure by looking for evidence of colligative properties of solutions.

Materials needed

Note: Materials will depend on the method(s) chosen for determining evaporation rate.

- Liquids: water, ethanol, acetone, pentane
- Solids: NaCl, sugar, wax (paraffin)
- NaCl solution (350 g/L)
- Sugar solution (350 g/L)
- Spotting plates or watch glasses
- Hot plates
- Beakers
- Erlenmeyer flask with stopper
- Thermometers or temperature probes
- Balances
- Timers
- Vacuum pump
- Toothpicks
- Attached worksheets

Instructional activity

Content/Teacher Notes

Before undertaking this activity, students should have a basic understanding of ionic versus covalent bonding, molarity, basic kinetics, and solubility. This lesson will focus on the conceptual understanding of the influence of intermolecular forces on the magnitude of the vapor pressure of a liquid, the determination of the normal boiling point of a substance, and the effect of solute particles on the vapor pressure of a solution and the colligative properties that occur as a result of their presence.

Reviewing the concepts of polarity, dissociation, and “like dissolves like” is recommended. Students should be familiar with the units of temperature and pressure and the conversions between them.

Most of the activities are fairly straightforward; however, the evaporation-rate investigation is somewhat open-ended. For weaker classes, you might want to decide which procedure they should do, while for stronger classes, you may want to give them some choices. Students in honors classes might design their own procedures within given parameters of the materials available. If temperature probes are available, students will find it extremely interesting to see the evaporation-rate effect graphically as evaporative cooling versus time. *CAUTION! Keep procedures as much on a microscale as possible to reduce fumes.*

Some options for measuring evaporation rate are the following:

- If you have electronic balances, measure loss of mass versus time directly.
- With timers, measure the time it takes for a drop to evaporate, doing repeated trials and averaging the results.
- With temperature probes, dip them in the liquid and then remove them, graphing temperature versus time. This allows the measurement of the time needed to evaporate (return to room temperature), and it allows students to see that more volatile liquids have a greater cooling effect when evaporating.

The molecular mass of hexane is too high to be used in this comparison. Pentane is fairly inexpensive and as safe as hexane to use, but you should still exercise caution. Butane (lighter fluid) could be used in place of pentane, but it is a greater fire risk and evaporates so quickly that only spot testing is an option.

Introduction

1. Start this lesson by having students diagram the molecules of the target liquids and identify their dominant intermolecular force.
2. Have students predict the relative magnitude of the intermolecular forces from weakest to strongest by having them complete the attached “Vapor Pressure and Intermolecular Forces” worksheet.
3. Define *vapor pressure,* and discuss its relationship to intermolecular forces. Along with this discussion, either demonstrate vapor pressure or use a computer simulation of it (see Resources, if you do not have one), showing how it is affected by temperature. For example, if you have a temperature probe, insert it into a one-hole stopper that fits an Erlenmeyer flask containing a small amount of water. Seal the flask, and start recording temperature versus time as you gently warm the flask so that the students can see the pressure increase as the temperature increases.
4. Ask students at what temperature water boils. Then ask them whether all substances that are liquid at room temperature boil at this temperature. Ask them whether water always boils at this temperature. Have them explain their reasoning.
5. Demonstrate water boiling at different temperatures. Place 100 mL of water into each of three small beakers, and label them “50°C,” “65°C,” and “80°C.” Raise the temperature of the water in each beaker to the labeled temperature. Have a student check the temperatures, which do not need to be exactly the labeled temperatures, and record the measured values. Place all three beakers on the plate of the vacuum pump, seat the lid, and turn on the pump. Ask students to watch and record the order in which they boil as the pressure decreases.
6. Release the vacuum, and discuss the relationship between air pressure and boiling point. Ask the students if they think it is possible to make ice water boil. If you have a good vacuum pump, you should be able to demonstrate this; if not, you may wish to show a video clip of it. As an alternative, under Resources there is a demo listed that shows water boiling by using ice.
7. Hand out the “Vapor Pressure versus Temperature” worksheet, and have students focus on curve A. Have them explain where the liquid and gas phases for substance A would be located and how many boiling points liquid A would have. *(Many)* Ask them where they would find the normal boiling point. Have them identify the normal boiling point of each of the four substances.

Procedure

Part 1: Determining Relative Vapor Pressures by Measuring Evaporation Rates

1. Ask students to explain the relationship between vapor pressure and evaporation rate. Encourage them to use their knowledge of intermolecular forces in this discussion.
2. Have students brainstorm some ways to measure evaporation rates with available lab equipment.
3. Divide the class into groups, and have each group choose a method of measuring evaporation rates.
4. Have each group design an experiment by writing a research question and hypothesis and identifying variables and constants. (Students are not always good at predicting all of the constants needed.) Check and approve the written questions and hypotheses before giving students the go-ahead to write detailed procedures.
5. Have the groups do multiple trials of their experiments and adjust their procedures as needed.
6. Have the groups organize their data into a table and graph their average values.

Part 2: Comparing Vapor Pressures

1. Hand out the "Vapor Pressure and Intermolecular Forces" worksheet.
2. Have students complete the worksheet.
3. Check the completed worksheets for accuracy.

