

Identifying Functional Groups in Common Organic Materials

Provide the molecular formula, common name, IUPAC name, and list the functional groups present for each of the following organic compounds. You may want to refer to Appendix C2 — Common Chemicals, or search through Unit 1 of your text.

	Formula	Common name/use	IUPAC name	Functional groups present
1.	$\text{CH}_4(\text{g})$			
2.			methanol	
3.		lighter fluid		
4.			ethyne	
5.	$\text{C}_2\text{H}_4(\text{g})$			
6.		beverage alcohol		
7.			benzene	
8.		vinegar		
9.		formaldehyde		
10.	$\text{C}_6\text{H}_5\text{OH}(\text{l})$			
11.			propanone	
12.			octane	
13.		xylene solvents		
14.	$(-\text{C}_2\text{F}_4-)_n$			
15.			dichlorodifluoromethane	
16.		radiator antifreeze		
17.			Polyethene	
18.	$\text{C}_6\text{H}_{12}\text{O}_{6(\text{s})}$			
19.	X	Vitamin C		
20.	X		acetylsalicylic acid	

Note: There are more common names and uses than listed here.

Identifying Functional Groups in Common Organic Materials, Solution

Provide the molecular formula, common name, IUPAC name, and list the functional groups present for each of the following organic compounds. You may want to refer to Appendix C2 — Common Chemicals or search through Unit 1 of your text.

	Formula	Common name/use	IUPAC name	Functional groups present
1.	$\text{CH}_4(\text{g})$	natural gas	methane	none
2.	$\text{CH}_2\text{OH}(\text{l})$	ww & gl antifreeze; methyl hydrate	methanol	C-OH, hydroxyl group
3.	$\text{C}_4\text{H}_{10(\text{g})}$	lighter fluid	butane	none
4.	$\text{C}_2\text{H}_{2(\text{g})}$	acetylene (torch)	ethyne	carbon–carbon triple bond
5.	$\text{C}_2\text{H}_{4(\text{g})}$	ethylene, monomer	ethene	carbon–carbon double bond
6.	$\text{C}_2\text{H}_5\text{OH}(\text{l})$	beverage alcohol	ethanol	C-OH, hydroxyl group
7.	$\text{C}_6\text{H}_{6(\text{l})}$	in gasoline; solvent	benzene	alternating carbon–carbon double bonds
8.	$\text{CH}_3\text{COOH}_{(\text{aq})}$	vinegar	ethanoic (acetic) acid	-COOH, carboxyl group
9.	$\text{H}_2\text{CO}(\text{l})$	formaldehyde	methanal	C=O, carbonyl group
10.	$\text{C}_6\text{H}_5\text{OH}(\text{l})$	phenol; in spring H₂O	hydroxybenzene	C-OH, hydroxyl group
11.	$(\text{CH}_3)_2\text{CO}(\text{l})$	acetone; solvent	propanone	C=O, carbonyl group
12.	$\text{C}_8\text{H}_{18(\text{l})}$	gasoline component	octane	none
13.	$\text{C}_6\text{H}_4(\text{CH}_3)_2(\text{l})$	xylene solvents	#, #-dimethylbenzene	carbon–carbon double bonds
14.	$(-\text{C}_2\text{F}_4-)_n$	Teflon®	polytetrafluoroethylene	C-F
15.	$\text{CCl}_2\text{F}_{2(\text{g})}$	CFC; refrigerant	dichlorodifluoromethane	C-Cl and C-F
16.	$\text{C}_2\text{H}_4(\text{OH})_{2(\text{l})}$	radiator antifreeze ethylene glycol	1,2-ethanediol	two C-OH, hydroxyl groups
17.	$(-\text{C}_2\text{H}_4-)_n$	polyethylene; plastic	polyethene	none
18.	$\text{C}_6\text{H}_{12}\text{O}_{6(\text{s})}$	by photosynthesis	glucose	C=O, carbonyl group, C-OH, hydroxyl groups
19.	X	Vitamin C	ascorbic acid	-COOH, carboxyl group
20.	X	ASA; pain killer	acetylsalicylic acid	-COOH, carboxyl group

Note: There are more common names and uses than listed here.

Review of Hydrocarbons: Naming Aliphatic and Aromatic Hydrocarbons

1. For each of the following IUPAC names, draw a structural diagram.

(a) 2-methylpentane

(b) octane

(c) 2,2,3-trimethylpentane

(d) ethene

(e) propyne

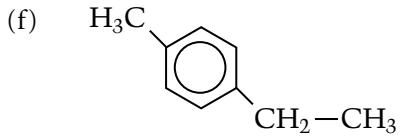
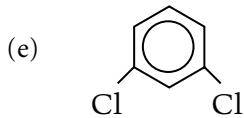
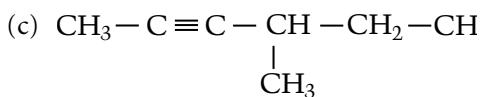
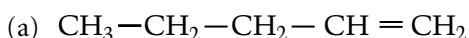
(f) methylpropyne

(g) cyclohexane

(h) 1,2-dimethylbenzene

(i) ethylbenzene

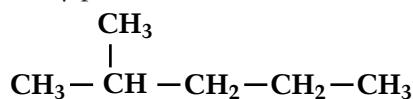
2. For each of the following structural diagrams, write the IUPAC name.



Review of Hydrocarbons: Naming Aliphatic and Aromatic Hydrocarbons, Solution

1. For each of the following IUPAC names draw a structural diagram.

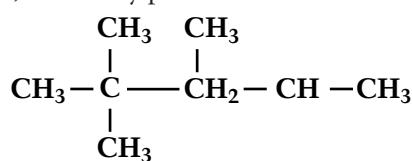
(a) 2-methylpentane



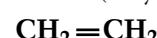
(b) octane



(c) 2,2,3-trimethylpentane



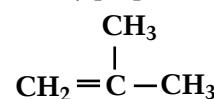
(d) ethene (ethylene)



(e) propyne



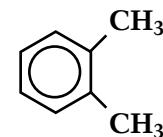
(f) methylpropene



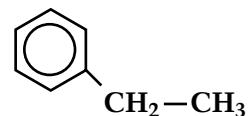
(g) cyclohexane



(h) 1,2-dimethylbenzene



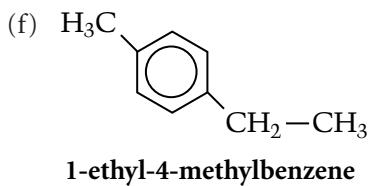
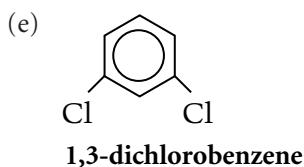
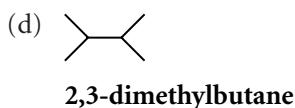
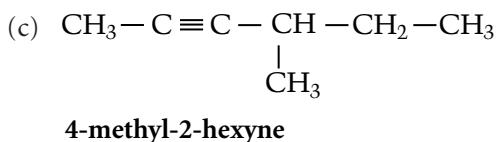
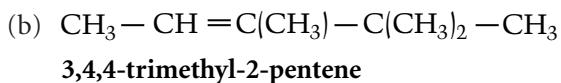
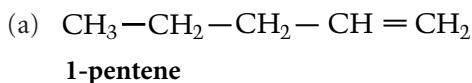
(i) ethylbenzene



(continued)

LSM 1.2-3

2. For each of the following structural diagrams, write the IUPAC name.



Try This Activity: Building Hydrocarbons

We can use molecular models to help us build, visualize, and manipulate organic compounds. From a molecular model kit, obtain 6 carbons and 14 hydrogens.

Pre-lab Questions

1. Define the term isomer.

2. What does each line in a structural diagram represent?

3. What does each stick in a molecular model represent?

4. What advantages do molecular models have over drawings?
 - (a) Build, draw, and name as many hydrocarbons as you can, using all pieces in each model. In a table like the one shown, record the molecular formula, draw the structural formula and give the name of the compound for all of your molecules

Molecular formula	Structural formula	Name of compound

- (b) Now, put away 2 hydrogen atoms, and build and name as many different structures as possible, using all of the remaining atoms. Again, record the information in a table like the one shown above.

(continued)

- (c) Repeat (b) as many times as possible, until you can no longer construct any hydrocarbon molecules.
- (d) From your information, list all the structural isomers that you built.

Post-Lab Questions

5. How many structural isomers are possible for methane, ethane, propane, butane, and pentane?

Methane:

Ethane:

Propane:

Butane:

Pentane:

6. Is cyclopentane an isomer of pentane? Explain.

7. What are the bond angles in ethane? ethene? ethyne?

Reactions of Hydrocarbons, Extra Exercises

Write an equation for each of the following reactions. Use molecular and structural formulas and classify the reaction as combustion, addition, substitution, hydrogenation, or hydration.

(a) Octane burns with oxygen gas.

(b) 2-methyl-1-pentene reacts with hydrogen.

(c) Acetylene and oxygen react.

(d) 3-methyl-1-butyne reacts with excess hydrogen.

(e) Chlorine reacts with ethane.

(f) Water and ethane are placed in concentrated sulfuric acid with a catalyst.

(continued)

(g) Bromine combines with benzene.

(h) Propene and HBr combine.

(i) Cyclohexene reacts with hydrogen gas.

(j) Acetylene reacts with excess hydrogen chloride.

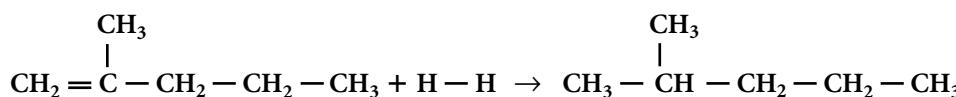
Reactions of Hydrocarbons, Extra Exercises, Solution

Write an equation for each of the following reactions. Use molecular and structural formulas and classify the reaction as combustion, addition, substitution, hydrogenation, hydrohalogenation, or hydration.

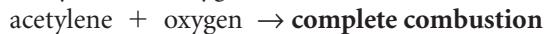
- (a) Octane burns with oxygen gas.



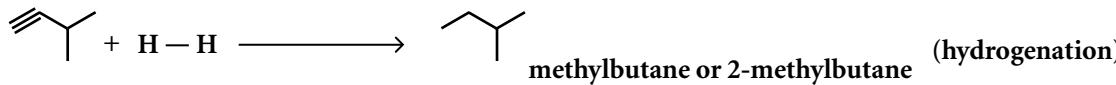
- (b) 2-methyl-1-pentene reacts with hydrogen.



- (c) Acetylene and oxygen react.



- (d) 3-methyl-1-butyne reacts with excess hydrogen.



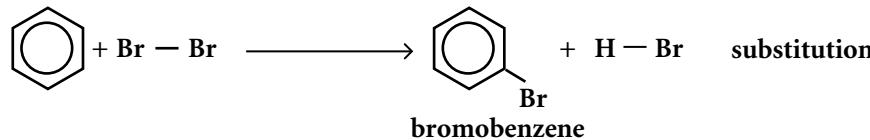
- (e) Chlorine reacts with ethane.



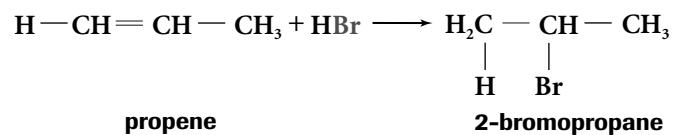
- (f) Water and ethane are placed in concentrated sulfuric acid with a catalyst.



- (g) Bromine combines with benzene.



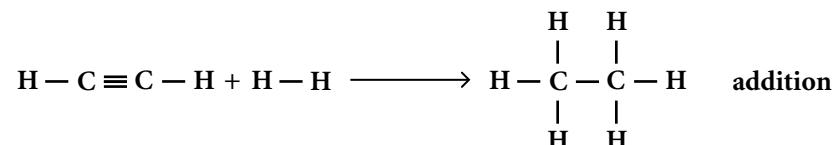
(continued)

LSM 1.3-2(h) Propene and HBr combine. addition - hydrohalogenation

(i) Cyclohexene reacts with hydrogen gas.



(j) Acetylene reacts with excess hydrogen chloride.



Identification and Naming of Organic Compounds, Extra Exercises

1. In the following questions, the IUPAC names of a variety of organic compounds are provided. Draw a structural diagram for each compound and identify the organic family to which it belongs.

(a) 2-bromopentane

(b) 1,4-dichlorobenzene

(c) butanoic acid

(d) ethyl methanoate

(e) 1-butanol

(f) propanal

(g) 1,1 dichloro-2,2-difluoroethane

(h) trimethylamine

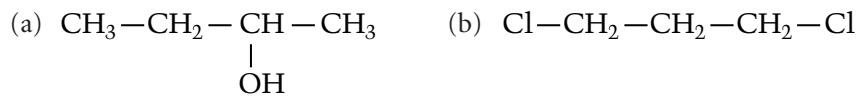
(i) 2-methyl-2-propanol

(j) propanamide

(continued)

LSM 1.4-1

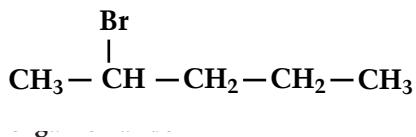
2. Give the IUPAC name for the following organic compounds represented by the given structural formula.



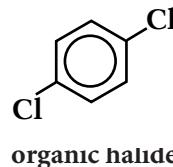
Identification and Naming of Organic Compounds, Extra Exercises, Solution

1. In the following questions, the IUPAC names of a variety of organic compounds are provided. Draw a structural diagram for each compound and identify the organic family to which it belongs.

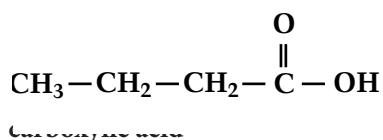
(a) 2-bromopentane



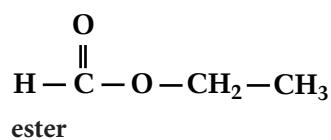
(b) 1,4-dichlorobenzene



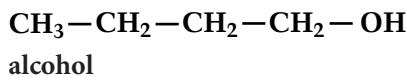
(c) butanoic acid



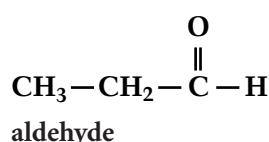
(d) ethyl methanoate



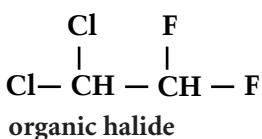
(e) 1-butanol



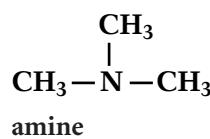
(f) propanal



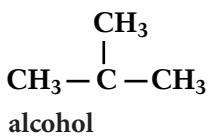
(g) 1,1 dichloro-2,2-difluoroethane



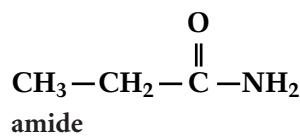
(h) trimethylamine



(i) 2-methyl-2-propanol

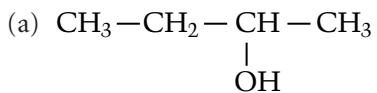


(j) propanamide

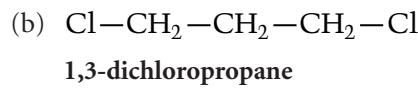
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LSM 1.4-2

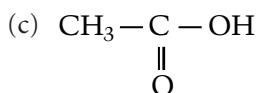
2. Give the IUPAC name for the following organic compounds represented by the given structural formula.



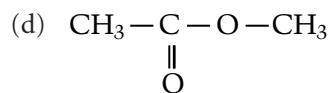
2-butanol



1,3-dichloropropane



ethanoic acid



methyl ethanoate

Organic Reactions, Extra Exercises

For each of the following questions, state the organic reaction type, draw structural diagrams for all reactants and products, and name all organic products. Do not balance the equations.

1. Propane reacts with fluorine.
 2. Chloroethane reacts with hydroxide ions.
 3. Ethanol, present in gasohol, burns in an automobile engine.
 4. 1-butanol reacts in the presence of concentrated sulfuric acid.
 5. Chloromethane is produced by reacting methane and chlorine.
 6. Bromine and ethene react to form an alkyl halide.

(continued)

LSM 1.5-1

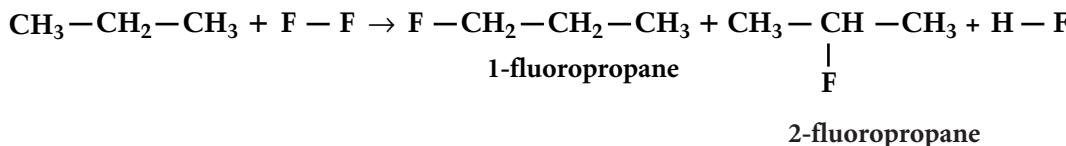
7. Hydrogen chloride and ethene react to produce an alkyl halide.

8. Acetic acid and ethanol react to produce a solvent used in nail polish remover.

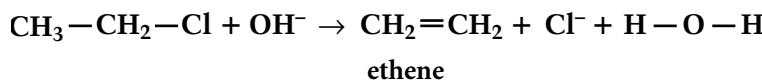
Organic Reactions, Extra Exercises, Solution

For each of the following questions, state the organic reaction type, draw structural diagrams for all reactants and products, and name all organic products. Do not balance the equations.

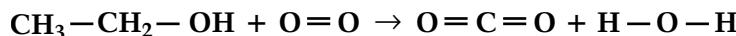
1. Propane reacts with fluorine.
substitution



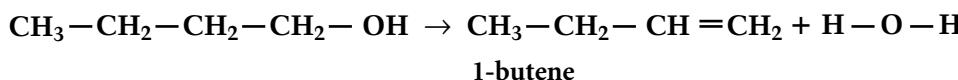
2. Chloroethane reacts with hydroxide ions.
elimination



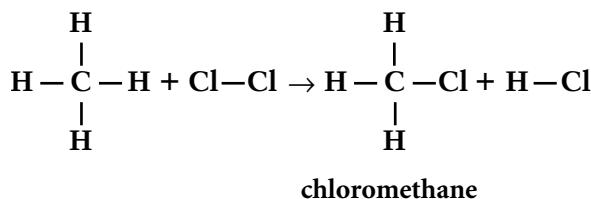
3. Ethanol, present in gasohol, burns in an automobile engine.
combustion



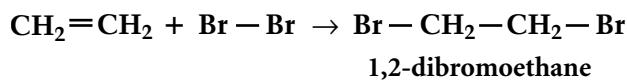
4. 1-butanol reacts in the presence of concentrated sulfuric acid.
elimination



5. Chloromethane is produced by reacting methane and chlorine.
substitution



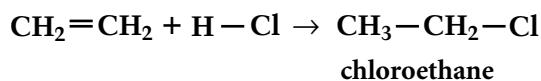
6. Bromine and ethene react to form an alkyl halide.
addition



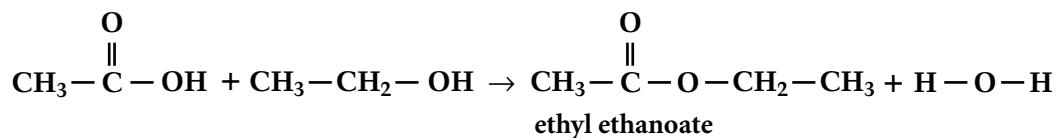
(continued)

LSM 1.5-2

7. Hydrogen chloride and ethene react to produce an alkyl halide.
addition



8. Acetic acid and ethanol react to produce a solvent used in nail polish remover.
condensation (esterification)



Try This Activity: How Many Can You Build?

1. From a molecular model kit, obtain 2 carbons, 6 hydrogens, and 1 oxygen.
Build two different structures that use all 9 atoms.

(a) In the table below, name these compounds and their functional groups.

Structural formula	Name	Functional group

2. Obtain one additional carbon atom and build as many different structures as possible, using all 10 atoms.

(b) In the table below, name these compounds and their functional groups.

Structural formula	Name	Functional group

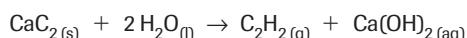
(c) Which of the compounds, listed in the tables, are isomers?

Lab Exercise 1.3.1: Preparation of Ethyne

INQUIRY SKILLS

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|---|---|
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| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input checked="" type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input type="radio"/> Communicating |
| <input type="radio"/> Conducting | |

Ethyne is the simplest and most widely used alkyne. It can be produced from the reaction of calcium carbide with water. Calcium hydroxide is the only other product.



Knowing the quantity of reactants and the balanced chemical equation for a reaction, we can predict the theoretical yield of a product. How close does the actual yield come to this theoretical yield? The difference in these two calculations can be used to assess the purity of the reactants.

Purpose

The purpose of this lab exercise is to determine the yield of ethyne from a synthesis procedure.

Question

What mass of ethyne is produced from a known mass of calcium carbide, and what is the purity of the calcium carbide?

Prediction

- (a) Predict the mass of ethyne and calcium hydroxide produced from the hydration reaction of 6.78 g of calcium carbide.

Experimental Design

A known mass of calcium carbide is reacted with an excess of water. The amount of calcium hydroxide formed is determined by titration of the entire volume of $\text{Ca(OH)}_{2\text{(aq)}}$ with 1.00 mol/L HCl. The amount of ethyne produced is then calculated stoichiometrically.

(continued)

Evidence

mass of CaC_2 reacted: 6.78 g

volume of 1.00 mol/L $\text{HCl}_{(\text{aq})}$ added to neutralize the $\text{Ca}(\text{OH})_2$ solution: 100.0 mL

Analysis

(b) Write a balanced chemical equation for the neutralization reaction between $\text{Ca}(\text{OH})_{2(\text{aq})}$ and $\text{HCl}_{(\text{aq})}$.

(c) Calculate the amount of $\text{Ca}(\text{OH})_{2(\text{aq})}$ actually produced in the reaction of $\text{CaC}_{2(\text{aq})}$ and water.

