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Catalog No. AP6467

Publication No. 6467

Colorful Electrolysis

Chemical Demonstration Kit

Introduction

Demonstrate simple electrolysis in a very colorful and dramatic way on an overhead projector.

Concepts

- Electrolysis
- Cathode
- Anode
- pH

Materials (for each demonstration)

Sodium sulfate solution, Na_2SO_4, 0.5 M, 40 mL
Universal indicator solution, 10 mL
Water, distilled
Battery clip with alligator ends
Battery, 9-V
Beaker, 100-mL
Graduated cylinder, 10 mL
Graduated cylinder, 50 mL
Overhead projector
Pencil leads
Petri dish, disposable
Stirring rod

Safety Precautions

To extend the life of the battery, avoid touching the positive and negative terminals to each other. Universal indicator is an alcohol-based solution and is flammable; do not use near an open flame. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information. Wash hands thoroughly with soap and water before leaving the laboratory.

Preparation

To prepare the demonstration solution, add 10 mL of universal indicator to 40 mL of 0.5 M sodium sulfate solution and stir.

Procedure

1. Place the two halves of a Petri dish on the projection stage of an overhead projector.
2. Pour enough demonstration solution into each half of the Petri dish to just cover the bottom of each half dish. Adjust the overhead so that the dishes are in clear focus. Each half dish should appear to be a rich, transparent green color.
3. Break a pencil lead in half. Attach the leads to opposite sides of the Petri dish with the alligator clips. Make sure the tip of each lead is submerged in the green solution and the alligator clips remain out of the solution.

4. To start the demonstration, clip the 9-volt battery into the snaps on the battery clip. See Figure 1.

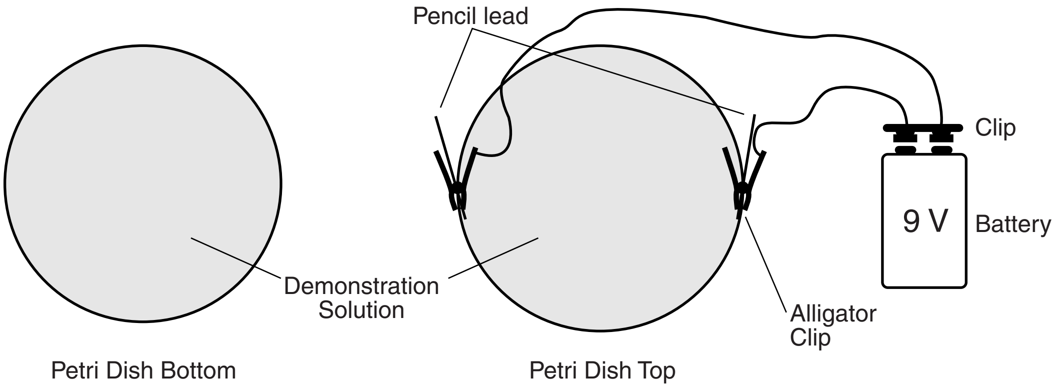

Figure 1. Demonstration Setup.

5. Let the demonstration run for 5–10 minutes and note the changing colors over time. (A purple color will appear at the cathode very quickly. An orange color at the anode will appear more slowly. Over time, the entire spectrum of universal indicator colors will appear.)
6. Discuss the results as the demonstration continues. Discuss the various colors as well as why the "extra" dish was included in the demonstration.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures governing the disposal of laboratory waste. All materials may be disposed of according to Flinn Suggested Disposal Method #26b.

Tips

- The reaction at the anode and cathode seem to proceed at the same pace. There is bubbling at both the anode and the cathode. A deep purple forms immediately at the cathode and an obvious red/orange at the anode. Both of these color changes occur with similar intensity. Purple, blue, green, yellow, and red/orange are all visible after five minutes.
- The demonstration can be repeated using a pH 7 buffer to show the effects of buffering a solution. (When the solution is buffered, no color changes result.)
- The sulfate ion is an extremely weak base (the pK_a for its conjugate acid, HSO_4^-, is 2.0). The initial indicator color for the electrolysis solution may be more blue-green rather than green. According to the *Merck Index,* the pH of a sodium sulfate solution is 6.0–7.6. Test a small amount of the sodium sulfate stock solution with bromthymol blue indicator before class—the solution should turn green. If the solution is blue, add one drop of 1 M hydrochloric acid to the stock solution. If the solution is yellow, add one drop of 1 M sodium hydroxide to the stock solution.
- The odor of chlorine may be observed at the anode if sodium chloride is substituted for sodium sulfate as the electrolyte in the electrolysis solution. Based on their standard reduction potentials, oxidation of chloride ion to chlorine ($E° = –1.36$ V) is less favorable than oxidation of water to oxygen ($E° = –1.23$ V). However, there is a significant overvoltage for oxidation of water, and oxidation of chloride competes with oxidation of water under typical electrolysis conditions. Although the cause of the overvoltage is poorly understood, it is generally believed to be due to a kinetically slow reaction at the anode.

Discussion

When an electric current is passed through an aqueous solution containing an electrolyte (Na_2SO_4), the water molecules break apart or decompose into their constituent elements, hydrogen and oxygen. The overall reaction occurs as two separate, independent half-reactions. Reduction of the hydrogen atoms to elemental hydrogen (H_2) occurs at the cathode (–), while oxidation of the oxygen atoms in water to elemental oxygen (O_2) occurs at the anode (+). Each half-reaction is accompanied by the production of OH^- or H^+ ions as shown below:

$$\text{Cathode: } 4e^- + 4H_2O \rightarrow 2H_2(g) + 4OH^-$$

$$\text{Anode: } 2H_2O \rightarrow O_2(g) + 4H^+ + 4e^-$$

At the cathode the excess OH^- will cause the pH to increase, resulting in a color change of the universal indicator solution from green (neutral, pH 7) to purple (basic, $pH \geq 10$).

At the anode the excess H^+ will cause the pH to decrease, resulting in a color change of the universal indicator solution from green to an orange/red color (acidic, $pH \leq 4$). The electrolysis half-reactions can also be followed by observing the production of gas bubbles at the cathode (H_2) and anode (O_2).

Universal indicator is an acid–base indicator that is different colors at different pH values. All colors will be visible in the Petri dish as electrolysis progresses and as the pH conditions continually change due to diffusion and neutralization.

pH	Color
4	red
5	orange
6	yellow
7	green
8	blue/green
9	dark blue
10	purple

Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

Unifying Concepts and Processes: Grades K–12
Systems, order, and organization
Evidence, models, and explanation
Content Standards: Grades 5–8
Content Standard A: Science as Inquiry
Content Standard B: Physical Science, properties and changes of properties in matter, transfer of energy
Content Standards: Grades 9–12
Content Standard A: Science as Inquiry
Content Standard B: Physical Science, structure and properties of matter, chemical reactions, interactions of energy and matter

Acknowledgment

Special thanks to Mike Shaw, West Stokes High School, King, NC for this demonstration idea.

The *Colorful Electrolysis Demonstration Kit* is available from Flinn Scientific, Inc.

Catalog No.	Description
AP6467	Colorful Electrolysis Demonstration Kit

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.

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Publication No. 10261

Upset Tummy? MOM to the Rescue!

A Colorful Antacid Demonstration

Introduction

Mix milk of magnesia (MOM) with universal indicator and observe the dramatic rainbow of colors as the antacid dissolves in the simulated stomach acid! This is a great demonstration to teach concepts of acids and bases, solubility, K_{sp} and "antacid-testing" consumer chemistry.

Concepts

- Acid–base neutralization
- Solubility and K_{sp}
- Antacids

Materials

Hydrochloric acid, HCl, 3 M, approximately 20 mL
Milk of magnesia, 20 mL
Universal indicator solution, 1%, 4–5 mL
Water, distilled or deionized, 800 mL
Beaker, 1-L (or other large beaker)
Beral-type pipets, 2
Graduated cylinder, 25-mL or 50-mL
Ice, crushed (or ice cubes)
Magnetic stir bar
Magnetic stir plate (or stirring rod)

Safety Precautions

Milk of magnesia is intended for laboratory use only; it has been stored with other non–food-grade laboratory chemicals and is not meant for human consumption. Hydrochloric acid solution is toxic by ingestion and inhalation and is corrosive to skin and eyes. Universal indicator solution is an alcohol-based flammable solution. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Please review current Safety Data Sheets for additional safety, handling, and disposal information.

Procedure

1. Measure 20 mL of milk of magnesia using a graduated cylinder and pour it into a 1-L beaker.
2. Place the 1-L beaker on a magnetic stir plate. Add a magnetic stir bar to the beaker.
3. Add water and crushed ice (or ice cubes) to give a total volume of approximately 800 mL. Turn on the stir plate so as to create a vortex in the mixture.
4. Add about 4–5 mL (about 2 pipets full) of universal indicator solution. Watch as the white suspension of milk of magnesia turns to a deep purple color. The color indicates that the solution is basic.
5. Add 2–3 mL (1 pipet full) of 3 M HCl. The mixture quickly turns red and then goes through the entire universal indicator color range back to purple.
6. Repeat this process, adding HCl one pipet full at a time, waiting after each addition until the mixture turns back to blue–purple.
7. The process can be repeated a number of times before all of the $Mg(OH)_2$ has dissolved and has reacted with the HCl. As more acid is added, the color changes will occur more rapidly and eventually the suspension will be completely dissolved. This will be apparent by a clear, red solution.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. The final solution may be neutralized according to Flinn Suggested Disposal Method #24b. Excess milk of magnesia may be placed in the trash according to Flinn Suggested Disposal Method #26a.

Discussion

The active ingredient in milk of magnesia is magnesium hydroxide, $Mg(OH)_2$. Magnesium hydroxide forms a suspension in water since it has a very low solubility—0.0009 g/100 mL in cold water and 0.004 g/100 mL in hot water.

Initially in the demonstration, the solution is basic due to the small amount of $Mg(OH)_2$ that goes into solution. The universal indicator gives the entire solution a violet color, indicating a pH of about 10. (See Universal Indicator Color Chart below.) Upon adding hydrochloric acid (the simulated "stomach acid"), the mixture quickly turns red because the acid disperses throughout the beaker, first neutralizing the small amount of dissolved $Mg(OH)_2$, and then turning the solution acidic from the excess acid that is present.

Universal Indicator Color Chart

Color	**Red**	**Orange**	**Yellow**	**Green**	**Green–blue**	**Blue**	**Violet**
pH	4	5	6	7	8	9	10

The excess acid causes more $Mg(OH)_2$ from the suspension to gradually dissolve. As more of the $Mg(OH)_2$ goes into solution, the acid is neutralized and eventually the solution becomes basic again from the excess $Mg(OH)_2$ that is present. The addition of universal indicator allows this process to be observed. During the process, the color of the mixture goes through the entire universal indicator color range—from red to orange to yellow to green to blue and finally back to violet. By adding more "stomach acid," the process can be repeated several times before all of the $Mg(OH)_2$ is dissolved and eventually neutralized.

Connecting to the National Standards

This laboratory activity relates to the following National Science Education Standards (1996):

Unifying Concepts and Processes: Grades K–12
Systems, order, and organization
Evidence, models, and explanation
Constancy, change, and measurement
Content Standards: Grades 5–8
Content Standard A: Science as Inquiry
Content Standard B: Physical Science, properties and changes of properties in matter
Content Standards: Grades 9–12
Content Standard A: Science as Inquiry
Content Standard B: Physical Science, structure and properties of matter, chemical reactions
Content Standard C: Life Science; matter, energy, and organization in living systems

Acknowledgments

Special thanks to Annis Hapkiewicz, Okemos High School, Okemos, MI and to Penney Sconzo, Westminster School, Atlanta, GA for separately bringing this demonstration to our attention.

References

Summerlin, L. R.; Borgford, C. L.; Ealy, J. B. *Chemical Demonstrations: A Sourcebook for Teachers,* Vol. 2; American Chemical Society: Washington, DC. 1988; p 173.

Materials for the *Upset Tummy? MOM to the Rescue!* are available from Flinn Scientific, Inc.

Catalog No.	Description
H0034	Hydrochloric Acid Solution, 3 M, 500 mL
U0001	Universal Indicator Solution, 100 mL
AP5934	Upset Tummy? MOM to the Rescue!—Chemical Demonstration Kit
AP8184	Magnetic Stirrer, Flinn

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.

POGIL AP* Chemistry Activities

Chemical Reactions and Stoichiometry

- Mass Spectroscopy
- Empirical Formulas
- Combustion Analysis
- Net Ionic Equations

Atomic Structure and The Periodic Table

- Advanced Periodic Trends
- Photoelectron Spectroscopy

Bonding

- Types of Bonds
- Polar and Nonpolar Molecules
- Properties of Covalent Bonds
- Lattice Energy
- Types of Solids
- Alloys

Properties of Gases

- Partial Pressures of Gases
- Deviations from the Ideal Gas Law
- Maxwell-Boltzmann Distributions

Thermodynamics and Kinetics

- Heats of Formation
- Rates of Reaction
- Method of Initial Rates
- Free Energy
- Work, Equilibrium and Free Energy

Equilibrium Systems

- Reaction Quotient
- Common Ion Effect on Solubility
- Fractional Precipitation
- Common Ion Effect on Acid Ionization
- Buffers
- Strength of Acids
- Titration Curves
- Polyprotic Acids

Oxidation and Reduction

- Electrochemical Cell Voltage
- Faraday’s Law

POGIL Chemistry Activities

Introduction to Chemistry

- Safety First
- Fundamentals of Experimental Design
- Organizing Data
- Significant Digits and Measurement
- Significant Zeros
- Classification of Matter

Atomic and Electron Structure

- Isotopes
- Ions
- Average Atomic Mass
- Coulombic Attraction
- Electron Energy and Light
- Electron Configurations

The Periodic Table

- Cracking the Periodic Table Code
- Periodic Trends

Ionic and Molecular Compounds

- Naming Ionic Compounds
- Polyatomic Ions
- Naming Molecular Compounds
- Naming Acids
- Molecular Geometry

Chemical Reactions and Stoichiometry

- Types of Chemical Reactions
- Relative Mass and the Mole
- Mole Ratios
- Limiting and Excess Reactants

Properties of Gases

- Gas Variables

Solubility and Solutions

- Saturated and Unsaturated Solutions
- Solubility
- Molarity

Thermochemistry

- Calorimetry
- Bond Energy

Equilibrium

- Equilibrium

Acids and Bases

- Acids and Bases
- Strong versus Weak Acids
- Calculating pH

Oxidation and Reduction

- Oxidation and Reduction
- The Activity Series
- Batteries

Reaction Quotient

How do you predict which direction a reaction will proceed to reach equilibrium?

Why?

When a reaction reaches equilibrium there must be some non-negligible amount of every species in the reaction, otherwise the reaction cannot react in both directions. Knowing this, it is very easy to predict which direction a reaction will go to reach equilibrium when one of the components of the reaction has an initial concentration of zero. Many of the problems you have worked with thus far have some component at zero concentration, but real life does not work that way. Most of the time, the reaction in question has some measureable amount of every species. Deciding which way the reaction will go to reach equilibrium then becomes more challenging.

Model 1 – A Theoretical Equilibrium

Trial 1	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	1.000 M	1.000 M	1.000 M
Change	←		
Equilibrium	1.464 M	1.464 M	0.536 M

Trial 2	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	2.000 M	0.500 M	0.500 M
Change			
Equilibrium	2.150 M	0.650 M	0.350 M

Trial 3	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	1.000 M	0.500 M	1.500 M
Change			
Equilibrium	1.864 M	1.364 M	0.636 M

Trial 4	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	1.600 M	1.000 M	0.400 M
Change			
Equilibrium	1.600 M	1.000 M	0.400 M

Trial 5	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	1.400 M	1.200 M	0.400 M
Change			
Equilibrium	1.388 M	1.188 M	0.412 M

Trial 6	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	0.750 M	2.000 M	0.250 M
Change			
Equilibrium	0.675 M	1.925 M	0.325 M

1. Examine Model 1.
   a. Write the theoretical chemical reaction that is used in the trials of Model 1.

   b. If 0.50 M of reactant A reacts, predict the change in concentration of B and C.

2. What variables were changed in the different trials shown in Model 1?

3. In Trial 1 of Model 1 there is an arrow in the "change" section of the table.
    a. Explain what that arrow represents.

    b. What evidence is present in the table to indicate the direction the arrow should be pointing?

4. With your group, determine which direction each of the other trials in Model 1 reacted to reach equilibrium. Indicate that direction with an arrow in the "change" section of the table.
5. Is it true that there are equal concentrations of reactants and products when all of the reactions in Model 1 reach equilibrium? Justify your answer with evidence from Model 1.

6. According to Model 1, are the final concentrations of all species in the reaction the same when the reaction reaches equilibrium, regardless of the initial concentration?

7. Does the reaction in Model 1 always proceed in the forward direction when there are more reactants than products? Justify your answer with evidence from Model 1.