Part 3: Effect of Solutes on the Vapor Pressure of Solutions

1. Have each student group find the boiling point of water at the prevailing pressure by heating the water on a hot plate and using a temperature probe.
2. Have the groups write a research question and hypothesis concerning the effect of sodium chloride (NaCl) on the boiling point of water and identify variables and constants.
3. Give each group 100 mL of NaCl solution (350 g/L), and have them find its boiling point. Make sure they understand that they need to record the temperature when it levels off. Have them graph the result.
4. Have each group repeat steps 2 and 3, using a sugar solution (also 350 g/L).

Part 4: Effect of Bond Type on Solubility

1. Give student groups the four liquids — water, ethanol, acetone, and pentane *(CAUTION! Exercise care when using pentane)* — and the three solids — sodium chloride (NaCl), sugar, and wax (paraffin).
2. Have the groups determine the solubility of each liquid in each of the other liquids, and the solubility of each solid in each liquid.
3. Have the groups design an appropriate data table before they start their experimentation.
4. Have the groups carry out each test by placing a small amount of each substance in the same well, stirring with a clean toothpick, and recording whether the substances are miscible or not.

Observations and Conclusions

Hold a class discussion, or give an assessment, in which the students do the following:

1. Define *vapor pressure,* and explain how it is measured.
2. Explain the relationship between intermolecular forces and vapor pressure.
3. Explain why water, with its low molar mass, has a much higher vapor pressure than carbon dioxide (CO_2), which has a much higher molar mass.

4. Read about colligative properties in the text, and explain how and why the addition of a solute affects the vapor pressure of the solvent.
5. Using the masses dissolved per liter of solution in Part 3 above, determine the molarity of the salt and sugar solutions. Did you have an equal number of particles in each solution? If not, which solution had a higher particle concentration?
6. Explain whether the number of particles in solution causes elevation of boiling point. Should it?
7. Tell which would have a higher particle concentration — a 1.0 *M* solution of salt or a 1.0 *M* solution of sugar? Why? Which solute would have greater effect on the vapor pressure of water? Why?
8. Explain why $CaCl_2$ melts ice on sidewalks more efficiently than does NaCl.
9. Explain why we cannot use wax particles to de-ice our sidewalks. What solubility rules determine this? How does your data from Part 4 support these rules? Give examples.
10. What is the normal boiling point of each liquid used in this lab?
11. Complete the questions on the "Vapor Pressure versus Temperature" worksheet.

Sample assessment

- Have students explain why it takes longer to boil an egg in Denver than in Richmond.
- Have students explain the function of antifreeze in a car and the danger of adding plain water when the antifreeze gets low. (The explanation should include both boiling point elevation and freezing point depression.)

Follow-up/extension

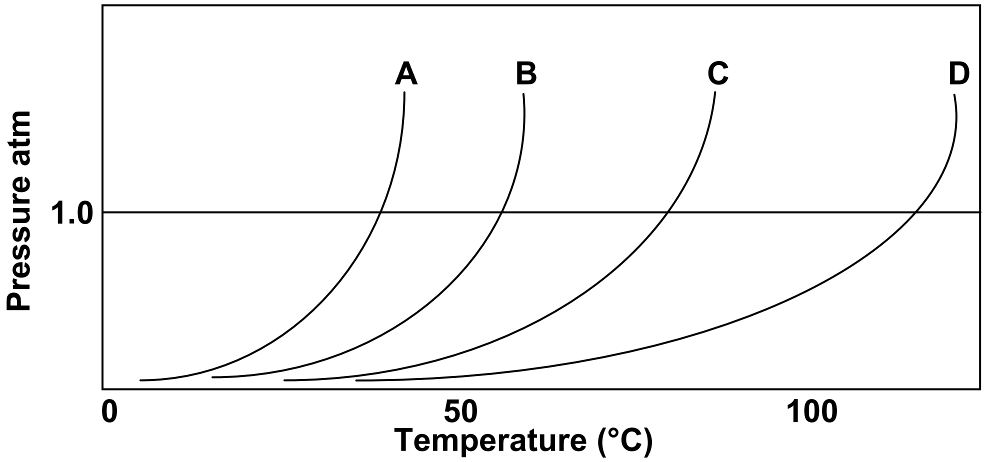
- Have students research how refrigerators and heat pumps work. Ask them to include an explanation of how vapor pressure must be considered when finding replacements for chlorofluorocarbons (CFCs).
- Have students explain why sweating cools the body and why rubbing alcohol and acetone feel cold on the skin. Make clear they need to demonstrate an understanding of vapor pressure in their answers.
- Have students explain how a pressure cooker works.

Resources

- *Chemistry Comes Alive! Boiling by Cooling.* http://jchemed.chem.wisc.edu/JCESoft/CCA/CCA2/MAIN/BOILCOL1/CD2R1.HTM. This Web site shows water boiling by using ice.
- *Colligative Properties.* http://neon.chem.uidaho.edu/~honors/collig.html. This Web site has a good, clear explanation of how colligative properties affect vapor pressure.
- *Consumer Information: How a Heat Pump Works.* http://www.ari.org/consumer/howdoesitwork/howHPwks.html. This Web site gives the basics of how an air conditioner works.
- *Evaporative Cooling.* http://www.colorado.edu/physics/2000/bec/evap_cool.html. This Web site gives a simple explanation of evaporative cooling.
- *Phase Changes: Vapor Pressure.* http://www.chm.davidson.edu/ChemistryApplets/PhaseChanges/VaporPressure.html. At the bottom of this Web page is a good simulation applet for modeling vapor pressure.