(d) Calculate the actual yield of ethyne, in grams.

(e) Calculate the percent purity of the CaC_2 .

Evaluation

(f) What assumptions did you make in your calculations?

Investigation 1.5.1: Comparison of Three Isomers of Butanol

INQUIRY SKILLS

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| <input checked="" type="radio"/> Conducting | |

The reactivity of alcohols can be accounted for by their molecular structure—particularly by the attachment of their hydroxyl functional group. The isomers of butanol are used as examples of 1°, 2°, and 3° alcohols to examine this relationship.

Purpose

The purpose of this investigation is to test our theories of how the molecular structure of an organic molecule affects its properties. To do this, we will determine and compare the chemical properties of three isomers of butanol.

Question

Does each alcohol undergo halogenation and controlled oxidation?

Prediction

- (a) Make predictions about your observations, with reasons.

Experimental Design

Each of the three isomers of butanol is mixed with concentrated $\text{HCl}_{(\text{aq})}$. The presence of an alkyl halide product is indicated by cloudiness of the mixture, as the halides are only slightly soluble in water. Each alcohol is also mixed with dilute $\text{KMnO}_{4(\text{aq})}$ solution, which provides conditions for controlled oxidation. Any colour change of the permanganate solution indicates that an oxidation reaction has taken place.

Materials

- lab apron
eye protection
gloves
1-butanol (pure)
2-butanol (pure)
2-methyl-2-propanol (pure)
concentrated $\text{HCl}_{(\text{aq})}$ (12 mol/L)
3 test tubes
test-tube rack
4 eye droppers
 $\text{KMnO}_{4(\text{aq})}$ solution (0.01 mol/L)
10-mL graduated cylinder



Concentrated hydrochloric acid is corrosive and the vapour is very irritating to the respiratory system. Avoid contact with skin, eyes, clothing, and the lab bench. Wear eye protection and a laboratory apron.



All three alcohols are highly flammable. Do not use near an open flame.

(continued)

Procedure

1. Place 3 test tubes in a test-tube rack. Using a clean eye dropper for each alcohol, place 2 drops of 1-butanol in the first tube; in the second, place 2 drops of 2-butanol; and in the third, place 2 drops of 2-methyl-2-propanol.
2. Carry the test-tube rack and tubes to the fume hood, and use a clean pipet to add 10 drops of concentrated $\text{HCl}_{(\text{eq})}$ to each of the three test tubes. Shake the mixture very gently and carefully. Return to your lab bench with the test-tube rack and tubes.
3. Allow the tubes to stand for 1 min and observe for evidence of cloudiness.
4. Follow your teacher's instructions for the disposal of the contents of the test tubes and for cleaning the test tubes.
5. Set up three test tubes as described in step 1, this time using 4 drops of each alcohol.
6. To each tube, carefully add 2 mL of 0.01 mol/L $\text{KMnO}_{4(\text{aq})}$ solution. Shake the mixture carefully.
7. Allow the tubes to stand for 5 min, with occasional shaking. Observe and record the colour of the solution in each tube.

Evidence

- (b) Prepare a table in which to record the observed properties of the three alcohols. Include structural diagrams for each alcohol.

Analysis

- (c) Answer the Question.

Evaluation

- (d) Evaluate the theory that you used to make your predictions.

(continued)

Synthesis

- (e) How can the results of this investigation be accounted for in terms of intermolecular forces?
- (f) Write a structural diagram equation to represent the reaction between each alcohol and $\text{HCl}_{(\text{aq})}$. Where no reaction occurred, write the starting materials and the words “no reaction.”
- (g) Write a structural diagram equation to represent the controlled oxidation of each alcohol in $\text{KMnO}_{4(\text{aq})}$ solution. Where no reaction occurred, write the starting materials and the words “no reaction.”
- (h) Summarize in a few sentences the halogenation and controlled oxidation reactions of 1° , 2° , and 3° alcohols.

Investigation 1.5.2: Trends in Properties of Alcohols

INQUIRY SKILLS

- | | |
|--|--|
| <input type="radio"/> Questioning | <input checked="" type="radio"/> Recording |
| <input checked="" type="radio"/> Hypothesizing | <input type="radio"/> Analyzing |
| <input type="radio"/> Predicting | <input type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input checked="" type="radio"/> Conducting | |

We might expect to see a trend in properties within a chemical family, such as alcohols. Is there a link, in the first four primary alcohols, between molecular size and physical properties? In this investigation we will first use our knowledge of intermolecular forces of the hydrocarbon components and the hydroxyl functional group to predict trends. We will then test our predictions experimentally.

Purpose

The purpose of this investigation is to create a theoretical model of 1° alcohols, that will explain their properties.

Question

What are the melting points, boiling points, solubilities, and acidity of the three 1° alcohols?

Experimental Design

Reference sources are used to look up the melting point and boiling point of several alcohols. The solubility of each alcohol is determined by mixing with a non-polar solvent (cyclohexane) and with water, and observing miscibility. The acidity of each alcohol in aqueous solution is tested with litmus paper.

Materials

- lab apron
- blue and pink litmus paper
- eye protection
- 3 test tubes
- ethanol
- test-tube rack
- 1-propanol
- test-tube holder
- 1-butanol
- graduated cylinder (10 mL)
- cyclohexane



All three alcohols are highly flammable. Do not use near an open flame.

Procedure

1. Place 1 mL of each alcohol in a separate test tube. To each alcohol, add 1 mL of cyclohexane. Observe and record any evidence of miscibility of the two liquids.
2. Follow your teacher's instructions for the disposal of the contents of the test tubes and for cleaning the test tubes.
3. Repeat step 1, using water in place of cyclohexane.

(continued)

4. Before disposing of the contents of the test tubes, add a small piece of blue and red litmus paper to each mixture. Record the colour of the litmus paper.

Evidence

- (a) Prepare a table with the following headings: alcohol name, structural diagram, melting point, boiling point, solubility in cyclohexane, solubility in water, colour with litmus.

- (b) Draw in the structural diagrams of each alcohol.

- (c) Using a print or electronic reference source, find the melting point and boiling point of each of the alcohols listed, and record the information in your table.



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Analysis

- (d) Your completed table provides answers to the Question. Summarize the trends in properties that you discovered.

Synthesis

- (e) Use your answers in (d) to develop a model of the primary alcohols. Represent your model in some way, and hypothesize how its features affect its properties.

Investigation 1.7.1: Properties of Carboxylic Acids

INQUIRY SKILLS

- Questioning
- Hypothesizing
- Predicting
- Planning
- Conducting
- Recording
- Analyzing
- Evaluating
- Communicating



Figure 1 Stearic acid is used to harden soaps, particularly those made with vegetable oils, that otherwise tend to be very soft.

A carboxylic acid is identified by the presence of a carboxyl group. The physical properties and reactivity of carboxylic acids are accounted for by the combination of their polar functional group and their nonpolar hydrocarbon components. We will look at some of the following properties of carboxylic acids in this investigation: melting and boiling points; solubility; acidity; reaction with $\text{KMnO}_{4(\text{aq})}$; and reaction with $\text{NaHCO}_{3(\text{aq})}$. The carboxylic acids we will be using are ethanoic (acetic) acid and octadecanoic (stearic) acid (**Figure 1**).

Purpose

The purpose of this investigation is to test our theoretical prediction of some chemical properties of carboxylic acids.

Question

- Devise a question that this investigation will allow you to answer.

Prediction

- Use your theoretical knowledge of the structure and functional groups of carboxylic acids to predict answers to the Question. (In some cases, you will only be able to compare the acids and give relative answers.)

Experimental Design

Several properties of two carboxylic acids of different molecular size are determined and compared. The melting points and boiling points of acetic acid and stearic acid are obtained from reference resources. The solubility of each acid in polar and nonpolar solvents is determined by mixing each acid with water and with oil. The reactions, if any, of each acid with $\text{NaHCO}_{3(\text{aq})}$, a base, and with $\text{KMnO}_{4(\text{aq})}$, an oxidizing agent, are observed and compared.

Materials

- ethanoic acid (glacial acetic acid)
- dilute ethanoic acid (8% acetic acid, vinegar)
- octadecanoic acid (stearic acid, solid)
- water
- vegetable oil
- pH meter or universal indicator
- $\text{KMnO}_{4(\text{aq})}$ (0.01 mol/L aqueous solution)
- $\text{NaHCO}_{3(\text{aq})}$ (saturated aqueous solution)
- 2 test tubes
- test-tube holder
- test-tube rack
- pipet and bulb



Glacial acetic acid is corrosive. Avoid contact with skin and eyes. It is also volatile, so you must be careful to avoid inhalation. Wear eye protection and a laboratory apron.

(continued)

**Figure 2**

eye dropper
graduated cylinder (10 mL)

Procedure

1. Add 5 mL of water to one test tube and 5 mL of oil to another test tube. In the fume hood, using an eye dropper, add one drop of glacial acetic acid to each tube (**Figure 2**). Shake each tube very carefully to mix. Make and record observations on the miscibility of the contents of each tube.
2. Still in the fume hood, add a drop of pH indicator to each of the test tubes in step 1, or use a pH meter to measure the pH. Record the results.
3. Follow your teacher's instructions to dispose of the contents of each test tube, and clean the test tubes.
4. Repeat steps 1, 2, and 3, using a small amount of solid stearic acid (enough to cover the tip of a toothpick). These steps do not need to be performed in the fume hood.
5. Place about 2 mL of saturated $\text{NaHCO}_3\text{(aq)}$ solution in each of two test tubes. Add 2 mL of dilute acetic acid to one tube, and a small amount of solid stearic acid to the other tube. Shake the tubes gently to mix and observe for formation of bubbles. Record your observations.
6. Place about 2 mL of $\text{KMnO}_4\text{(aq)}$ solution in each of two test tubes. Add 2 mL of dilute acetic acid to one tube, and a small amount of solid stearic acid to the other tube. Shake the tubes gently to mix and observe for any change in colour. Record your observations.

Evidence

- (c) Prepare a table to record structural diagrams, molar mass, and the properties under investigation for acetic acid and for stearic acid. Complete the information in the table for acid name, structural diagram, and molar mass.
- (d) Using print or electronic resources, obtain the melting point and boiling point for acetic acid and stearic acid, and record them in the table.



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(continued)

Analysis

- (e) Your table of evidence should contain your answer to the Question. Compare the melting points and boiling points of acetic acid and stearic acid. Account for the differences in these properties in terms of the molecular structure and intermolecular forces of each acid.
- (f) Compare the solubilities of acetic acid and stearic acid in water and in oil. Account for any differences in their solubilities in terms of molecular structure and intermolecular forces.
- (g) Do acetic acid and stearic acid appear to be organic acids in this investigation? Explain why or why not, with reference to experimental reactants and conditions.
- (h) Write an equation to illustrate any reaction between the acids and $\text{NaHCO}_3\text{(aq)}$.
- (i) Do acetic acid and stearic acid undergo controlled oxidation reactions? Explain why or why not. Draw a structural diagram equation to illustrate your answer.

Evaluation

- (j) Did the Experimental Design enable you to collect appropriate evidence?
- (k) Compare the answer obtained in your Analysis to your Prediction. Account for any differences.
- (l) Did your theoretical model of carboxylic acids enable you to correctly predict the chemical properties of these acids? Give reasons for your answer.

Activity 1.7.2: Synthesis of Esters

Many esters are found naturally in fruits, and are responsible for some of their pleasant odours. Synthetic esters are produced from condensation reactions between alcohols and carboxylic acids. In this activity you will synthesize, and detect the odours of, several esters.

Materials

lab apron	test-tube rack
eye protection	wax pencils
ethanol	test-tube holder
2-propanol	500-mL beaker
1-pentanol	hot plate
glacial acetic acid	pipet and bulb
graduated cylinder	evaporating or petri dish
2 mL concentrated $\text{H}_2\text{SO}_4(\text{aq})$	balance
3 test tubes	



Glacial acetic acid is corrosive. Avoid contact with skin and eyes. It is also volatile, so you must be careful to avoid inhalation. Wear eye protection and a laboratory apron.



All three alcohols and cyclohexane are highly flammable. Do not use near an open flame.

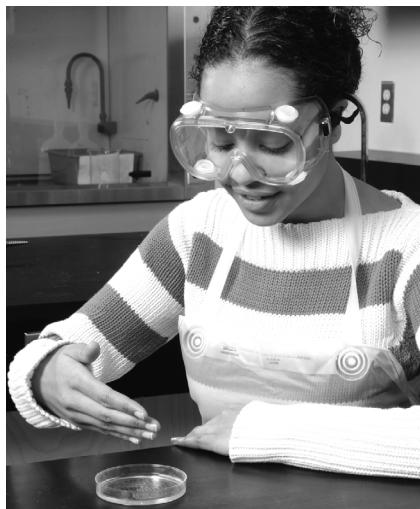


Figure 3

Table 1 Contents of Test Tubes for Synthesizing Esters

Contents	Tube #1	Tube #2	Tube #3
Alcohol (1 mL)	ethanol	2-propanol	1-pentanol
Acid (1 mL or 1 g)	glacial acetic acid	glacial acetic acid	glacial acetic acid
Catalyst (0.5 mL)	conc. $\text{H}_2\text{SO}_4(\text{aq})$	conc. $\text{H}_2\text{SO}_4(\text{aq})$	conc. $\text{H}_2\text{SO}_4(\text{aq})$

- Carefully return the test tubes, to your lab bench, and place all three of them in the hot-water bath. Be careful not to point the test tubes at anybody. After about 5 min of heating, remove the test tubes from the heat and put them back in the rack.
- Pour the contents of the first test tube into an evaporating dish half filled with cold water, and identify the odour of the ester as instructed by your teacher (Figure 3). Repeat for each ester.

- (a) For each of the three reactions, identify the odours of the esters.

(continued)

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- (b) Draw structural diagram equations to represent each of the three esterification reactions in this investigation. Write the IUPAC name of each reactant and product.
- (c) What was the function of the concentrated sulfuric acid in these reactions?
- (d) What evidence is there that the carboxylic acids used in this investigation are soluble or insoluble in aqueous solution? Explain this evidence in terms of molecular structure of the acids.
- (e) What evidence is there that the esters synthesized in this investigation are soluble or insoluble in aqueous solution? Explain this evidence in terms of molecular structure of the esters.

Activity 1.9.1: Building Molecular Models to Illustrate Reactions

In this activity, you will build molecular models of organic compounds and use them to illustrate a variety of chemical reactions. This should help you to see the differences in molecular structure between reactants and products. Your teacher may assign you to work in pairs or small groups.

Materials

molecular model kit

Part 1: Plan Your Own Reactions

- (a) Copy **Table 2** and fill in the reactant and product columns with the chemical names and formulas of each compound as you complete the molecular models for each reaction. Supply the name of any missing reaction type.
- For the first reaction indicated in **Table 2**, build a molecular model of a reactant (alkane) of your choice, and its product.
 - Compare the models of the reactants and products, and examine how the structure changes as a result of the reaction.
 - Disassemble the models and repeat the process for the next reaction. Continue for all the reactions listed.
- (b) Classify each of the named substances as aliphatic, cyclic, or aromatic.

Table 2 Reactions to be Modelled

Reaction #	Reactant	Reaction type	Product
1	alkane	substitution	
2	alkene	addition	
3	alcohol		aldehyde
4			ether
5	aldehyde	controlled oxidation	
6			chlorobenzene
7		controlled oxidation	ketone
8		esterification	
9			amide
10	an aromatic compound		aromatic halide

(continued)

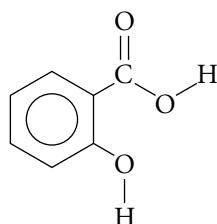
Part 2: Modelling the Synthesis of Acetic Acid

(c) Draw a flow chart to show the synthesis of acetic acid, starting with ethene.

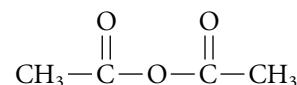
- Build a molecular model of ethene and make changes to the model to illustrate the reactions in the flow chart, until you finally make a model of acetic acid.

Activity 1.9.2: Preparation of an Ester — Aspirin

The analgesic properties of willow bark led chemists to the isolation of its active ingredient, acetylsalicylic acid, commonly called Aspirin or ASA. ASA is synthesized from the reaction between salicylic acid and acetic anhydride, the structures of which are shown. In this activity, you will prepare Aspirin and calculate the percentage yield.



salicylic acid



acetic anhydride

Materials

- lab apron
- eye protection
- 4 g salicylic acid
- 8 mL acetic anhydride
- 4 drops concentrated $\text{H}_2\text{SO}_4\text{(aq)}$
- 250-mL conical flask
- 600-mL beaker
- eye dropper
- glass stirring rod
- hot plate
- 20 g crushed ice
- ice water
- ice bath (large container of crushed ice and water)
- filter funnel
- ring stand with ring clamp
- filter paper
- balance
- Prepare a hot-water bath by heating 300-mL of water in a 600 mL beaker on a hot plate until boiling.
- Obtain approximately 4 g of salicylic acid and determine its mass to the nearest 0.01 g. Transfer the sample to a 250-mL conical flask.
- In a fume hood, add to the conical flask 8.0 mL of acetic anhydride, and stir with a glass stirring rod until all the solid has dissolved.
- Remain in the fume hood and, using an eye dropper, carefully add 4 drops of concentrated $\text{H}_2\text{SO}_4\text{(aq)}$. Stir the mixture.
- Return to the lab bench and place the conical flask in the hot-water bath for 15 min, to allow the reaction to proceed.



Acetic anhydride is a severe eye irritant and must only be handled in a fume hood.

Concentrated sulfuric acid is corrosive. Wear eye protection and a laboratory apron.

Chemicals used or produced in a chemistry lab must never be consumed; they may contain toxic impurities.

(continued)

**Figure 4**

- Remove the conical flask from the hot-water bath and add to it about 20 g of crushed ice and about 20 mL of ice water. Place the flask in the ice bath and stir for about 10 min (**Figure 4**).
- Determine the mass of a piece of filter paper to the nearest 0.01 g, and set up the filtration apparatus. Carefully filter the contents of the conical flask and wash the crystals with 10 mL of ice water.
- Allow the filter paper and crystals to dry completely. Determine the mass of the dry filter paper and crystals.
 - (a) Calculate the percentage yield of the ASA obtained. Comment on the laboratory procedure, and on your lab skills.
 - (b) Explain why ASA tablets have a sour taste.
 - (c) The acidic character of ASA may be used to quantify the amount of pure ASA in a sample. From your knowledge of acids and bases, suggest a procedure that may be performed in a school laboratory to test the purity of the ASA sample that you prepared. Include a description of the apparatus needed and how the results may be interpreted, as well as any safety precautions and emergency laboratory procedures.

Investigation 2.1.1: Identification of Plastics

INQUIRY SKILLS

- | | |
|-----------------|-----------------|
| ○ Questioning | ● Recording |
| ○ Hypothesizing | ● Analyzing |
| ○ Predicting | ● Evaluating |
| ○ Planning | ● Communicating |
| ● Conducting | |



Figure 1

The triangular symbols on plastic containers allow us to identify the kind of plastic from which they are made.

Identification of Plastics

In recent times, the massive quantities of plastics we throw away have spurred communities to implement recycling programs in an effort to reduce the amount of waste going to landfill sites. As different types of plastics are made of different components, effective recycling requires a systematic identification of each type of polymer. To aid this identification, a resin identification coding system was established in 1988 by the Society of the Plastics Industry, Inc. (SPI) (**Figure 1**). **Table 1** gives the properties and end products of resins identified by their SPI codes.

We will differentiate among several unknown samples of plastics by their density, flame colour, solubility in acetone, and resistance to heating; then we will use the results to identify the plastics and their SPI codes.

Purpose

The purpose of this investigation is to use the properties of plastics to identify unknown samples and their SPI codes.

Question

What is the SPI resin code for each of the six unknown plastic samples?

Experimental Design

The different composition and structures of the plastics allow us to differentiate them by their properties. The samples are placed in liquids of different densities to separate them by flotation. Some samples are identified by their solubility in acetone. Other samples are identified by their resistance to heating. A flame test with a hot copper wire reveals plastics that contain chlorine atoms. The chlorine reacts with the copper wire on heating to produce copper chloride, which burns with a green flame.

Materials

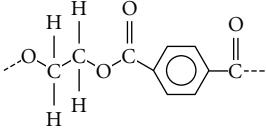
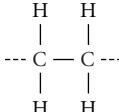
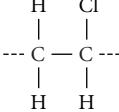
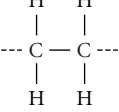
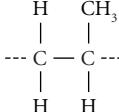
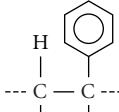
- | | |
|---|------------------------|
| lab apron | 100-mL beaker |
| eye protection | glass stirring rod |
| 1 cm × 1 cm samples (unknown) | copper wire, 15 cm |
| of each of the 6 categories of plastics,
each cut into an identifiable shape | cork or rubber stopper |
| water | tongs |
| 60 g 2-propanol (70%) | paper towel |
| corn oil (e.g., Mazola) | hot plate |
| 50 mL acetone | lab burner |
| three 250-mL beakers | |



Acetone and 2-propanol are highly flammable liquids and must be kept well away from open flames.