STOP

8. Write the equilibrium constant expression for the reaction in Model 1.

9. Discuss with your group how you could determine the equilibrium constant, K_{eq}, for the reaction in Model 1. Divide the work among group members. Use data from multiple trials to calculate the equilibrium constant for the reaction and determine the average. Show all work.

Read This!

The key to knowing which direction a reaction will need to proceed in order to reach equilibrium is knowing if you have too much reactant or too much product compared to the equilibrium state. Keep in mind, however, that there are many combinations of reactant and product concentrations that constitute an equilibrium state.

Model 2 – Comparing *Q* and K_{eq}

Trial	Reaction Quotient, *Q*	Equilibrium Constant, K_{eq}	*Q* versus K_{eq}	Direction to Equilibrium
1				
2				
3				
4				
5				
6				

10. Fill in the Equilibrium Constant column in Model 2 using data from Model 1.
11. Fill in the Direction to Equilibrium column in Model 2 using data from Model 1.

Read This!

The **reaction quotient** for a reaction is the ratio of products to reactants, similar to the equilibrium constant. The difference is you calculate the ratio with initial conditions.

$$K_{eq} = \frac{[C]_{eq}}{[A]_{eq}[B]_{eq}} \qquad Q = \frac{[C]_{initial}}{[A]_{initial}[B]_{initial}}$$

12. Calculate the reaction quotient for each of the trials in Model 1 and record the data in Model 2 in the appropriate column. Divide the work among group members. Show your work below.

STOP

13. Consider how the concentration values in the reaction quotient change when the reaction proceeds in the forward direction.

    a. Does the numerator increase or decrease?

    b. Does the denominator increase or decrease?

    c. Overall, does the Q ratio increase or decrease when the reaction proceeds in the forward direction?

14. Consider how the concentration values in the reaction quotient change when the reaction proceeds in the reverse direction.

    a. Does the numerator increase or decrease?

    b. Does the denominator increase or decrease?

    c. Overall does the Q ratio increase or decrease when the reaction proceeds in the forward direction?

15. Fill in the Q versus K_{eq} column in the table in Model 2. Write $Q > K_{eq}$, $Q < K_{eq}$ or $Q = K_{eq}$.

16. Complete the following statements.

    a. When the reaction quotient is **greater than** the equilibrium constant, the reaction proceeds more in (the forward, the reverse, neither) direction to reach equilibrium.

    b. When the reaction quotient is **less than** the equilibrium constant, the reaction proceeds more in (the forward, the reverse, neither) direction to reach equilibrium.

    c. When the reaction quotient is **equal to** the equilibrium constant, the reaction proceeds more in (the forward, the reverse, neither) direction to reach equilibrium.

17. Consider the following reaction.

$$2SO_2\,(g) \; + \; O_2\,(g) \;\longleftrightarrow\; 2SO_3\,(g) \qquad K_{eq} \; = \; 3900 \text{ at } 2000 \text{ K}$$

a. Write the equilibrium constant expression for the reaction.

b. Write the reaction quotient expression for the reaction.

c. A reaction vessel contains 0.150 M sulfur dioxide, 0.150 M oxygen and 2.000 M sulfur trioxide. Predict the direction the reaction must shift to reach equilibrium. Show a calculation to justify your answer.

18. Consider the following reaction.

$$2HI\,(g) \;\longleftrightarrow\; H_2\,(g) \; + \; I_2\,(g) \qquad K_{eq} \; = \; 0.25 \text{ at } 25\,°\text{C}$$

A reaction vessel contains 0.500 M hydrogen, 0.500 M iodine vapor and 0.750 M hydrogen iodide. Predict the direction the reaction must shift to react equilibrium. Show a calculation to justify your answer.

Extension Questions

19. Consider the following reaction.

$$Cd^{2+}(aq) + 4Cl^{1-}(aq) \longleftrightarrow CdCl_4^{2-}(aq) \qquad K_{eq} = 108 \text{ at } 25\,°C$$

a. Write the reaction quotient expression for the reaction.

b. A reaction vessel contains 0.100 M cadmium ion, 0.500 M chloride and 0.250 M tetrachlorocadmate ion. Predict the direction the reaction must shift to reach equilibrium. Show a calculation to justify your answer.

20. Consider the following reaction.

$$C(s) + H_2O(g) \longleftrightarrow H_2(g) + CO(g) \qquad K_{eq} = 4.09 \times 10^{-3} \text{ at } 2000 \text{ K}$$

a. A 1.000-L reaction vessel contains 4.00 M of water vapor, 0.040 M hydrogen and 0.040 M carbon monoxide. Calculate the reaction quotient and state which direction the quotient predicts the reaction must shift to reach equilibrium.

b. Discuss why the prediction above is wrong in this case.

21. LeChâtelier's principle states that an equilibrium system will shift in the direction that reduces the stress put on the system. Is this consistent with the predictions made by calculating the reaction quotient and comparing Q to the equilibrium constant? Provide an example to support your answer.

Reaction Quotient

How do you predict which direction a reaction will proceed to reach equilibrium?

Why?

When a reaction reaches equilibrium there must be some non-negligible amount of every species in the reaction, otherwise the reaction cannot react in both directions. Knowing this, it is very easy to predict which direction a reaction will go to reach equilibrium when one of the components of the reaction has an initial concentration of zero. Many of the problems you have worked with thus far have some component at zero concentration, but real life does not work that way. Most of the time, the reaction in question has some measureable amount of every species. Deciding which way the reaction will go to reach equilibrium then becomes more challenging.

Model 1 – A Theoretical Equilibrium

Trial 1	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	1.000 M	1.000 M	1.000 M
Change	⟵		
Equilibrium	1.464 M	1.464 M	0.536 M

Trial 2	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	2.000 M	0.500 M	0.500 M
Change	⇠		
Equilibrium	2.150 M	0.650 M	0.350 M

Trial 3	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	1.000 M	0.500 M	1.500 M
Change	⇠		
Equilibrium	1.864 M	1.364 M	0.636 M

Trial 4	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	1.600 M	1.000 M	0.400 M
Change	*No change*		
Equilibrium	1.600 M	1.000 M	0.400 M

Trial 5	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	1.400 M	1.200 M	0.400 M
Change	⇢		
Equilibrium	1.388 M	1.188 M	0.412 M

Trial 6	$A(g) + B(g) \longleftrightarrow C(g)$		
Initial	0.750 M	2.000 M	0.250 M
Change	⇢		
Equilibrium	0.675 M	1.925 M	0.325 M

1. Examine Model 1.
   a. Write the theoretical chemical reaction that is used in the trials of Model 1.

   $A(g) + B(g) \longleftrightarrow C(g)$

   b. If 0.50 M of reactant A reacts, predict the change in concentration of B and C.

   Since all of the coefficients in the reaction are one, the concentration of B will decrease by 0.50 M and the concentration of C will increase by 0.50 M.

2. What variables were changed in the different trials shown in Model 1?

   Each trial in Model 1 has different initial concentrations of the three species.

3. In Trial 1 of Model 1 there is an arrow in the "change" section of the table.

   a. Explain what that arrow represents.

   The arrow indicates the direction that the reaction shifts in order to reach equilibrium.

   b. What evidence is present in the table to indicate the direction the arrow should be pointing?

   The concentration of the product, C, decreased and the concentrations of the reactants, A and B, increased. That indicates the reaction shifted in the reverse direction.

4. With your group, determine which direction each of the other trials in Model 1 reacted to reach equilibrium. Indicate that direction with an arrow in the "change" section of the table.

   See Model 1.

5. Is it true that there are equal concentrations of reactants and products when all of the reactions in Model 1 reach equilibrium? Justify your answer with evidence from Model 1.

   No. Even if the reactant concentrations are added together, they do not equal the product concentrations in any of the trials.

6. According to Model 1, are the final concentrations of all species in the reaction the same when the reaction reaches equilibrium, regardless of the initial concentration?

   No, each of the trials in Model 1 has a different equilibrium concentration.

7. Does the reaction in Model 1 always proceed in the forward direction when there are more reactants than products? Justify your answer with evidence from Model 1.

   No. In Trial 2 of Model 1, there are far more reactants than products, and the reaction went in the reverse direction.

STOP

8. Write the equilibrium constant expression for the reaction in Model 1.

   $$K_{eq} = \frac{[C]}{[A][B]}$$

9. Discuss with your group how you could determine the equilibrium constant, K_{eq}, for the reaction in Model 1. Divide the work among group members. Use data from multiple trials to calculate the equilibrium constant for the reaction and determine the average. Show all work.

   The equilibrium data from each trial in Model 1 can be used to calculate the equilibrium constant.

   Trial 1

   $$K_{eq} = \frac{(0.536\ M)}{(1.464\ M)(1.464\ M)}$$

   $$K_{eq} = 0.250$$

   Trial 2

   $$K_{eq} = \frac{(0.350\ M)}{(2.150\ M)(0.650\ M)}$$

   $$K_{eq} = 0.250$$

   Trial 3

   $$K_{eq} = \frac{(0.636\ M)}{(1.864\ M)(1.364\ M)}$$

   $$K_{eq} = 0.250$$

   Average = *0.250*

Read This!

The key to knowing which direction a reaction will need to proceed in order to reach equilibrium is knowing if you have too much reactant or too much product compared to the equilibrium state. Keep in mind, however, that there are many combinations of reactant and product concentrations that constitute an equilibrium state.

Model 2 – Comparing Q and K_{eq}

Trial	Reaction Quotient, Q	Equilibrium Constant, K_{eq}	Q versus K_{eq}	Direction to Equilibrium
1	*1.000*	*0.250*	>	←
2	*0.500*	*0.250*	>	←
3	*3.00*	*0.250*	>	←
4	*0.250*	*0.250*	=	*No change*
5	*0.238*	*0.250*	<	→
6	*0.167*	*0.250*	<	→

10. Fill in the Equilibrium Constant column in Model 2 using data from Model 1.

    See Model 2.

11. Fill in the Direction to Equilibrium column in Model 2 using data from Model 1.

    See Model 2.

Read This!

The **reaction quotient** for a reaction is the ratio of products to reactants, similar to the equilibrium constant. The difference is you calculate the ratio with initial conditions.

$$K_{eq} = \frac{[C]_{eq}}{[A]_{eq}[B]_{eq}} \qquad Q = \frac{[C]_{initial}}{[A]_{initial}[B]_{initial}}$$

12. Calculate the reaction quotient for each of the trials in Model 1 and record the data in Model 2 in the appropriate column. Divide the work among group members. Show your work below.

    Trial 1

    $$Q = \frac{(1.000\ M)}{(1.000\ M)(1.000\ M)}$$

    $$Q = 1.000$$

    Trial 2

    $$Q = \frac{(0.500\ M)}{(2.000\ M)(0.500\ M)}$$

    $$Q = 0.500$$

    Trial 3

    $$Q = \frac{(1.500\ M)}{(1.000\ M)(0.500\ M)}$$

    $$Q = 3.000$$

13. Consider how the concentration values in the reaction quotient change when the reaction proceeds in the forward direction.

   a. Does the numerator increase or decrease?

   The numerator represents the products. In a forward reaction, the products increase. Therefore, the numerator of Q *will increase as the reaction proceeds in the forward direction.*

   b. Does the denominator increase or decrease?

   The denominator represents the reactants. In a forward reaction, the reactants decrease. Therefore, the denominator of Q *will decrease as the reaction proceeds in the forward direction.*

   c. Overall, does the Q ratio increase or decrease when the reaction proceeds in the forward direction?

   The numerator increases and the denominator decreases, so Q *will increase overall.*

14. Consider how the concentration values in the reaction quotient change when the reaction proceeds in the reverse direction.

   a. Does the numerator increase or decrease?

   The numerator represents the products. In a reverse reaction, the products decrease. Therefore, the numerator of Q *will decrease as the reaction proceeds in the reverse direction.*

   b. Does the denominator increase or decrease?

   The denominator represents the reactants. In a reverse reaction, the reactants increase. Therefore, the denominator of Q *will increase as the reaction proceeds in the reverse direction.*

   c. Overall does the Q ratio increase or decrease when the reaction proceeds in the forward direction?

   The numerator decreases and the denominator increases, so Q *will decrease overall.*

15. Fill in the Q versus K_{eq} column in the table in Model 2. Write $Q > K_{eq}$, $Q < K_{eq}$ or $Q = K_{eq}$.

   See Model 2.

16. Complete the following statements.

   a. When the reaction quotient is **greater than** the equilibrium constant, the reaction proceeds more in (the forward, [✓ the reverse,] neither) direction to reach equilibrium.

   b. When the reaction quotient is **less than** the equilibrium constant, the reaction proceeds more in ([✓ the forward,] the reverse, neither) direction to reach equilibrium.

   c. When the reaction quotient is **equal to** the equilibrium constant, the reaction proceeds more in (the forward, the reverse, [✓ neither)] direction to reach equilibrium.

17. Consider the following reaction.

$$2SO_2(g) + O_2(g) \longleftrightarrow 2SO_3(g) \qquad K_{eq} = 3900 \text{ at } 2000\text{ K}$$

a. Write the equilibrium constant expression for the reaction.

$$K_{eq} = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$$

b. Write the reaction quotient expression for the reaction.

$$Q = \frac{[SO_3]^2_{initial}}{[SO_2]^2_{initial}[O_2]_{initial}}$$

c. A reaction vessel contains 0.150 M sulfur dioxide, 0.150 M oxygen and 2.000 M sulfur trioxide. Predict the direction the reaction must shift to reach equilibrium. Show a calculation to justify your answer.

$$Q = \frac{(2.000\ M)^2}{(0.150\ M)^2(0.150\ M)} = 1185$$

Since Q *is less than* K_{eq}*, the reaction will shift forward to reach equilibrium.*

18. Consider the following reaction.

$$2HI(g) \longleftrightarrow H_2(g) + I_2(g) \qquad K_{eq} = 0.25 \text{ at } 25\,°\text{C}$$

A reaction vessel contains 0.500 M hydrogen, 0.500 M iodine vapor and 0.750 M hydrogen iodide. Predict the direction the reaction must shift to react equilibrium. Show a calculation to justify your answer.

$$Q = \frac{[H_2]_{initial}[I_2]_{initial}}{[HI]^2_{initial}}$$

$$Q = \frac{(0.500\ M)(0.500\ M)}{(0.750\ M)^2} = 0.444$$

Since Q *is greater than* K_{eq}*, the reaction will shift in the reverse direction to reach equilibrium.*

Extension Questions

19. Consider the following reaction.

$$Cd^{2+}(aq) + 4Cl^{1-}(aq) \longleftrightarrow CdCl_4^{2-}(aq) \qquad K_{eq} = 108 \text{ at } 25\,°C$$

a. Write the reaction quotient expression for the reaction.

$$Q = \frac{[CdCl_4^{2-}]_i}{[Cd^{2+}]_i[Cl^{1-}]_i}$$

where i = initial.

b. A reaction vessel contains 0.100 M cadmium ion, 0.500 M chloride and 0.250 M tetrachlorocadmate ion. Predict the direction the reaction must shift to reach equilibrium. Show a calculation to justify your answer.

$$Q = \frac{(0.250\ M)}{(0.100\ M)(0.500\ M)^4} = 40$$

Since Q *is less than* K_{eq}*, the reaction would shift in the forward direction to reach equilibrium.*

20. Consider the following reaction.

$$C(s) + H_2O(g) \longleftrightarrow H_2(g) + CO(g) \qquad K_{eq} = 4.09 \times 10^{-3} \text{ at } 2000\ K$$

a. A 1.000-L reaction vessel contains 4.00 M of water vapor, 0.040 M hydrogen and 0.040 M carbon monoxide. Calculate the reaction quotient and state which direction the quotient predicts the reaction must shift to reach equilibrium.

$$Q = \frac{[H_2]_i[CO]_i}{[H_2O]_i}$$

$$Q = \frac{(0.040\ M)(0.040\ M)}{4.00\ M} = 4.00 \times 10^{-3}$$

Based on the Q *value, the reaction would need to proceed in the forward direction to reach equilibrium because* Q *is less than* K_{eq}.

b. Discuss why the prediction above is wrong in this case.

According to the initial conditions stated in the problem, there is no carbon solid present in the reaction vessel. Therefore, the reaction must react in the reverse direction to reach equilibrium. The reaction quotient does not take solids into account and therefore cannot predict the direction of this equilibrium.

21. LeChâtelier's principle states that an equilibrium system will shift in the direction that reduces the stress put on the system. Is this consistent with the predictions made by calculating the reaction quotient and comparing *Q* to the equilibrium constant? Provide an example to support your answer.

Yes. Suppose that a system at equilibrium has some extra reactant added to the reaction vessel. LeChâtelier predicts the reaction will move in the forward direction to reduce the added substance. The added reactant will also increase the value of the denominator of Q*, making it lower than* K_{eq} *for the reaction. This would also result in a prediction of a forward reaction to reach equilibrium. Therefore, the two methods of predicting the direction a reaction shifts are consistent.*

Teacher Resources – Reaction Quotient

Learning Objectives

1. Write the reaction quotient, *Q*, for a chemical equilibrium.
2. Calculate the reaction quotient ratio using initial concentrations of all gaseous species in a chemical equilibrium.
3. Predict the direction in which the reaction will appear to proceed to reach equilibrium by comparing the reaction quotient to the equilibrium constant, K_{eq}.