Vapor Pressure versus Temperature

Name: ____ **Date:** ________________________



1. Every point on each line is a boiling point for that substance. At what temperature does substance B boil at 0.5 atm? _______________
2. Normal boiling point occurs at 1.0 atm. What is the normal boiling point for substance C? _____________________
3. Which substances would be liquids at room temperature and 1.0 atm? _______________
4. Which substances would be gases at room temperature and 1.0 atm? _______________
5. What would be the physical state of substance D at 120°C and 0.50 atm? _______
6. Which substance has the strongest intermolecular forces? Why?

7. Which substance has the weakest intermolecular forces? Why?

8. If you have a mixture of A and C, which substance would boil first? Why?

9. If all of these substances were nonpolar, which would have the lowest molecular mass? Why?

10. If all of these substances were polar, which one is most likely to have hydrogen bonding? Why?

11. The substances on the graph above are the four liquids used in this lab. Identify each liquid.

A _______________________ B ___________________________
C _______________________ D ___________________________

Vapor Pressure and Intermolecular Forces

Name: ____ **Date:** ________________________

Part 1: Complete the table below.

Substance	Lewis Dot Diagram	Structural Diagram	Bond Type	Polarity
Water				
Ethanol				
Acetone				
Pentane				
NaCl				N/A
Sugar				
Wax				

Part 2: Answer the questions below.

1. Which liquids should be able to dissolve wax? ____

   Why? ____

2. Which liquids should be able to dissolve salt? ____________________

   Why? ____________________

3. Which liquids should be able to dissolve sugar? ____________________

   Why? ____________________

4. Which type of IMF is predominant in water?____________________

   Why? ____________________

5. Which type of IMF is predominant in pentane?____________________

   Why? ____________________

6. Which type of IMF is predominant in acetone? ____________________

   Why? ____________________

7. Place the solvents (liquids) in order of increasing IMF. Explain your reasoning.

   ____

   ____

8. Both salt and sugar dissolve in water. Which one also dissociates? ____________________

   Why? ____________________

9. Write an equation to show how each substance (salt and sugar) dissolves, then explain the difference in the number of particles produced per mole of compound.

   ____

   ____

Soap, Slime, and Creative Chromatography

Organizing Topic Kinetic Theory

Overview Students explore the intermolecular interactions of matter.

Related Standards of Learning CH.3; CH.5f

Objectives

The students will

- recognize that forces of attraction between molecules determine the physical changes of state;
- understand that polar substances dissolve ionic or polar substances, and nonpolar substances dissolve nonpolar substances.

Materials needed

- Cottonseed oil
- Ethanol
- 20% NaOH solution
- Plastic pipettes
- Watch glasses
- 250-mL beakers
- Saturated NaCl solution
- Paper towels
- Scoop
- Vegetable oil
- 1% $CaCl_2$ solution
- 4% polyvinyl alcohol solution
- Food coloring
- Plastic medicine cups
- Wooden splints
- 4% sodium tetraborate
- White glue
- Filter paper
- Markers with water-soluble ink in black, purple, green, yellow, red, orange, blue, and brown
- Isopropyl alcohol
- Distilled water

Instructional activity

Content/Teacher Notes

Fats and oils are members of the class of compounds referred to as “lipids.” Fats and oils may be hydrolyzed by the action of a strong base. Hydrolysis of a fat or oil with NaOH results in the sodium salt of the fatty acids. Such salts are more commonly called “soaps,” and the reaction is commonly referred to as “saponification.” Hence, soaps are metallic salts of fatty acids. When common NaCl is added, the soap will precipitate. When soap is used in hard water, insoluble calcium salts of the fatty acid and other precipitates are formed. These precipitates are often referred to as “soap scum” or “bathtub ring.”

In Experiment 1, cottonseed oil will be hydrolyzed with NaOH to form a soap, and the properties of the soap will be investigated briefly.

A *polymer* is a large molecule formed by joining together repeating units of small molecules. The individual molecules forming the polymer are called “monomers.” The monomers link together like chains. Chemists usually represent a polymer with notation that shows the simplest repeating unit and the total number of repeating units (*n*): for example, $(C_6H_{12})_{n.}$ A typical polymer, Styrofoam, may have about 25,000 carbon atoms in its chain and would have a molar mass of about 350,000 grams. The molecules in a polymer can bend and twist and often tangle like a plate of spaghetti. Polymers can be made stiffer by introducing chemical bonds between the polymer chains, a process called “crosslinking.” The greater the number of crosslinks in a polymer, the more rigid the material will be.

Chromatography uses differences in the molecular structure and polarity of molecules to separate the components of a mixture. Chromatography uses a stationary phase and a mobile phase. Some compounds will be attracted to the stationary phase and will not travel or spread very much. Other compounds will be attracted to the mobile phase and will be carried along with it. Whether a compound is attracted to the stationary phase or to the mobile phase depends on the intermolecular attractions between the compound and the different phases.

In this experiment, students will use inks made of water-soluble, colored pigments. The cellulose of the filter paper is the stationary phase, and water and isopropyl alcohol are the mobile phases. As the mobile phase travels up the paper, it carries the different pigments along with it. Different-colored pigments are carried along at different rates, some traveling farther and faster than others. The speed of the pigment depends on the size of the pigment molecule and on how strongly the pigment is attracted to the paper. Since the mobile phase carries the different pigments at different rates, the colored inks will separate to reveal the colors that were mixed to make them.