(continued)

Table 1 Codes on Everyday Plastics

SPI resin code	Structure of monomer	Density (g/cm ³)	Properties	End products
 PETE polyethylene terephthalate		1.38–1.39	Transparent, strong, impermeable to gas and to oils, softens at approximately 100°C	Bottles for carbonated drinks, containers for peanut butter and salad dressings
 HDPE high-density polyethylene		0.95–0.97	Naturally milky white in colour, strong and tough, readily moulded, resistant to chemicals, permeable to gas	Containers for milk, water, and juice, grocery bags, toys, liquid detergent bottles
 PVC vinyl (polyvinyl chloride)		1.16–1.35	Transparent, stable over long time, not flammable, tough, electrical insulator	Construction pipe and siding, carpet backing and window-frames, wire and cable insulation, floor coverings, medical tubing
 LDPE low-density polyethylene		0.92–0.94	Transparent, tough, and flexible, low melting point, electrical insulator	Dry cleaning bags, grocery bags, wire and cable insulation, flexible containers and lids
 PP polypropylene		0.90–0.91	Excellent chemical resistance, strong, low density, high melting point	Ketchup bottles, yogurt and margarine containers, medicine bottles
 PS polystyrene		1.05–1.07	Transparent, hard and brittle, poor barrier to oxygen and water vapour, low melting point, may be in rigid or foam form, softens in acetone	Cases for compact discs, knives, spoons, forks, trays; cups, grocery store meat trays, fast-food sandwich containers

(continued)

Procedure

Part 1: Testing for Density

1. Obtain one sample of each of the 6 plastic materials.
2. Place all 6 samples in a 250-mL beaker containing 100 mL of water and stir with a stirring rod. Allow the samples to settle. Use tongs to separate the samples that float from the samples that sink, and dry each sample with paper towel.
3. Prepare an alcohol solution by weighing out 60 g of 2-propanol (rubbing alcohol) in a 250-mL beaker and adding water to make a total of 100 g. Mix well.
4. Take any samples that float in water and place them in the alcohol solution. Stir and allow the samples to settle. Use tongs to separate the samples that float from the ones that sink, and dry each sample with paper towel.
5. Take any samples that float in the alcohol solution and place them in a 250-mL beaker containing 100 mL of corn oil. Stir and allow the samples to settle for a few minutes. Note any samples that float and any that sink.

Part 2: Testing for Flame Colour

6. Take the samples that sank in water, and test each for flame colour in a fume hood, using a 15-cm length of copper wire attached to a cork or rubber stopper. Holding the cork, heat the free end of the copper wire in a lab burner flame until the wire glows. Touch the hot end of the copper wire to a sample so that a small amount melts and attaches to the wire. Heat the melted sample that is attached to the copper wire in a flame. Record the colour of the flame.

Part 3: Testing with Acetone

7. Ensure that all open flames are extinguished. Of the samples tested in step 6, obtain a fresh sample of any material that did not burn with a green flame. Use tongs to test the samples for softness. Place each of these fresh samples in a 100-mL beaker containing 50 mL of nail polish remover. Watch the samples for a few minutes and note any colour changes. Remove the samples with tongs and test each sample for increased softness.

Part 4: Testing for Resistance to Melting

8. Heat a 250-mL beaker half filled with water on a hot plate until the water comes to a rolling boil. Place any sample that remained unchanged in step 7 into the boiling water and keep the water at a boil for a few minutes. Note any change in shape and softness in the sample.

(continued)

Evidence

- (a) Make a table in which to record the observations you expect from each test.

Analysis

- (b) From the information given in **Table 1** and your own observations, identify and give possible SPI codes for each of the 6 samples tested.

Evaluation

- (c) Obtain the actual SPI resin codes for each sample from your teacher, and evaluate the reliability of your results. Suggest any changes to the procedure that would improve the reliability.

(continued)

Synthesis

- (d) From your analysis and your understanding of the properties of the plastics tested, design an effective process for a recycling plant to separate plastic products that are collected. Draw a flow chart for such a process and briefly describe the steps involved, any problems you can foresee, and safety precautions required.

Activity 2.1.1: Making Guar Gum Slime

Guar gum, a vegetable gum derived from the guar plant, has a molar mass of about 220 000–250 000 g/mol. It is used as a protective colloid, a stabilizer, thickener, and film-forming agent for cheese, in salad dressing, ice cream, and soups; as a binding and disintegrating agent in tablet formulations; and in suspensions, emulsions, lotions, creams, and toothpastes. In short, it is a very useful polymer.

In this activity, you will make “slime” by creating a reversible crosslinked gel made from guar gum. The crosslinking is accomplished by adding sodium borate, $\text{Na}_2\text{B}_4\text{O}_7 \cdot 10 \text{ H}_2\text{O}$, commonly called borax.

Materials

guar gum—food grade from health food stores

water

saturated sodium borate (borax) solution

vinegar

graduated cylinder or measuring spoons

balance or measuring spoons

Popsicle stick or glass rod, for stirring

glass or disposable cup, or beaker

food colouring

sealable plastic bags for storing slime

small funnel and funnel support

- Measure 80 mL of water into the cup.
- Add one to two drops of food colouring if desired.
- Measure 0.5 g of guar gum (1/8 tsp). Add it to the water and stir until dissolved. Continue stirring until the mixture thickens (approximately 1 or 2 min).
- Add 15 mL (3 tsp) of saturated borax solution, and stir. The mixture will gel in 1–2 min. Let the slime sit in the cup for a few minutes to gel completely. (To store the slime for more than a few minutes, put it into a plastic bag and seal the top.)
 - (a) Describe what happens to the slime when you pull it slowly.
 - (b) Describe what happens to the slime when you pull it sharply.
- Put some slime on a smooth, hard surface and hit it with your hand.
- (c) Describe what happens to the slime.



Guar gum, if not food grade, is not safe to taste.

Sodium borate is moderately toxic in quantities of more than 1 g/1 kg of body weight. Wash any borax from hands with water. Wash hands after handling the slime.

Slime may stain or mar clothing, upholstery, or wood surfaces. Clean up spills immediately by wetting with vinegar, followed by soapy water.

(continued)

LSM 2.LA-2

- Place a funnel on a funnel support. Put some slime into the funnel. Push it through the funnel.
 - (d) Describe what happens as the slime comes out of the hole.
- Take about 20 mL of the slime and add about 5 mL of vinegar.
 - (e) Describe any changes in the properties of the slime.
- (f) Using your knowledge of the structure and bonding of polymers, explain each of your observations.
- Dispose of the slime as directed by your teacher.

Investigation 2.2.1: Preparation of a Polyester — A Condensation Polymer

INQUIRY SKILLS

- Questioning
 - Hypothesizing
 - Predicting
 - Planning
 - Conducting
 - Recording
 - Analyzing
 - Evaluating
 - Communicating

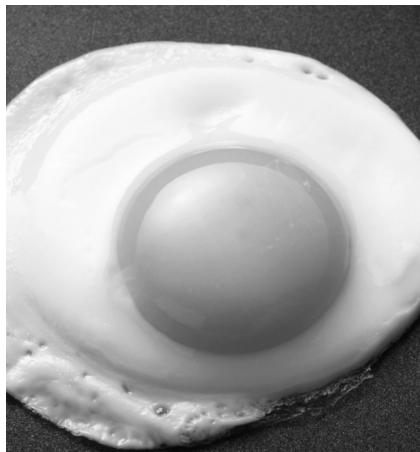


Figure 2

Egg white, when cooked, changes its structure irreversibly.

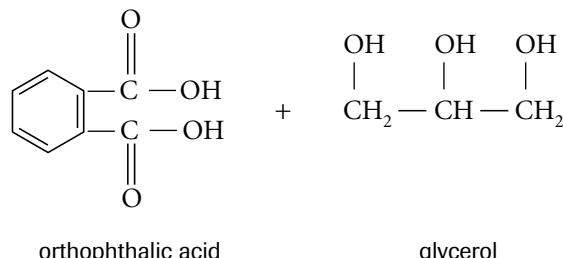
Condensation polymers are produced by reactions between functional groups on monomers, with the elimination of a small molecule such as water.

The reaction between an alcohol with two or more hydroxyl groups (such as glycerol) and an acid with two or more carboxyl groups produces a polyester known as an alkyd resin, commonly used in making paints and enamels. One of these is the Glyptal resin.

Glyptal is a thermoset polymer, meaning that it solidifies or “sets” irreversibly when heated. The heating causes a crosslinking reaction between polymer molecules. (This is the effect seen when proteins such as egg whites (**Figure 2**), another thermoset polymer, are cooked.) In Glyptal resin, phthalic anhydride forms crosslinks with other glycerol molecules, holding them together in a polymer.

Purpose

The purpose of this investigation is to test the principle of combining monomers to produce a thermoset polymer: glyptal.



Prediction

- (a) Write the equation for the reaction you predict will occur.

(b) Predict the properties that you will observe, if the polymerization reaction takes place.

(continued)

Experimental Design

The component monomers, glycerol and orthophthalic acid (in the form of phthalic anhydride, are mixed and heated together to form the thermoset polymer, Glyptal. The reaction is an esterification reaction, and any water produced is removed by boiling the mixture. The product is then tested for solubility in a non-polar solvent.

Materials

lab apron
 eye protection
 gloves
 ethylene glycol (1,2-ethanediol), 2 g
 glycerol, 2 g
 phthalic anhydride powder, 3 g
 solvent (paint thinner or nail polish remover), 5 mL
 two 100-mL beakers
 glass stirring rod
 beaker tongs
 watch glass to fit beaker
 hot plate
 small metal container (e.g., aluminum pie dish)



Phthalic anhydride is toxic and a skin irritant. Handle with care, and wear gloves, eye protection, and a lab apron.



The solvent is flammable, so must be kept well away from any open flame.

Procedure

- Place 2 g of glycerol and 3 g of phthalic anhydride in a 100-mL beaker and mix with a glass stirring rod.
- Heat gently on a hot plate, while stirring, to dissolve all the solid.
- Cover with a watch glass and continue heating gently until the mixture boils, and then boil for five minutes.
- Carefully pour the solution into a metal container. Let the plastic cool completely at room temperature.
- Observe and record the properties of the plastic formed.
- Place about 5 mL of the solvent supplied in a beaker and try to dissolve a portion of the plastic in the solvent.
- Allow the solvent containing any dissolved plastic to evaporate and observe any residue formed.
- Dispose of materials according to your teacher's instructions.
- Repeat steps 1 to 8 using ethylene glycol in place of glycerol.

Analysis

- Describe the properties of the product that you observed. Do they match your prediction?
- What makes this product suitable for use in household paints?

(continued)

Evaluation

- (e) Evaluate the concept that heating two component monomers together will produce a polymer with the desired properties.

Synthesis

- (f) Suppose this experimental design were repeated using 1,2-ethanediol in place of glycerol. Predict the reaction and properties of its product. Explain why your predictions are different from those made in (a) and (b), illustrating your answer with molecular structures.

Activity 2.7.1: Making Soap

Fats and oils are triglycerides of glycerol and fatty acids. The ester linkages in the triglycerides are broken in the presence of a strong base, such as NaOH, and heat. This reaction is called saponification and the sodium salt of the fatty acids produced is called soap. The other product formed is glycerol.

Materials

 **Ethanol is flammable. Ensure that there are no open flames.**

 **Sodium hydroxide pellets are extremely corrosive to eyes and skin. They must be handled with forceps. If NaOH comes into contact with skin, rinse with copious amounts of cold water. If it is splashed in the eyes, flush with water at an eyewash station for at least 10 minutes, then get medical attention.**

Procedure

- Put on a lab apron and eye protection.
- Set up a 100-mL beaker and label it Beaker A. (See **Table 2** for a summary of the contents of Beaker A and other beakers used in this activity.) Using forceps, add 18 pellets of solid NaOH to Beaker A. Do not allow the pellets to touch your skin. Add 10 mL of distilled water to the NaOH pellets and stir with a glass rod to dissolve. Set this beaker aside.
- Set up a 250-mL beaker and label it Beaker B. Add 15 g of a fat, such as lard or shortening, or an oil such as corn oil or olive oil to Beaker B. Add 15 mL of ethanol to the fat or oil and warm the mixture very gently on a hot plate, stirring with a glass rod to dissolve.
- Pour the contents of Beaker A into Beaker B and heat the mixture gently on the hot plate (low setting). Stir the mixture continuously for at least 20 min. If the mixture bubbles or splatters, use tongs to remove the beaker from the hot plate. Then return the beaker to the hot plate when the mixture has cooled slightly.
- When the reaction is complete, the mixture should thicken and have the appearance and consistency of a creamy pudding. Remove the beaker from the heat and allow to cool. You may add a drop or two of food colouring to the mixture at this stage to colour the soap.
- Set up another 100-mL beaker and label it Beaker C. Add 4 g of NaCl and 20 mL of cold distilled water. Stir to dissolve.
- Add the cold salt solution from Beaker C to the soap mixture in Beaker B. This should cause the soap to precipitate from the solution.
- Add 10 mL of vinegar to the mixture to neutralize any excess NaOH. Pour off any liquid into the sink.
- Add 10 mL of distilled water to wash the excess vinegar off the soap. Pour off any liquid into the sink.
- Set up a filter funnel and filter paper. Pour the soap mixture into the funnel, taking care not to puncture the filter paper. If desired, you may add a few drops of perfume or scent to the soap at this stage.

(continued)

- The soap remains in the filter funnel and can be left in the filter paper to dry, or it can be taken out of the filter paper and shaped and left to dry on a paper towel.

Table 2 Contents of Beakers A, B, and C

Beaker A (100 mL)	Beaker B (250 mL)	Beaker C (100 mL)
18 pellets of NaOH _(s)	15 g of fat or oil	4 g NaCl
10 mL distilled water	15 mL of ethanol	20 mL cold distilled water
stir to dissolve	warm gently on hot plate to dissolve	stir to dissolve

Analysis

- Why is a saponification reaction considered to be the reverse of an esterification reaction?
- Compare the soaps made by other students in the class. Is there any difference in hardness of the soaps made from different fats and oils? Explain your answer.
- Are soap molecules polar or nonpolar? Draw a sketch of a soap molecule to illustrate your answer.
- From your knowledge of fatty acids, what are some possible fatty acids present in the fat or oil you used? Draw structural diagrams for a possible triglyceride in the fat or oil used to make your soap, and for a possible soap molecule in your soap.
- What substances may be in the filtrate after the soap was filtered out?

Rydberg and Planck Equations, Extra Exercises

1. In general, a human eye cannot detect photons of frequency greater than 750 THz (7.50×10^{14} Hz). Use the Planck equation to find the maximum photon energy necessary to stimulate receptors in the eye.
 2. Blue-violet lines of the hydrogen emission spectrum are shown on page 175. Recall that *all* of the visible photons emitted by hydrogen atoms involve electron energy-level transitions from higher levels down to a final level of $n_f = 2$.
 - (a) Use the Rydberg equation to show which of the blue-violet lines on the left side of the spectrum on page 175 has an initial energy level of 5.
 - (b) Use algebraic substitution to combine the wave equation ($c = f\lambda$, where c is the speed of light, 3.00×10^8 m/s) and the Planck equation to determine the energy of photons of this blue-violet light in a single combined calculation.

Rydberg and Planck Equations, Extra Exercises, Solution

1. In general, a human eye cannot detect photons of frequency greater than 750 THz (7.50×10^{14} Hz). Use the Planck equation to find the maximum photon energy necessary to stimulate receptors in the eye.

$$E = hf$$

$$h = 6.63 \times 10^{-34} \text{ J/Hz}$$

$$f = 7.50 \times 10^{14} \text{ Hz}$$

$$E = \frac{6.63 \times 10^{-34} \text{ J}}{1 \text{ Hz}} \times 7.50 \times 10^{14} \text{ Hz}$$

$$E = 4.97 \times 10^{-19} \text{ J}$$

The maximum photon energy that will stimulate receptors in this eye is 4.97×10^{-19} J.

2. Blue-violet lines of the hydrogen emission spectrum are shown on page 175 of the Student Text. Recall that *all* of the visible photons emitted by hydrogen atoms involve electron energy-level transitions from higher levels down to a final level of $n_f = 2$.

- (a) Use the Rydberg equation to show which of the blue-violet lines on the left side of the spectrum on page 175 has an initial energy level of 5.

$$\frac{1}{\lambda} = R_H \left(\frac{1}{n_f^2} - \frac{1}{n_i^2} \right)$$

$$R_H = 1.10 \times 10^7 / \text{m}$$

$$n_f = 2 \quad n_i = 5$$

$$\frac{1}{\lambda} = \frac{1.10 \times 10^7}{1 \text{ m}} \left(\frac{1}{2^2} - \frac{1}{5^2} \right)$$

$$\lambda = 4.33 \times 10^{-7} \text{ m} = 433 \text{ nm}$$

The calculated wavelength of emitted light is 433 nm, which best matches the second blue-violet line from the left.

(continued)

- (b) Use algebraic substitution to combine the wave equation ($c = f\lambda$, where c is the speed of light, 3.00×10^8 m/s) and the Planck equation to determine the energy of photons of this blue-violet light in a single combined calculation.

$$h = 6.63 \times 10^{-34} \text{ J/Hz} = 6.63 \times 10^{-34} \text{ J}\cdot\text{s}$$

$$c = 3.00 \times 10^8 \text{ m/s}$$

$$E = hf$$

$$c = f\lambda \quad \text{therefore} \quad f = \frac{c}{\lambda}$$

$$\text{and substituting, } E = \frac{hc}{\lambda}$$

$$E = \frac{6.63 \times 10^{-34} \text{ J}\cdot\text{s} \times 3.00 \times 10^8 \frac{\text{m}}{\text{s}}}{4.33 \times 10^{-7} \text{ m}}$$

$$E = 4.59 \times 10^{-19} \text{ J}$$

The energy of each photon of blue-violet light is 4.59×10^{-19} J.

Electron Configurations, Extra Exercises

1. State the symbol for each of the four energy sublevels (types of orbitals), the number of orbitals in each sublevel, and the maximum number of electrons in that sublevel.

2. State Hund's rule in your own words.

3. Write the full electron configuration for each of the following:
 - (a) aluminum atom

 - (b) cobalt atom

 - (c) phosphide ion

4. What is the shorthand form (sometimes called the “kernel method”) for writing electron configurations?

5. Write the chemical symbol for the atom corresponding to each of the following descriptions:
 - (a) [Ar] $4s^2$

 - (b) [Kr] $5s^2\ 4d^{10}\ 5p^1$

 - (c) [Xe] $6s^2\ 4f^{14}\ 5d^5$

6. List the chemical symbols and names for six ions isoelectronic with an argon atom.

7. State the similarity in electron configurations among the atoms of Group 16.

Electron Configurations, Extra Exercises, Solution

1. State the symbol for each of the four energy sublevels (types of orbitals), the number of orbitals in each sublevel, and the maximum number of electrons in that sublevel.

s, 1 orbital, 2 electrons

p, 3 orbitals, 6 electrons

d, 5 orbitals, 10 electrons

f, 7 orbitals, 14 electrons

2. State Hund's rule in your own words.

No electron pairing takes place in p, d, or f orbitals until each orbital of the given set contains one electron.

3. Write the full electron configuration for each of the following:

(a) aluminum atom

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$

(b) cobalt atom

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$

(c) phosphide ion

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

4. What is the shorthand form (sometimes called the “kernel method”) for writing electron configurations?

It is an abbreviation of the electronic configuration of the atom of an element using, as a starting point, an atomic symbol to represent the electron configuration of the noble gas atom that immediately precedes the specified element in the periodic table.

5. Write the chemical symbol for the atom corresponding to each of the following descriptions:

(a) [Ar] $4s^2$

Ca

(b) [Kr] $5s^2\ 4d^{10}\ 5p^1$

In

(c) [Xe] $6s^2\ 4f^{14}\ 5d^5$

Re

6. List the chemical symbols and names for six ions isoelectronic with an argon atom.

P³⁻ phosphide ion

K⁺ potassium ion

S²⁻ sulfide ion

Ca²⁺ calcium ion

Cl^{-} chloride ion}

Sc³⁺ scandium(III) ion

7. State the similarity in electron configurations among the atoms of Group 16.

All atoms in Group 16 have an electron configuration that ends in ns² np⁴.

Investigation 3.1.1: The Nature of Cathode Rays

INQUIRY SKILLS

- | | |
|---|--|
| <input type="radio"/> Questioning | <input checked="" type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
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When cathode rays were first discovered in about 1860, their nature was a mystery. One hypothesis was that they were a form of electromagnetic radiation, like visible light or ultraviolet light. An obvious test would be to check some properties of cathode rays and compare them to the known properties of light. The purpose of this experiment is to test the hypothesis relating cathode rays to light.

Question

What effect do electric charges and magnets have on the direction of motion of a cathode ray?

Experimental Design

A cathode ray is tested separately with a bar magnet and with two charged, parallel plates (independent variables). In each case, the dependent variable is the deflection of the cathode ray. Controlled variables include the charged plates and the magnet. As a control, the same tests are repeated with a thin beam of light, such as a laser.

Materials

cathode-ray tube
power supply with variable voltage
connecting wires
bar magnet
2 lab stands
2 clamps
laser light source (e.g., laser pointer)



Serious shock hazard may result from the use of a high-voltage supply. Unplug or turn off power supply when connecting and disconnecting wires. Do not shine a laser beam into anyone's eyes.

Procedure

1. Clamp the cathode tube on a lab stand so that it is horizontal.
2. Connect the cathode-ray tube to the power supply. Check all connections before plugging the power supply into the electrical outlet. Turn on the power supply.
3. Once the cathode ray is visible (either inside the tube or on the screen at the end of the tube), bring a bar magnet near the tube and note any effect on the cathode ray.
4. Connect the two parallel plates to the same or different power supply and slowly increase the potential difference (voltage) on the plates. Note the effect on the cathode ray. (The positively charged plate is at the red connection; the negatively charged plate at the black connection.)
5. Turn off the power supply and disconnect the cathode-ray tube.
6. Clamp the laser light source horizontally on a lab stand so that the end of the beam is visible on a wall or screen.
7. Hold the bar magnet near the beam and note any effect by observing the dot on the screen.

(continued)

8. Aim the laser beam so that it shines between the parallel plates inside the cathode-ray tube and the end of the beam is again visible.
9. Connect the parallel plates as in step 4 and note any effect of the charged plates on the laser beam.

Evidence

- (a) Create a table to record your observations.