Prerequisites

1. Students should be able to write an equilibrium constant expression for a chemical equilibrium.
2. Students should understand that reversible reactions react in both directions until equilibrium is reached. The equilibrium state may be different for different trials if the initial conditions differ.

Assessment Questions

1. Write the reaction quotient expression for the following chemical equilibrium.

$$N_2\,(g) \;+\; 3Cl_2\,(g) \;\longleftrightarrow\; 2NCl_3\,(g)$$

2. Which of the following is true for a reaction where the reaction quotient, *Q*, is greater than the equilibrium constant, K_{eq}?
   - *a.* The reaction reacts only in the forward direction.
   - *b.* The reaction reacts only in the reverse direction.
   - *c.* The reaction reacts in both directions, but faster in the forward direction.
   - *d.* The reaction reacts in both directions, but faster in the reverse direction.
3. Consider the following equilibrium. The equilibrium constant for this reaction is 0.030 at 250 °C. Predict which direction the reaction will shift in order to reach equilibrium if the reaction vessel initially contains 0.100 M PCl_5, 0.500 M PCl_3 and 0.010 M Cl_2.

   $PCl_5\,(g) \;\longleftrightarrow\; PCl_3\,(g) \;+\; Cl_2\,(g)$

Assessment Target Responses

1. $Q = \dfrac{[NCl_3]_i^2}{[N_2]_i[Cl_2]_i^3}$
2. *d*
3. $Q > K_{eq}$, *therefore the reaction will shift in the reverse direction.*

Teacher Tips

- Throughout this activity be mindful of how students are talking about the chemical equilibria systems. It is very easy for students to say the reaction "reacts forward" or "reacts backward," but do they realize that the equilibrium systems are always reacting in BOTH directions? Ask students to articulate this understanding at least once during the activity and then agree upon common language that you find acceptable. Should the students talk about how the reaction "shifts," or do you prefer they say the reaction reacts faster in one direction compared to the other?
- Using a visual model or simulation to illustrate how a reaction reaches equilibrium from a variety of initial states can be very helpful to students.
- The main portion of this activity uses only gas phase equilibria examples. Depending on when you approach this topic in your study of equilibrium systems, you may or may not want to use the extension questions that use heterogeneous equilibria examples. If students are not aware of how to deal with pure solids or liquids in an equilibrium system, you may want to avoid the extension questions.

Alignment with AP Chemistry Framework

- **Essential knowledge 6.A.2:** The current state of a system undergoing a reversible reaction can be characterized by the extent to which reactants have been converted to products. The relative quantities of reaction components are quantitatively described by the reaction quotient, Q.
- **Essential knowledge 6.A.3:** When a system is at equilibrium, all macroscopic variables, such as concentrations, partial pressures and temperature, do not change over time. Equilibrium results from an equality between the rates of the forward and reverse reactions, at which point $Q = K$.

  Learning Objective 6.4 The student can, given a set of initial conditions (concentrations or partial pressures) and the equilibrium constant, K, use the tendency of Q to approach K to predict and justify the prediction as to whether the reaction will proceed toward products or reactants as equilibrium is approached.

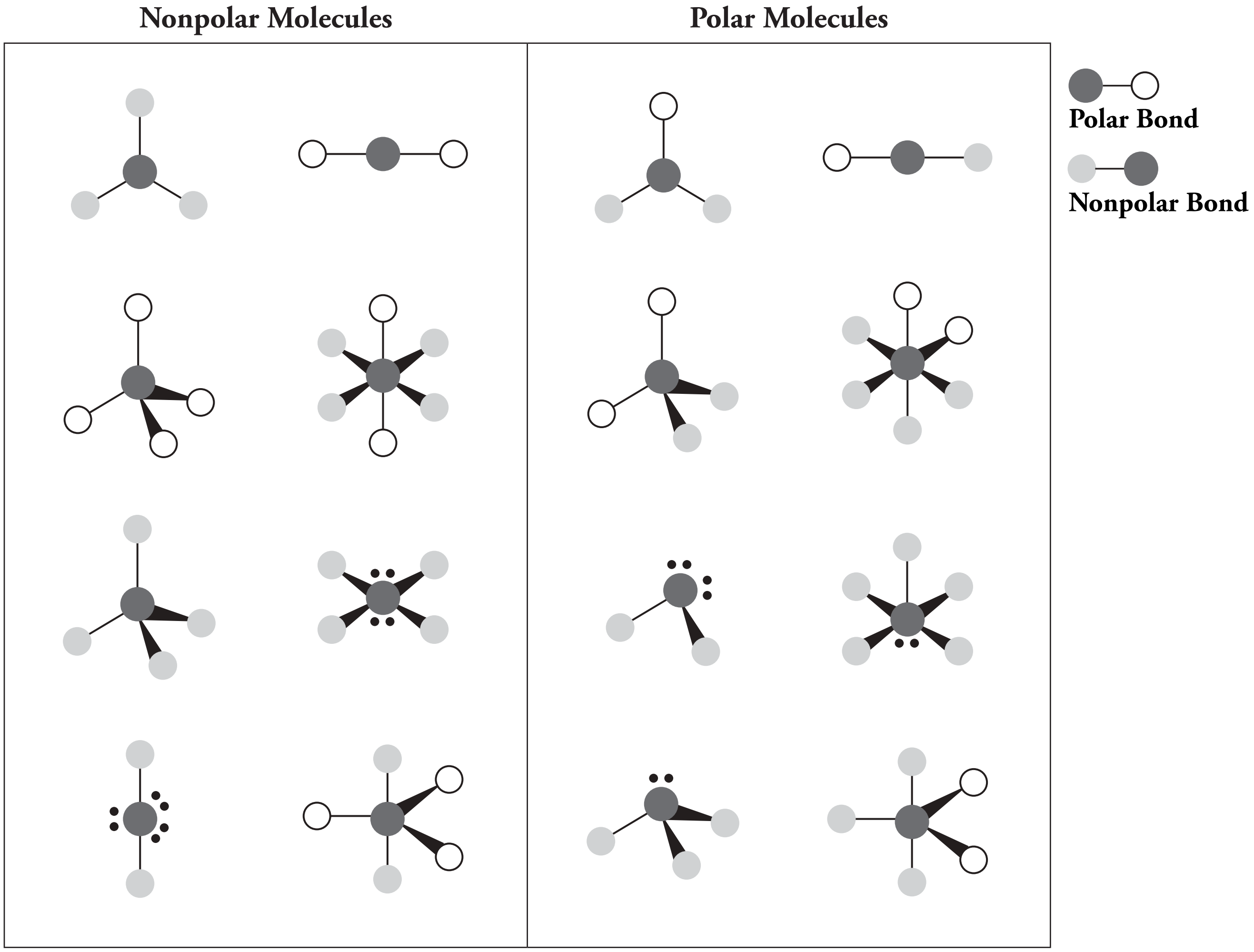
Polar and Nonpolar Molecules

What makes a molecule polar?

Why?

The physical properties of a substance are dictated in part by whether or not a molecule is polar. For example, oil and water do not mix because water is polar whereas oil is nonpolar. Another example is carbon dioxide and water. At room temperature, carbon dioxide is a gas while water is a liquid because carbon dioxide is nonpolar while water is polar. In this activity, you will explore the factors that contribute to a molecule's polarity or nonpolarity.

Model 1 – Examples of Nonpolar and Polar Molecules



1. Consider Model 1. How is a polar bond differentiated from a nonpolar bond?

2. Formaldehyde has the chemical formula CH_2O, and it is trigonal planar. Draw this molecule using open and shaded circles as it might be shown in Model 1.

3. Label each diagram in Model 1 with the three-dimensional electronic shape that it represents.

4. According to Model 1, can the shape of a molecule explain polarity? Justify your reasoning.

5. Refer to Model 1. Circle the correct word to complete each sentence.
   a. Nonpolar molecules (never, may, always) contain polar bonds.
   b. Polar molecules (never, may, always) contain polar bonds.

6. A student states "Polar molecules are just molecules that contain polar bonds. If there are no polar bonds, then the molecule is nonpolar." Do you agree or disagree with this statement? Justify your reasoning using evidence from Model 1.

7. How is a lone pair of electrons illustrated in Model 1?

8. Refer to Model 1. Circle the correct word to complete the sentence.
   a. Nonpolar molecules (never, may, always) contain lone pairs of electrons.
   b. Polar molecules (never, may, always) contain lone pairs of electrons.

9. Is the presence or absence of a lone pair of electrons sufficient to explain the polarity of molecules? Justify your reasoning using evidence from Model 1.

10. Refer to Model 1.
    a. When polar bonds are present in a nonpolar molecule, how are they arranged around the center atom—on the same side of the molecule or on opposite sides of the molecule?

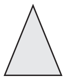
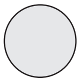
    b. When polar bonds are present in a polar molecule, how are they arranged around the center atom—on the same side of the molecule or on opposite sides of the molecule?

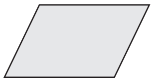
11. Refer to Model 1.

    a. When lone pairs of electrons are present in a nonpolar molecule, how are they arranged around the center atom—on the same side of the molecule or on opposite sides of the molecule?

    b. When lone pairs of electrons are present in a polar molecule, how are they arranged around the center atom—on the same side of the molecule or on opposite sides of the molecule?

12. Consider the term "symmetry" as it is used in the English language. As a group, rank the following shapes from most symmetric to least symmetric.



13. Which set in Model 1, the nonpolar molecules or polar molecules, contains molecules that are symmetric about multiple planes of symmetry?

Read This!

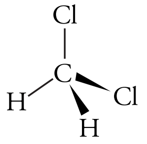
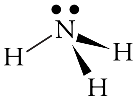
Polar molecules have an unequal distribution of charge around the molecule as a whole. This could be due to a bond being polar or a lone pair of electrons being present. Both would cause a concentration of negative charge on one side or in one area of the molecule due to an unequal distribution of electrons among nuclei. If a molecule has more than one polar bond or lone pair of electrons and they are symmetrically arranged, there is no unequal distribution of charge and the molecule is nonpolar.

14. Suppose a molecule contained two bonds around the center atom, and both those bonds were polar. What would the angle between those two bonds need to be for the molecule to be considered nonpolar? (Assume there are no lone pairs on the center atom.)

15. Suppose a molecule contained four identical bonds around the center atom, and all four of those bonds were polar. What would the angle between any of the four bonds need to be for the molecule to be considered nonpolar? (Assume there are no lone pairs on the center atom.)

16. If a molecule contains only one polar bond and no lone pairs, is there an arrangement that will make that molecule nonpolar? Justify your reasoning.

17. Fill in the table below for each molecule to determine if the molecule is polar or nonpolar. You may want to refer to a table of electronegativity values.

Name	**Dichloromethane**	**Carbon dioxide**	**Ammonia**
Structure		O═C═O	
Does the molecule have polar bonds? If yes, which ones?			
Does the molecule have lone pairs on the center atom?			
If there are polar bonds and/or lone pairs, are they symmetric?			
Is the molecule polar or nonpolar?			

18. Draw the three-dimensional structures for each molecule. Fill in the table for each molecule to determine if the molecule is polar or nonpolar.

Name	**Methane (CH_4)**	**Formaldehyde (CH_2O)**	**Xenon tetrafluoride (XeF_4)**
Structure			
Does the molecule have polar bonds? If yes, which ones?			
Does the molecule have lone pairs on the center atom?			
If there are polar bonds and/or lone pairs, are they symmetric?			
Is the molecule polar or nonpolar?			

19. Determine if the following molecules are polar or nonpolar.

a. Carbon tetrafluoride (CF_4)

b. Water (H_2O)

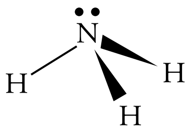
c. Sulfur dioxide (SO_2)

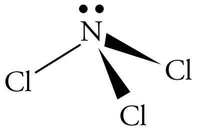
d. Phosphorous pentachloride (PCl_5)

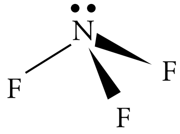
e. Hydrogen sulfide (H_2S)

Extension Questions

Model 2 – A Series of Polar Molecules







Dipole Moment 1.42 D 0.60 D 0.23 D

20. The polarity of a molecule can be quantified as the dipole moment. The dipole moment is either found indirectly through experiment or calculated theoretically. The dipole moment is measured in debyes.

    a. What is the abbreviation for the unit “debye”?

    b. Which molecule in Model 2 is the most polar according to the dipole moments given?

21. Discuss as a group the dipole moment values provided in Model 2. Propose some possible explanations for the changes in the dipole moments between the three molecules. In particular note anything that seems unusual in the data.

22. The dipole moment of a bond or molecule is often represented with an arrow that points toward the more negative side of the bond.

    a. Determine which bonds in Model 2 are polar bonds using a table of electronegativity values. A large difference in electronegativity is characteristic of a polar bond.

    b. For any bond shown in Model 2 that is significantly polar, draw an arrow next to the bond to show the direction of polarity.

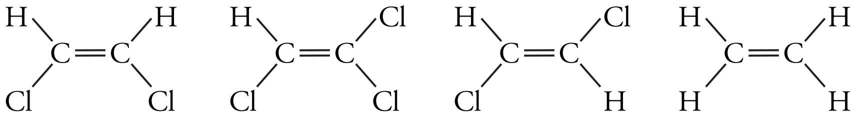
23. The lone pairs in a molecule also provide a dipole. Draw an arrow that points away from the center atom for any lone pairs on the molecules in Model 2.

24. Consider the arrows you have drawn in Model 2. Explain why NH_3 is the most polar of the three molecules.

25. Consider the arrows you have drawn in Model 2. Explain why NF_3 is the least polar of the three molecules.

26. The phosphorus trifluoride molecule (PF_3) has a dipole of 1.03 D. Compare this molecule to the nitrogen trifluoride molecule in Model 2. Propose a reason for the larger dipole in the molecule containing phosphorus.

27. Rank the following molecules in order of polarity from least to most. Use arrows on each molecule to support your answer.



28. Select the molecule in each pair that has the largest dipole moment. Include diagrams and complete sentences to justify your reasoning.

a. CS_2 vs. SO_2

b. H_2O vs. OF_2

c. CF_4 vs. SF_4

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Electrochemistry: Build Your Own Hand-held Battery

AP* Chemistry Investigation

A Guided-Inquiry Wet/Dry Experiment

Introduction

Experience and learn the concepts you need to help you succeed on the AP* Chemistry exam with this guided-inquiry activity! This activity covers topics from Big Idea 3 in the AP Chemistry curriculum. Complete a thorough homework set before lab day to delve into the basic principles of electrochemistry, involving the similarities and differences between galvanic and electrolytic cells. Then, you are tasked with the challenge to build your very own hand-held battery out of a few simple materials. The object of the challenge is to successfully assemble the battery materials so the illumination of a red LED component completes the task. You'll love this safe and fun activity while gaining a deeper understanding of electrochemistry and its real-world connection to batteries!

Concepts

- Half-cell reaction
- Galvanic cell vs. electrolytic cell
- Oxidation–reduction reactions
- Standard reduction potential

Background

Galvanic Cells

An electrochemical cell results when an oxidation reaction and a reduction reaction occur, and the resulting electron transfer between the two processes occurs through an external wire. The oxidation and reduction reactions are physically separated from each other and are called *half-cell reactions*. A half-cell can be prepared from almost any metal in contact with a solution of its ions. Since each element has its own electron configuration, each element develops a different electrical potential, and different combinations of oxidation and reduction half-cells result in different voltages for the completed electrochemical cell.

The *standard reduction potential* is the voltage that a half-cell, under standard conditions (1 M, atm, 25 °C), develops when it is combined with the standard hydrogen electrode, that is arbitrarily assigned a potential of zero volts. A chart of reduction half-cell reactions, arranged in order of decreasing standard reduction potential, shows the relative ease of reduction of each substance listed. The more positive the reduction potential, the easier the reduction. A spontaneous cell (a battery) can be constructed if two half-cells are connected internally using a salt bridge, and externally using a metallic connector. In an electrochemical cell, the reaction listed in the standard reduction potential chart with the more positive voltage occurs as a reduction, and the reaction listed with the less positive voltage reverses and occurs as an oxidation reaction. The cell voltage can be found by adding the voltages listed in the table, with the value of the voltage for the oxidation reaction becoming the negative of its reduction reaction voltage.