Procedure

Experiment 1: Saponification (Soap Making)

1. Obtain a plastic pipette, the tip of which has been cut to approximately 1 inch in length.
2. On a watch glass, combine 15 drops of cottonseed oil, 10 drops of ethanol, and 10 drops of 20% NaOH solution. Stir the reagents to mix completely.
3. Carefully suck the mixture on the watch glass into the cut-off pipette. Invert the pipette, and place it into a small beaker. Allow the pipette to stand for a few minutes. After this time, the liquid in the pipette will have separated into two layers — a yellow oil layer and a colorless water layer.
4. Put a 250-mL beaker half full of water on a ring stand, and heat the water to about 85–90°C. This water will be used to heat the oil/NaOH/ethanol mixture gently.
5. Transfer the pipette containing the mixture, *stem end up*, to the hot water bath. Heat the mixture in the water bath, occasionally shaking the pipette, until the two layers coalesce into a single layer. This typically takes 5–10 minutes, depending on the temperature of the water and the amount of shaking done.
6. Obtain approximately 25 mL of a saturated NaCl solution. When the mixture in the pipette has formed a single layer, remove the pipette from the water bath, and squirt its contents into the saturated NaCl solution. Stir the mixture. This should cause a small quantity of white soap to form at the surface of the solution.
7. Using a clean scoop, skim the flakes of soap from the surface of the NaCl solution, and transfer them to a paper towel. Allow the soap to dry briefly on the paper towel.
8. Place about 100 mL of distilled water in a clean Erlenmeyer flask. Scrape the soap flakes from the paper towel into the water in the flask. Stopper the flask, and shake to dissolve the soap. Record what happens regarding an important property of the soap.

9. Add 2 drops of vegetable oil to a test tube containing about 5 mL of water. Note that the oil forms a separate layer. Stopper the test tube and shake. Allow the test tube to stand for a few minutes: the water and oil will separate into layers again. Repeat the test, using 5 mL of your soap solution in place of the water. The soap allows an emulsion to form between oil and water. Record your observations.
10. Place about 5 mL of your soap solution in a test tube. Add 10–15 drops of 1% $CaCl_2$ solution to the soap solution. (Ca^{2+} is one of the ions found in hard water; it forms a precipitate [soap scum] with common soaps. Mg^{2+} ions also result in hard water.) Record your observations.

Experiment 2: Polymerization (Formation of a Polymer)

Procedure A: Slime

1. Mix 10 mL of 4% polyvinyl alcohol solution with 2 drops of food coloring in a small plastic medicine cup. Stir to mix completely.
2. Add 2 mL of 4% sodium tetraborate solution *slowly* while stirring vigorously with a wooden splint. Stir until the mixture has "gelled."
3. Take the mixture out of the container, and test some of its properties. (The more crosslinking chemical [sodium tetraborate] you add, the stiffer the polymer will be.)

Procedure B: Silly Putty

1. Combine 5 mL of white glue and 2 drops of food coloring in a plastic cup. Stir to mix completely.
2. Add an equal amount of 4% sodium tetraborate solution *slowly* while stirring vigorously with a wooden splint. Stir the mixture until it has begun to thicken.
3. Wash the lump under running water while squeezing it with your hands to remove any excess glue. Keep working the lump under running water until it is clean and smooth.
4. Test some of the properties of this polymer.

Experiment 3: Creative Chromatography

In this experiment, you will make designs with different colored markers on a piece of filter paper. The colors will spread and overlap as they react with the different mobile phases. Experimenting with different mobile phases will show which gives a better separation of pigments. You may use pure water or pure isopropyl alcohol, or you may make a mixture of each in any ratio that you wish.

1. Take a piece of filter paper, and draw a design on it (dots, lines, squiggles, swirls, or whatever you want), using a variety of colored markers.
2. Place the filter paper on top of a beaker to catch any excess liquid that is not absorbed by the filter paper.
3. Use a pipette to drop your solvent of choice slowly, drop by drop, in the middle of the filter paper. Observe what happens to your design as the solvent moves outward from the center of the filter paper.
4. Repeat your experiment several times with other designs and other mixtures of mobile phases. Feel free to be creative. You may experiment with dropping the solvent in locations other than the middle of the paper. Do not use too much solvent, as excess solvent will tend to wash away your design.
5. Explain which mobile phases you decided to use and the reasons for your choices.

Observations and Conclusions

1. Explain how soap allows an emulsion to form between oil and water.
2. Explain how the Ca^{2+} ions react with the soap to form soap scum.
3. Describe some of the common properties of the "slime" and the Silly Putty.
4. Which mobile phase did you decide to use for the first chromatography lab? Why?
5. Explain some differences between water and isopropyl alcohol that led to different interactions with the inks.

Follow-up/extension

- Have students research advanced chromatographic procedures and the uses of these procedures in industry.
- Have students research the advanced chemical reactions that occur during saponification.

Organizing Topic — Acids, Bases, and Electrolytes

Standards of Learning

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include

a) designated laboratory techniques;
b) safe use of chemicals and equipment; and
c) proper response to emergency situations.

CH.4 The student will investigate and understand that quantities in a chemical reaction are based on molar relationships. Key concepts include

g) acid/base theory: strong electrolytes, weak electrolytes, and nonelectrolytes; dissociation and ionization; pH and pOH; and the titration process.