Analysis

- (b) Based on your evidence, does it appear that cathode rays are like electromagnetic radiation such as visible light? Justify your answer.

Evaluation

- (c) Are there any obvious flaws in this experiment? Suggest some improvements in the Experimental Design, Materials, and/or Procedure to improve the quality or quantity of the evidence.

(continued)

- (d) Evaluate the hypothesis that cathode rays are a form of electromagnetic radiation.

Synthesis

- (e) Use your knowledge about attraction and repulsion of electric charges. What does the bending of the cathode ray when passing near electrically charged plates suggest about the composition of the cathode rays?
- (f) Based on the evidence collected, what is the sign of the charge of the particles in a cathode ray?

Activity 3.1.1: Rutherford's Gold Foil Experiment

The purpose of this activity is to illustrate Rutherford's famous alpha-particle-scattering experiment, using a computer simulation.

Materials

Nelson Chemistry 12 CD; PC

- Start the “Rutherford” simulation from the Nelson Chemistry 12 CD on your computer.
- Follow the instructions and choose the initial settings suggested by your teacher.
- Record general and specific observations while the experiment is running.
- If requested, change the settings and run the experiment again.
 - (a) According to the Thomson model of the atom, what result is predicted for a stream of alpha particles striking a layer of gold atoms?
 - (b) Summarize the main evidence from Rutherford’s experiment.
 - (c) Compare the relative numbers of alpha particles that travelled relatively undeflected to the number that recoiled. What does this suggest about the relative size of the nucleus?
 - (d) Evaluate the prediction from the Thomson model and the Thomson model itself.
 - (e) Use your modern knowledge of the components of an atom. How would the results of Rutherford’s experiment with an aluminum foil be similar to his experiment with a gold foil? How would the results be different?

Activity 3.3.1: Hot Solids

All hot solids, liquids, and gases produce some form of light, occasionally in the visible region but often in other invisible regions. The purpose of this activity is to study the light produced by a hot solid.

Materials

Variac (variable power supply)	diffraction grating
clear lamp with vertical filament	(e.g., 6000–7500
overhead projector	lines/cm)
2 pieces of heavy paper	tape

- Set up a clear incandescent lamp with a vertical filament. Plug the cord into a Variac that is plugged into an electrical outlet. Set the Variac at its lowest setting.
 - (a) Switch on the Variac and turn off all lights. Observe the filament as the voltage is slowly increased from 0 to about 110 V and describe the change in colour.
 - (b) People sometimes describe hot objects using terms such as “red hot” or “white hot.” Which do you think is hotter?
 - (c) State two objects in your home which may be red hot at certain times.
- An overhead projector bulb produces a very white, bright light when the filament is white hot. Set up the projector to shine its light on a white wall or board. Two straight-edge pieces of heavy paper are placed about 5 mm apart on the horizontal stage of the projector so that they block all of the light except for a thin strip visible on the wall or board (**Figure 1**).

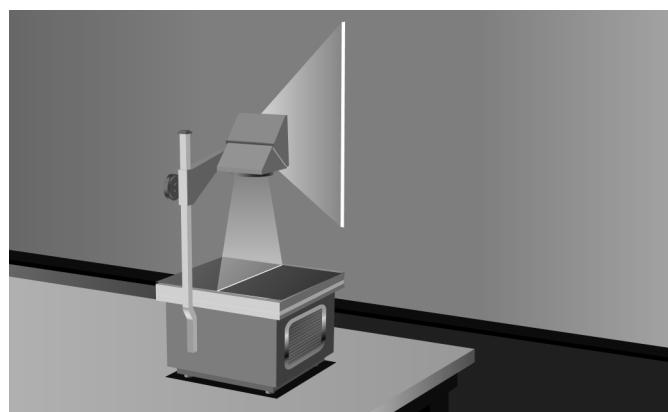


Figure 1

(continued)

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- A diffraction grating is used to separate the light into its various components.
Hang the grating in the centre of the final lens using a piece of plastic tape.
Shut off all room lights and observe.
- (d) Sketch and label the main colours in the visible spectrum that you see on the right-hand side of the central white line. This spectrum is known as a continuous spectrum. (The same spectrum, but reversed, appears on the left side.)
- (e) If your eyes were constructed differently, you might be able to see beyond the blue/violet end and beyond the red end of the spectrum. What are these regions called? Label them on your diagram.
- (f) Which do you think is more dangerous to you—being exposed to a red lamp or an ultraviolet lamp? Justify your choice.
- (g) Assuming the danger is related to the energy of different types of light, label your spectrum from (d) with an arrow going from low to high energy.

Investigation 3.3.1: The Photoelectric Effect

INQUIRY SKILLS

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| <input type="radio"/> Questioning | <input checked="" type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input checked="" type="radio"/> Conducting | |

An electroscope is a device that is used to detect and measure electric charges.

Purpose

The purpose of Investigation 3.3.1 is to create an initial theoretical understanding of the photoelectric effect.

Question

What effect does light have on a negatively charged metal plate?

Experimental Design

An electroscope with a zinc plate is charged negatively using a vinyl strip rubbed with paper towel, and then observed when white light and ultraviolet light are each shone onto the zinc. The control is the charged electroscope under normal room lighting.

Materials

electroscope (with a flat top)
vinyl plastic strip
paper towel
zinc plate (about 3–5 cm sized square)
steel wool
lamp with 100-W bulb
ultraviolet light source

Procedure

1. Set the electroscope on a flat surface and make certain the vane or leaves can move freely.
2. Rub the vinyl strip several times with a piece of paper towel to negatively charge the strip.
3. Touch the vinyl strip one or more times to the top of the electroscope until the vane or leaves move and remain in their new position. (If nothing happens, have your teacher check your electroscope.)
4. Observe the electroscope for about 1 min.
5. Touch the electroscope with your finger to discharge or neutralize it. Note the change in the electroscope.
6. Scrub the zinc plate with the steel wool until the plate is shiny. Place the zinc plate on the top of the electroscope.
7. Charge the electroscope again (steps 2–3).
8. Plug in the lamp with the 100-W bulb, turn it on, and shine the light onto the zinc plate from a distance of about 10 cm.

(continued)

9. Observe the electroscope for about 1 min.
10. If necessary, charge the electroscope to the same angle as before. Repeat steps 8 and 9 using the ultraviolet light.

Evidence

- (a) Create a table to record your observations.

Analysis

- (b) What effect did a bright white light from a normal lamp have on the negatively charged zinc plate? Answer relative to the control — the charged electroscope sitting for 1 min under room-light conditions.
- (c) What effect, compared to the control, did the ultraviolet light have on the charged zinc plate?
- (d) If the intensity (brightness) of the light is responsible for the effect, which light should work better? How certain are you of your answer? Provide your reasoning.
- (e) Explain, in terms of electrons and protons, the existence of a negatively charged zinc plate.
- (f) Did the electroscope become more or less charged when illuminated with ultraviolet light compared to visible light?
- (g) Using your answers to (e) and (f), suggest an initial explanation for the effect of the ultraviolet light on a negatively charged zinc plate.

(continued)

Evaluation

- (h) What other test could be done with the electroscope and zinc plate to make the interpretation in (g) more certain?

- (i) What other general improvement to the materials could be made to improve the quality of the evidence?

- (j) How certain are you about your answer to the question in part (g)? Provide reasons.

Synthesis

- (k) Ordinary glass absorbs ultraviolet light and does not allow it to pass through. Predict the results of this experiment if a glass plate were placed between the ultraviolet light and the charged zinc plate.

- (l) Would direct sunlight through an open or a closed window discharge the electroscope? Provide your reasoning.

Activity 3.4.1: Line Spectra

All substances absorb or emit some part of the electromagnetic spectrum. Some substances absorb some of the visible spectrum, while other substances may absorb in the ultraviolet region or other regions. The spectrum produced is called a dark-line or absorption spectrum. Substances can also emit light in different parts of the electromagnetic spectrum. Under certain conditions, the light emitted appears as bright lines. The spectrum produced is called a bright-line spectrum. The purpose of this activity is to illustrate the two main types of line spectra—bright-line and absorption.

Materials

overhead projector
2 pieces of heavy paper
diffraction grating (600 lines/mm)
2 large beakers
flat glass plate to cover beaker
pure water
tiny crystals of potassium permanganate and of iodine
spectrum tube power supply
hydrogen gas tube
spectroscope



Serious shock hazard may result from the use of a high-voltage supply.

- An overhead projector is set up to shine its light on a white wall or board. Two straight-edged pieces of heavy paper are placed about 5 mm apart on the horizontal stage of the projector so that they block all of the light except for a thin strip.
- A diffraction grating (containing about 6000–7500 lines/cm) is hung in the centre of the vertical projecting lens using a piece of plastic tape.
- Place a large clean beaker on the slit formed by the two pieces of paper. Shut off the room lights and observe the continuous visible spectrum.
- Now fill the beaker with some pure water and observe the spectrum again.
 - (a) Do either glass or water change the colours in the visible spectrum? Do these substances absorb visible light?
- Add a few crystals of potassium permanganate to the water and stir. Observe the visible spectrum.
- (b) What effect does aqueous potassium permanganate have on the visible spectrum?
- Warm some solid iodine in a beaker in a fume hood or by some other safe method. Cover the beaker with a flat glass plate. Place the beaker containing the iodine vapour on the overhead projector and observe the spectrum.

(continued)

- (c) Compare the spectrum obtained to the one for the potassium permanganate solution in step 5. Can gases also absorb electromagnetic radiation?

- Observe the spectrum of sunlight recorded on Earth's surface (**Figure 2**).

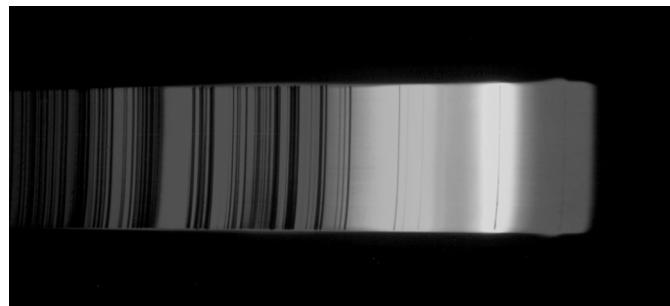


Figure 2

The spectrum of sunlight recorded on the surface of Earth is a dark-line (absorption) spectrum due to gases in the atmospheres of Earth and the Sun absorbing specific parts of the sunlight.

- (d) What evidence is there that some light is being absorbed? Suggest some possible gases that might be responsible.
- Set up a hydrogen gas tube in a gas discharge power supply. Switch on the power and turn off all lights. Observe the spectrum with a handheld spectroscope. Draw and label a diagram of this spectrum. If time permits, observe the spectra of other gases.
 - (e) Under what conditions do gases produce light that is in the visible region of the electromagnetic spectrum?
 - (f) In this case, is the spectrum produced a bright-line or continuous spectrum?
 - (g) How are line spectra used in chemical analysis?

Activity 3.4.2: The Hydrogen Line Spectrum and the Bohr Theory

The bright-line or emission spectrum of hydrogen (**Figure 3**) has been known since the mid-1800s. The position or wavelength of each of the coloured lines in the visible region was precisely measured by the Swedish scientist, Anders Ångström in 1862. (A unit of length was named after Ångström, $1 \text{ \AA} = 0.1 \text{ nm}$, exactly.) The purpose of this activity is to use a computer simulation to demonstrate some of the main concepts of the Bohr theory and to relate these to the hydrogen emission spectrum.

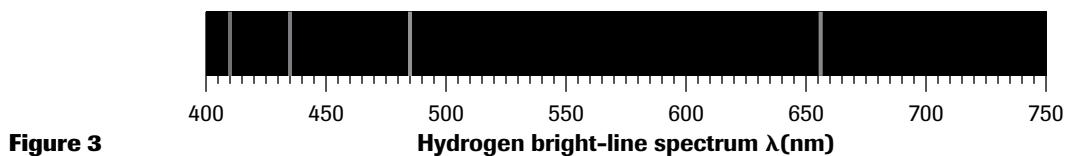


Figure 3

Materials

Nelson Chemistry 12 CD, PC

- Using **Figure 3**, estimate the wavelength of each of the four lines in the visible region of the hydrogen spectrum. These visible lines belong to the group of lines known as the Balmer series.
 - Start the “Bohr” simulation from the Nelson Chemistry 12 CD on your computer. Under the “Series” menu, select “Balmer.”
 - Set the “New State” at 3 and click the “Photon” button.
 - With the electron now in $n_i = 3$, set the “New State” to 2.
- Note and record the wavelength of the light to the nearest nanometre. Click the “Photon” button.
- Is some light (a photon) absorbed or released in this transition?
- To which line in the Balmer series does this transition correspond?
 - Repeat the simulation using the following settings:
 $n_i = 4, n_f = 2$
 $n_i = 5, n_f = 2$
 $n_i = 6, n_f = 2$

(continued)

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- Answer questions (b) to (d) for each of these transitions.
- (e) How do your answers from (a) using **Figure 3** compare with your answers to (d) using the computer simulation? Is this surprising? Explain briefly.
- (f) If some light (a photon) is absorbed by an electron, what happens to the electron? Try this with the simulation program.
- (g) How does the wavelength of light corresponding to the transition from $n_i = 3$ to $n_f = 2$ compare with $n_i = 2$ to $n_f = 3$? Explain briefly why this is necessary, according to the Bohr theory.
- (h) An electron cannot undergo a transition from $n_i = 1$ to $n_f = 2.5$. According to the Bohr theory, why is this not possible?

Investigation 3.5.1: Paramagnetism

INQUIRY SKILLS

- | | |
|--|---|
| <ul style="list-style-type: none"><input type="radio"/> Questioning<input type="radio"/> Hypothesizing<input type="radio"/> Predicting<input checked="" type="radio"/> Planning<input checked="" type="radio"/> Conducting | <ul style="list-style-type: none"><input checked="" type="radio"/> Recording<input checked="" type="radio"/> Analyzing<input checked="" type="radio"/> Evaluating<input type="radio"/> Communicating |
|--|---|

Paramagnetism was first investigated and named by Michael Faraday in the mid-1800s. At this time, before the discovery of the electron, there was no theoretical explanation of the cause of paramagnetism. According to modern atomic theory, paramagnetism is believed to be caused by the presence of unpaired electrons in an atom or ion.

Purpose

The scientific purpose of this investigation is to determine experimentally which substances are paramagnetic.

Question

Which substances containing calcium, zinc, copper(II), and manganese(II) ions are paramagnetic?

Experimental Design

Test tubes containing the sulfates of each of the ions are suspended by threads from a support. Evidence for any attraction of each test tube toward a strong magnet is observed.

- (a) Identify the independent, dependent, and controlled variables.

Materials

- eye protection
- 4 small test tubes
- stirring rod
- thread
- laboratory stand
- clamp
- horizontal bar
- strong magnet (e.g., neodymium)
- a few grams of the solids:
 - calcium sulfate
 - zinc sulfate
 - copper(II) sulfate
 - manganese(II) sulfate

(continued)

Procedure

- (b) Write a complete procedure for this experiment. Include safety precautions with respect to handling and disposal of the chemicals used. Have the procedure checked by your teacher before you proceed.

Analysis

- (c) Answer the Question based on the evidence collected.

Evaluation

- (d) Are there any flaws or possible improvements in the Experimental Design? Describe briefly.

- (e) Suggest some improvements to the Materials and Procedure.

- (f) How certain are you about the evidence obtained? Include possible sources of error or uncertainty.

Lab Exercise 3.6.1: Quantitative Paramagnetism

In Investigation 3.5.1, you obtained some preliminary evidence for a possible connection between unpaired electrons (as determined by the electron configuration) and paramagnetism. The purpose of this lab exercise is to test this hypothesis with a quantitative experiment.

Question

What effect does the number of unpaired electrons have on the strength of the paramagnetism of metal salts?

Prediction/Hypothesis

- (a) Write a prediction and provide your reasoning based on electron configurations.

Experimental Design

A sensitive electronic balance is used to measure the attraction between a powerful magnet and a test tube containing a metal salt. The balance is tared (zeroed) before the test tube is lowered (**Figure 4**). The mass reading is taken just before contact of the test tube with the magnet. Several ionic compounds containing different metal ions are individually tested using the same mass of each compound.

- (b) Identify the independent, dependent, and controlled variables.

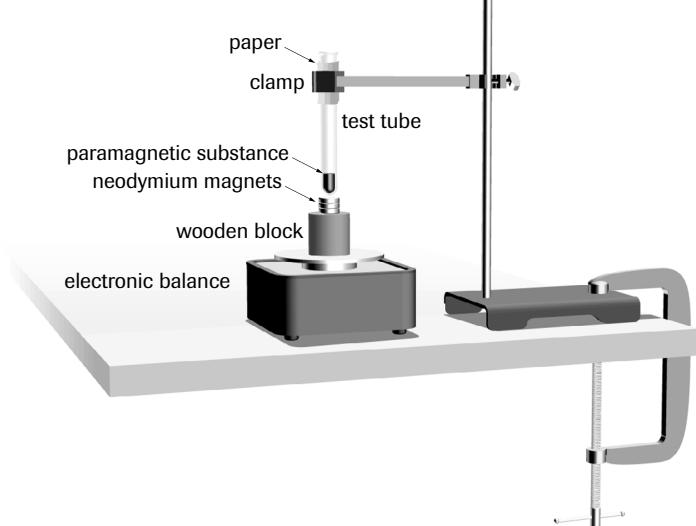


Figure 4 A strong magnet or magnets (such as neodymium magnets) and a paramagnetic substance attract each other. This means that the magnet and block are slightly lifted toward the fixed test tube.

Evidence

Table 1 Change in Mass in a Strong Magnetic Field

Ionic compound	Mass reading, Δm (g)
$\text{CaSO}_{4(\text{s})}$	0.00
$\text{Al}_2(\text{SO}_4)_{3(\text{s})}$	0.00
$\text{CuCl}_{(\text{s})}$	0.00
$\text{CuSO}_4 \cdot 5\text{H}_2\text{O}_{(\text{s})}$	-0.09
$\text{NiSO}_4 \cdot 7\text{H}_2\text{O}_{(\text{s})}$	-0.22
$\text{CoCl}_2 \cdot 6\text{H}_2\text{O}_{(\text{s})}$	-0.47
$\text{FeSO}_4 \cdot 7\text{H}_2\text{O}_{(\text{s})}$	-0.51
$\text{MnSO}_4 \cdot \text{H}_2\text{O}_{(\text{s})}$	-1.26
$\text{FeCl}_3 \cdot 6\text{H}_2\text{O}_{(\text{s})}$	-0.95

mass of each compound in test tube = 3.00 g

(continued)

Analysis

- (c) What is the significance of a zero-mass reading for some substances and negative-mass readings for other substances?
- (d) How does this change in mass relate to the paramagnetic strength of the substance? (Each of the compounds has a different molar mass and therefore a different amount in moles in the controlled mass of 3.00 g. In order to make a valid comparison, you need to know the change in mass per mole of the substance.)
- (e) Create a table with headings ionic compound, molar mass, number of moles. Create and complete another table with the following headings: metal ion, electron configuration, number of unpaired electrons, mass decrease per mole.
- (f) Plot a graph of the number of unpaired electrons (x-axis) and mass decrease per mole (y-axis). Draw a best-fit line.
- (g) Answer the Question asked at the beginning of this investigation.

(continued)

Evaluation

- (h) Evaluate the Experimental Design. Are there any obvious flaws? Any improvements?
- (i) Suggest some improvements to the materials and procedure that would improve the quality and quantity of the evidence collected.
- (j) How confident are you with the experimental answer to the question?
- (k) Evaluate the Prediction (verified, falsified, or inconclusive). State your reasons.
- (l) Does the hypothesis appear to be acceptable based on your evaluation of the prediction?

Activity 3.7.1: Modelling Standing Electron Waves

A mechanical model of Schrödinger's standing waves associated with electrons can be made using a thin, stiff, loop of wire which is vibrated with a variable frequency mechanical oscillator. The mechanical oscillator is like a heavy-duty speaker cone with a rod attached to its centre. As the cone and rod move up and down, whatever is attached to the rod oscillates up and down. Vibrating one point in the loop sets up waves in the wire. This is like holding the edge of a long spring, oscillating one end back and forth, and generating waves that move along the spring. When returning waves meet they interfere with each other, either constructively (increasing the amplitude) or destructively (decreasing the amplitude). Standing waves are a special case of wave interference that results in apparently stationary nodes (zero amplitude points) and antinodes (maximum amplitude points).

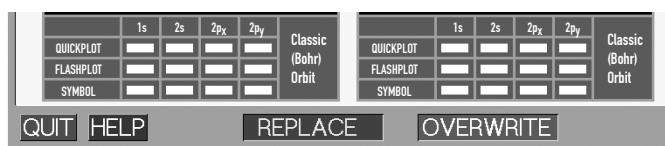
- Secure the oscillator on a sturdy stand. Attach the plug containing the loop of wire and adjust so that the plane of the loop is horizontal.
- Set the frequency to its lowest setting. Plug in the oscillator and turn it on.
- Slowly increase the frequency and observe the results.
- Continue increasing the frequency until no further observations are possible because the nodes and antinodes are no longer visible.
- Slowly decrease the frequency back down to its lowest setting and view the changes in reverse order.
- Repeat this procedure, if necessary, to complete your observations.
 - (a) Describe, in general, the appearance of the nodes and antinodes.
 - (b) Do all frequencies produce standing wave patterns? Discuss briefly.
 - (c) List the number of antinodes from the lowest possible to as many as you were able to observe.
 - (d) How does this physical model relate to the wave mechanics model of the atom? What are some limitations of this model?