As an example, consider a cell made up of copper and aluminum half-cells.

$$\mathrm{Cu^{2+}(aq)} + 2\,e^- \rightarrow \mathrm{Cu(s)} \qquad E^\circ = 0.34\ \mathrm{V}$$

$$\mathrm{Al^{3+}(aq)} + 3\,e^- \rightarrow \mathrm{Al(s)} \qquad E^\circ = -1.66\ \mathrm{V}$$

The copper reaction has the more positive potential and remains a reduction reaction. The aluminum reaction with the less positive (more negative) potential is reversed and becomes an oxidation reaction. Its potential is now an oxidation potential:

$$\mathrm{Al(s)} \rightarrow \mathrm{Al^{3+}(aq)} + 3\,e^- \qquad E^\circ = +1.66\ \mathrm{V}$$

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The reduction potential and the oxidation potential are added to find the cell voltage:

$$3Cu^{2+}(aq) + 2Al(s) \rightarrow 3Cu(s) + 2Al^{3+}(aq)$$

$$E^\circ_{cell} = E^\circ_{reduction} + E^\circ_{oxidation}$$

$$E^\circ_{cell} = 0.34\ V + 1.66\ V = 2.00\ V$$

A positive value for E°_{cell} indicates the oxidation–reduction reaction, as written, is *spontaneous*.

A cell representation such as the following: $Zn(s)\ |\ Zn^{2+}(1.0\ M)\ \|\ Cu^{2+}(0.0010\ M)\ |\ Cu(s)$ means that a cell is constructed of zinc metal dipping into a 1.0 M solution of Zn^{2+}. The symbol "|" refers to a phase boundary. The symbol "‖" indicates a salt bridge between the zinc ion solution and the copper ion solution. The second half-cell is copper metal dipping into a 0.0010 M solution of copper ions. The anode is on the left (where oxidation occurs) and the cathode is on the right (where reduction occurs).

Electrolytic Cells

When an electric current is passed through an aqueous solution containing an electrolyte (Na_2SO_4), the water molecules break apart or decompose into their constituent elements, hydrogen and oxygen. The overall reaction occurs as two separate, independent half-reactions. Reduction of the hydrogen atoms to elemental hydrogen (H_2) occurs at the cathode (–), while oxidation of the oxygen atoms in water to elemental oxygen (O_2) occurs at the anode (+). Each half-reaction is accompanied by the production of OH^- or H^+ ions as shown below:

$$\text{Cathode: } 4e^- + 4H_2O \rightarrow 2H_2(g) + 4OH^-$$

$$\text{Anode: } 2H_2O \rightarrow O_2(g) + 4H^+ + 4e^-$$

Experiment Overview

Gain pre-lab preparation by completing the following homework set to gain conceptual understanding of galvanic and electrolytic cells. Then, hit the ground running on lab day and build your own hand-held battery from simple components. You will have time to do a post hand-held battery build analysis. Draw and label your observations in your lab notebook. Did you build an electrolytic cell or a galvanic cell? Prove it.

Pre-Lab Homework Assignment

Complete the following homework set and turn in any graphs or figures you were asked to create. Use a separate sheet of paper, if necessary.

Redox Reactions

1. Identify equations *a–d* as redox or nonredox. Explain.
   a. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$
   b. $Zn^{2+}(aq) + Al(s) \rightarrow Zn(s) + Al^{3+}(aq)$
   c. $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$
   d. $HF(g) + H_2O(l) \rightleftharpoons H_3O^+ + F^-(aq)$
2. Identify the oxidation half-reaction and the reduction half-reaction in *a–c*.
   a. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$
   b. $2Al(s) + 3CuCl_2(aq) \rightarrow 2AlCl_3(aq) + 3Cu(s)$
   c. $Zn(s) + V^{3+}(aq) \rightarrow V^{2+}(aq) + Zn^{2+}(aq)$
3. Go back to 2, *a–c* and identify:
   a. The oxidizing and reducing agents.
   b. The species that is oxidized and the species that is reduced.

Electrolytic Cell

A student setup a Hoffman apparatus electrolysis experiment as shown in Figure 1. The Hoffman electrolysis apparatus is a type of electrolytic cell where an electric current passes through an aqueous solution containing an electrolyte. As a result, water molecules decompose into their constituent elements. Two independent half-reactions are observed at the cathode and the anode. For a free demonstration video on the Hoffman apparatus, visit flinnsci.com.

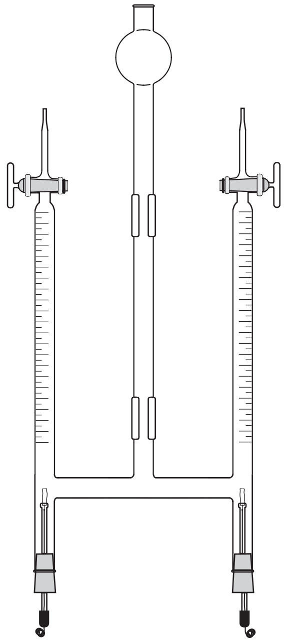

Figure 1.

4. The student added 0.5 M sodium sulfate electrolytic solution to the mouth of the Hoffman apparatus and then connected the battery leads to the electrodes. She immediately witnessed bubble formation and collected two gases on each side of the assembly. Answer questions *a–d*.
   a. Why did the student choose 0.5 M sodium sulfate for this experiment? Are there other electrolytic solutions she may test?
   b. Report the overall reaction and the independent half reactions. Identify the gases collected. Predict the volumes of each gas collected in the apparatus.
   c. Identify the cathode and anode and report where each occurs in the half reactions from question 4b.
   d. Can the student use the Hoffman apparatus without the battery? In other words, will the decomposition (formation of gases) spontaneously occur?
5. The student carefully opened each stopcock and collected each gas in separate test tubes. She inserted a lit wood splint into each.
   a. Predict what occurred to the lit wood splint of the gas collected at the anode.
   b. Predict what occurred to the lit wood splint of the gas collected at the cathode.
6. As an extension, 1-mL of universal indicator solution was added to the sodium sulfate solution while the Hoffman apparatus was connected. Predict the observations. *Hint:* Look up the exhibited color changes of universal indicator solution at various pH values.

Galvanic Cell

See Figure 2 for the second experiment setup, a galvanic cell. In a galvanic cell, a spontaneous chemical reaction releases energy in the form of electricity (moving electrons).

7. Using arrows, label the parts of the galvanic cell in Figure 2 and answer questions *a–d*. Use this list of key words: electrodes, electrode storage/compartment, cathode, anode, and salt bridge.

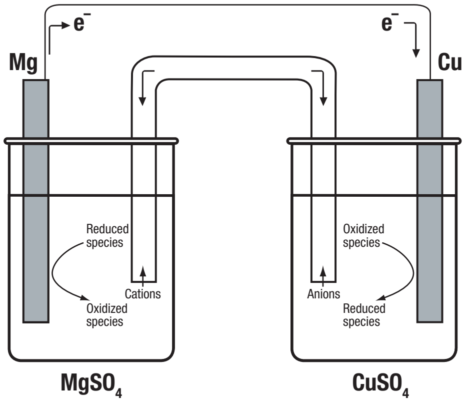

Figure 2. Galvanic Cell

   a. What would happen to the cell if the salt bridge was removed?

   b. Write the half-reactions that are taking place.

   c. Write the cell notation.

   d. Calculate the standard cell potential, E°.

8. Seek educational resources and provide an example of a replacement electrode if the copper electrode was not available in Figure 2.

9. The materials to build your very own hand-held battery include: 2 small squares of filter paper, 1 M copper(II) sulfate solution, 1 M sodium sulfate solution, magnesium ribbon, and an LED with copper tape attached to the positive terminal. Again, the challenge is to build a hand-held battery by successfully arranging the components to light the LED. *Helpful tips:*

   a. Think safety, first. Make sure you have the proper PPE available to perform this lab, i.e., goggles, apron, and gloves.

   b. Make a list of the equipment and glassware needed for this lab.

   c. Once you successfully complete the experiment, draw a hand-held battery figure in your notebook and label the parts of your battery, i.e. the cathode, anode etc. Write the reactions occurring.

   d. How does this battery differ from those practiced in the homework set? How is it similar?

   e. What are the half-cell reactions?

   f. Inspect the separate components of the hand-held battery after connecting it and lighting the LED. Describe any observations.

Safety Precautions

The copper(II) sulfate solution is harmful if swallowed and causes serious skin and eye irritation. The sodium sulfate solution may be harmful if in contact with skin. Magnesium ribbon is a flammable solid. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory. Please follow all laboratory safety guidelines.

Teacher's Notes

Electrochemistry: Build Your Own Hand-held Battery
AP* Chemistry Investigation
A Guided-Inquiry Wet/Dry Experiment

Materials Included in Kit *(for 12 groups of students)*

Copper(II) sulfate solution, $CuSO_4$, 1 M, 500 mL

Sodium sulfate solution, $NaSO_4$, 1 M, 375 mL

Copper foil conductive adhesive, Cu, 12″ piece, 2

Filter paper, 100 sheets

LEDs, clear, red, 24

Magnesium ribbon, Mg, 6 ft, 1

Sand paper, 1 (shared)

Additional Materials Required *(for each lab group)*

Weigh boats, medium, 24

Beakers, 50-mL, 24

Tweezers, 12

Scissors, 12

Graduated cylinders, 10-mL, 12

Deionized or distilled water

Time Required

This laboratory activity was specifically written, per teacher request, to be completed in one 50-minute class period. It is important to allow time between the *Pre-Lab Homework Assignment* and the *Lab Activity*. Prior to beginning the homework, show the students the hand-held battery materials—this will get the procedure thought process rolling.

Pre-Lab Preparation and Complete *Build Your Own Hand-held Battery* Procedure

1. Gently polish both LED terminals with the sand paper on all LEDs to be used by the students.
2. Cut 24 copper conductive adhesive tape pieces 2 cm in length.
3. Cover the positive terminal (the longer terminal) with the 2 cm piece of the adhesive conductive tape of each LED to be used by the students.
4. Cut 24 pieces of rectangular shaped filter paper. The sodium sulfate filter paper is the salt bridge and should be bigger in size than the copper(II) sulfate filter paper.
   a. Cut the 12 pieces to be submerged in copper(II) sulfate in about ½ cm^2 in size.
   b. Cut 12 pieces to be submerged in the sodium sulfate solution in about 1 cm^2 in size.

Build Your Own Hand-held Battery Procedure

Lead the students through steps 1–3 after providing the materials. At step 4, allow them to determine the correct component arrangements to successfully light the LED.

1. Measure 2 mL, each of the 1 M copper(II) sulfate solution and the 1 M sodium sulfate solution using a 10-mL graduated cylinder. Pour into separate medium sized weigh boats or small beakers (50-mL).
2. Using tweezers, dip the larger (1 cm^2) pre-cut filter paper into the sodium sulfate solution and the smaller (1/2 cm^2) pre-cut filter paper into the copper(II) sulfate solution. Dip long enough to completely coat each filter paper (about 10 seconds).
3. Place both on a separate weigh boat to let dry until the filter papers are damp with solution, not dripping. Students may hold each with tweezers and gently wave to decrease drying time.
4. While filter paper is drying, students make predictions. In their notebooks, students illustrate the experiment and identify the anode, cathode, the salt bridge, and evidence of electron flow; the lit LED. Students should include the half-cell reactions taking place.

5. See Figure 3 on how to correctly arrange the components. Lead the students into squeezing the LED between the thumb and index finger for component contact. Add a drop of DI water if necessary, but not too much to drown the battery. Darken the room or cup hand over LED to watch it light.

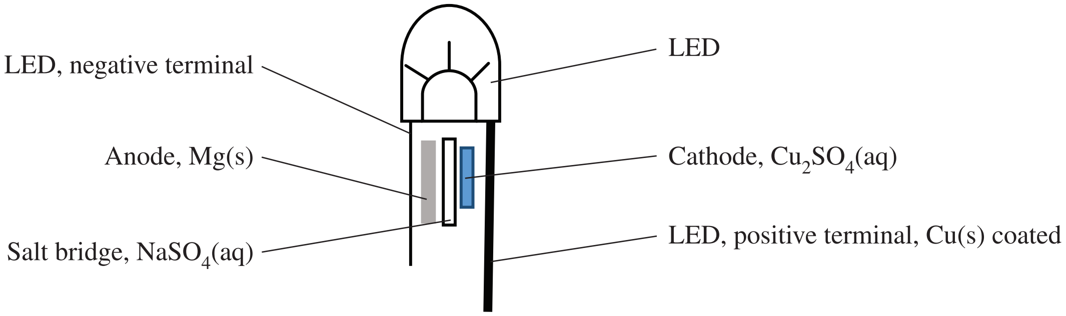

Figure 3. The correct arrangement of the hand-held battery. Squeeze with thumb and index finger on each side of the LED terminals to light the LED.

Safety Precautions

The copper(II) sulfate solution is harmful if swallowed and causes serious skin and eye irritation. The sodium sulfate solution may be harmful if in contact with skin. Magnesium ribbon is a flammable solid. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory. Please follow all laboratory safety guidelines. Remind students to wash their hands thoroughly with soap and water before leaving the laboratory. Please review current Safety Data Sheets for additional safety, handling, and disposal information.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. You may save all materials, including solutions, for future labs. Any leftover copper(II) sulfate and sodium sulfate solutions may be flushed down the drain with excess water according to Flinn disposal method #26b.

Alignment with AP* Chemistry Curriculum Framework—Big Idea 3

Enduring Understandings and Essential Knowledge

Chemical reactions can be classified by considering what the reactants are, what the products are, or how they change from one into the other. Classes of chemical reactions include synthesis, decomposition, acid–base, and oxidation–reduction reactions. (Enduring Understanding 3B)

3B3: In oxidation–reduction (redox) reactions, there is a net transfer of electrons. The species that loses electrons is oxidized, and the species that gains electrons is reduced.

Chemical and physical transformations may be observed in several ways and typically involve a change in energy. (Enduring Understanding 3C)

3C3: Electrochemistry shows the interconversion between chemical and electrical energy in galvanic and electrolytic cells.

Chemical equilibrium is a dynamic, reversible state in which rates of opposing processes are equal. (Enduring Understanding 6A)

6A4: The magnitude of the equilibrium constant, K, can be used to determine whether the equilibrium lies toward the reactant side or product side.

Chemical equilibrium plays an important role in acid–base chemistry and in solubility. (Enduring Understanding 6C)

6C3: The solubility of a substance can be understood in terms of chemical equilibrium.

Learning Objectives

3.8 The student is able to identify redox reactions and justify the identification in terms of electron transfer.

3.12 The student can make qualitative or quantitative predictions about galvanic or electrolytic reactions based on half-cell reactions and potentials and/or Faraday's laws.

3.13 The student can analyze data regarding galvanic or electrolytic cells to identify properties of the underlying redox reactions.

6.7 The student is able, for a reversible reaction that has a large or small K, to determine which chemical species will have very large versus very small concentrations at equilibrium.

6.21 The student can predict the solubility of a salt, or rank the solubility of salts, given the relevant K_{sp} values.

Science Practices

2.2 The student can apply mathematical routines to quantities that describe natural phenomena.

2.3 The student can estimate numerically quantities that describe natural phenomena.

4.3 The student can collect data to answer a particular scientific question.

5.1 The student can analyze data to identify patterns or relationships.

Lab Hints

- A pre-lab prep alternative to cutting filter paper squares: you may first dip the filter paper in the sulfate solutions, allow to dry, then cut into squares for use by the students.
- Depending the inquiry level of your students, they may perform all of the set-up steps from the *Pre-Lab Preparation.*

Teaching Tips

- Flinn Scientific has excellent video resources that enhance the teaching experience! Simply type in the key word electrolysis or Hoffman apparatus to pull up some great videos.
- The Colorful Electrolysis Demonstration is a great extension to this lab! This demonstration kit is available from Flinn (Catalog No. AP6467).

Answers to Pre-Lab Homework Assignment *(Student answers will vary.)*

Redox Reactions

1. Identify equations *a–d* as redox or nonredox. Explain.

Electrons are gained or lost in redox reactions; a species is oxidized or reduced. 1 c and d are nonredox due to being acid base reactions.

a. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

Redox

b. $Zn^{2+}(aq) + Al(s) \rightarrow Zn(s) + Al^{3+}(aq)$

Redox

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

Nonredox

d. $HF(g) + H_2O(l) \rightleftharpoons H_3O^+ + F^-(aq)$

Nonredox

2. Identify the oxidation half-reaction and the reduction half-reaction in *a–c*.

a. $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ *(oxidation)*

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ *(reduction)*

b. $2Al(s) + 3CuCl_2(aq) \rightarrow 2AlCl_3(aq) + 3Cu(s)$

$Al(s) \rightarrow Al^{3+} + 2e^-$ *(oxidation)*

$Cu^{2+} + 2e^- \rightarrow Cu(s)$ *(reduction)* *c.* $Zn(s) + V^{3+}(aq) \rightarrow V^{2+}(aq) + Zn^{2+}(aq)$

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ (oxidation)

$V^{3+}(aq) + 2e^- \rightarrow V^{2+}$ (reduction)

3. Go back to 2, *a–c* and identify:

    a. The oxidizing and reducing agents.

    2a. Zn is the reducing species and Cu^{2+} is the oxidizing species.
    2b. Al is the reducing species and Cu^{2+} is the oxidizing species.
    2c. Zn is the reducing species and V^{3+} is the oxidizing species.

    b. The species that is oxidized and the species that is reduced.