Essential Understandings, Knowledge, and Skills	Correlation to Textbooks and Other Instructional Materials
The student will use hands-on investigations, problem solving activities, scientific communication, and scientific reasoning to	
• recognize that acids and bases are defined by several theories;	____________________
• explain the emergence of modern theories based on historical development;	____________________
• state the characteristics of acids and bases based on the	
° Arrhenius Theory;	____________________
° Bronsted-Lowry Theory;	____________________
• molarity equals moles/dm^3 or moles/L of solution, and [ ] refers to molar concentration;	____________________
• define pH and pOH;	____________________
• understand the relationship between pH and pOH;	____________________
• explain that strong electrolytes dissociate completely and weak electrolytes dissociate partially;	____________________
• utilize acid-base titration and pH indicators in the laboratory;	____________________
• recognize neutralization reactions;	____________________
• understand and demonstrate ° MSDS warnings; ° safety rules for science; ° laboratory safety cautions; ° safe techniques and procedures.	____________________

A Study of Acids and Bases

Organizing Topic Acids, Bases, and Electrolytes

Overview Students explore some of the properties and reactions of acid and bases.

Related Standards of Learning CH.1a, b, c; CH.3e

Objectives

The students will
- understand the reactions between acids and bases;
- understand and demonstrate
  - ° MSDS warnings
  - ° safety rules for science
  - ° laboratory safety cautions
  - ° safe techniques and procedures.

Materials needed

- 10-mL glass transfer pipettes and pipette bulbs
- 1.0 *M* solutions of HCl and of NaOH (or prepared 0.1 *M* solutions of each)
- Water
- 100-mL volumetric flasks
- 3-mL plastic syringes
- 50-mL beakers
- Phenolphthalein indicator
- Clean, dry evaporating dishes
- Heat source
- Scales
- Solution of HCl with an unknown molarity between 0.05 *M* and 0.1 *M*
- 4 acid and/or base solutions with a variety of pH values (For use of universal indicator, the pH values 4, 6, 7, 8, and 9 work best.)
- 24-well plates
- Universal indicator
- Indicator color chart
- Goggles
- Vinegar solution

Instructional activity

Content/Teacher Notes

In this activity, students will be exposed to the fundamental reactions of acid and bases, as well as the calculations necessary to determine pH, pOH, $[H^+]$, and $[OH^-]$. The first part of the lab requires the use of 100-mL volumetric flasks and 10-mL glass transfer pipettes and pipette bulbs. If you do not have these in lab quantities, you can do this part as a teacher demonstration or omit this step and prepare the 0.1 *M* solutions for the students.

Properties of Acids:
- Aqueous solutions of acids have a sour taste.
- Acids change the color of acid-base indicators.

- Acids react with active metals to release hydrogen: $Zn(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2(g)$.
- Acids react with bases to produce salts and water (neutralization): $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$.
- Aqueous solutions of acids conduct electric current (they are electrolytes).

Properties of Bases:
- Aqueous solutions of bases have a bitter taste.
- Bases change the color of acid-base indicators.
- Dilute aqueous solutions of bases feel slippery.
- Bases react with acids to produce salts and water.
- Aqueous solutions of bases conduct electric current (they are electrolytes).

Arrhenius Acids and Bases:
- Svante Arrhenius, Swedish chemist (1859–1927)
- Arrhenius Acid — a chemical compound that increases the concentration of hydrogen ions, H^+, in aqueous solution
- Arrhenius Base — a substance that increases the concentration of hydroxide ions, OH^-, in aqueous solution

Aqueous Solutions of Acids:
- Acids are molecular compounds that ionize in solution
  - $HNO_3 \rightarrow H^+ + NO_3^-$
  - $HCl \rightarrow H^+ + Cl^-$
  - $H_2SO_4 \rightarrow H^+ + SO_4^{2-}$

Strength of Acids:
- Strong acids ionize completely (100%) in solution and are strong electrolytes. Examples of strong acids are sulfuric acid (H_2SO_4), hydrochloric acid (HCl), and nitric acid (HNO_3).
- Weak acids ionize only slightly (<1%) and are weak electrolytes. An example of a weak acid is acetic acid ($HC_2H_3O_2$).

Bronsted-Lowry Acids and Bases:
- Bronsted-Lowry Acid — a molecule or ion that is a proton donor (a proton is a hydrogen ion H^+)
- Bronsted-Lowry Base — a molecule or ion that is a proton acceptor
- Bronsted-Lowry Acid-Base Reaction — a reaction in which protons are transferred from the acid to the base

Amphoteric Compounds:
- Any species that can react as either an acid or a base
  - Water acts as both an acid and a base: $H_2O + H_2O \rightarrow H_3O^+ + OH^-$.
  - H_3O+ is called the "hydronium ion" and is what actually exists in an acidic solution. There are no free H^+ ions.

Strong Acid-Strong Base Neutralization:
- Neutralization — reaction of an acid with a base to produce water and a salt, such as the reaction of hydronium ions and hydroxide ions to form water molecules, or as $KOH(aq) + HNO_3(aq) \rightarrow KNO_3(aq) + H_2O(l)$.