Activity 3.7.2: Simulation of Electron Orbitals

Materials

Nelson Chemistry 12 CD, PC

- Start the SIRs program on the Nelson Chemistry 12 CD.
- Select “SIR Orbital.” Read and then click on the SIR Orbital title screen.
- You should now see two screens that will allow comparisons of various $n = 1$ and $n = 2$ orbitals. The “Bohr” choice is representative only and does not present corresponding Bohr orbits. It serves as a general reminder of the view of an electron in an atom, according to the Bohr theory.
- There are a number of options available.



“Overwrite”—By selecting this option, you can overlay several diagrams onto one on the screen.

“Replace”—When this is selected, the existing diagram is erased and replaced by a new diagram.

“Quickplot”—This produces a summary of many calculations of the location of an electron at an instant in time. This final diagram represents the electron probability density for an electron of energy corresponding to the selected orbital.

“Flashplot”—Selecting this option initiates individual calculations performed by your computer. Each flashing “star” (which leaves a yellow dot) indicates the instantaneous location for an electron of energy corresponding to the selected orbital. This illustrates how the final probability density (as seen in Quickplot) is generated.

“Symbol”—Chemists find it convenient to reduce the probability density diagram to a figure that contains at least a 90% probability of the electron being located within the bounds of the figure. These kinds of diagrams are very useful for describing chemical bonding.

Note: Orbitals are three dimensional. The SIR program eliminates one dimension (z) to produce a two-dimensional diagram (like a slice or cross section of the full 3-D view).

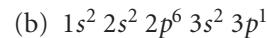
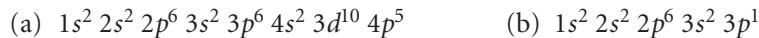
- In the “Flashplot” choices, select “1s” and let this run until it is finished.
- At the bottom right of the right window, select “Classic (Bohr) Orbit.”
 - Compare the views of the electron according to the quantum mechanics theory (left window) and the Bohr theory (right window).

(continued)

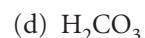
- Select “End” for the Classic (Bohr) Orbit.
- Click “End.” Select the “Overwrite” and “Figure” boxes, then click “1s” in the left window.
- In the right window, use “Replace” and “Flashplot” for the 2s orbital. When the simulation is finished, select “Overwrite” and “Figure” for 2s.
 - (b) Compare the electron probability density for the 1s and 2s orbitals.
- (c) How do the sizes of these orbitals compare?
- (d) How does the energy of an electron in these orbitals compare?
- Using the functions of this program (e.g., “Quickplot” and “Symbol”) create representations of 2s and 2px and then 2px and 2py.
 - (e) Compare the 2s and 2p orbitals and the 2px and 2py orbitals.
- (f) Which p orbital is missing? What would be the relative size, shape, and orientation of this orbital?
- In the left window, select the “Overwrite” and “Symbol” boxes, then click “1s”.
 - (g) Assuming full 1s and 2s orbitals, and half-filled 2px and 2py, what atom would this represent?
- (h) In 3-D, what would be the appearance of this combined electron probability density?
- The view of the complete atom is not very useful. Now overwrite this view using the “Symbol” for each of the four orbitals.
 - (i) Sketch this diagram and label each of the four orbitals.

Lewis Structures, Extra Exercises

1. Draw the Lewis symbol for the atom corresponding to each of the following electronic configurations:



2. Draw the Lewis structure for each of the following molecules or ions:

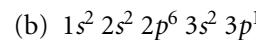
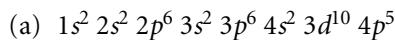


3. Write a balanced chemical equation for the following reaction, using Lewis structures to represent the reactants and products.

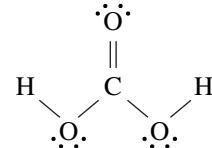
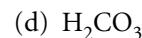
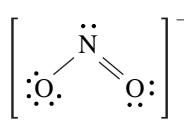
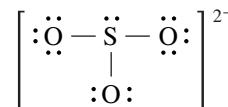
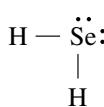
Ethane and chlorine are reacted to form chloroethane and hydrogen chloride.

Lewis Structures, Extra Exercises, Solution

1. Draw the Lewis symbol for the atom corresponding to each of the following electronic configurations:

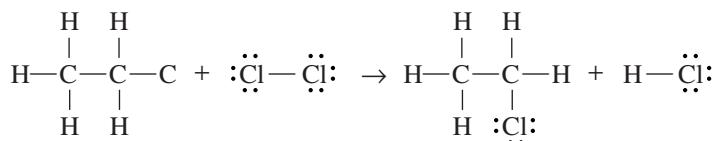


2. Draw the Lewis structure for each of the following molecules or ions:



3. Write a balanced chemical equation for the following reaction, using Lewis structures to represent the reactants and products.

Ethane and chlorine are reacted to form chloroethane and hydrogen chloride.



Bond Orbital Hybridization, Extra Exercises

1. According to the valence bond theory, what are the requirements for the formation of a covalent bond between two atoms?

2. (a) Write the electron configuration for a carbon atom.

(b) According to this configuration, how many half-filled orbitals does a carbon atom have?

(c) According to the valence bond theory, how many covalent bonds are predicted for a carbon atom? Why?

(d) Using a simple, common example, show that the prediction in (c) is clearly false.

(e) Describe briefly how this problem was solved to explain the evidence that carbon atoms form four single covalent bonds with atoms such as hydrogen.

3. For the atom identified in each of the following molecules, identify the number and type of hybrid orbitals of the atom, state how many sigma and pi bonds are formed by this atom, and state the shape and bond angle around the atom.
 - (a) the C atom in chloroform, CHCl_3

 - (b) the “1” carbon atom in 1-butyne, $\text{CH}_3\text{CH}_2\text{CCH}$

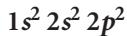
 - (c) the C atom in formaldehyde, HCHO

Bond Orbital Hybridization, Extra Exercises, Solution

1. According to the valence bond theory, what are the requirements for the formation of a covalent bond between two atoms?

Two atoms must each have a half-filled orbital which can overlap to produce a new combined orbital containing two paired electrons.

2. (a) Write the electron configuration for a carbon atom.



- (b) According to this configuration, how many half-filled orbitals does a carbon atom have?

two ($2p^2$)

- (c) According to the valence bond theory, how many covalent bonds are predicted for a carbon atom? Why?

According to the valence bond theory, covalent bonds are formed from the overlap of half-filled orbitals. Therefore, a carbon atom should form two covalent bonds because its configuration shows two half-filled orbitals.

- (d) Using a simple, common example, show that the prediction in (c) is clearly false.

CH_4 , methane (and not CH_2)

- (e) Describe briefly how this problem was solved to explain the evidence that carbon atoms form four single covalent bonds with atoms such as hydrogen.

A $2s$ electron is promoted to the empty $2p$ orbital to produce four, half-filled orbitals (sp^3) that can overlap and form four covalent bonds.

3. For the atom identified in each of the following molecules, identify the number and type of hybrid orbitals of the atom, state how many sigma and pi bonds are formed by this atom, and state the shape and bond angle around the atom.

- (a) the C atom in chloroform, CHCl_3

Four hybrid sp^3 orbitals form during bonding. The C atom forms four sigma bonds. This should result in a tetrahedral shape around this atom, with bond angles of 109.5° .

- (b) the “1” carbon atom in 1-butyne, $\text{CH}_3\text{CH}_2\text{CCH}$

Two hybrid sp orbitals form during bonding. The C atom forms two sigma bonds and two pi bonds. This should result in a linear shape around this C atom, with bond angles of 180° .

- (c) the C atom in formaldehyde, HCHO

Three hybrid sp^2 orbitals form during bonding. The C atom forms three sigma bonds and one pi bond. This should result in a trigonal planar shape around this C atom, with bond angles of 120° .

Shapes of Molecules, Extra Exercises

For each of the following molecules, draw a Lewis structure and use VSEPR theory to predict the shape around each central atom. Include a structural diagram to represent the molecular shape.

(Your predictions can be tested using a molecular model kit.)

1. hydrogen iodide, HI

2. silane, SiH₄

3. phosphine, PH₃

4. formaldehyde, H₂CO

5. hydrogen peroxide, H₂O₂

6. carbon dioxide, CO₂

7. phosphorus pentafluoride, PF₅

(continued)

8. nitrosyl chloride, NOCl

9. phosphorus oxychloride, POCl_3

10. uranium hexafluoride, UF_6

Shapes of Molecules, Extra Exercises, Solution

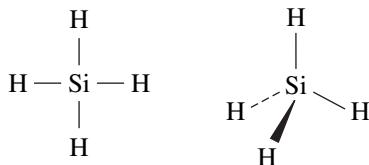
For each of the following molecules, draw a Lewis structure and use VSEPR theory to predict the shape around each central atom. Include a structural diagram to represent the molecular shape.

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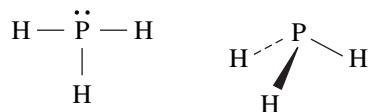
1. hydrogen iodide, HI
linear



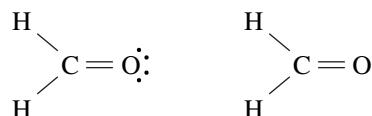
2. silane, SiH_4
tetrahedral



3. phosphine, PH_3
pyramidal



4. formaldehyde, H_2CO
trigonal planar



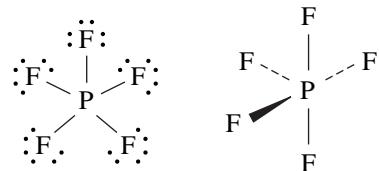
5. hydrogen peroxide, H_2O_2
V-shaped around each O



6. carbon dioxide, CO_2
linear



7. phosphorus pentafluoride, PF_5
trigonal bipyramidal

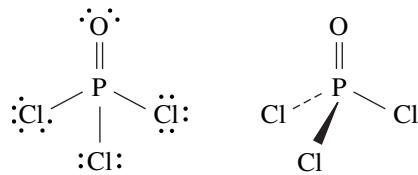


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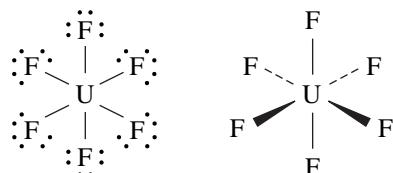
8. nitrosyl chloride, NOCl
V-shaped



9. phosphorus oxychloride, POCl_3
tetrahedral



10. uranium hexafluoride, UF_6
octahedral



A Defence of the VSEPR Model, Extra Exercises

In the early 1970s, the validity of VSEPR theory was questioned in the scientific literature by R. S. Drago. The article reproduced below is the response of R. J. Gillespie defending his theory. This type of dialogue is very common in all sciences. The purpose of this exercise is to investigate the nature of science as illustrated by the language used by one well-known scientist. The specific content, especially references to unfamiliar concepts, is not important here. Read the following excerpt and underline all examples of language depicting the nature of science; e.g., characteristics of science, appeals to evidence, reference to theories, uncertainty, or tentativeness.

A Defence of the Valence Shell Electron Pair Repulsion (VSEPR) Model

Ronald James Gillespie

Drago (J. Chem. Ed., 50, 244, 1973) has claimed that the factors that influence the geometry of molecules are more complicated than the VSEPR model would lead one to believe, and he has proposed a set of empirical rules which he claims are intellectually more satisfying than the VSEPR model. It is obvious that the prediction of the structure of a polyatomic molecule from first principles is a very complicated problem which at present necessitates the expenditure of very large amounts of computer time. Qualitative theories which enable one to "understand" and to predict the geometry of molecules are therefore still useful and the VSEPR model is such a qualitative theory. It does not pretend to give a complete explanation for the geometry of a molecule, but it does attempt to provide a simple model for the prediction of molecular geometry which is soundly based on quantum mechanics and which is in fact, in the author's opinion, more intellectually satisfying than a set of entirely empirical rules.

... It should be noted that the VSEPR theory makes no use of atomic orbitals or hybrid orbitals formed from these atomic orbitals, and it thus avoids the possibility of misleading a student into thinking that hybrid orbitals give an explanation for molecular geometry when, in fact, they provide nothing more than an approximate description of the bonding when one already knows the geometry. This is recognized by Drago who points out that after deciding the structure of a molecule by his rules "one can then *guess* at which hybrid orbitals would be used to point towards the atoms surrounding the central atom." It should be emphasized that the description of electrons in molecules in terms of atomic orbitals and hybrid orbitals formed from them is an approximation, which is not necessarily any more "correct" or any more useful than the approximation of representing each pair of electrons by a sphere.

... Bearing in mind that no approximate theory can be perfect, and that exceptions are always to be expected, it nevertheless seems reasonable to claim that the VSEPR theory is a useful, simple theory for the prediction of molecular geometry and that, in particular, it is more successful, more soundly based, and more intellectually satisfying than either Drago's empirical set of rules or the hybrid orbital theory. Drago's claim that the VSEPR theory involves a set of arbitrary rules which have to be memorized by the student is not justified, as these so-called rules are in fact a logical set of deductions based on the assumption that the fundamental

(continued)

property of electrons described by the Pauli exclusion principle is the most important factor in the determination of molecular geometry. The overall success of the theory seems to justify the conclusion that, in the vast majority of cases, the dominant factor in the determination of molecular shape is indeed the interactions between the largely localized electron pairs in the valence shell of a central atom. The justification of any theory is its reliability and usefulness for the prediction of new facts, and the VSEPR theory is more reliable for predicting molecular shapes than any other theory we have at present. In this sense it is the best theory we have, and it should surely therefore be taught to students, if only as one of several alternatives. Moreover, a logical and consistent theory is undoubtedly more satisfying and more stimulating for the student than any empirical set of rules, such as those proposed by Drago. It is surely not a valid criticism of a theory to say that it is incomplete — many successful theories are admittedly approximate and incomplete....

Reprinted, with permission, from the *Journal of Chemical Education* 15 (6 June 1974).

A Defence of the VSEPR Model, Extra Exercises, Solution

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(Answers may vary, but students should be able to defend their choices.)

A Defence of the Valence Shell Electron Pair Repulsion (VSEPR) Model

Ronald James Gillespie

Drago (J. Chem. Ed., 50, 244, 1973) has claimed that the factors that influence the geometry of molecules are more complicated than the VSEPR model would lead one to believe, and he has proposed a set of empirical rules which he claims are intellectually more satisfying than the VSEPR model. It is obvious that the prediction of the structure of a polyatomic molecule from first principles is a very complicated problem which at present necessitates the expenditure of very large amounts of computer time. Qualitative theories which enable one to "understand" and to predict the geometry of molecules are therefore still useful and the VSEPR model is such a qualitative theory. It does not pretend to give a complete explanation for the geometry of a molecule, but it does attempt to provide a simple model for the prediction of molecular geometry which is soundly based on quantum mechanics and which is in fact, in the author's opinion, more intellectually satisfying than a set of entirely empirical rules.

... It should be noted that the VSEPR theory makes no use of atomic orbitals or hybrid orbitals formed from these atomic orbitals, and it thus avoids the possibility of misleading a student into thinking that hybrid orbitals give an explanation for molecular geometry when, in fact, they provide nothing more than an approximate description of the bonding when one already knows the geometry. This is recognized by Drago who points out that after deciding the structure of a molecule by his rules "one can then *guess* at which hybrid orbitals would be used to point towards the atoms surrounding the central atom." It should be emphasized that the description of electrons in molecules in terms of atomic orbitals and hybrid orbitals formed from them is an approximation, which is not necessarily any more "correct" or any more useful than the approximation of representing each pair of electrons by a sphere.

... Bearing in mind that no approximate theory can be perfect, and that exceptions are always to be expected, it nevertheless seems reasonable to claim that the VSEPR theory is a useful, simple theory for the prediction of molecular geometry and that, in particular, it is more successful, more soundly based, and more intellectually satisfying than either Drago's empirical set of rules or the hybrid orbital theory. Drago's claim that the VSEPR theory involves a set of arbitrary rules which have to be memorized by the student is not justified, as these so-called rules are in fact a

(continued)

logical set of deductions based on the assumption that the fundamental property of electrons described by the Pauli exclusion principle is the most important factor in the determination of molecular geometry. The overall success of the theory seems to justify the conclusion that, in the vast majority of cases, the dominant factor in the determination of molecular shape is indeed the interactions between the largely localized electron pairs in the valence shell of a central atom. The justification of any theory is its reliability and usefulness for the prediction of new facts, and the VSEPR theory is more reliable for predicting molecular shapes than any other theory we have at present. In this sense it is the best theory we have, and it should surely therefore be taught to students, if only as one of several alternatives. Moreover, a logical and consistent theory is undoubtedly more satisfying and more stimulating for the student than any empirical set of rules, such as those proposed by Drago. It is surely not a valid criticism of a theory to say that it is incomplete — many successful theories are admittedly approximate and incomplete....

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Bonding and the State of Matter of Pure Substances, Extra Exercises

1. Complete the following blanks by writing in the type of bond and likely state(s) of matter at SATP, for each example.

Bonded particles	Bond type	Nature of bonding	State at SATP	Structure of matter
metal atoms	(a) _____	variable strengths, nondirectional bonds	(b) _____	continuous, close-packed crystals
metalloid atoms	(c) _____	very strong directional bonds	(d) _____	continuous, network crystals
ions	(e) _____	strong nondirectional bonds	(f) _____	continuous, regular crystals
nonmetal atoms	(g) _____	very strong directional bonds	*****	formation of single molecules
molecules	(h) _____ (i) _____ (j) _____	Relatively weak bonds: nondirectional bonds; somewhat directional bonds; directional bonds	(k) _____	continuous, often irregular arrangements

2. For each of the following substances, list all the types of chemical bonds believed to be present according to our current theories and predict the most likely state of matter at SATP.

- (a) Ge
- (b) Co
- (c) CO
- (d) MnO₂
- (e) CH₃COOH

- (f) SiC
- (g) C₂H₅Cl
- (h) CH₃NH₂

- (i) CuSO₄·5H₂O

- (j) Extension: UF₄ and UF₆

Bonding and the State of Matter of Pure Substances, Extra Exercises, Solution

1. Complete the following blanks by writing in the type of bond and likely state(s) of matter at SATP, for each example.

Bonded particles	Bond type	Nature of bonding	State at SATP	Structure of matter
metal atoms	(a) metallic	variable strengths, nondirectional bonds	(b) solid	continuous, close-packed crystals
metalloid atoms	(c) covalent	very strong directional bonds	(d) solid	continuous, network crystals
ions	(e) ionic	strong nondirectional bonds	(f) solid	continuous, regular crystals
nonmetal atoms	(g) covalent	very strong directional bonds	*****	formation of single molecules
molecules	(h) London forces (i) dipole-dipole (j) hydrogen	Relatively weak bonds: nondirectional bonds; somewhat directional bonds; directional bonds	(k) s, l, or g	continuous, often irregular arrangements

2. For each of the following substances, list all the types of chemical bonds believed to be present according to our current theories and predict the most likely state of matter at SATP.

- | | |
|---|---|
| (a) $\text{Ge}_{(s)}$ | covalent bonds |
| (b) $\text{Co}_{(s)}$ | metallic bonds |
| (c) $\text{CO}_{(g)}$ | covalent bond, London forces, dipole-dipole forces |
| (d) $\text{MnO}_{2(s)}$ | ionic bonds |
| (e) $\text{CH}_3\text{COOH}_{(l)}$ | covalent bonds, London forces, dipole-dipole forces, hydrogen bonds |
| (f) $\text{SiC}_{(s)}$ | covalent bonds |
| (g) $\text{C}_2\text{H}_5\text{Cl}_{(g \text{ or } l)}$ | covalent bond, London forces, dipole-dipole forces |
| (h) $\text{CH}_3\text{NH}_{2(g \text{ or } l)}$ | covalent bonds, London forces, dipole-dipole forces, hydrogen bonds |
| (i) $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}_{(s)}$ | ionic bonds, covalent bonds, London forces, dipole-dipole forces, hydrogen bonds |
| (j) $\text{UF}_{4(s)}$ | ionic bonds
<i>[Prediction of an ionic compound is seemingly correct in this case. This compound has a very high melting point of 2375°C.]</i> |
| $\text{UF}_{6(s)}$ | ionic bonds
<i>[Prediction appears to be false. This anomalous substance behaves like a molecular substance. It melts at about 64°C.]</i> |

Activity 4.3.1: Shapes of Molecules

Chemists use molecular models to study the shapes of molecules. Since the models are built to reflect the theory, they can be used to test your understanding of the theory. The purpose of this activity is to use VSEPR theory to predict the stereochemistry of some common molecules.

Materials

molecular models kit
a legend of colour codes for models of atoms

(a) Use VSEPR theory to predict the shape of the following molecules: CCl_4 , C_2H_4 , C_2F_2 , NCl_3 , OF_2 , and NH_2OH .

- Assemble molecular models for the chemicals listed in the question.
- (b) Sketch a 3-D diagram of each molecule assembled and classify its shape.
- (c) Evaluate the predictions that you made. How does your understanding of VSEPR theory have to be revised?

Investigation 4.4.1: Testing for Polar Molecules

In the winter when the humidity is relatively low, it is not unusual to acquire an electric charge by walking across a carpet in stocking feet or by pulling off a sweater.

A nonmetal rod or strip can also easily acquire a positive or negative charge by friction. As a diagnostic test, if you bring a charged rod or plastic strip near a liquid stream, and the stream is attracted to the rod or strip, then the molecules in the liquid are polar.

Purpose

The purpose of this investigation is to test the empirical rules provided in **Table 1**, Section 4.4 of the Student Text.

Question

Which of the liquids provided have polar molecules?

Prediction

- (a) Use the empirical rules in **Table 1**, Section 4.4 to predict whether the liquids provided contain polar molecules.



Check the MSDS sheets for all liquids used and follow appropriate safety precautions.

Experimental Design

A thin stream of each liquid is tested by holding a positively or negatively charged rod or plastic strip near the liquid (**Figure 1**, Section 4.4).