    2a. Zn is oxidized and Cu^{2+} is reduced.
    2b. Al is oxidized and Cu^{2+} is reduced.
    2c. Zn is oxidized and V^{3+} is reduced.

Electrolytic Cell

A student setup a Hoffman apparatus electrolysis experiment as shown in Figure 1. The Hoffman electrolysis apparatus is a type of electrolytic cell where an electric current passes through an aqueous solution containing an electrolyte. As a result, water molecules decompose into their constituent elements. Two independent half-reactions are observed at the cathode and the anode. For a free demonstration video on the Hoffman apparatus, visit flinnsci.com.

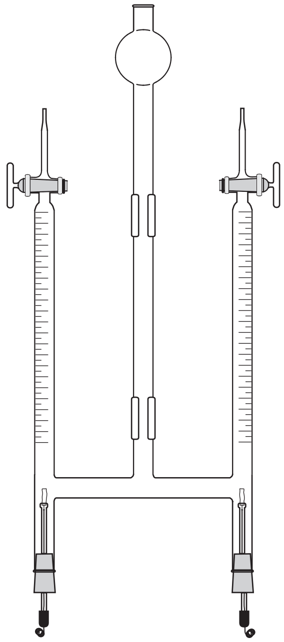

Figure 1.

4. The student added 0.5 M sodium sulfate electrolytic solution to the mouth of the Hoffman apparatus and then connected the battery leads to the electrodes. She immediately witnessed bubble formation and collected two gases on each side of the assembly. Answer questions *a–d*.

    a. Why did the student choose 0.5 M sodium sulfate for this experiment? Are there other electrolytic solutions she may test?

    Sodium sulfate is an electrolyte, it is conductive. The student can use solutions of potassium iodide or potassium chloride.

    b. Report the overall reaction and the independent half reactions. Identify the gases collected. Predict the volumes of each gas collected in the apparatus.

    Oxygen and hydrogen gases are produced/collected. Since hydrogen gas forms at twice the molar volume in the overall reaction, double the amount of hydrogen gas collects in the apparatus vs. oxygen gas.

    $4H_2O \rightarrow 4H_2 + 2O_2$

    Cathode: $4e^- + 4H_2O \rightarrow 2H_2(g) + 4OH^-$

    Anode: $2H_2O \rightarrow O_2(g) + 4H^+ + 4e^-$

    c. Identify the cathode and anode and report where each occurs in the half reactions from question 4b.

    See 4b.

    d. Can the student use the Hoffman apparatus without the battery? In other words, will the decomposition (formation of gases) spontaneously occur?

    No, electrolysis reactions require external power where a non-favorable redox reaction occurs.

5. The student carefully opened each stopcock and collected each gas in separate test tubes. She inserted a lit wood splint into each.

    a. Predict what occurred to the lit wood splint of the gas collected at the anode.

    Oxygen gas is flammable, so it ignited the wood splint.

    b. Predict what occurred to the lit wood splint of the gas collected at the cathode.

Hydrogen gas is also flammable, so it ignited the wood splint as well.

6. As an extension, 1-mL of universal indicator solution was added to the sodium sulfate solution while the Hoffman apparatus was connected. Predict the observations. Hint: Look up the exhibited color changes of universal indicator solution at various pH values.

   Beautiful colors will result. At the cathode: solution will be purple/blue due to the production of basic hydroxide ions. At the anode: solution will be pink/red due to the production of acidic protons.

Galvanic Cell

See Figure 2 for the second experiment setup, a galvanic cell. In a galvanic cell, a spontaneous chemical reaction releases energy in the form of electricity (moving electrons).

7. Using arrows, label the parts of the galvanic cell in Figure 2 and answer questions a–d. Use this list of key words: electrodes, electrode storage/compartment, cathode, anode, and salt bridge.

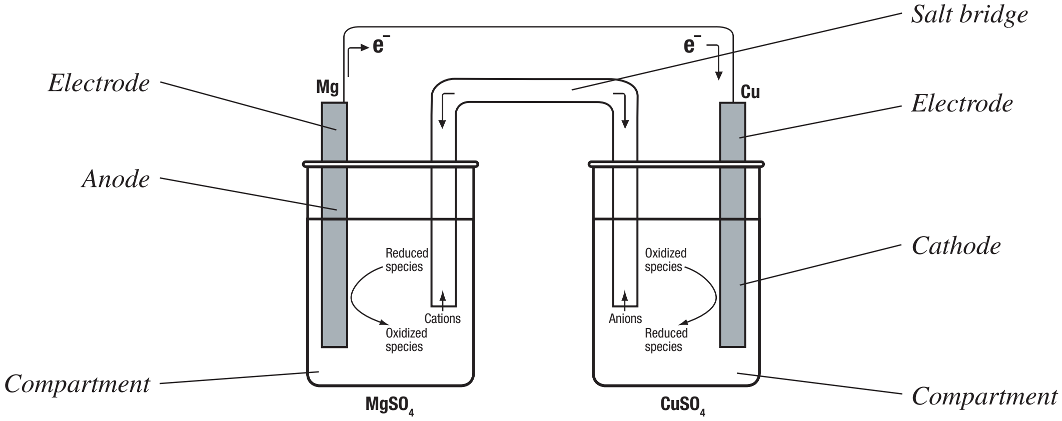

Figure 2. Galvanic Cell

a. What would happen to the cell if the salt bridge was removed?

Build-up of positive and negative charges would occur; salt bridges are needed for chemical neutrality.

b. Write the half-reactions that are taking place.

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$

c. Write the cell notation.

$Mg(s)\,|\,Mg^{2+}(1\ M)\,\|\,Cu^{2+}\ (1\ M)\,|\,Cu(s)$

d. Calculate the standard cell potential, E°.

$E^\circ_{cell} = E^\circ_{cathode} - E^\circ_{anode}$

$E^\circ_{cell} = 0.34 - (-2.37) = 2.71\ V$

8. Seek educational resources and provide an example of a replacement electrode if the copper electrode was not available in Figure 2.

   A standard, non-conductive electrode such as platinum or graphite are good options.

9. The materials to build your very own hand-held battery include: 2 small squares of filter paper, 1 M copper(II) sulfate solution, 1 M sodium sulfate solution, magnesium ribbon, and an LED with copper tape attached to the positive terminal. Again, the challenge is to build a hand-held battery by successfully arranging the components to light the LED. *Helpful tips:*

   a. Think safety, first. Make sure you have the proper PPE available to perform this lab, i.e., goggles, apron, and gloves.

   b. Make list of the equipment and glassware needed for this lab.

c. Once you successfully complete the experiment, draw a hand-held battery figure in your notebook and label the parts of your battery, i.e. the cathode, anode etc. Write the reactions occurring.

See Figure 3

d. How does this battery differ from those practiced in the homework set? How is it similar?

It is a galvanic cell—it does not require external power to work. The filter papers are the cell's compartments and there is not a magnesium sulfate solution present.

e. What are the half-cell reactions?

Cu tape was applied to the positive LED terminal of the battery as the conductive material to prevent reaction between Mg(s) and the bare LED terminal.

$Mg(s) \rightarrow Mg^{2+} + 2e^-$ *(oxidation, anode)*

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ *(reduction, cathode)*

f. Inspect the separate components of the hand-held battery after connecting it and lighting the LED. Describe any observations.

Student answers will vary. Dark spots are seen on the copper(II) sulfate filter paper, which is copper metal. Without the conductive tape, the LED terminals can darken due to oxidation of Mg(s).

References

Eggen, P.; Skaugrud, B. *An Easy-to-Assemble Three-Part Galvanic Cell. J. Chem. Educ.* **2015**, *92 (6)*, 1053–1055.

AP Chemistry Guided-Inquiry Experiments: Applying the Science Practices*; The College Board: New York, NY, **2013.**

The *Electrochemistry: Build Your Own Hand-held Battery AP* Chemistry Investigation—A Guided-Inquiry Wet/Dry Experiment* is available from Flinn Scientific, Inc.

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AP8496	Electrochemistry: Build Your Own Hand-held Battery AP* Chemistry Investigation—A Guided-Inquiry Wet/Dry Experiment

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Designing a Hand Warmer

AP* Chemistry Big Idea 5, Investigation 12

An Advanced Inquiry Lab

Introduction

Put your chemistry skills to commercial use! From instant cold packs to flameless ration heaters and hand warmers, the energy changes accompanying physical and chemical transformations have many consumer applications. The backbone of these applications is calorimetry—measuring heat transfer. Investigate the energy changes accompanying the formation of solutions for common laboratory salts, and then apply the results to design a hand warmer that is reliable, safe and inexpensive.

Concepts

- Enthalpy change
- Calorimetry
- Heat of solution
- Specific heat
- Exothermic versus endothermic
- System and surroundings

Background

Hand warmers are familiar cold weather gear used to quickly provide warmth to frigid fingers. Many commercial hand warmers consist of a plastic package containing a solid and an inner pouch filled with water. When the pack is activated, the solid dissolves in water and produces a large temperature change.

The energy or *enthalpy change* associated with the process of a solute dissolving in a solvent is called the *heat of solution* (ΔH_{soln}). At constant pressure, this enthalpy change, ΔH_{soln}, is equal in magnitude to the heat loss or gain, q, to the surroundings. In the case of an ionic solid dissolving in water, the overall energy change is the net result of three processes—the energy required to break the attractive forces between ions in the crystal lattice (ΔH_1 = +C kJ/mole), the energy required to disrupt intermolecular forces between water molecules (ΔH_2 = +D kJ/mole), and the energy released when the dissociated (free) ions form ion-dipole attractive forces with the water molecules (ΔH_3 = –F kJ/mole). The overall process can be represented by the following equations.

$$M_aX_b(s) \rightarrow aM^{+b}(aq) + bX^{-a}(aq) \qquad \Delta H_{soln} = \Delta H_1 + \Delta H_2 + \Delta H_3 = (+C + D - F)\ \text{kJ/mole}$$

If the amount of energy *released* in the formation of hydrated ions (ΔH_3) is greater than the amount of energy *required* to separate the solute and solvent particles ($\Delta H_1 + \Delta H_2$), then the sum (ΔH_{soln}) of the energy changes will be negative and the solution process exothermic (releases heat). If the amount of energy released in the formation of hydrated ions is less than the amount of energy required to separate the solute and solvent particles, then the sum of the energy changes will be positive and the solution process endothermic (absorbs heat).

Heats of solution and other enthalpy changes are generally measured in an insulated vessel called a *calorimeter* that reduces or prevents heat loss to the atmosphere outside the reaction vessel. The process of a solute dissolving in water may either release heat into the resulting aqueous solution or absorb heat from the solution, but the amount of heat exchanged between the calorimeter and the outside surroundings should be minimal. When using a calorimeter, the reagents being studied are mixed directly in the calorimeter and the temperature is recorded both before and after the reaction has occurred. The amount of heat transfer (q) may be calculated using the heat energy equation:

$$q = m \times s \times \Delta T \qquad \textit{Equation 1}$$

where m is the **total mass of the solution** (solute plus solvent), s is the specific heat of the solution, and ΔT is the observed temperature change. The specific heat of the solution is generally assumed to be the same as that of water, namely, 4.18 J/g·°C.

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When measuring the heat transfer for an *exothermic* heat of solution using a calorimeter, most of the heat released is absorbed by the aqueous solution (q_{aq}). A small amount of the heat will be absorbed by the calorimeter itself (q_{cal}). The overall heat transfer (q_{soln}) for the reaction (the system) then becomes:

$$q_{soln} = -(q_{aq} + q_{cal}) \qquad \textit{Equation 2}$$

In order to determine the correction factor q_{cal} for heat of solution calculations, the heat capacity of the calorimeter, also called the calorimeter constant, must be determined experimentally. The calorimeter constant has units J/°C. This calibration experiment is done by mixing equal volumes of hot and cool water in the calorimeter and measuring the temperature after 20 seconds. The resulting value is assumed to be the instantaneous mixing temperature, T_{mix}. The average temperature T_{avg} of the initial hot (T_H) and cool water (T_C) is also calculated:

$$T_{avg} = (T_H + T_C)/2$$

The difference between T_{avg} and T_{mix} is due to the heat lost by the water and absorbed by the calorimeter. The heat lost by the water, q_{water}, is:

$$q_{water} = (\text{mass of water}) \times (\text{specific heat of water}) \times (T_{mix} - T_{avg}) \qquad \textit{Equation 3}$$

where the mass is the total mass of hot and cool water. The heat gained by the calorimeter, q_{calor}, is equal to that lost by the water, but opposite in sign. The calorimeter constant, C_{cal}, is calculated as follows:

$$C_{cal} = \frac{q_{calor}}{(T_{mix} - T_{initial})} \qquad \textit{Equation 4}$$

where $T_{initial}$ is the initial temperature of the calorimeter containing cool water.

To calculate the correction factor q_{cal} for use in Equation 2 above—to determine the heat of solution or heat of reaction for any system—the calorimeter constant is multiplied by the change in temperature of that solution:

$$q_{cal} = \Delta T\,(^\circ\mathrm{C}) \times C_{cal}\,(\mathrm{J/^\circ C})$$

Experiment Overview

The purpose of this advanced inquiry lab is to design an effective hand warmer that is inexpensive, nontoxic and safe for the environment. The investigation begins with an introductory activity to become familiar with the principles of calorimetry and heat of solution calculations. The results provide a model for the guided-inquiry challenge, which is to design an optimum hand warmer for consumer applications. Working in groups of four, each student group will be provided six different solids, along with their costs and individual Material Safety Data Sheets (MSDS). Determine the heat of solution for each solid and analyze the cost and safety information to propose a design for the best all-around hand warmer.

Pre-Lab Questions

1. When chromium chloride, $CrCl_2$, is dissolved in water, the temperature of the water decreases.
   a. Is the heat of solution exothermic or endothermic?
   b. Which is stronger—the attractive forces between water molecules and chromium and chloride ions, or the combined ionic bond strength of $CrCl_2$ and intermolecular forces between water molecules? Explain.
2. A solution was formed by combining 25.0 g of solid A with 60.0 mL of distilled water, with the water initially at 21.4 °C. The final temperature of the solution was 25.3 °C. Calculate the heat released as the solid dissolved, q_{soln}, assuming no heat loss to the calorimeter (see Equation 1).
3. In Question 2 above, the calorimeter was found to have a heat capacity of 8.20 J/°C. If a correction is included to account for the heat absorbed by the calorimeter, what is the heat of solution, q_{soln}?
4. The solid in Question 2 was aluminum sulfate, $Al_2(SO_4)_3$. Calculate the molar heat of solution, ΔH_{soln}, for aluminum sulfate. *Hint:* The units for molar heat of solution are kilojoules per mole (kJ/mole). First determine the heat released per gram of solid.

Materials (for each lab group)

Ionic solids:

Set A: Ammonium chloride, NH_4Cl, 15 g

Calcium chloride, anhydrous, $CaCl_2$, 15 g

Sodium acetate, $NaCH_3CO_2$, 15 g

Set B: Sodium chloride, NaCl, 15 g

Lithium chloride, LiCl, 15 g

Sodium carbonate, Na_2CO_3, 15 g

Magnesium sulfate, anhydrous, $MgSO_4$, 5 g

Water, deionized or distilled

Balance, 0.01-g precision (shared)

Beaker, 250-mL

Calorimeter (two nested polystyrene cups)

Graduated cylinder, 100-mL

Heat-resistant gloves

Hot plate

Magnetic stirrer and stir bar, or stirring rod

Paper towels

Support stand and ring clamp

Thermometer, digital

Timer or stopwatch

Weighing dishes

Safety Precautions

Lithium chloride is moderately toxic by ingestion. Calcium chloride and ammonium chloride are slightly toxic. Magnesium sulfate is a body tissue irritant. Sodium acetate is a body tissue and respiratory tract irritant. Avoid contact of all chemicals with eyes and skin. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory. Please follow all laboratory safety guidelines.

Introductory Activity

Part A. Heat Capacity of the Calorimeter

1. Working in pairs, set up a calorimeter consisting of two nested polystyrene cups in a ring clamp attached to a support stand.
2. Place a magnetic stirrer below the calorimeter, then lower the ring clamp until the bottom of the cup just sits on the surface of the magnetic stirrer (see Figure 1).
3. Measure 100.0 mL of distilled water in a 100-mL graduated cylinder and transfer the water into the calorimeter.

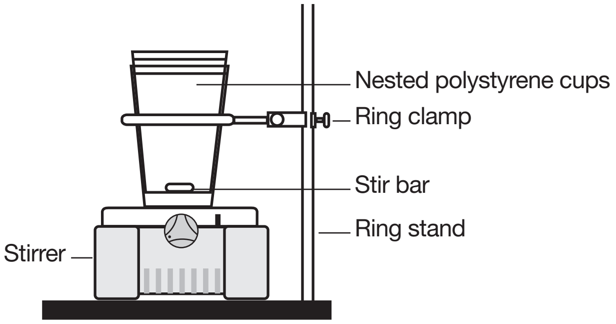


Figure 1.