Hydronium Ions and Hydroxide Ions:
- Self-ionization of water: $H_2O(l) + H_2O(l) \rightarrow H_3O^+(aq) + OH^-(aq)$
- Neutral, acidic, and basic solutions
  - Neutral $[H_3O^+] = [OH^-]$ (pH = 7)
  - Acidic $[H_3O^+] > [OH^-]$ (pH < 7)

  - Basic $[H_3O^+] < [OH^-]$ (pH > 7)

pH Calculations:

- $pH = -\log [H^+]$ and $[H^+] = 10^{-pH}$
- $pOH = -\log [OH^-]$ and $[OH^-] = 10^{-pOH}$
- $[H^+] = [OH^-] = 1 \times 10^{-7}$ (in a neutral solution)
- $K_W = [OH^-] [H^+] = 1 \times 10^{-14}$
- pH + pOH = 14

Introduction

1. Explain to students that in this activity, they will study some of the properties and reactions of acids and bases by conducting the following four experiments:
   - Preparing a dilute solution from a more concentrated one
   - Performing a neutralization reaction
   - Analyzing a solution by the technique of titration
   - Determining the pH, pOH, $[H^+]$ and $[OH^-]$ of acid or base solutions

Procedure

Experiment 1: Preparing a Dilute Solution from a More Concentrated One

1. Prepare the acid and base solutions as follows:
   - Pipette 10 mL of 1.0 *M* HCl into a clean, dry 100-mL volumetric flask. Add water to the mark on the volumetric flask. Label the flask.
   - Repeat the process, using the 1.0 *M* NaOH.
2. Calculate the new molarity of the acid and base solutions.
3. Calculate the pH and pOH for each of the prepared solutions.

Experiment 2: Neutralization Reaction

1. Obtain 2 plastic 3-mL syringes, and label one "Acid" and the other "Base."
2. Use the Acid syringe to pull up 3 mL of the prepared acid solution and place it in a clean, dry 50-mL beaker. Add 2 drops of phenolphthalein indicator.
3. Use the Base syringe to pull up 3 mL of the prepared base solution and slowly add it drop by drop to the beaker containing the acid solution. Swirl after each addition. Keep adding the base until the solution turns a very pale pink and the color remains for at least 30 seconds.
4. Transfer the solution to a pre-weighed, clean, dry evaporating dish, and heat gently to remove the liquid. Do not allow the solid to splatter.
5. When the solid is dry and the dish is cool, weigh the dish, and product and record the mass of dish and solid. Subtract to find the mass of the solid.
6. Write a balanced equation for the reaction that took place between the HCl and the NaOH.
7. Calculate the theoretical amount of solid that could be produced from the 3 mL of acid used in the reaction.
8. Weigh and record the amount of NaCl that you produced in the lab. Calculate your percent yield in the experiment.

Experiment 3: Analysis of an Unknown Acid — Titration

1. Obtain 3 mL of the unknown acid solution, and place it in a clean, dry 50-mL beaker.
2. Add 2 drops of phenolphthalein indicator.

3. Fill a syringe with your prepared base solution, and record the initial volume of base solution in the syringe.
4. Slowly add the base solution in the syringe, drop by drop, to the acid solution in the beaker, swirling after each addition. Keep adding the base solution until the acid solution turns a very pale pink. Record the final volume of the base solution remaining in the syringe.
5. Calculate the ΔV for base solution in the syringe.
6. Using the volume and molarity of the known base solution, calculate the molarity of the unknown acid solution, using this formula $M_A V_A = M_B V_B$

Experiment 4: Determination of pH, pOH, $[H^+]$, and $[OH^-]$

1. Place several drops of each of the labeled, unknown acid and/or base solutions in a well plate. Add 1 drop of universal indicator to each well. From the color of the indicator and using the indicator color chart, determine the pH of each solution.
2. Calculate the pH, pOH, $[H^+]$, and the $[OH^-]$ for each unknown acid and/or base solution, and record.

Observations and Conclusions

1. Use the student answers to the calculations within each experiment to lead discussions that will prompt correct conclusions about the behavior of acids and bases.

Sample assessment

- Ask the students to analyze the molarity of a vinegar solution, using the technique of titration that they used in the lab.

Follow-up/extension

- Have students research the following common industrial acids:
  - sulfuric acid
  - nitric acid
  - phosphoric acid
  - hydrochloric acid
  - acetic acid.

Acid-Base Theory

Organizing Topic Acids, Bases, and Electrolytes

Overview Students distinguish between acids and bases by means of physical and chemical properties.

Related Standards of Learning CH.1; CH.4g

Objectives

The students will

- recognize that acids and bases are defined by several theories;
- identify Arrhenius acids and Arrhenius bases;
- explain pH;
- read pH scale;
- calculate pH for an acidic or basic solution, and calculate the concentration of hydronium ions (H_3O^+) and hydroxide ions (OH^-) from pH;
- understand the relationship between pH and pOH;
- utilize acid-base titration and pH indicators in the laboratory.

Materials needed

- 1 lemon
- Bottle of household cleaner containing ammonium hydroxide
- Bottle of vinegar
- Bar of soap
- Roll of antacid tablets
- Litmus paper
- 6 acid or base solutions
- Tape
- Sodium hydroxide (NaOH) pellets
- Water
- Balances
- Glass stirring rods
- Corks or stoppers (wax paper and rubber bands) for flasks
- Phenolphthalein indicator
- 500-mL beakers or Erlenmeyer flasks
- 0.1 *M* HCl solution
- 0.1 *M* NaOH solution (2 g of NaOH dissolved in 500 mL of water)
- Medicine droppers
- 10-mL graduated cylinders (or test tubes)
- Stoppers for graduated cylinders (or test tubes)
- Test tube rack (or cardboard)
- Ruler

Instructional activity

Content/Teacher Notes

This lesson may require more than one class period or block.