Materials

lab apron
safety glasses
medicine dropper or 50-mL buret
clamp and stand (if buret is used)
buret funnel (if buret is used)
400-mL beaker
acetate strip (marked +)
vinyl strip (marked -)
paper towel
various liquids

(continued)

Procedure

1. Fill the dropper/buret with one of the liquids.
2. Rub the acetate strip back and forth several times with a paper towel.
3. Allow drops or a thin stream of the liquid to pour into the waste beaker.
4. Hold the charged acetate strip close to the liquid stream and observe any effect or none.
5. Repeat steps 1 through 4 with the charged vinyl strip.
6. Move to the next station and repeat steps 1 through 5 using the next liquid provided.

Evidence

- (b) Prepare and complete a table for your observations.

Analysis

- (c) Answer the Question.

Evaluation

- (d) Evaluate the experimental design, materials, procedure, and skills employed.

- (e) Evaluate your prediction and the empirical rules used to make the prediction.

Lab Exercise 4.5.1: Boiling Points and Intermolecular Forces

INQUIRY SKILLS

- | | |
|--|--|
| <input type="radio"/> Questioning | <input type="radio"/> Recording |
| <input checked="" type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input checked="" type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input type="radio"/> Conducting | |

In all liquids, intermolecular forces are important, but these forces become negligible in the gas state for the conditions at which liquids boil. Therefore, we are looking at a situation where intermolecular forces must be overcome by adding energy, but no new bonds are formed. The temperature at which a liquid boils reflects the strength of the intermolecular forces present among the molecules. Higher temperatures mean more energy has been added and the intermolecular forces must have been stronger.

Purpose

The purpose of this lab exercise is to test the theory and rules for London and dipole–dipole forces.

Question

What is the trend in boiling points of the hydrogen compounds of elements in groups 14–17 (**Table 1**)?

Hypothesis/Prediction

- (a) Based upon dipole–dipole and London forces, write a prediction for the trend in boiling points within and between groups. Your prediction could include a general sketch of a graph of boiling point versus number of electrons per molecule. Provide your reasoning.

Analysis

- (b) Complete a graph of the evidence by plotting boiling point versus number of electrons per molecule.

(continued)

- (c) Answer the Question.

Evidence

Table 1 Boiling Points of the Hydrogen Compounds of Elements in Groups 14–17

Group	Hydrogen compound	Boiling point (°C)
14	CH _{4(g)}	−162
	SiH _{4(g)}	−112
	GeH _{4(g)}	−89
	SnH _{4(g)}	−52
15	NH _{3(g)}	−33
	PH _{3(g)}	−87
	AsH _{3(g)}	−55
	SbH _{3(g)}	−17
16	H ₂ O _{l(l)}	100
	H ₂ S _(g)	−61
	H ₂ Se _(g)	−42
	H ₂ Te _(g)	−2
17	HF _(g)	20
	HCl _(g)	−85
	HBr _(g)	−67
	HI _(g)	−36

Evaluation

- (d) Assuming that the evidence is valid, evaluate the Prediction and the concept of intermolecular forces used to make the prediction.
- (e) Are there any anomalies (unexpected evidence) in the evidence presented? Suggest an explanation.

Investigation 4.5.1: Hydrogen Bonding

INQUIRY SKILLS

- | | |
|--|--|
| <input type="radio"/> Questioning | <input checked="" type="radio"/> Recording |
| <input checked="" type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input checked="" type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input checked="" type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input checked="" type="radio"/> Conducting | |

Exothermic and endothermic physical and chemical processes are explained by comparing the energy required to break bonds and the energy released when bonds are formed. The net effect of bonds broken and bonds formed is either an exothermic (energy released) or an endothermic (energy absorbed) change. Thermochemical changes can be physical (phase changes) or chemical (chemical changes)—or solution formation, which does not seem to fit the physical–chemical classification.

Purpose

The purpose of this investigation is to test the concept of hydrogen bonding.

Question

How does the temperature change for the mixing of ethanol with water compare with the mixing of glycerol in water?

Hypothesis/Prediction

- (a) Write a prediction, complete with reasoning, based upon the concept of hydrogen bonding.

Experimental Design

Equal volumes of ethanol and water, and glycerol and water are mixed and the change in temperature is recorded.

- (b) Identify the variables.

Materials

lab apron	two 10-mL graduated cylinders
eye protection	nested pair of polystyrene cups
distilled water	cup lid with centre hole
ethanol	two thermometers
glycerol	250-mL beaker (for support)

(continued)

Procedure

- (c) Write a complete procedure including disposal instructions.

Analysis

- (d) Answer the Question.

Evaluation

- (e) Evaluate the evidence by judging the experimental design, materials, and procedure. Note any flaws and improvements.
- (f) How certain are you about the evidence obtained? Justify your answer, including possible experimental errors and uncertainties.
- (g) If you think that the evidence is of suitable quality, evaluate the prediction and reasoning used. If not, discuss your reasons.

Investigation 4.6.1: Classifying Mystery Solids

INQUIRY SKILLS

- | | |
|--|--|
| <ul style="list-style-type: none"><input type="radio"/> Questioning<input type="radio"/> Hypothesizing<input type="radio"/> Predicting<input checked="" type="radio"/> Planning<input checked="" type="radio"/> Conducting | <ul style="list-style-type: none"><input checked="" type="radio"/> Recording<input checked="" type="radio"/> Analyzing<input checked="" type="radio"/> Evaluating<input checked="" type="radio"/> Communicating |
|--|--|

When analyzing an unknown solid, physical properties can help to quickly narrow down the possibilities to a particular class of solids—ionic, metallic, molecular, and covalent network.

Question

To what class of solids do the four mystery solids belong?

Experimental Design

- (a) Write a general plan to answer the question.

Materials

- (b) Using your design and commonly available materials, prepare a list of the materials to be used.

Procedure

- (c) List all steps in the appropriate order to answer the question. Include safety precautions.

(continued)

Analysis

- (d) Identify the class of each of the solids.

Evaluation

- (e) Discuss the quality of your evidence and how certain you are about the answer obtained. Suggest some improvements to increase the certainty of the classification.

Heat Calculations, Extra Exercises

1. Calculate the quantity of heat required to warm 1.25 L of water from 22.0°C to 98.0°C in an electric kettle.
 2. What mass of aluminum in a car engine will absorb 1.00 MJ of heat when the temperature rises from 22°C to 102°C after the car is started?
 3. In a laboratory experiment, 2.00 kJ of heat flowed to a 100 g sample of a liquid solvent, causing a temperature increase from 15.40°C to 21.37°C. Calculate the specific heat capacity of the liquid.
 4. A human body loses about 360 kJ of heat every hour. Assuming that an average human body is equivalent to about 60 kg of water, what temperature decrease would this heat transfer cause? (Of course, this heat is replaced by body metabolism.)

Heat Calculations, Extra Exercises, Solution

1. Calculate the quantity of heat required to warm 1.25 L of water from 22.0°C to 98.0°C in an electric kettle.

$$q = mc\Delta T$$

$$q = 1250 \text{ g} \times 4.19 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \times (98.0 - 22.0)^\circ\text{C}$$

$$q = 398 \text{ kJ}$$

2. What mass of aluminum in a car engine will absorb 1.00 MJ of heat when the temperature rises from 22°C to 102°C after the car is started?

$$q = mc\Delta t$$

$$1.00 \text{ MJ} = m \times 0.900 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \times (102 - 22)^\circ\text{C}$$

$$m = 14 \text{ kg}$$

3. In a laboratory experiment, 2.00 kJ of heat flowed to a 100 g sample of a liquid solvent, causing a temperature increase from 15.40°C to 21.37°C. Calculate the specific heat capacity of the liquid.

$$q = mc\Delta t$$

$$2.00 \text{ MJ} = 100 \text{ g} \times c \times (21.37 - 15.40)^\circ\text{C}$$

$$c = 3.35 \text{ J}/(\text{g}\cdot^\circ\text{C})$$

4. A human body loses about 360 kJ of heat every hour. Assuming that an average human body is equivalent to about 60 kg of water, what temperature decrease would this heat transfer cause? (Of course, this heat is replaced by body metabolism.)

$$q = mc\Delta t$$

$$360 \text{ kJ} = 60 \text{ kg} \times 4.19 \frac{\text{J}}{\text{g}\cdot^\circ\text{C}} \times \Delta t$$

$$\Delta t = 1.4^\circ\text{C}$$

Enthalpy Changes and Calorimetry, Extra Exercises

1. Calculate the enthalpy change for the melting of a 30 g ice cube.
 2. A reference gives a value of +39.23 kJ/mol for the molar enthalpy of vaporization for methanol. What enthalpy change occurs in the evaporation of 10.0 g of methanol?
 3. An experiment produces evidence that the evaporation of 4.00 g of liquid butane, $C_4H_{10(l)}$, requires a gain in enthalpy of 1.67 kJ. Find the molar enthalpy of vaporization for butane from this evidence.
 4. A calorimeter has a heat capacity of 40.00 kJ/ $^{\circ}C$. Complete combustion of 1.00 g of hydrogen in this calorimeter causes a temperature increase of $3.54^{\circ}C$. Calculate the molar enthalpy of combustion for hydrogen from this evidence.
 5. Combustion of 3.50 g of ethanol, $C_2H_5OH_{(l)}$, in a calorimeter with a heat capacity of 15.2 kJ/ $^{\circ}C$ causes a temperature increase from $19.88^{\circ}C$ to $26.18^{\circ}C$. Find the molar enthalpy of combustion for ethanol from this evidence.

Enthalpy Changes and Calorimetry, Extra Exercises, Solution

1. Calculate the enthalpy change for the melting of a 30 g ice cube.

$$\begin{aligned}\Delta H &= n\Delta H_{\text{fus}} \\ &= 30 \text{ g} \times \frac{1 \text{ mol}}{18.02 \text{ g}} \times \frac{6.03 \text{ kJ}}{\text{mol}} \\ &= 10 \text{ kJ}\end{aligned}$$

2. A reference gives a value of +39.23 kJ/mol for the molar enthalpy of vaporization for methanol. What enthalpy change occurs in the evaporation of 10.0 g of methanol?

$$\begin{aligned}\Delta H &= n\Delta H_{\text{vap}} \\ &= 10 \text{ g} \times \frac{1 \text{ mol}}{32.05 \text{ g}} \times \frac{39.23 \text{ kJ}}{\text{mol}} \\ &= 12.2 \text{ kJ}\end{aligned}$$

3. An experiment produces evidence that the evaporation of 4.00 g of liquid butane, $\text{C}_4\text{H}_{10(\text{l})}$, requires a gain in enthalpy of 1.67 kJ. Find the molar enthalpy of vaporization for butane from this evidence.

$$\begin{aligned}\Delta H &= n\Delta H_{\text{vap}} \\ 1.67 \text{ kJ} &= 4.00 \text{ g} \times \frac{1 \text{ mol}}{58.14 \text{ g}} \times H_{\text{vap}} \\ \Delta H_{\text{vap}} &= 24.3 \text{ kJ/mol}\end{aligned}$$

4. A calorimeter has a heat capacity of 40.00 kJ/ $^{\circ}\text{C}$. Complete combustion of 1.00 g of hydrogen in this calorimeter causes a temperature increase of 3.54°C . Calculate the molar enthalpy of combustion for hydrogen from this evidence.

$$\begin{aligned}\Delta H &= q \\ \text{H}_2 &\quad (\text{calorimeter}) \\ n\Delta H_c &= C\Delta t \\ 1.00 \text{ g} \times \frac{1 \text{ mol}}{2.02 \text{ g}} \times \Delta H_c &= 40.00 \frac{\text{kJ}}{^{\circ}\text{C}} \times 3.54^{\circ}\text{C} \\ \Delta H_c &= 286 \text{ kJ/mol}\end{aligned}$$

5. Combustion of 3.50 g of ethanol, $\text{C}_2\text{H}_5\text{OH}_{(\text{l})}$, in a calorimeter with a heat capacity of 15.2 kJ/ $^{\circ}\text{C}$ causes a temperature increase from 19.88°C to 26.18°C . Find the molar enthalpy of combustion for ethanol from this evidence.

$$\begin{aligned}\Delta H &= q \\ \text{C}_2\text{H}_5\text{OH} &\quad (\text{calorimeter}) \\ n\Delta H_c &= C\Delta t \\ 3.50 \text{ g} \times \frac{1 \text{ mol}}{46.08 \text{ g}} \times \Delta H_c &= 15.2 \frac{\text{kJ}}{^{\circ}\text{C}} \times (26.18 - 19.88)^{\circ}\text{C} \\ \Delta H_c &= 1.26 \text{ MJ/mol}\end{aligned}$$

Representing Enthalpy Changes, Extra Exercises

1. Iron(II) sulfide ore is roasted according to the following chemical equation.



(a) Rewrite this chemical equation including the energy as a term in the balanced equation.

(b) What is the molar enthalpy for iron(II) sulfide in this reaction?

(c) What is the molar enthalpy for iron(III) oxide in this reaction?

2. Boron reacts with hydrogen to form diboron hexahydride (diborane) gas. The molar enthalpy of reaction for boron is +15.7 kJ/mol.

(a) Write the balanced chemical equation using whole number coefficients and including the energy change as a ΔH_r .

(b) Write the balanced chemical equation using whole number coefficients and including the energy change as a term in the balanced equation.

3. The molar enthalpy of combustion for octane, $\text{C}_8\text{H}_{18(\text{l})}$, is reported to be -1.3 MJ/mol.

(a) Write the balanced chemical equation using whole number coefficients and including the energy change as a ΔH_r .

(b) Write the balanced chemical equation using whole number coefficients and including the energy change as a term in the balanced equation.

(continued)

LSM 5.3-2

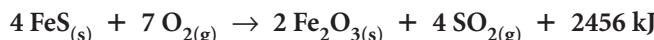
4. Draw potential energy diagrams to communicate the following chemical reactions. Assume SATP conditions.
- (a) the formation of chromium(III) oxide
 - (b) the simple decomposition of silver iodide
 - (c) the formation of carbon disulfide

Representing Enthalpy Changes, Extra Exercises, Solution

1. Iron(II) sulfide ore is roasted according to the following chemical equation.



- (a) Rewrite this chemical equation including the energy as a term in the balanced equation.



- (b) What is the molar enthalpy for iron(II) sulfide in this reaction?

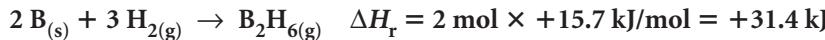
$$\frac{H_c}{\text{FeS}} = \frac{-2456 \text{ kJ}}{4 \text{ mol}} = -614 \text{ kJ/mol}$$

- (c) What is the molar enthalpy for iron(III) oxide in this reaction?

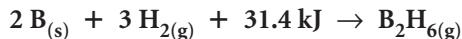
$$\frac{H_c}{\text{Fe}_2\text{O}_3} = \frac{-2456 \text{ kJ}}{2 \text{ mol}} = -1228 \text{ kJ/mol}$$

2. Boron reacts with hydrogen to form diboron hexahydride (diborane) gas. The molar enthalpy of reaction for boron is 115.7 kJ/mol.

- (a) Write the balanced chemical equation using whole number coefficients and including the energy change as a ΔH_r .

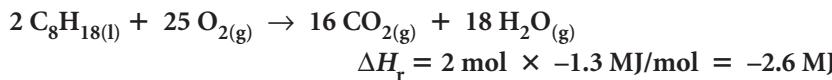


- (b) Write the balanced chemical equation using whole number coefficients and including the energy change as a term in the balanced equation.



3. The molar enthalpy of combustion for octane, $\text{C}_8\text{H}_{18(\text{l})}$, is reported to be -1.3 MJ/mol .

- (a) Write the balanced chemical equation using whole number coefficients and including the energy change as a ΔH_r .



- (b) Write the balanced chemical equation using whole number coefficients and including the energy change as a term in the balanced equation.

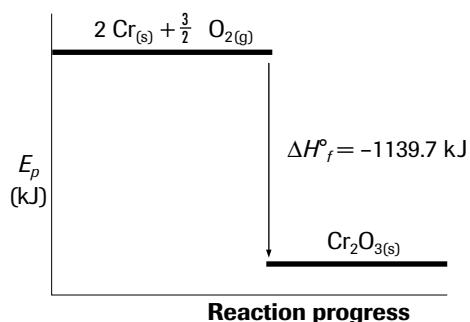


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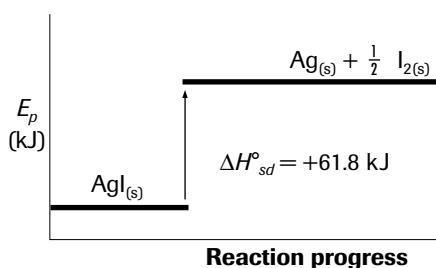
4. Draw potential energy diagrams to communicate the following chemical reactions. Assume SATP conditions.

- the formation of chromium(III) oxide
- the simple decomposition of silver iodide
- the formation of carbon disulfide

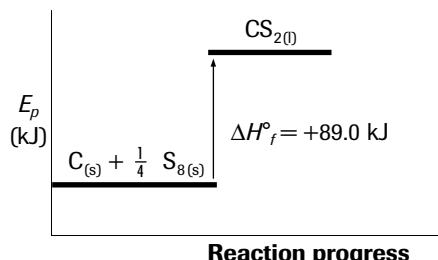
(a) The Formation of Chromium(III) Oxide



(b) The Simple Decomposition of Silver Iodide

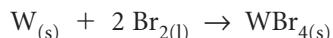


(c) The Formation of Carbon Disulfide

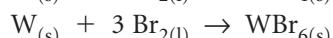


Predicting ΔH Using Hess's Law, Extra Exercises

1. The enthalpy changes for the formation of two wolfram bromides are shown below.

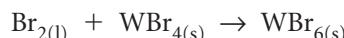


$$\Delta H_1^\circ = -146.7 \text{ kJ}$$



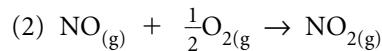
$$\Delta H_2^\circ = -184.4 \text{ kJ}$$

Calculate the standard enthalpy change for the following reaction.



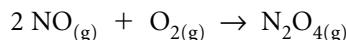
2. Given: (1) $N_2O_{4(g)} \rightarrow 2 NO_{2(g)}$

$$\Delta H_1^\circ = +58 \text{ kJ}$$

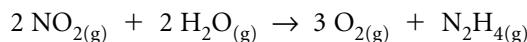


$$\Delta H_2^\circ = -56 \text{ kJ}$$

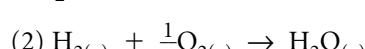
Calculate the standard enthalpy change for the following reaction.



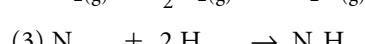
3. Use the following reactions and enthalpy changes to predict the standard enthalpy change for



$$\Delta H_1^\circ = +33.2 \text{ kJ}$$



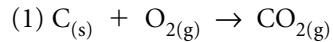
$$\Delta H_2^\circ = -241.8 \text{ kJ}$$



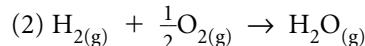
$$\Delta H_3^\circ = +47.6 \text{ kJ}$$



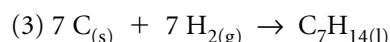
4. Use the following formation reaction evidence to calculate the standard enthalpy change for the complete combustion of cycloheptane.



$$\Delta H_1^\circ = -393.5 \text{ kJ}$$



$$\Delta H_2^\circ = -241.8 \text{ kJ}$$



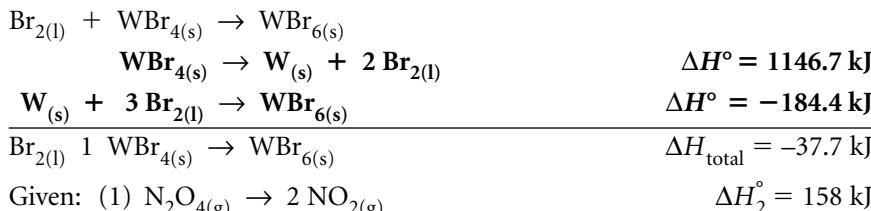
$$\Delta H_3^\circ = +115.0 \text{ kJ}$$

Predicting ΔH Using Hess's Law, Extra Exercises, Solution

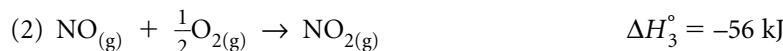
1. The enthalpy changes for the formation of two wolfram bromides are shown below.



Calculate the standard enthalpy change for the following reaction.



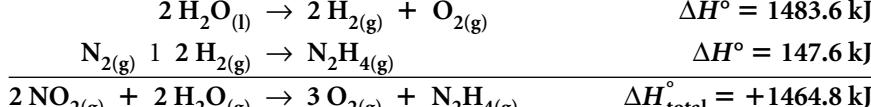
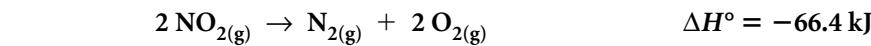
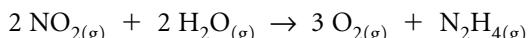
2. Given: (1) $N_2O_{4(g)} \rightarrow 2 NO_{2(g)}$ $\Delta H_1^\circ = 158 \text{ kJ}$



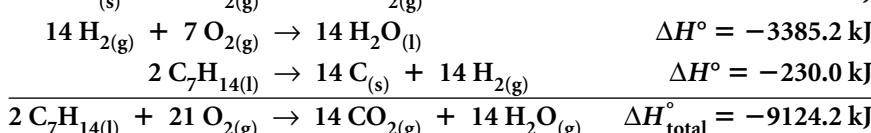
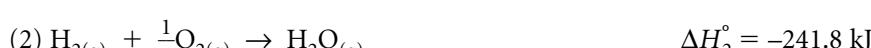
Calculate the standard enthalpy change for the following reaction.