4. Add a magnetic stirring bar to the calorimeter, and set the bar spinning slowly. If a magnetic stirrer is not available, use a stirring rod. Do not remove the stirring rod from the calorimeter.
5. Measure and record the initial temperature of the water.
6. Heat approximately 125 mL of distilled water to 60–70 °C in a 250-mL beaker.
7. Using heat-resistant gloves, measure 100.0 mL of the hot water in a 100-mL graduated cylinder.
8. Measure and record the temperature of the hot water.
9. Immediately pour the hot water into the room temperature water in the calorimeter.
10. Insert the thermometer, and stir the water.
11. Record the mixing temperature T_{mix} after 20 seconds.
12. Empty the calorimeter and dry the inside.
13. Calculate the calorimeter constant, C_{cal}, using T_{mix} and Equations 3 and 4 from the *Background* section.

Part B. Calorimetry Procedure

Working in pairs, examine the heat energy change for the following solution.

$$MgSO_4(s) + H_2O(l) \rightarrow Mg^{2+}(aq) + SO_4^{2-}(aq)$$

1. Measure 100.0 mL of distilled or deionized water in a 100-mL graduated cylinder and transfer to the calorimeter.
2. Measure and record the initial temperature of the water.
3. Measure 5.00 g of anhydrous magnesium sulfate in a weighing dish.
4. Put a magnetic stir bar or stirring rod into the calorimeter and slowly stir the water.
5. Quickly add the 5.00 g of anhydrous magnesium sulfate to the calorimeter and insert the thermometer.
6. Monitor the temperature and record the highest or lowest temperature reading.
7. Calculate the molar heat of solution for magnesium sulfate. Include the correction due to the heat capacity of the calorimeter.

Guided-Inquiry Design and Procedure

Form a working group with other students and discuss the following questions.

1. Review the calorimetry procedure:
   - *a.* What data is needed to calculate the enthalpy change for a reaction?
   - *b.* Identify the variables that will influence the experimental data.
   - *c.* What variables should be controlled (kept constant) during the procedure?
   - *d.* The independent variable in an experiment is the variable that is changed by the experimenter, while the dependent variable responds to or depends on the changes in the independent variable. Name the independent and dependent variables in a calorimetry experiment to determine the molar heat of solution.
   - *e.* Discuss the factors that will affect the precision of the experimental results.
2. One pair of students in the group should study the three solids in Set A, while the other pair studies Set B.
3. Working collaboratively with the general procedure provided in the *Introductory Activity,* design and carry out experiments to determine the heat of solution for each solid. Be sure to review all safety precautions with your instructor before starting.
4. Extrapolating from the information collected, predict which solid(s) could be used in an effective hand warmer meeting the following requirements:
   - The hand warmer must contain 10 g of an ionic solid and an inner pouch filled with 40 mL of water.
   - Activating the hand warmer must increase the temperature of the resulting solution by at least 20 °C.
   - The solid should be nontoxic, safe for the environment, and economical.
5. Review the cost information shown below and consult the MSDS for each potential hand warmer. Propose the optimum design for the most cost-effective hand warmer that is nontoxic and least harmful to the environment.

Solid	NH_4Cl	$CaCl_2$	LiCl	$NaCH_3CO_2$	Na_2CO_3	NaCl
Cost ($)/Kilogram	21.90	10.80	68.30	27.30	5.95	4.25

6. With your instructor's permission, verify the design and demonstrate the use of your hand warmer.

Opportunities for Inquiry

Instant cold packs are used to treat sports and other injuries when ice is unavailable. Research the properties of commercial cold packs and select a "cold pack solid" from among the solids provided in this activity. Propose and test a design (quantity of solid and volume of water) for an effective instant cold pack.

AP Chemistry Review Questions

Integrating Content, Inquiry and Reasoning

Review the following data from a calorimetry experiment to determine the heat of fusion of ice. After shaking off any excess water, several ice cubes were added to 99 g of warm water contained in a calorimeter. The initial temperature of the warm water was 46.8 °C. The ice–water mixture was stirred until the temperature reached a stable, minimum value, which was 1.1 °C. Any *unmelted* ice remaining at this point was immediately and carefully removed using tongs and the mass of the water in the calorimeter was measured—154 g.

1. Use the heat energy equation to calculate the amount of heat in joules released by the warm water as it cooled.

2. Based on the law of conservation of energy, what amount of heat was absorbed by the ice as it melted?

3. Determine the amount of energy absorbed per gram of ice as it melted.

4. Calculate the heat of fusion (the heat required to melt ice) in units of kilojoules/mole.

5. The literature value for the heat of fusion of ice is 6.02 kJ/mole. What is the percent error for the experimentally determined heat of fusion?

6. When a mixture of ice and water originally at 0 °C is heated, the temperature remains constant (within experimental error) until all of the ice melts. Explain what happens to the heat energy that is absorbed during this time while the temperature does not change.

Teacher's Notes

Investigation 12—Designing a Hand Warmer

Part I. Lab Preparation

Part II. Teacher Guidance

Part III. Sample Data, Results, and Analysis

Part I. Lab Preparation

Materials Included in Kit (for 24 students working in pairs)

Ammonium chloride, NH_4Cl, 120 g
Calcium chloride, anhydrous, $CaCl_2$, 120 g
Lithium chloride, LiCl, 120 g
Magnesium sulfate, anhydrous, $MgSO_4$, 120 g
Sodium acetate, anhydrous, $NaCH_3CO_2$, 120 g
Sodium carbonate, anhydrous, Na_2CO_3, 125 g
Sodium chloride, NaCl, 120 g
Cups, polystyrene, 8 oz., 24

Additional Materials Required (for each lab group)

Water, deionized or distilled
Balance, 0.01-g precision (shared)
Beaker, 250-mL
Graduated cylinder, 100-mL
Heat-resistant gloves
Hot plate (shared)
Magnetic stirrer and stir bar, or stirring rod
Paper towels
Support stand and ring clamp
Thermometer, digital
Timer or stopwatch
Weighing dishes

Time Required

This laboratory activity can be completed in two 50-minute class periods. It is important to allow time between the *Introductory Activity* and the *Guided-Inquiry Activity* for students to discuss and design the guided-inquiry procedures. Also, all student-designed procedures must be approved for safety before students are allowed to implement them in the lab. *Pre-Lab Questions* may be completed before lab begins the first day and the results and analysis may be completed the day after the lab or as homework. An additional lab period would be needed for students to complete an optional inquiry investigation (see *Opportunities for Inquiry*). Students may find the (M)SDS from the Flinn Scientific website, www.flinnsci.com.

Safety Precautions

Lithium chloride is moderately toxic by ingestion. Calcium chloride and ammonium chloride are slightly toxic. Magnesium sulfate is a body tissue irritant. Sodium acetate is a body tissue and respiratory tract irritant. Avoid contact of all chemicals with eyes and skin. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Remind students to wash their hands thoroughly with soap and water before leaving the laboratory. Please review current Material Safety Data Sheets for additional safety, handling, and disposal information.

Disposal

Please consult your current *Flinn Scientific Catalog/Reference Manual* for general guidelines and specific procedures, and review all federal, state and local regulations that may apply, before proceeding. The solid samples may be stored for future use or placed in the trash according to Flinn Suggested Disposal Method #26a. The experimental solutions may be rinsed down the drain with plenty of excess water according to Flinn Suggested Disposal Method #26b.

Part II. Teacher Guidance

Alignment to AP Chemistry Curriculum Framework

Enduring Understandings and Essential Knowledge

Forces of attraction between particles (including the noble gases and also different parts of some large molecules) are important in determining many macroscopic properties of a substance, including how the observable physical state changes with temperature. (2B)

2B3: Intermolecular forces play a key role in determining the properties of substances, including biological structures and interactions.

The type of bonding in the solid state can be deduced from the properties of the solid state. (2D)

2D1: Ionic solids have high melting points, are brittle, and conduct electricity only when molten or in solution.

Energy is neither created nor destroyed, but only transformed from one form to another. (5B)

5B2: When two systems are in contact with each other and are otherwise isolated, the energy that comes out of one system is equal to the energy that goes into the other system. The combined energy of the two systems remains fixed. Energy transfer can occur though either heat exchange or work.

5B3: Chemical systems undergo three main processes that change their energy: heating/cooling, phase transitions, and chemical reactions.

5B4: Calorimetry is an experimental technique that is used to measure the change in energy of a chemical system.

Learning Objectives

2.15 The student is able to explain observations regarding the solubility of ionic solids and molecules in water and other solvents on the basis of particle views that include intermolecular interactions and entropic effects.

2.23 The student can create a representation of an ionic solid that shows essential characteristics of the structure and interactions present in the substance.

2.24 The student is able to explain a representation that connects properties of an ionic solid to its structural attributes and to the interactions present at the atomic level.

5.4 The student is able to use conservation of energy to relate the magnitudes of the energy changes occurring in two or more interacting systems, including identification of the systems, the type (heat versus work), or the direction of energy flow.

5.5 The student is able to use conservation of energy to relate the magnitudes of the energy changes when two nonreacting substances are mixed or brought into contact with one another.

5.6 The student is able to use calculations or estimations to relate energy changes associated with heating/cooling a substance to the heat capacity, relate energy changes associated with a phase transition to the enthalpy of fusion/vaporization, relate energy changes associated with a chemical reaction to the enthalpy of the reaction, and relate energy changes to PΔV work.

5.7 The student is able to design and/or interpret the results of an experiment in which calorimetry is used to determine the change in enthalpy of a chemical process (heating/cooling, phase transition, or chemical reaction) at constant pressure.

Science Practices

1.1 The student can create representations and models of natural or man-made phenomena and systems in the domain.

1.4 The student can use representations and models to analyze situations or solve problems qualitatively and quantitatively.

2.2 The student can apply mathematical routines to quantities that describe natural phenomena.

2.3 The student can estimate numerically quantities that describe natural phenomena.

4.2 The student can design a plan for collecting data to answer a particular scientific question.

5.1 The student can analyze data to identify patterns or relationships.

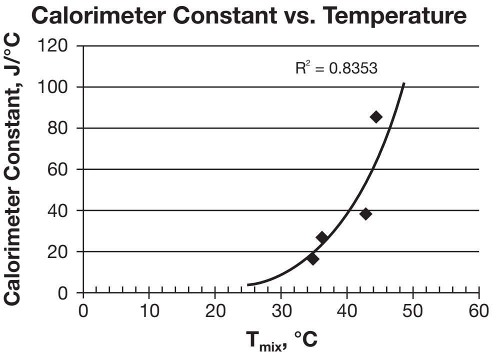
6.2 The student can construct explanations of phenomena based on evidence produced through scientific practices.

7.1 The student can connect phenomena and models across spatial and temporal scales.

Lab Hints

- The best thermometers to use are digital electronic thermometers (such as Flinn Scientific Catalog No. AP8716) or temperature sensors connected to a computer-based interface system such as LabPro. Flinn digital thermometers are reasonably inexpensive, update every second, and are precise to the nearest 0.1 °C. Temperature measurements may be a significant source of error in calorimetry experiments.
- Two polystyrene cups nested together provide better insulation and thermal stability than one cup. If two cups are used, students can easily run two trials without rinsing and drying the cup between trials. Simply have the students interchange the actual solution cup and the bottom cup between measurements.
- We report a calorimeter constant value of 17 J/°C for the *Introductory Activity* mixing hot and cold water where the mixing temperature (T_{mix}) was 34.8 °C. The value of C_{cal} depends on the average or mixing temperature of the hot and cold water. Intuitively, it makes sense that C_{cal} will depend on temperature, and indeed we found this to be the case. (You would expect the calorimeter constant to be zero if you mix room temperature water with room temperature water at 25 °C.)

  On the left is a graph of C_{cal} versus T_{mix} for four different mixing temperatures from 34 to 45 °C. For greatest accuracy, we recommend that the calorimeter constant determination be carried out by mixing room temperature water with **hot water** at 45–55 °C, for an expected T_{mix} in the 33–38 °C range. The maximum temperature in the exothermic heat of solution experiments would fall in this temperature range.



Teaching Tips

- Students may need help with the calorimetry calculations. It is important to keep track of the signs associated with the heat changes, and to understand the notion of the system versus the surroundings. A loss of heat is assigned a negative value, a gain of heat a positive value. If a reaction is exothermic, it releases or loses heat and *q* has a negative value. The same quantity of heat is absorbed by the solution and the heat change for the solution therefore has the same value but with a positive sign.
- A slight error is incorporated into the calorimeter constant calibration to make the calculations simpler and the data collection less time-consuming. Please request *Thermodynamics: Enthalpy of Reaction and Hess's Law* (Flinn Scientific Publication No. 8832) for the full calorimeter constant determination procedure.

- Another type of hand warmer uses a pouch of supersaturated sodium acetate trihydrate. The crystallization of sodium acetate trihydrate from its supersaturated solution is a spontaneous physical process. The Gibbs free energy expressions for this process are:

  $$\Delta G = \Delta H - T\Delta S \qquad \textit{Equation 5}$$

  $$\Delta G < 0 \qquad \textit{Equation 6}$$

  Because the process results in a more ordered state,

  $$\mathrm{NaCH_3CO_2(aq) + 3H_2O(l) \rightarrow NaCH_3CO_2 \cdot 3H_2O(s)} \qquad \textit{Equation 7}$$

  the change in entropy, ΔS, is negative. This makes the value of $-T\Delta S$ positive. Since ΔG is negative, ΔH must also be negative and its absolute value must be greater than the value of $-T\Delta S$. The crystallization reaction is highly exothermic. Depressing a metal button in the solution pouch essentially starts a chain reaction that causes the entire solution to crystallize. The liquid becomes a solid and releases so much heat that it "freezes"! This hand warmer is available from Flinn Scientific, Catalog No. AP1933.

- There are two conceptual hurdles students must overcome in thermochemistry—the difference between *heat* and *temperature* and the definition of the *system* versus the *surroundings*. For a reaction taking place in solution, students must realize that the liquid, the solvent, is not directly involved in the reaction. It is part of the medium, the surroundings. Reactions are generally classified as exothermic or endothermic based on the temperature change in the surroundings, which is opposite in sign to that of the system. Thus, if the temperature of the surroundings increases, it is because the energy of the system has decreased. The temperature of the system itself is often inaccessible.

- One of the more stubborn student misconceptions is the idea that if the reaction mixture gets cold, it must have lost heat, therefore the reaction must be exothermic. This misconception may be traced to a lack of understanding of the system versus the surroundings. The temperature change that is measured in a typical coffee-cup calorimeter experiment is that of the surroundings. A heat of solution experiment is probably more confusing on this point than a heat of neutralization or heat of combustion experiment, because water is involved in the reaction. Also, using the combined mass of the solute and the solvent in the heat equation to calculate the heat change tends to blur the traditional distinction between the reactants and products versus the solvent.

- Enthalpy is an abstract concept that is often difficult for students to understand. Most textbooks include diagrams of enthalpy versus "reaction coordinate" (reactants and products) that help students visualize the difference in the sign of ΔH for exothermic and endothermic reactions. Qualitative laboratory activities such as "Thermodynamics in a Bag" (Flinn Catalog No. AP4779) and "Discovering Instant Cold Packs" (Flinn Catalog No. AP6375) are very helpful in teaching enthalpy because they allow students to see and feel the real effects of enthalpy changes.

- The most important element for success in an inquiry-based activity is student preparation. Sufficient time should be alotted for students to think through the measurements that must be made, how they will be made, the variables that will influence the measurements, and how the variables can be controlled, if necessary.

Part III. Sample Data, Results, and Analysis

Answers to Pre-Lab Questions *(Student answers will vary.)*

1. When chromium chloride, $CrCl_2$, is dissolved in water, the temperature of the water decreases.
   a. Is the heat of solution exothermic or endothermic?
   b. Which is stronger—the attractive forces between water molecules and chromium and chloride ions, or the combined ionic bond strength of $CrCl_2$ and intermolecular forces between water molecules? Explain.
   a. The heat of solution is endothermic—the system absorbs heat from the surroundings.
   b. The energy released in the formation of hydrated ions is less than the energy required to break the ionic crystal lattice and intermolecular forces between water molecules.