Before beginning this lab, prepare a 0.1 molar solution of HCl from bottled hydrochloric acid. (If you purchase bottles of 1 molar HCl, dilute the solution with water 10:1.) Also, prepare a 0.1 molar solution of the base NaOH. (The atomic masses of Na, O, and H are 23, 16, and 1, respectively. One mole of NaOH is therefore 40 grams of NaOH. 0.05 of 40 g is 2 g. Mixing 2 g of NaOH in 500 mL of water yields a 0.1 molar solution of NaOH.)

Phenolphthalein, used as a base indicator in this lab, is available in solution through any laboratory supply house. Phenolphthalein is colorless below pH 7 and purple above pH 7. When titrating a base with an acid, it is convenient to stain the base with phenolphthalein so that the neutralization of the base is evident when the color disappears.

Introduction

1. Display a lemon, a bottle of household cleaner containing ammonium hydroxide, a bottle of vinegar, a bar of soap, and a roll of antacid tablets. Point out that the lemon contains an acid called citric acid. The presence of citric acid gives the family of citrus fruits their name. Another acid found in citrus fruits is ascorbic acid, commonly called “vitamin C.”
2. Emphasize that acids are common in our everyday lives. We even have acid in our stomachs — a powerful solution of hydrochloric acid — to help us digest foods.
3. Point out that the cleaner, the soap, and the antacid tablets contain substances called bases. Bases have a variety of uses but are mostly used to neutralize the acids employed in many different industrial and chemical procedures.
4. Write the following chart on the board, and ask students to compare the chemical formulas of the acids and bases listed:

Acid	Base
HCl – hydrochloric acid	NaOH – sodium hydroxide
H_2SO_4 – sulfuric acid	$Mg(OH)_2$ – magnesium hydroxide
H_2CO_3 – carbonic acid	NH_4OH – ammonium hydroxide
HNO_3 – nitric acid	KOH – potassium hydroxide

   Ask: “How are the acid formulas the same?” *(Each formula has an “H” for hydrogen at the front of it.)* All acids contribute a hydrogen ion (H^+) to water when in solution. Ask: “How are the base formulas the same?” *(Each formula has an OH for hydroxide at the end of it.)* All bases contribute a hydroxide ion (OH^-) to water when in solution.

5. Draw a KWL chart on the board, and have the class fill in the first two columns (“Know,” “Want to Know”) while comparing and contrasting the physical properties of acids and bases. Have the students copy the chart into their notebooks.
6. Ask a volunteer to bite into the lemon (or do it yourself) to demonstrate the meaning of the term sour. Something that taste sour makes us pucker our lips. Ask students what they would do if they bit into a bar of soap. Soap has a bitter taste that makes us want to expectorate.

Procedure

Pre-Lab: Preparing a Basic Solution

1. Inform students that in this lab, they will test for the presence of acids and bases and use an acid to neutralize a base. In order to do this, they will first prepare for use a 0.1 molar solution of a common base, sodium hydroxide (NaOH).

Chemical Symbol or Formula	**Atomic Mass (1 Mole)**
Na	
O	
H	
NaOH	
1 mole of NaOH × 0.05 = ____ grams of NaOH + 500 mL water = 0.1 M NaOH solution	

2. Assist students in preparing their own flask of a 0.1 M NaOH solution according to the following steps:
   a. Use the periodic table to find the atomic mass of sodium, oxygen, and hydrogen, and enter each into a table like the one shown at right. The atomic mass in grams of each of these elements is equal to 1 mole of the element.
   b. Add these three masses to get the mass of 1 mole of sodium hydroxide (NaOH).
   c. Use a balance to measure out 1/20 (0.05) of a mole of NaOH.
   d. Fill a 500-mL beaker or Erlenmeyer flask with water.
   e. Add the 0.05 mole of NaOH to the 500 mL of water. Since a 1 molar solution of any substance is the molecular weight of that substance in 1,000 mL of water, this mixture will be a 0.1 molar solution of NaOH.
   f. Stir gently until the NaOH is completely dissolved in the solution. Use wax paper and a rubber band (or a cork or rubber stopper) to cover your sodium hydroxide solution.
   g. Store the sodium hydroxide solution at room temperature for use later. Remind students that because the solution will be used at a later time, the flask must be labeled before storing.

Exercise: Calculating the pH of Acidic and Basic Solutions

1. Review the physical and chemical properties of acids and bases.
2. Explain that litmus is a dye obtained from lichen — i.e., a symbiotic community of algae and fungi that grows on rocks. Chemically, it is a mixture of several organic compounds (mostly carbon, hydrogen, and oxygen) that changes color in the presence of an acid or a base. The color created by a drop of acidic or basic solution on litmus paper reflects the concentration of hydrogen or hydroxide ions present in the solution. Litmus paper turns red in the presence of an acid and blue in the presence of a base.
3. Draw the pH scale on the board, and compare it to the color legend on the litmus paper roll. Explain how pH is the negative logarithm of hydrogen ion concentration.
4. A short review of logarithms may be useful for students. Point out that the logarithm of a number is the power to which 10 must be raised in order to equal the number. Have students create a pH scale based on the logarithms.
5. After the review of logarithms, have students practice calculating pH for an acidic or basic solution and calculating the concentration of hydronium ions (H_3O^+) and hydroxide ions (OH^-) from pH.