3. Use the following reactions and enthalpy changes to predict the standard enthalpy change for



4. Use the following formation reaction evidence to calculate the standard enthalpy change for the complete combustion of cycloheptane.



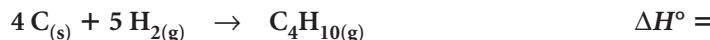
Hess's Law, Extra Exercises

Instructions

1. Cut the known equations into strips. You will have a known equation A on one side of the strip, and its reverse known equation B on the other side.
 2. Group the known equations with their proper target equations below.
 3. In each group, flip the known equations until the products and the reactants match those of the target equation.
 4. Glue the known equations on this sheet so the proper known equations are above the target equations.
 5. Number the known equations in each group.
 6. Multiply the known equations in each group until the stoichiometry of the known equations matches that of the target equation.
 7. Add the equations and enthalpies to solve for the target.
- (a) What is the enthalpy change for the formation of two moles of nitrogen monoxide from its elements?



- (b) What is the enthalpy change for the formation of one mole of butane (C_4H_{10}) gas from its elements?



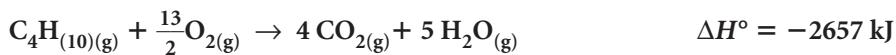
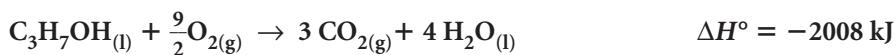
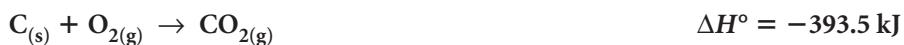
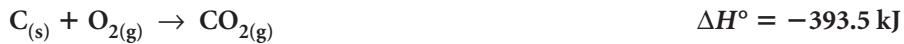
- (c) Determine the enthalpy change involved in the formation of two moles of liquid propanol.



(continued)

Known Equations A

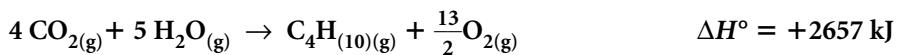
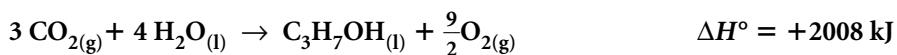
Note: Photocopy Known Equations A on one side of a sheet and Known Equations B on the reverse side of the same sheet.



(continued)

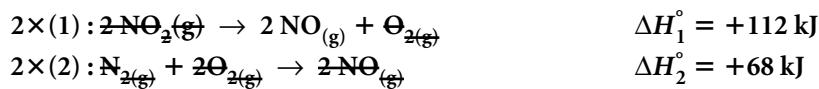
Known Equations B

Note: Photocopy Known Equations A on one side of a sheet and Known Equations B on the reverse side of the same sheet.

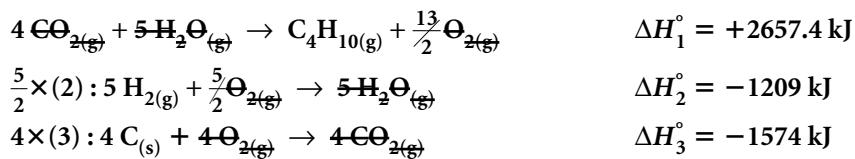


Hess's Law, Extra Exercises, Solution

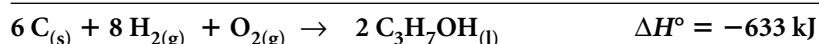
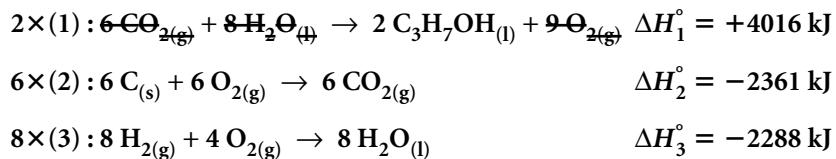
- (a) What is the enthalpy change for the formation of two moles of nitrogen monoxide from its elements?



- (b) What is the enthalpy change for the formation of one mole of butane (C_4H_{10}) gas from its elements?



- (c) Determine the enthalpy change involved in the formation of two moles of liquid propanol.



Predicting ΔH Using Standard Enthalpies of Formation, Extra Exercises

1. Laboratory quantities of ethylene can be prepared by an elimination reaction of ethanol using an acid catalyst. Calculate the enthalpy change for the conversion of ethanol into ethylene and water using standard enthalpies of formation.
 2. Calculate the enthalpy of combustion for acetic acid using standard enthalpies of formation.
 3. An initial step in the production of iron in a blast furnace involves the conversion of iron(III) oxide and carbon monoxide into iron(II, III) oxide and carbon dioxide. Calculate the enthalpy of this reaction using standard enthalpies of formation.

(continued)

LSM 5.5-2

4. The fertilizer urea is produced along with liquid water by the reaction of ammonia and carbon dioxide. Calculate the enthalpy of this reaction using standard enthalpies of formation.

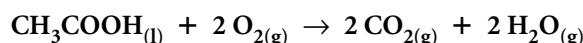
Predicting ΔH Using Standard Enthalpies of Formation, Extra Exercises, Solution

1. Laboratory quantities of ethylene can be prepared by an elimination reaction of ethanol using an acid catalyst. Calculate the enthalpy change for the conversion of ethanol into ethylene and water using standard enthalpies of formation.



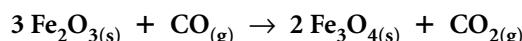
$$\begin{aligned}\Delta H &= \sum nH_f^\circ_{(\text{products})} - \sum nH_f^\circ_{(\text{reactants})} \\ &= \left(1 \text{ mol} \times \frac{+52.5 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \right) - \left(1 \text{ mol} \times \frac{-235.2 \text{ kJ}}{1 \text{ mol}} \right) \\ &= -233.3 \text{ kJ} - (-235.2 \text{ kJ}) \\ &= +1.9 \text{ kJ}\end{aligned}$$

2. Calculate the enthalpy of combustion for acetic acid using standard enthalpies of formation.



$$\begin{aligned}\Delta H &= \sum nH_f^\circ_{(\text{products})} - \sum nH_f^\circ_{(\text{reactants})} \\ &= \left(2 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} + 2 \text{ mol} \times \frac{-241.8 \text{ kJ}}{1 \text{ mol}} \right) \\ &\quad - \left(1 \text{ mol} \times \frac{-432.8 \text{ kJ}}{1 \text{ mol}} + 2 \text{ mol} \times \frac{0 \text{ kJ}}{1 \text{ mol}} \right) \\ &= -1270.6 \text{ kJ} - (-432.8 \text{ kJ}) \\ &= -837.8 \text{ kJ}\end{aligned}$$

3. An initial step in the production of iron in a blast furnace involves the conversion of iron(III) oxide and carbon monoxide into iron(II, III) oxide and carbon dioxide. Calculate the enthalpy of this reaction using standard enthalpies of formation.



$$\begin{aligned}\Delta H &= \sum nH_f^\circ_{(\text{products})} - \sum nH_f^\circ_{(\text{reactants})} \\ &= \left(2 \text{ mol} \times \frac{-1118.4 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} \right) \\ &\quad - \left(3 \text{ mol} \times \frac{-824.2 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-110.5 \text{ kJ}}{1 \text{ mol}} \right) \\ &= -2630.3 \text{ kJ} - (-2583.1 \text{ kJ}) \\ &= -47.2 \text{ kJ}\end{aligned}$$

LSM 5.5-3

4. The fertilizer urea is produced along with liquid water by the reaction of ammonia and carbon dioxide. Calculate the enthalpy of this reaction using standard enthalpies of formation.

$$\begin{aligned}2 \text{ NH}_{3(\text{g})} + \text{ CO}_{2(\text{g})} &\rightarrow \text{ CO(NH}_2\text{)}_{2(\text{s})} + \text{ H}_2\text{O}_{(\text{l})} \\ \Delta H &= \sum nH_f^\circ_{(\text{products})} - \sum nH_f^\circ_{(\text{reactants})} \\ &= \left(1 \text{ mol} \times \frac{-333.5 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \right) \\ &\quad - \left(2 \text{ mol} \times \frac{-45.9 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} \right) \\ &= -619.3 \text{ kJ} - (-485.3 \text{ kJ}) \\ &= -134.0 \text{ kJ}\end{aligned}$$

Multi-Step Energy Calculations, Extra Exercises

1. For the following combustion, what mass of carbon dioxide is produced when 1500 kJ of energy is released?



2. How much energy is released when 1.00 t of sulfur trioxide is produced by the following reaction?



3. In respiration, glucose is oxidized by oxygen gas to produce carbon dioxide gas, liquid water, and energy. What is the energy released when 18.0 g of glucose is consumed?

4. Methanol is burned in a bomb calorimeter. Liquid water is formed as a product. If 3.40 g of methanol reacts, what is the expected temperature change in a calorimeter with a heat capacity of 6.75 kJ/ $^{\circ}\text{C}$?

5. A waste heat exchanger is used to absorb the energy from the complete combustion of hydrogen sulfide gas. What volume of water undergoing a temperature change of 64°C is required to absorb all of the energy from the burning of 15 kg of hydrogen sulfide?

Multi-Step Energy Calculations, Extra Exercises, Solution

1. For the following combustion, what mass of carbon dioxide is produced when 1500 kJ of energy is released?



$$\Delta H_c = \frac{\Delta H}{n} = \frac{-2502 \text{ kJ}}{4 \text{ mol}} = -626 \text{ kJ/mol CO}_2$$

$$\Delta H = n\Delta H_c$$

$$1500 \text{ kJ} = m \times \frac{1 \text{ mol}}{44.01 \text{ g}} \times 626 \text{ kJ/mol}$$

$$m = 106 \text{ g}$$

2. How much energy is released when 1.00 t of sulfur trioxide is produced by the following reaction?

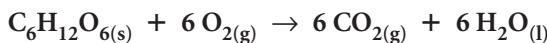


$$\Delta H_r = \frac{\Delta H}{n} = \frac{-192.8 \text{ kJ}}{2 \text{ mol}} = -96.4 \text{ kJ/mol SO}_3$$

$$\Delta H = n\Delta H_c$$

$$= 1.00 \text{ Mg} \times \frac{1 \text{ mol}}{80.06 \text{ g}} \times \frac{96.4 \text{ kJ}}{1 \text{ mol}} = 1.20 \text{ GJ}$$

3. In respiration, glucose is oxidized by oxygen gas to produce carbon dioxide gas, liquid water, and energy. What is the energy released when 18.0 g of glucose is consumed?



$$\Delta H = \sum nH_f^\circ_{\text{(products)}} - \sum nH_f^\circ_{\text{(reactants)}}$$

$$= \left(6 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} + 6 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \right) -$$

$$\left(1 \text{ mol} \times \frac{-1273.1 \text{ kJ}}{1 \text{ mol}} + 6 \text{ mol} \times \frac{0 \text{ kJ}}{1 \text{ mol}} \right) = -2802.7 \text{ kJ}$$

$$H_c = \frac{-2802.7 \text{ kJ}}{1 \text{ mol}} = -2802.7 \text{ kJ/mol C}_6\text{H}_{12}\text{O}_6$$

$$\Delta H = n\Delta H_c$$

$$= 18.0 \text{ g} \times \frac{1 \text{ mol}}{180.18 \text{ g}} \times \frac{2802.7 \text{ kJ}}{1 \text{ mol}} = 280 \text{ kJ}$$

(continued)

4. Methanol is burned in a bomb calorimeter. Liquid water is formed as a product. If 3.40 g of methanol reacts, what is the expected temperature change in a calorimeter with a heat capacity of 6.75 kJ/°C?

$$\begin{aligned}
 \text{CH}_3\text{OH}_{(\text{l})} + \frac{3}{2}\text{O}_{2(\text{g})} &\rightarrow \text{CO}_{2(\text{g})} + 2\text{H}_2\text{O}_{(\text{l})} \\
 \Delta H &= \sum nH_f^\circ_{\text{(products)}} - \sum nH_f^\circ_{\text{(reactants)}} \\
 &= \left(1 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} + 2 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \right) \\
 &\quad - \left(1 \text{ mol} \times \frac{-239.1 \text{ kJ}}{1 \text{ mol}} + \frac{3}{2} \text{ mol} \times \frac{0 \text{ kJ}}{1 \text{ mol}} \right) = -726.0 \text{ kJ} \\
 \Delta H_c &= \frac{726.0 \text{ kJ}}{1 \text{ mol}} = -726.0 \text{ kJ/mol CH}_3\text{OH} \\
 \Delta H_c &= C\Delta t \\
 &\quad (\text{methanol}) \quad (\text{calorimeter}) \\
 3.40 \text{ g} \times \frac{1 \text{ mol}}{32.05 \text{ g}} \times \frac{726.0 \text{ kJ}}{1 \text{ mol}} &= \frac{6.75 \text{ kJ}}{\text{°C}} \times \Delta t \\
 \Delta t &= 11.4^\circ\text{C}
 \end{aligned}$$

5. A waste heat exchanger is used to absorb the energy from the complete combustion of hydrogen sulfide gas. What volume of water undergoing a temperature change of 64°C is required to absorb all of the energy from the burning of 15 kg of hydrogen sulfide?

$$\begin{aligned}
 \text{H}_2\text{S}_{(\text{g})} + \frac{3}{2}\text{O}_{2(\text{g})} &\rightarrow \text{SO}_{2(\text{g})} + \text{H}_2\text{O}_{(\text{g})} \\
 \Delta H &= \sum nH_f^\circ_{\text{(products)}} - \sum nH_f^\circ_{\text{(reactants)}} \\
 &= \left(1 \text{ mol} \times \frac{-296.8 \text{ kJ}}{1 \text{ mol}} + 1 \text{ mol} \times \frac{-241.8 \text{ kJ}}{1 \text{ mol}} \right) \\
 &\quad - \left(1 \text{ mol} \times \frac{-20.6 \text{ kJ}}{1 \text{ mol}} + \frac{3}{2} \text{ mol} \times \frac{0 \text{ kJ}}{1 \text{ mol}} \right) = -518.0 \text{ kJ} \\
 \Delta H_c &= \frac{-518.0 \text{ kJ}}{1 \text{ mol}} = \frac{-518.0 \text{ kJ}}{\text{mol H}_2\text{S}} \\
 \Delta H &= q \\
 &\quad \text{H}_2\text{S} \quad (\text{water}) \\
 n\Delta H_c &= mc\Delta t \\
 15 \text{ kg} \times \frac{1 \text{ mol}}{34.08 \text{ g}} \times \frac{518.0 \text{ kJ}}{1 \text{ mol}} &= m \times 4.19 \frac{\text{kJ}}{\text{g} \cdot \text{°C}} \times 64^\circ\text{C} \\
 m &= 850 \text{ g} \\
 v &= 0.85 \text{ L}
 \end{aligned}$$

Multi-Step Energy Calculations, Extra Exercises, Solution

1. For the following combustion, what mass of carbon dioxide is produced when 1500 kJ of energy is released?



$$\Delta H_c = \frac{\Delta H}{n} = \frac{-2502 \text{ kJ}}{4 \text{ mol}} = -626 \text{ kJ/mol CO}_2$$

$$\Delta H = n\Delta H_c$$

$$1500 \text{ kJ} = m \times \frac{1 \text{ mol}}{44.01 \text{ g}} \times 626 \text{ kJ/mol}$$

$$m = 106 \text{ g}$$

2. How much energy is released when 1.00 t of sulfur trioxide is produced by the following reaction?

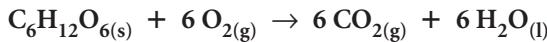


$$\Delta H_r = \frac{\Delta H}{n} = \frac{-192.8 \text{ kJ}}{2 \text{ mol}} = -96.4 \text{ kJ/mol SO}_3$$

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3. In respiration, glucose is oxidized by oxygen gas to produce carbon dioxide gas, liquid water, and energy. What is the energy released when 18.0 g of glucose is consumed?



$$\Delta H = \sum nH_f^\circ_{\text{(products)}} - \sum nH_f^\circ_{\text{(reactants)}}$$

$$= \left(6 \text{ mol} \times \frac{-393.5 \text{ kJ}}{1 \text{ mol}} + 6 \text{ mol} \times \frac{-285.8 \text{ kJ}}{1 \text{ mol}} \right) -$$

$$\left(1 \text{ mol} \times \frac{-1273.1 \text{ kJ}}{1 \text{ mol}} + 6 \text{ mol} \times \frac{0 \text{ kJ}}{1 \text{ mol}} \right) = -2802.7 \text{ kJ}$$

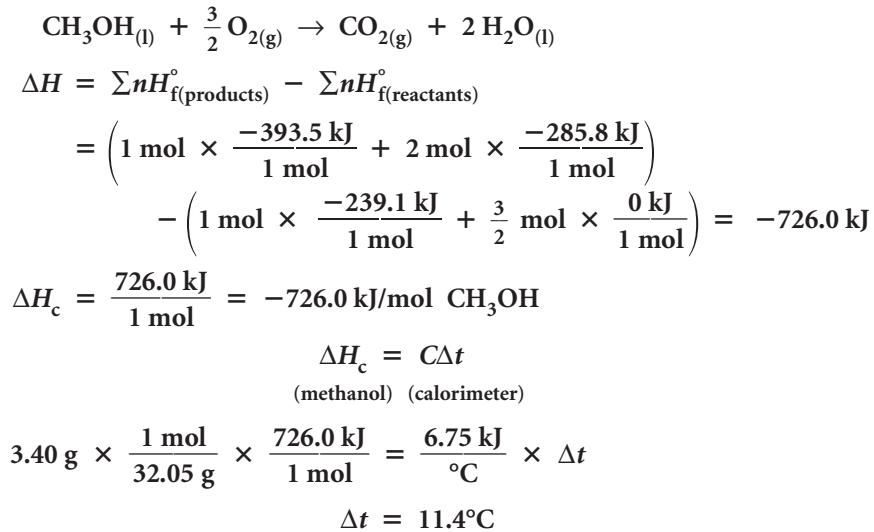
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$$\Delta H = n\Delta H_c$$

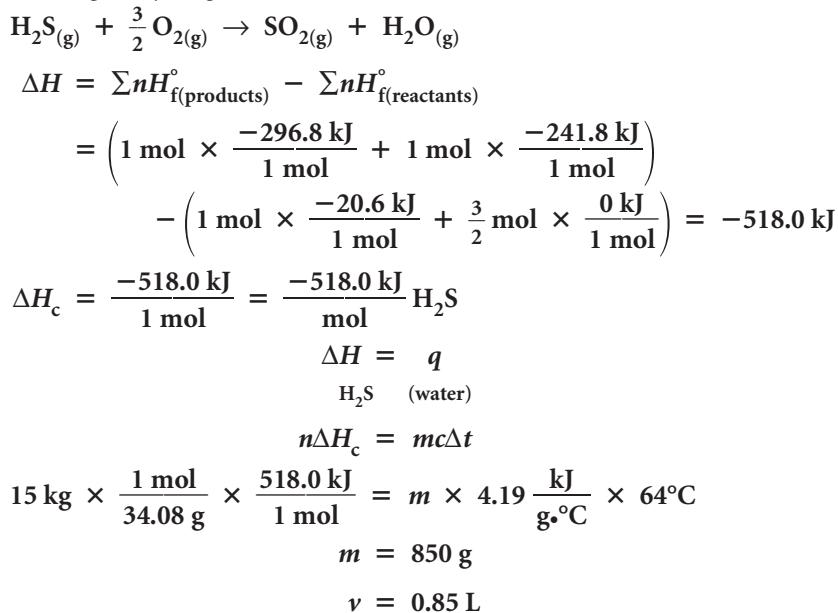
$$= 18.0 \text{ g} \times \frac{1 \text{ mol}}{180.18 \text{ g}} \times \frac{2802.7 \text{ kJ}}{1 \text{ mol}} = 280 \text{ kJ}$$

(continued)

4. Methanol is burned in a bomb calorimeter. Liquid water is formed as a product. If 3.40 g of methanol reacts, what is the expected temperature change in a calorimeter with a heat capacity of 6.75 kJ/°C?



5. A waste heat exchanger is used to absorb the energy from the complete combustion of hydrogen sulfide gas. What volume of water undergoing a temperature change of 64°C is required to absorb all of the energy from the burning of 15 kg of hydrogen sulfide?



Investigation 5.1.1: Medical Cold Packs

INQUIRY SKILLS

- | | |
|--|---|
| <input type="radio"/> Questioning
<input type="radio"/> Hypothesizing
<input type="radio"/> Predicting
<input checked="" type="radio"/> Planning
<input checked="" type="radio"/> Conducting | <input checked="" type="radio"/> Recording
<input checked="" type="radio"/> Analyzing
<input checked="" type="radio"/> Evaluating
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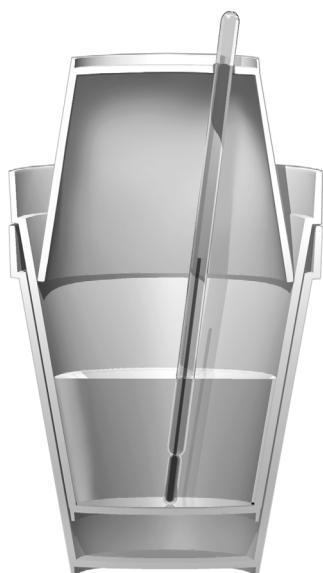


Figure 1

A simple laboratory calorimeter consists of an insulated container made of three nested polystyrene cups, a measured quantity of water, and a thermometer. The chemical system is placed in or dissolved in the water of the calorimeter. Energy transfers between the chemical system and the surrounding water are monitored by measuring changes in the temperature of the water.