2. A solution was formed by combining 25.0 g of solid A with 60.0 mL of distilled water, with the water initially at 21.4 °C. The final temperature of the solution was 25.3 °C. Calculate the heat released as the solid dissolved, q_{soln}, assuming no heat loss to the calorimeter (see Equation 1).

$$\begin{aligned} q_{soln} &= -(m \cdot s \cdot \Delta T) = -(25.0 + 60.0)\text{g} \times 4.18\text{ J/g}\cdot{}^\circ\text{C} \times (25.3 - 21.4)^\circ\text{C} \\ &= -(85.0\text{ g} \times 4.18\text{ J/g}\cdot{}^\circ\text{C} \times 3.9\ ^\circ\text{C}) \\ &= -1390\text{ J} \end{aligned}$$

3. In Question 2 above, the calorimeter was found to have a heat capacity of 8.20 J/°C. If a correction is included to account for the heat absorbed by the calorimeter, what is the heat of solution, q_{soln}?

$$q_{soln} = -(m \cdot s \cdot T + C_{cal}\,\Delta T) = 1390\text{ J} + [8.20\text{ J/}^\circ\text{C} \times (25.3 - 21.4)^\circ\text{C}]$$

$$q_{soln} = -(1390\text{ J} + 32\text{ J}) = -1420\text{ J}$$

4. The solid in Question 2 was aluminum sulfate, $Al_2(SO_4)_3$. Calculate the molar heat of solution, ΔH_{soln}, for aluminum sulfate. *Hint:* The units for molar heat of solution are kilojoules per mole (kJ/mole). First determine the heat released per gram of solid.

$$\Delta H_{soln} = -(1420\text{ J/25.0 g}) \times 342.15\text{ g/mol} = -19400\text{ J/mol} = -19.4\text{ kJ/mol}$$

Sample Data for Heat Capacity of the Calorimeter

Volume of deionized water, cold	100.0 mL
Temperature, cold water	22.5 °C
Volume of deionized water, hot	100.0 mL
Temperature, hot water	47.6 °C
Final temperature (T_{mix})	34.8 °C
Temperature change, cold water	12.3 °C
Temperature change, hot water	–12.8 °C
Enthalpy change, cold water, q_{cold}(J)	5146 J
Enthalpy change, hot water, q_{hot}(J)	–5356 J
Temperature change, calorimeter	12.3 °C
Enthalpy change, calorimeter, q_{cal}(J)	208 J
Calorimeter constant, (J/°C)	17 J/°C

Calorimeter Constant Calculation

$$C_{cal} = \frac{q_{calor}}{(T_{mix} - T_{initial})}$$

$$T_{avg} = (T_H + T_C)/2 = (47.6 + 22.5)^\circ\text{C}/2 = (70.1/2)^\circ\text{C} = 35.1\ ^\circ\text{C}$$

$$q_{calor} = -q_{water} = (\text{mass of water}) \times (\text{specific heat of water}) \times (T_{mix} - T_{avg})$$

Mass of water = 200 mL × 0.994 g/mL (density of water at T_{avg}, 35 °C)

$$q_{calor} = -199\text{ g} \times 4.18\text{ J/g}\cdot{}^\circ\text{C} \times (-0.25)^\circ\text{C}$$

$$q_{calor} = 208\text{ J}$$

$$C_{cal} = \frac{q_{calor}}{(T_{mix} - T_{initial})} = 208\text{ J/}(34.8 - 22.5)^\circ\text{C} = 17\text{ J/}^\circ\text{C}$$

Sample Data for Magnesium Sulfate Calorimetry Procedure

Magnesium Sulfate, Anhydrous	
Volume of deionized water	100.0 mL
Density H_2O	0.9980 g/mL
Mass of H_2O	99.80 g
Mass of $MgSO_4$	5.00 g
Initial temperature	21.4 °C
Final temperature	29.3 °C
Temperature change	7.9 °C

Molar Heat of Solution of Magnesium Sulfate:

$$\text{Molar heat of solution} = \left(\frac{q_{soln}}{\text{grams solute}}\right) \times \text{molar mass}$$

where $q_{soln} = -(q_{aq} + q_{cal})$

$$\text{Molar heat of soln} = \frac{-[(104.8\text{ g})(4.18\text{ J/g}\bullet{}^\circ\text{C})(7.9\ {}^\circ\text{C}) + (17\text{ J/}{}^\circ\text{C})(7.9\ {}^\circ\text{C})]}{5.0\text{ grams}} \times \frac{120.39\text{ g}}{\text{mole}} =$$

$$\frac{-(3461\text{ J} + 134\text{ J})}{5.0\text{ grams}} \times \frac{120.39\text{ g}}{\text{mole}} \times \frac{1\text{ kJ}}{1000\text{ J}} = -87\text{ kJ/mole (rounded to two significant figures)}$$

Answers to Guided-Inquiry Discussion Questions

1. Review the calorimetry procedure and answer the following questions:

   a. What data is needed to calculate the enthalpy change for a reaction?

   Data for the three terms involved in the heat energy equation ($q = m \times s \times \Delta T$) must be known or measured. The mass (m) is the mass of the solution after the solid has dissolved. The specific heat capacity (s) is assumed to be the same as the specific heat capacity of water (4.18 J/g·°C). The temperature change (ΔT) is equal to the difference between the final and initial temperatures ($T_{final} - T_{initial}$). ***Note to teachers:*** *Assuming the specific heat capacity of the solution is the same as that of water may be a major source of error in the heat calculations.*

   b. Identify the variables that will influence the experimental data.

   Some of the critical variables include: (1) the mass of the solute; (2) the volume (mass) of the solvent; (3) whether all of the solute dissolves in the solvent; (4) the heat-insulating properties of the reaction container; (5) how well the reaction mixture is stirred; (6) how stable the initial temperature reading is.

   c. What variables should be controlled (kept constant) during the procedure?

   The following variables should be held constant during the procedure: the volume (mass) of the solvent; the type of reaction container that is used (two insulating foam cups nestled one inside the other will provide better insulation than one cup); continuous stirring of the reaction mixture.

   d. The independent variable in an experiment is the variable that is changed by the experimenter, while the dependent variable responds to or depends on the changes in the independent variable. Name the independent and dependent variables in the calorimetry procedure.

   In a calorimetry experiment, the mass of the solute in grams is the independent variable and will be varied in different trials. The temperature change that is produced depends on the mass of the solute and is thus the dependent variable in a calorimetry experiment.

e. Discuss the factors that will affect the precision of the experimental results.

Many factors will influence the precision of the results:

- *The precision of the balance used to measure the mass of solute.*
- *The precision of the graduated cylinder used to measure the volume of solvent.*
- *The precision of the thermometer used to measure the temperature of the reaction mixture.*
- *The number of times the experiment is repeated to average the effects of random errors.*
- *The type of vessel that is used as the calorimeter—how much heat is gained or lost by the calorimeter itself.*

The first three measurements should be made with the most precise glassware and equipment available in the lab—centigram balances (at least), appropriate size graduated cylinders, and digital thermometers, if possible. One important way to improve the precision of the experimental results is to average data obtained over several runs or trials. A minimum of 2–3 trials is recommended. Alternatively, class data may be averaged to eliminate outlying results.

Sample Data and Results for Hand Warmer Design

Sample data to determine the heat of solution for each solid were determined using 5 g of solid and 45 mL of water.

Sodium Chloride, NaCl	
Volume of deionized water	45.0 mL
Mass of NaCl	5.038 g
Initial temperature	23.3 °C
Final temperature	21.9 °C
Temperature change	–1.4 °C

Calcium Chloride, $CaCl_2$	
Volume of deionized water	45.0 mL
Mass of $CaCl_2$	5.075 g
Initial temperature	22.2 °C
Final temperature	36.4 °C
Temperature change	14.2 °C

Sodium Acetate, $NaCH_3CO_2$	
Volume of deionized water	45.0 mL
Mass of $NaCH_3CO_2$	5.118 g
Initial temperature	23.4 °C
Final temperature	28.1 °C
Temperature Change	4.7 °C

Sodium Carbonate, Na_2CO_3	
Volume of deionized water	45.0 mL
Mass of Na_2CO_3	4.061 g
Initial temperature	22.9 °C
Final temperature	27.9 °C
Temperature change	5.0 °C

Lithium Chloride, LiCl	
Volume of deionized water	45.0 mL
Mass of LiCl	5.001 g
Initial temperature	23.7 °C
Final temperature	42.9 °C
Temperature Change	19.2 °C

Ammonium Chloride, NH_4Cl	
Volume of deionized water	45.0 mL
Mass of NH_4Cl	5.006 g
Initial temperature	23.9 °C
Final temperature	16.5 °C
Temperature change	–7.4 °C

Sample calculations to determine the heat of solutions (assume 1.000 g/mL as H_2O density):

For $CaCl_2$:

$$q_{soln} = -(q_{aq} + q_{cal})$$
$$q_{soln} = -(mC\Delta T + C_{cal}\Delta T)$$
$$q_{soln} = -[(50.075\text{ g})(4.18\text{ J/g}\bullet{}^\circ\text{C})(14.2\ ^\circ\text{C}) + (17\text{ J/}^\circ\text{C})(14.2\ ^\circ\text{C})]$$
$$q_{soln} = -(2972 + 241)\text{ J}$$
$$q_{soln} = -3213\text{ J}$$

$$\Delta H_{soln} = \frac{-3213\text{ J}}{5.075\text{ g}} \times \frac{1\text{ kJ}}{1000\text{ J}} \times \frac{111\text{ g}}{\text{mole}} = \frac{-70.2\text{ kJ}}{\text{mole}}$$

Solid	Temperature change, ΔT, °C	Calorimeter contents energy change, q_{aq} (J)	Calorimeter energy change, q_{cal} (J)	Internal Energy Change, q_{soln} (J)	ΔH_{soln} (kJ/mole)
NaCl	–1.4	–300	–24	324	3.8
$CaCl_2$	14.2	2972	241	–3213	–70.2
$NaCH_3CO_2$	4.7	986	80	–1065	–17
Na_2CO_3	5.0	1026	85	–1111	–29
LiCl	19.2	4017	327	–4343	–36.8
NH_4Cl	–7.4	–1547	–126	1673	18

Extrapolate from the information collected and estimate the temperature change for each solid when 10 g combine with 40 mL of water. The ratio of 50/45 is used to predict the expected temperature change (ΔT) with 40 mL of water because the total mass of water and solid is used in the heat energy equation (see the sample calculation). The observed ΔT was measured with approximately 50 g total mass (45 g water plus 5 g solid). The predicted ΔT corresponds to 45 g total mass (40 g water plus 5 g solid).

- ΔT, °C (45 mL) is the observed temperature change with 45 mL of water and experimental mass of solid.
- ΔT, °C (40 mL) is the estimated temperature change with 40 mL of water and experimental mass of solid.

  = 50/45 × [ΔT, °C (45 mL)]; for $CaCl_2$ = (50/45) × 14.2 °C = 16.0 °C
- 10 g/40 mL ΔT is the predicted temperature increase for a hand warmer containing 10 g of solid and 40 mL of water.

  = (10 g/experimental mass of solid) × [(ΔT, °C (40 mL)]

for $CaCl_2$ = (10/5.075) × 16.0 °C = 31.5 °C

Solid	Cost($)/g	Δ*T*, °C (45 mL)	Δ*T*, °C (40 mL)	10 g/40 mL Δ*T*	Total Cost of Solid
NaCl	0.0079	–1.4	–1.6	N/A	N/A
$CaCl_2$	0.0131	14.2	16.0	31.5	0.131
$NaCH_3CO_2$	0.0258	4.7	5.2	10.2	0.258
Na_2CO_3	0.0123	5.0	5.6	13.8	0.123
LiCl	0.0655	19.2	21.3	42.6	0.655
NH_4Cl	0.0219	–7.4	–8.3	N/A	N/A

Conclusion

The best all-around hand warmer would contain calcium chloride. It produces the required temperature change of at least 20 °C and is less expensive and less toxic than lithium chloride.

Answers to AP Chemistry Review Questions

Review the following data from a calorimetry experiment to determine the heat of fusion of ice. After shaking off any excess water, several ice cubes were added to 99 g of warm water contained in a calorimeter. The initial temperature of the warm water was 46.8 °C. The ice–water mixture was stirred until the temperature reached a stable, minimum value, which was 1.1 °C. Any *unmelted* ice remaining at this point was immediately and carefully removed using tongs and the mass of the water in the calorimeter was measured—154 g.

1. Use the heat energy equation to calculate the amount of heat in joules released by the warm water as it cooled.

   $\Delta T = T_{final} - T_{initial} = 1.1 - 46.8\ °C = -45.7\ °C$

   $q(\text{warm water}) = (4.18\ J/g{\cdot}°C) \times 99\ g \times (-45.7\ °C) = -18{,}900\ J$

2. Based on the law of conservation of energy, what amount of heat was absorbed by the ice as it melted?

   $q(\text{ice}) = -q(\text{warm water}) = +18{,}900\ J$

3. Determine the amount of energy absorbed per gram of ice as it melted.

   Mass of ice melted = 154 g – 99 g = 55 g

   Amount of energy absorbed per gram of ice as it melted = q(ice)/mass of ice = 18,900 J/55 g = 340 J/g

4. Calculate the heat of fusion (the heat required to melt ice) in units of kilojoules/mole.

   Heat of fusion (kJ/mole) = 340 J/g × 18 g/mole × 1 kJ/1000 J = 6.1 kJ/mole

5. The literature value for the heat of fusion of ice is 6.02 kJ/mole. What is the percent error for the experimentally determined heat of fusion?

   $$\text{Percent error} = \frac{|6.1 - 6.02|}{6.02} \times 100\% = 2\%$$

6. When a mixture of ice and water originally at 0 °C is heated, the temperature remains constant (within a few degrees Celsius) until all of the ice melts. Explain what happens to the heat energy that is absorbed during this time while the temperature does not change.

   The energy absorbed breaks down the attractive forces that hold the water molecules in the rigid structure of ice.

Reference

AP Chemistry Guided-Inquiry Experiments: Applying the Science Practices;* The College Board: New York, NY, 2013.

Designing a Hand Warmer—Advanced Inquiry Laboratory Kit and supporting materials are available from Flinn Scientific, Inc.

Catalog No.	Description
AP7654	Designing a Hand Warmer—Advanced Inquiry Laboratory Kit

Consult your *Flinn Scientific Catalog/Reference Manual* for current prices.

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Explain and Predict — Practice Free Response Questions for AP* Chemistry

Question 8

A sample of N_2O_5 gas is introduced into a cylinder with a piston. N_2O_5 decomposes and reaches equilibrium with its decomposition products according to the equation below: **(10 points)**

$$N_2O_5(g) \longleftrightarrow N_2O_3(g) + O_2(g)$$

a. Assume that N_2O_5 is brown and N_2O_3 is clear/colorless. How, via macroscopic observation, is it possible to determine the time necessary for the system to reach equilibrium? Justify your answer (**2 points**).

b. Write the K_p expression for the equilibrium (**1 point**).

c. The K_p for the equilibrium is 2.3×10^{-5}. Determine the partial pressures of all gaseous species present if the initial N_2O_5 pressure is 0.12 atm (**2 points**).

d. The piston is withdrawn slowly and then held in place. Will the value of K_p increase, decrease, or stay the same before returning to its original value? Justify your answer (**2 points**).

e. In another experiment, the flask is charged with N_2O_5, N_2O_3, and O_2 at the following partial, non-equilibrium pressures: 1.1 atm, 0.8 atm, and 2.4 atm, respectively. Predict whether the amount of gaseous oxygen in the container will increase, decrease, or stay the same. Justify your prediction (**2 points**).

f. Explain, with reference to microscale events, why the pressure of the system increases when the equilibrium shifts to the right (**1 point**).

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Photoelectron Spectroscopy

Evidence for Electronic Structure

Guided-Inquiry Learning Activity for AP* Chemistry

Introduction

The chemical properties of elements are based on the number of electrons in the neutral atom and the arrangement of electrons into shells and subshells reflecting specific or quantized energy levels. Photoelectron spectroscopy and ionization energy measurements provide direct evidence for the electronic structure of atoms.

Concepts

- Photoelectric effect
- Coulomb's law
- Electronic structure
- Planck's law
- Ionization energy
- Atomic orbitals
- Quantum theory
- Photoelectron spectroscopy

Background

Photoelectron spectroscopy is defined as the measurement of the relative number of electrons of different energies that are ejected from atoms when they are bombarded with high-energy electromagnetic radiation. Photoelectron spectra are generated based on the *photoelectric effect*, which was first observed for metals in the 19th century. When light of the appropriate wavelength and energy is shined on a metal, electrons may be given off or ejected from the metal surface. If the light is below a certain minimum or threshold frequency, no electrons are produced, regardless of the intensity of the light source. This apparent paradox could not be explained by the laws of classical physics. In 1905, Albert Einstein applied the new quantum theory of light to explain the photoelectric effect. Einstein's explanation of the nature of the photoelectric effect represented a watershed event in the history of science. The development of quantum physics and quantum mechanics in the years that followed ultimately transformed and led to our current understanding of atomic and electron structure.

Experiments with the photoelectric effect led to two major findings: 1) light must be above a certain frequency for electron emission to occur, and 2) as the frequency of light increases above this minimum, the average kinetic energy of the dislodged particles increases. These findings were interpreted by Albert Einstein using Planck's law. This law describes light in terms of photons possessing a *quantum* or bundle of energy, where the energy (E) of a photon is proportional to its frequency (v).

$$E = hv \qquad \textit{Planck's law}$$

The photoelectric effect is only observed when the frequency of light is greater than a specific threshold value, which depends on the nature of the metal. If the energy of a photon is greater than the work function, defined as the minimum energy needed to displace an electron from a metal, the metal will give up an electron. If more photons having this energy strike the metal, additional electrons will be lost, but their kinetic energy will not change. The kinetic energy of photoelectrons ejected from a metal depends on the frequency (energy) of the light source and the work function of the metal. Albert Einstein was awarded the Nobel Prize for Physics in 1921 for his work on the photoelectric effect, which bridged the gap between the laws of classical and quantum physics.