Experiment 1: Measuring the pH of Acids and Bases

1. Distribute six litmus paper strips to each student, and have them label the six strips “Test A” through “Test F.” Have the students attach each strip to a sheet of paper in their lab book and next to each strip, write: “Solution: ________________/pH___________.”
2. Have the students test a single drop of six solution samples, using a medicine dropper.
3. Instruct the students to compare the color created by the sample drop against the color legend on the litmus paper dispenser, and record the approximate pH of each solution. Have them use tape to secure the samples to the page, making sure to label each solution used.
4. Have the students answer the following questions:
   - What ion is characteristic of aqueous solutions of all acids? *(The hydronium ion (H_3O^+) is characteristic of all acids, while the hydroxide ion is characteristic of all bases.)*

- What is the difference between a strong acid and a weak acid? *(A strong acid is completely ionized; a weak acid is only partially ionized.)*
- What are three characteristics of an acid solution? *(It tastes sour, it conducts electricity, and it reacts with metals above hydrogen in the activity series.)*
- What is the equation for the reaction of a strong acid, HBr, with water? ($HBr + H_2O \rightarrow H_3O^{+} + Br^{-}$)

Demonstration: Titration

1. Begin the discussion by asking students what they commonly do when they have an upset stomach. What do they take to help it feel better? Explain that antacids contain hydroxides that neutralize acid like the HCl in stomach acid. This can be explained by the following general chemical equation:

$$\underset{\text{acid}}{HX} + \underset{\text{base}}{YOH} \rightarrow \underset{\text{salt}}{XY} + \underset{\text{water}}{HOH}$$

2. Draw the following chart on the board, one row at a time, to show how a variety of common acids and bases are neutralized to form salt and water.

Acid +	**Base** →	**Salt** +	**Water**
HCl Stomach acid	NaOH Lye for soap	NaCl Table salt	HOH H_2O = water
H_2CO_3 Soda pop acid	NaOH Lye for soap	Na_2CO_3 Washing soda	HOH H_2O = water
H_2SO_4 Industrial acid	$Mg(OH)_2$ Milk of Magnesia	$MgSO_4$ Epsom salt	HOH H_2O = water
H_2CO_3 Soda pop acid	$Ca(OH)_2$ Antacid tablets	$CaCO_3$ Limestone chalk	HOH H_2O = water
HNO_3 Fertilizer	KOH Potash	KNO_3 Saltpeter	HOH H_2O = water

3. Have the students carefully observe as you add 100 mL of 0.1 *M* NaOH solution to an "empty" beaker that actually contains a drop of phenolphthalein placed there without the students' knowledge. Keep the identity of the clear liquid secret. Seeing the clear liquid turn purple as you pour it into the "empty" beaker will cause some surprise.
4. Then, display the purple fluid and pour a few drops of it into another clear, unidentified liquid — 100 mL of 0.1 *M* HCl (or slightly more concentrated) solution. This will also raise eyebrows when the liquid remains clear.
5. Pour the HCl into the NaOH until the color disappears. Now, identify all of your solutions, and explain that the purple fluid, phenolphthalein, can indicate the presence of a base only when there is base present. When the acid neutralizes the base to form salt and water, the base is no longer present, and the phenolphthalein becomes colorless.

Experiment 2: Titration

1. Obtain a 10-mL graduated cylinder. (If one is not available, the alternative is to use a test tube rack or a piece of cardboard with a hole cut in it to hold a test tube upright in a beaker.)

2. Retrieve your flask of previously prepared 0.1 *M* NaOH solution, and pour 3 mL of it into the 10-mL graduated cylinder. (Alternatively, use a ruler to measure 3 cm of the solution as you pour it into the test tube.)
3. Add one drop of phenolphthalein indicator to your 3mL of 0.1 *M* NaOH solution. The solution will turn purple, indicating the presence of a base.
4. Obtain 20 mL of a 0.1 *M* HCl solution from your teacher, and add a dropper full of the HCl solution to the base.
5. Use a small dropper or cork stopper to cap the tube or cylinder and turn it upside down, then right side up. This simple action will mix the solution thoroughly.
6. Repeat steps 7 and 8 until the base solution clears completely.
7. Determine the amount of acid that was needed to neutralize the base by finding the final amount of solution in the graduated cylinder and subtracting from it the original amount. (Alternatively, use the ruler to measure the final amount of solution, and subtract the original amount from it.)
8. Graph the gathered data, and describe the composition of the titration mixture represented by the graph.

Sample assessment

- Use the students' lab reports for assessment.
- Have the students write a step-by-step description of the process of titration.

Follow-up/extension

- Have students read the list of ingredients in five common acidic household substances. Have them create a table listing those five common substances and the acids they contain.
- Have students read the list of ingredients in five common basic household substances. Have them create a table listing those five common substances and the bases they contain.
- Have students explore how certain flowers and shrubs thrive (or change color) in acidic vs. basic soils.

Resources

- *Acid-Base Tea Party: A Physical Science Activity.* http://galileo.phys.virginia.edu/Education/outreach/8thgradesol/AcidBaseFrm.htm.
- *Flower Indicators: A Physical Science Activity.* http://galileo.phys.virginia.edu/Education/outreach/8thgradesol/FlowerIndicatorsFrm.htm.