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Photoelectron spectroscopy (PES) utilizes the principles of the photoelectric effect to generate spectra revealing the electronic structure (energy levels) of atoms. A photoelectron spectrophotometer consists of three basic parts: a vacuum chamber for the sample, a radiation or light source, and an electron analyzer to separate electrons based on their kinetic energies (Figure 1).

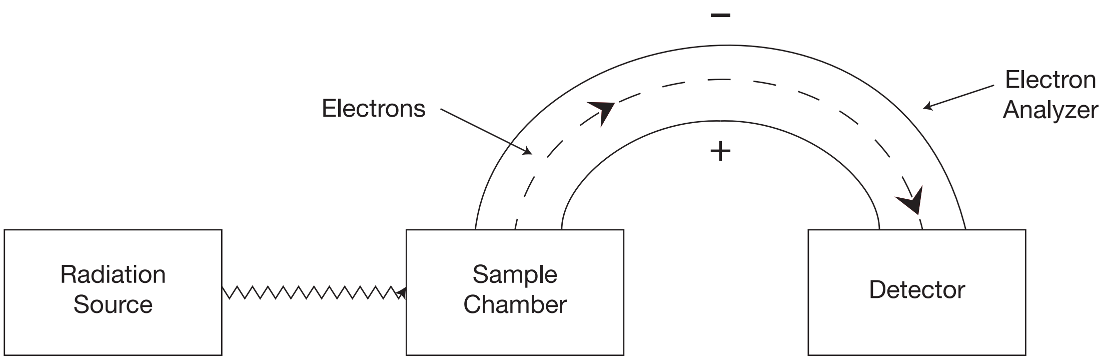


Figure 1. Basic features of a photoelectron spectrophotometer.

Samples of gas-phase atoms are exposed to electromagnetic radiation of sufficient energy, typically shortwave ultraviolet light or X-rays, to eject electrons from the atoms. Although each atom can emit only a single photoelectron using this technique, a sample always contains a multitude of atoms. Different atoms within the sample will emit electrons from different energy levels. The kinetic energy (KE) of the resulting photoelectrons is measured with an electron analyzer. PES uses monochromatic (single frequency) electromagnetic radiation. Based on the law of conservation of energy, the kinetic energy (KE) of a photoelectron is equal to the energy of the incident light ($h\nu$) minus the ionization energy (IE), which is the energy required to remove an electron from a particular energy level within an atom.

$$\text{KE} = h\nu - \text{IE} \qquad \textit{Equation 1}$$

From Equation 1, the ionization energies of the electrons within an atom can be calculated. Because the incident light is of sufficient energy to ionize both core and valence electrons, there is an equal probability that each electron in an atom will be ejected. The number of electrons ejected is proportional to the number of electrons present at each energy level. A *photoelectron spectrum* displays the signal intensity, corresponding to the relative number of electrons ejected at a particular energy, versus the ionization energy. For a spectrum displaying two or more peaks, a comparison of the peak heights allows for the relative number of electrons at each energy level to be determined.

Overview of the Activity

The purpose of this guided-inquiry learning activity is to examine experimental evidence for the electron configuration of neutral atoms. The activity is divided into three parts:

A. Review Coulomb's law to predict the relative ionization energies of electrons in an atom.

B. Investigate the basic principles of photoelectron spectroscopy and identify the way information is presented in a photoelectron spectrum.

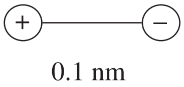
C. Analyze the photoelectron spectra of elements 1–20 to predict and explain their electron configurations.

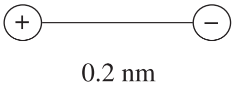
Part A. Coulomb's Law and the Ionization Energy of Electrons

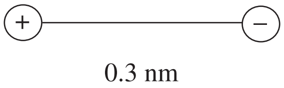
Analysis of atomic emission spectra leads to the conclusion that electrons in an atom reside in quantized energy levels or shells. The shells have discrete energies based on the distance separating the electrons from the nucleus and the charge of the nucleus. These energies may be qualitatively compared using Coulomb's law, which states that the force (F) experienced by two charged objects, q_1 and q_2, is inversely related to the square of the distance (r) separating them (Equation 2). If the two objects have opposite charges, then the force between them is attractive; if the charges are the same then the force is repulsive.

$$F \propto \frac{q_1 q_2}{r^2} \qquad \textit{Equation 2}$$

1. Rank the diagrams A, B and C representing the distances between a proton (+) and an electron (–) in order of smallest force of attraction to largest. Explain how you made your determination.

A.


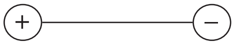
B.


C.


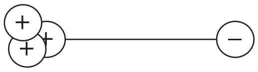
2. Which electron in diagrams A–C would require the most energy to be removed from the vicinity of the proton?

3. If the distance between the proton and electron in diagram C were increased to 0.60 nm, would you predict the force of attraction to be stronger or weaker than the original?

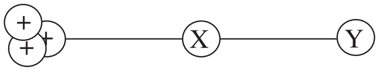
4. Consider the following diagrams representing different numbers of protons in an atomic nucleus. Which electron would require the most energy to be removed?

D.


E.


F.


5. Describe the types of forces experienced by electrons X and Y in the following diagram.



6. Which electron would require more energy to be removed from the system of charges?

Consider the following "shell diagram" for a hypothetical atom (Figure 2).

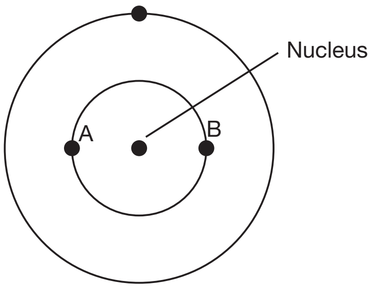


Figure 2. Shell diagram of electron structure.

7. Which electrons in Figure 2 are in the same energy level?

8. What type of force is experienced between the electrons and the nucleus?

9. What type of force is experienced between different electrons in Figure 2?

10. Compare the force of the nucleus on the electrons in both shells in Figure 2. Based on Coulomb's law, which electron(s) experience a greater force? Explain.

11. Predict which electron(s) would require the least amount of energy to be removed (ionized) from the atom. Explain your reasoning.

12. Suppose photoelectron spectroscopy yielded ionization energy values of 0.76 MJ/mol and 5.25 MJ/mol for the electrons in this hypothetical atom. How much energy, in MJ, is required to remove electron B from one mole of atoms? *Note:* 1 MJ = 1 Megajoule = 1 × 10^6 J.

Part B. Principles of Photoelectron Spectroscopy

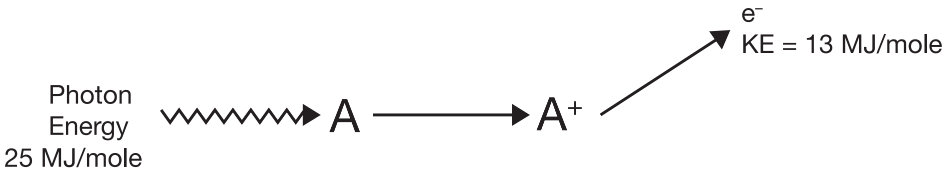


Figure 3. The photoelectric process.

1. Figure 3 represents the basic physical process that takes place in photoelectron spectroscopy. What is the energy per mole of the photon in this example?

2. What is the kinetic energy per mole of the photoelectron that is released from A?

3. Determine the ionization energy per mole of the ejected electron.

4. The photoelectron spectrum is usually plotted as shown below in Figure 4 for A. *Ionization energy* is on the *x*-axis versus *signal intensity* on the *y*-axis. Note the position of zero on the *x*-axis. *(a)* Write a sentence describing how the value of the ionization energy increases along the *x*-axis in a photoelectron spectrum. *(b)* Which direction on the *x*-axis corresponds to greater kinetic energy of an emitted photoelectron?

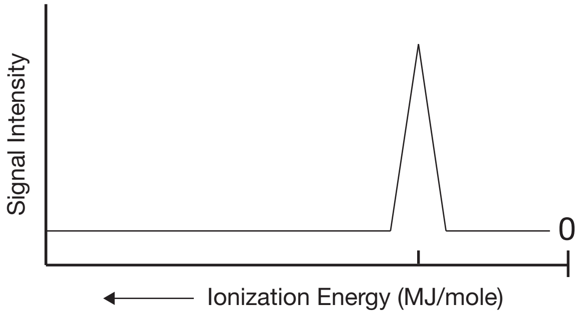


Figure 4. A simulated photoelectron spectrum.

5. What determines the position of the peak in the spectrum? What is the numerical value of the peak for atom (element) A?

6. How many energy levels are represented by the photoelectron spectrum in Figure 4? Explain your reasoning.

7. Is it possible to determine how many electrons are present in A based on this spectrum? Explain.

8. The simulated photoelectron spectrum for a second hypothetical element is shown below. What does the number of peaks indicate about the energy levels for the electrons in this element?

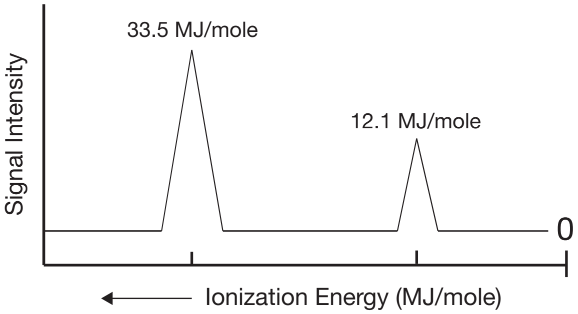


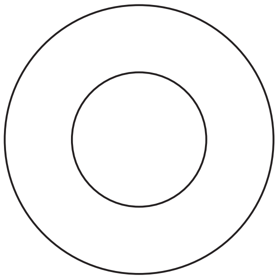
Figure 5. Simulated photoelectron spectrum.

9. What are the ionization energies for the electrons in Figure 5?

10. Which ionization energy value corresponds to an electron that is closer to the nucleus?

11. Compare the peak heights for the signals in this spectrum. Predict the relative number of electrons per atom originating from each energy level.

12. Complete the following shell diagram for this hypothetical element based on the photoelectron spectrum in Figure 5.



Part C. Photoelectron Spectra (PES) and Electron Configurations of the Elements

Table 1 summarizes PES data for elements 1–4 (hydrogen through beryllium).

Table 1.

Element	**Ionization Energy (MJ/mol)**	
	Peak A	**Peak B**
H	1.31	——
He	2.37	——
Li	6.26	0.52
Be	11.5	0.90

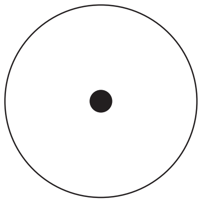
1. Compare the ionization energy values for hydrogen and helium. What is the main factor accounting for the difference in ionization energy of these two elements?

2. Helium has two electrons but only one peak in its photoelectron spectrum. Explain.

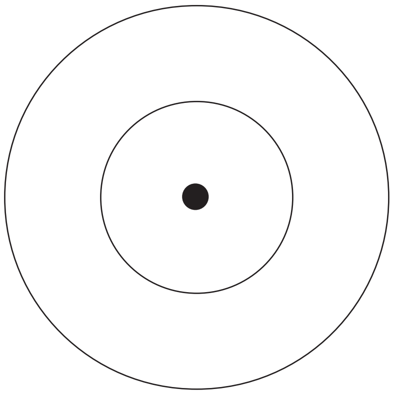
3. For each element in Table 1, the electrons giving rise to Peak A are all in the same energy level ($n = 1$). Explain how the charge on the nucleus and the distance of these electrons from the nucleus might (or might not) account for the trend in these ionization energy values for elements 1–4.

4. For each element, identify the peak corresponding to the first (lower) ionization energy. Account for the difference in energy for the first ionization of elements 1–4.

5. Compare the ionization energy for Peak A in the photoelectron spectrum of hydrogen versus Peak B for lithium. Draw simple shell diagrams to illustrate the difference in ionization energy for these peaks.



Hydrogen



Lithium

6. The photoelectron spectrum of lithium is shown in Figure 6. Identify the peak representing the first ionization energy of lithium.

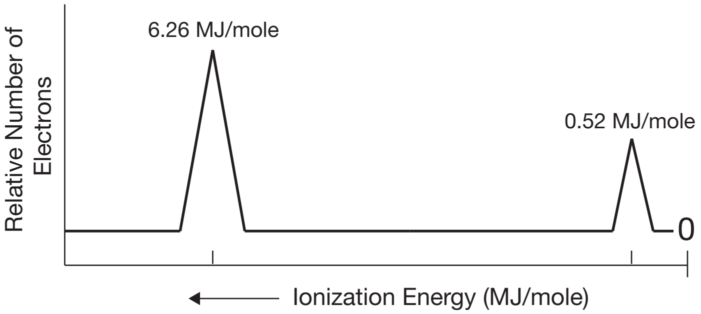


Figure 6. Photoelectron spectrum of lithium.

7. Compare the energy and relative height of the peaks in Figure 6. Write out the accepted electron configuration for lithium and assign the peaks in the PES of lithium to these atomic orbitals.

8. Consider the element boron with five electrons in a neutral atom. Predict the number of peaks, and their relative height, in the photoelectron spectrum of boron.

9. The actual photoelectron spectrum of boron is shown below. Describe any similarities and differences in the information obtained from this spectrum with the predictions you made in Question 8.

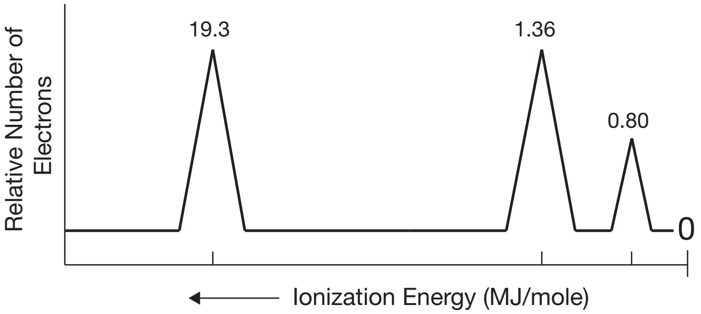


Figure 7. Photoelectron spectrum of boron.

10. The $n = 2$ energy level of an atom consists of two subshells (2s and 2p orbitals). What does the presence of a third peak in the PES of boron indicate about its electronic structure? Use the terms shells and subshells in your answer, and be as specific as possible concerning the difference in energy of the three peaks, as well as their relative height.

11. Assign each peak in Figure 7 to electrons and atomic orbitals in the accepted electron configuration of boron.

12. The 1s and 2s orbitals can each accommodate two electrons. How many electrons can occupy the 2p subshell in an atom?

13. Complete Table 2 to predict the number of peaks, and the relative number of electrons, for the remaining elements in row 2 of the periodic table. Arrange the peaks from lowest to highest ionization energy to show the relative number of electrons.

Table 2.

Element	Atomic Number	Number of Peaks	Relative Number of Electrons (from lowest to highest IE)
Li	3	2	1:2
Be			
B	5	3	1:2:2
C			
N			
O			
F			
Ne			

14. Compared to the elements in Table 2, the PES of sodium (element 11) contains a new, fourth peak. Would you expect this peak to have a lower or higher ionization energy value than the lowest IE peak for neon? Explain based on the arrangement of elements in the periodic table and the shell structure of atoms.

Table 3 summarizes atomic orbital assignments for the ionization energy values obtained from the photoelectron spectra of elements 11–20.

Table 3.

Element	Ionization Energy (MJ/mol)					
	1s	**2s**	**2p**	**3s**	**3p**	**4s**
Na	104	6.84	3.67	0.50		
Mg	126	9.07	5.31	0.74		
Al	151	12.1	7.19	1.09	0.58	
Si	178	15.1	10.3	1.46	0.79	
P	208	18.7	13.5	1.95	1.06	
S	239	22.7	16.5	2.05	1.00	
Cl	273	26.8	20.2	2.44	1.25	
Ar	309	31.5	24.1	2.82	1.52	
K	347	37.1	29.1	3.93	2.38	0.42
Ca	390	42.7	34.0	4.65	2.90	0.59

15. Based on the ionization energy data in Table 3, which orbitals appear to be closer in absolute energy—a 1s and 2s orbital, or a 2s and 3s orbital? Explain based on energy as a function of distance from the nucleus.

16. The potential energies of electrons in an atom are always negative, meaning they are more stable than isolated or "free" electrons. This leads to an apparently contradictory statement that "electrons that are higher in energy have lower ionization energies." Write a short 1–3 sentence explanation for ionization energy that eliminates potential confusion about these terms.

17. The relative energies for the atomic orbitals of Mg are shown below. Draw lines to show the approximate relative energies of the 1s, 2s, 2p, 3s, and 3p orbitals of Al.