We can study physical changes that involve liquids and aqueous solutions, using a polystyrene calorimeter like the one shown in **Figure 1**. Such materials can be applied to practical thermochemical systems. For example, when athletes are injured, they may immediately hold an “instant cold pack” against the injury. The medical cold pack operates on the principle that certain salts dissolve endothermically in water. The amount of heat per unit mass involved in the dissolving of a compound is a characteristic property of that substance. It is called the enthalpy of solution. **Table 1** lists some compounds that might be possible candidates for a cold pack because they absorb energy when they dissolve.

Table 1 Enthalpies of Solution for Compounds in a Medical Cold Pack

Salt	Enthalpy of solution
ammonium chloride, NH_4Cl	0.277 kJ/g
potassium nitrate, KNO_3	0.345 kJ/g
ammonium nitrate, NH_4NO_3	0.321 kJ/g
sodium acetate trihydrate, $\text{NaC}_2\text{H}_3\text{O}_2 \cdot 3 \text{H}_2\text{O}$	0.144 kJ/g
potassium chloride, KCl	0.231 kJ/g

Purpose

The purpose of this investigation is to use calorimetry to determine the enthalpy of solution of an unknown salt, and then to use that value to identify the salt in a medical cold pack.

Question

What is the identity of the unknown salt?

Experimental Design

You will be given a solid sample of about 10 g of a salt typically used in medical cold packs and listed in **Table 1**.

- Design an experiment to determine the heat transferred when a measured mass of the salt is dissolved in water.

You will then apply calorimetric calculations to discover the identity of the salt.

(continued)

LSM 5.LA-1

- (b) Write a detailed description of the calculations that you will use to determine the identity of the salt. Show all formulas and units that you will use. You will find it convenient to assign symbols (e.g., m_1 , m_2 , T_1 , etc.) to the measurements that you expect to make so that your calculations are clear.

Materials

lab apron
eye protection
centigram or analytical balance
8–10 g of an unknown salt (from the list in **Table 1**)
water
thermometer
3 Styrofoam cups
100-mL graduated cylinder

Procedure

- (c) Write your Procedure as a series of numbered steps, clearly identifying the masses, volumes, and specific equipment to be used (see Materials list), necessary safety and disposal considerations, and a table in which to record your observations.

1. Have your Experimental Design, calculations, and Procedure approved by your teacher before performing your experiment.

Analysis

- (d) Determine, by calculation, the identity of the unknown salt.

Evaluation

- (e) Calculate a percentage difference between your experimental value and the accepted value for enthalpy of solution of the identified salt.

(continued)

LSM 5.LA-1

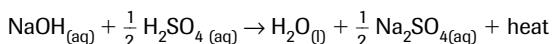
- (f) How confident are you in your identification? Suggest three sources of experimental error in this experiment.
- (g) If some heat were transferred to the air or Styrofoam cups, would your calculated enthalpy of solution of the unknown salt be too high or too low? Explain.
- (h) If some salt were accidentally spilled as it was transferred from the balance to the Styrofoam cup, would your calculated enthalpy of the unknown salt be too high or too low? Explain.

Investigation 5.2.1: Molar Enthalpy of a Chemical Change

INQUIRY SKILLS

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|---|---|
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<input type="radio"/> Hypothesizing
<input type="radio"/> Predicting
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<input checked="" type="radio"/> Conducting | <input checked="" type="radio"/> Recording
<input checked="" type="radio"/> Analyzing
<input checked="" type="radio"/> Evaluating
<input checked="" type="radio"/> Communicating |
|---|---|

When aqueous solutions of acids and bases react in a calorimeter, the solutions may act as both system and surroundings. The acid and base form the system as they react to form water and a dissolved salt. For example, sodium hydroxide and sulfuric acid react to form sodium sulfate and water:



The reactant solutions are mostly water containing dissolved and dispersed acid and base particles. Thus, a solution of an acid or a base may be regarded for *calorimetric* purposes to have the same specific heat capacity as water, at least to the degree of experimental accuracy that applies to simple calorimeters. For example, when 100 mL of a dilute acid solution reacts exothermically with 150 mL of a solution of a dilute base, the acid and base may be thought to react and release heat to a total of 250 mL of water. Assuming that the solutions have the same density as water (1.00 g/mL) also makes calculations simpler.

Purpose

The purpose of this investigation is to use calorimetry to obtain an empirical value for the molar enthalpy of neutralization of sodium hydroxide by sulfuric acid.

Question

What is the molar enthalpy of neutralization (ΔH_{neut}) of sodium hydroxide with sulfuric acid?

Experimental Design

A measured volume of sodium hydroxide solution of known concentration will be combined with excess sulfuric acid solution of known concentration. (The sodium hydroxide will be the limiting reagent.) Calorimetric measurements will be made to determine the heat produced by the reaction.

- (a) Demonstrate by calculation that the quantity of acid identified in the procedure will be enough to completely consume the base, and that the base will therefore be the limiting reagent for calculation purposes.



CAUTION: 1.0 mol/L sodium hydroxide solution is corrosive, and 1.0 mol/L sulfuric acid solution is an irritant.

Wear eye protection and a lab apron. Clean up any spills immediately. At these dilutions, the chemicals may be disposed of down the drain with lots of water.

Materials

- | | |
|-------------------------------------|--------------------------------|
| lab apron | thermometer |
| eye protection | polystyrene calorimeter |
| 1.0 mol/L sodium hydroxide solution | two 100-mL graduated cylinders |
| 1.0 mol/L sulfuric acid solution | |

(continued)

Procedure

1. Add 50 mL of 1.0 mol/L sodium hydroxide solution to a polystyrene calorimeter. Measure and record its temperature.
2. Measure and record the temperature of a 30-mL sample of 1.0 mol/L sulfuric acid solution in a graduated cylinder.
3. Carefully add the acid to the base, stirring slowly with the thermometer. Measure and record the maximum temperature obtained.

Analysis

(b) Calculate

(i) the masses of acid and base solutions;

(ii) the temperature changes in the acid and base solutions;

(iii) the total heat absorbed by the calorimetric liquids (acid and base);

(iv) the amount of base (in moles) that reacted; and

(v) the molar enthalpy of neutralization of the sodium hydroxide.

Evaluation

- (c) Calculate a percentage difference by comparing your experimental values and the accepted values for enthalpy of neutralization with respect to sodium hydroxide (-56 kJ/mol).
- (d) Suggest any sources of experimental error in this investigation. Evaluate the experimental design, and your skill in carrying it out.

(continued)

Synthesis

- (e) What would have been the effect on the calculated enthalpy of reaction if you had used
- (i) 100 mL of sulfuric acid solution? Explain.
- (ii) 20 mL of sulfuric acid solution? Explain.

Investigation 5.3.1: Combustion of Alcohols

INQUIRY SKILLS

- | | |
|---|--|
| <input type="radio"/> Questioning | <input checked="" type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input checked="" type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input checked="" type="radio"/> Conducting | |

When alcohols burn they produce heat. Do different alcohols produce different quantities of heat—do their molar enthalpies of combustion differ? In this investigation, you will link your study of thermochemistry to your knowledge of the molecular structure of alcohols to investigate energy relationships.

Purpose

The purpose of this investigation is to use calorimetry to determine the molar enthalpy change in the combustion of each of a series of alcohols.

Question

How do the enthalpies of combustion change as the alcohol molecules become larger (i.e., ethanol to butanol)?

Prediction

- (a) Predict what should happen to the enthalpies of combustion per mole of alcohol as the alcohol molecules become larger.

Experimental Design

You will burn measured masses of a series of alcohols, and calculate the amount of each alcohol burned. Assume that the energy produced is transferred to a measured volume of water, the temperature change of which is calculated. The enthalpy change involved in the combustion of one mole of each alcohol can then be calculated and compared.

Materials

- lab apron
- eye protection
- centigram or analytical balance
- alcohol burners containing ethanol, propanol, and butanol
- thermometer
- small tin can, open at one end, and cut under the rim on opposite sides
- stirring rod
- ring stand with iron ring
- large tin can, open at both ends, with vent openings cut around the rim at the base
- 100-mL graduated cylinder

(continued)

**Figure 2**

Apparatus for burning alcohols



Alcohols are highly flammable. Do not attempt to refuel the burners. If refuelling is necessary, ask your teacher to do so.

Do not move the burners after they are lit.

Procedure

1. Obtain an alcohol burner containing one of the alcohols.
2. Put on your eye protection.
3. Place 100 mL of cold water in a small can suspended over an alcohol burner surrounded by a larger can (**Figure 2**).
4. Measure the mass of the alcohol burner and the temperature of the water.
5. Burn the alcohol such that the heat from the reaction is transferred as efficiently as possible into the water.
6. Cease heating when the temperature has risen about 20°C.
7. Re-weigh the alcohol burner and remaining alcohol.
8. Repeat steps 3 to 7 for the remaining alcohols.

Analysis

- (b) Assume that 100 mL of water has a mass of 100 g. For each of the alcohols, calculate and tabulate:
 - (i) the heat absorbed by the water in the calorimeter, which equals the heat produced by alcohol burning;
 - (ii) the heat produced per gram of alcohol;
 - (iii) the amount (in moles) of alcohol burned;
 - (iv) the heat produced per mole of alcohol.
- (c) Write a thermochemical equation to represent the burning of 1 mol of each of the alcohols, including your experimental value for ΔH_{comb} .
- (d) Use the four methods described in Section 3.3 of the Student Text to represent the experimentally determined enthalpy change for the burning of 2 mol of ethanol.
- (e) Use your representations to answer the Question.

(continued)

Evaluation

- (f) What are the sources of experimental error in this experiment?
- (g) Accepted values for combustion of the three alcohols are: 1369 kJ/mol, 2008 kJ/mol, and 3318 kJ/mol. Calculate a percentage difference between your experimental value and the accepted values. Based on this calculation, was the Experimental Design an acceptable method to answer the Question?

Synthesis

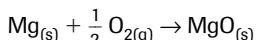
- (h) If you were in the business of buying, transporting, and storing alcohols for use as home-heating fuels, which of these alcohols would you choose to work with? Explain.

Investigation 5.4.1: Hess's Law

INQUIRY SKILLS

- | | |
|-----------------|-----------------|
| ○ Questioning | ● Recording |
| ○ Hypothesizing | ● Analyzing |
| ● Predicting | ● Evaluating |
| ○ Planning | ● Communicating |
| ● Conducting | |

The combustion of magnesium is very rapid and exothermic (**Figure 3**), and is represented by the equation

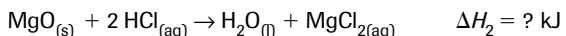


It is possible to observe and measure a series of reactions that enable us, with the use of Hess's law, to determine the enthalpy change for this reaction.

(1) Magnesium reacts in acid to form hydrogen gas and a salt:



(2) Magnesium oxide reacts in acid to form water and a salt:



The values ΔH_1 and ΔH_2 can be determined empirically using a simple calorimeter. The enthalpy of reaction for a third reaction can be found in a reference table.

(3) Hydrogen and oxygen gases react to form water:

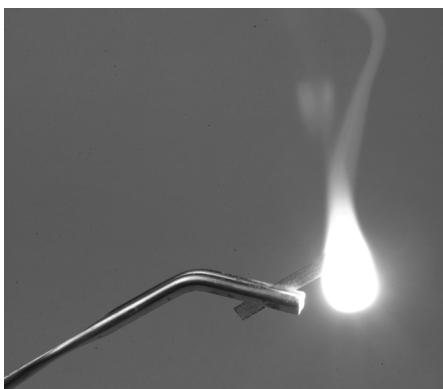
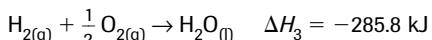


Figure 3

The reaction of magnesium in air is very exothermic.

Purpose

The purpose of this investigation is to use Hess's law to determine the molar enthalpy of combustion of magnesium, using calorimetry.

Question

What is the molar enthalpy of combustion, ΔH_c , of magnesium?

Experimental Design

Measured masses of magnesium and magnesium oxide will be added to measured volumes of hydrochloric acid solution of known concentration. The temperature changes will be determined. Calculated enthalpies of reaction will be combined, using Hess's law, to determine the enthalpy of combustion of magnesium.

Prediction

- Show how the three known equations and their enthalpies of reaction may be combined, using Hess's law, to yield the target equation and its enthalpy of combustion.

(continued)

Materials



Hydrochloric acid is corrosive. Eye protection and a lab apron should be worn. All spills should be cleaned up quickly and any skin that has come into contact with acid should be immediately and thoroughly rinsed with cold water.

lab apron
eye protection
steel wool
centigram or milligram balance
thermometer
polystyrene calorimeter
100-mL graduated cylinder
scoopula
10- to 15-cm strip of magnesium ribbon
magnesium oxide powder
1.00 mol/L hydrochloric acid



Magnesium ribbon is highly flammable and should be kept far from any source of ignition.

Procedure

1. Measure 100.0 mL of 1.00 mol/L hydrochloric acid into a polystyrene cup. Measure the initial temperature of the acid solution to the nearest 0.2°C.
2. Polish a length of magnesium ribbon with steel wool. Determine the mass (to ± 0.01 g) of approximately 0.5 g of magnesium metal. Add the solid to the solution, stir it, and record the maximum temperature that the solution attains.
3. Dispose of the products as directed by your teacher, and rinse and dry the equipment.
4. Repeat the first three steps, using approximately 1 g of magnesium oxide powder measured to ± 0.01 g.

Analysis

- (b) Were the changes exothermic or endothermic? Explain.
- (c) For the first reaction, calculate the enthalpy change per mole of magnesium.
- (d) Write a thermochemical equation for the reaction of magnesium in acid, including your experimental value.

(continued)

- (e) Repeat steps (c) and (d) for magnesium oxide.
- (f) Using Hess's law, the values you have found experimentally, and the given value for the enthalpy change for the formation of water from its elements, determine the molar enthalpy of combustion of magnesium.
- (g) Which of the measured values limited the precision of your value? Explain.

Evaluation

- (h) Explain why (and how) your calculated enthalpies of reaction would be inaccurate if
- (i) some heat were transferred to the air or Styrofoam cup;
 - (ii) the surface of the magnesium ribbon had a coating of MgO.
- (i) Suggest some other possible sources of experimental error in this investigation.
- (j) The accepted value for the molar enthalpy of combustion of magnesium is -601.6 kJ/mol . Calculate a percentage difference by comparing your experimental values and the accepted values. Comment on your confidence in the evidence.

(continued)

- (k) Based on your evaluation of the Experimental Design and the evidence, is Hess's law an acceptable method to calculate enthalpies of reaction?

Synthesis

- (l) Suggest an experimental technique that could be used to determine the enthalpy of combustion of magnesium directly.

Lab Exercise 5.5.1: Testing Enthalpies of Formation

INQUIRY SKILLS

- | | |
|---|--|
| <input type="radio"/> Questioning | <input type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input checked="" type="radio"/> Predicting | <input type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input type="radio"/> Communicating |
| <input type="radio"/> Conducting | |

Calorimetry is the basic experimental tool used to determine enthalpies of reaction. When carefully obtained calorimetric results are used to find an enthalpy of reaction, the calculations should be consistent with results obtained using standard enthalpies of formation.

Question

What is the molar enthalpy of combustion of methanol?

Prediction

- (a) Use the given values for standard enthalpy of formation to calculate the molar enthalpy of combustion of methanol. (Assume that the products of the reaction are gaseous carbon dioxide and liquid water only.)

Experimental Design

A known mass of methanol is burned in a calibrated bomb calorimeter. The enthalpy of combustion is also calculated using standard enthalpies of formation. The two values are compared to test the standard enthalpies.

Materials

methanol
bomb calorimeter
thermometer

Evidence

Table 2 Observations for Burning Methanol

Quantity	Measurement
mass of methanol reacted	4.38 g
heat capacity of bomb calorimeter	10.9 kJ/C°
initial temperature of calorimeter	20.4°C
final temperature of calorimeter	27.9°C

(continued)

Analysis

- (b) Use the experimental values to calculate, to the appropriate number of significant figures, the molar enthalpy of combustion of methanol.

Evaluation

- (c) Calculate the percentage difference between the experimental and predicted values. Does this experiment support the standard enthalpies?

Student Worksheet

LSM 6.3-2

Rate Law Calculations, Extra Exercises

1. What is the rate law for the reaction $2 \text{NO}_{(\text{g})} + 2 \text{H}_{2(\text{g})} \rightarrow \text{N}_{2(\text{g})} + 2 \text{H}_2\text{O}_{(\text{g})}$ which produces the following evidence when the reactant concentrations are varied and the initial rate of production of water vapor is measured?

Trial	Initial $[\text{NO}_{(\text{g})}]$ (mmol/L)	Initial $[\text{H}_{2(\text{g})}]$ (mmol/L)	Initial rate of $\text{H}_2\text{O}_{(\text{g})}$ production (mmol/(L·s))
1	100	100	1.37
2	100	200	2.75
3	200	100	5.47

2. Find the rate law for the hypothetical reaction $\text{A} + 6 \text{B} + 4 \text{C} \rightarrow \text{D} + 2 \text{E} + 3 \text{F}$ which produces the following evidence when reactant concentrations are varied and the initial rate of formation of product D is measured.

Trial	Initial $[\text{A}_{(\text{aq})}]$ (mmol/L)	Initial $[\text{B}_{(\text{aq})}]$ (mmol/L)	Initial $[\text{C}_{(\text{aq})}]$ (mmol/L)	Initial rate of D production (mmol/(L·s))
1	0.0100	0.0100	0.0200	0.002 42
2	0.0200	0.0100	0.0200	0.004 83
3	0.0100	0.0200	0.0200	0.019 36
4	0.0100	0.0100	0.0400	0.009 71

Rate Law Calculations, Extra Exercises, Solution

1. What is the rate law for the reaction $2 \text{NO}_{(\text{g})} + 2 \text{H}_{2(\text{g})} \rightarrow \text{N}_{2(\text{g})} + 2 \text{H}_2\text{O}_{(\text{g})}$ which produces the following evidence when the reactant concentrations are varied and the initial rate of production of water vapor is measured?

Trial	Initial $[\text{NO}_{(\text{g})}]$ (mmol/L)	Initial $[\text{H}_{2(\text{g})}]$ (mmol/L)	Initial rate of $\text{H}_2\text{O}_{(\text{g})}$ production (mmol/(L·s))
1	100	100	1.37
2	100	200	2.75
3	200	100	5.47

Comparing trials 1 and 2, $r \propto [\text{H}_{2(\text{g})}]^m$, where $m = 1$

Comparing trials 1 and 3, $r \propto [\text{NO}_{(\text{g})}]^n$, where $n = 2$

$$r \propto [\text{H}_{2(\text{g})}] [\text{NO}_{(\text{g})}]^2$$

$$\text{Therefore, } r = k [\text{H}_{2(\text{g})}] [\text{NO}_{(\text{g})}]^2$$

$$\text{From trial 1, } 0.001\ 37 \text{ mol/(L·s)} = k \times 0.100 \text{ mol/L} \times (0.100 \text{ mol/L})^2$$

$$k = \frac{0.001\ 37 \text{ mol/(L·s)}}{0.100 \text{ mol/L} \times (0.100 \text{ mol/L})^2} = 1.37 \text{ L}^2/(\text{mol}^2\cdot\text{s})$$

According to the evidence, the rate law is

$$r = 1.37 \text{ L}^2/(\text{mol}^2\cdot\text{s}) [\text{H}_{2(\text{g})}] [\text{NO}_{(\text{g})}]^2$$

Trial	Initial $[\text{A}_{(\text{aq})}]$ (mmol/L)	Initial $[\text{B}_{(\text{aq})}]$ (mmol/L)	Initial $[\text{C}_{(\text{aq})}]$ (mmol/L)	Initial rate of D production (mmol/(L·s))
1	0.0100	0.0100	0.0200	0.002 42
2	0.0200	0.0100	0.0200	0.004 83
3	0.0100	0.0200	0.0200	0.019 36
4	0.0100	0.0100	0.0400	0.009 71

2. Find the rate law for the hypothetical reaction $\text{A} + 6 \text{B} + 4 \text{C} \rightarrow \text{D} + 2 \text{E} + 3 \text{F}$ which produces the following evidence when reactant concentrations are varied and the initial rate of formation of product D is measured.

Comparing trials 1 and 2, $r \propto [\text{A}_{(\text{aq})}]^x$, where $x = 1$

Comparing trials 1 and 3, $r \propto [\text{B}_{(\text{aq})}]^y$, where $y = 3$

Comparing trials 1 and 4, $r \propto [\text{C}_{(\text{aq})}]^z$, where $z = 2$

$$r \propto [\text{A}_{(\text{aq})}] [\text{B}_{(\text{aq})}]^3 [\text{C}_{(\text{aq})}]^2$$

(continued)

LSM 6.3-3

$$\text{Therefore, } r = k [A_{(\text{aq})}] [B_{(\text{aq})}]^3 [C_{(\text{aq})}]^2$$

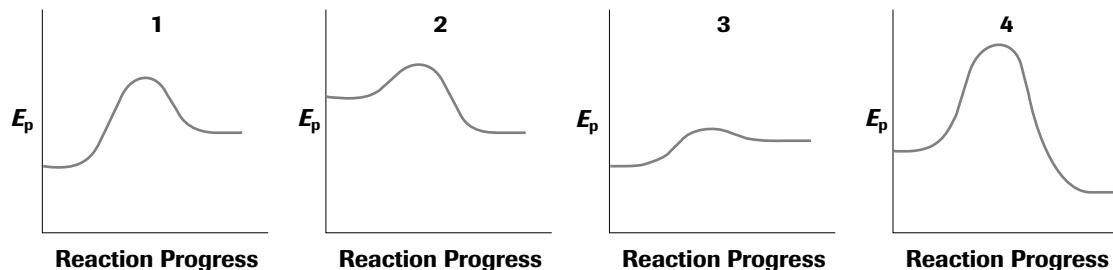
From trial 1,

$$0.002\ 42 \text{ mol}/(\text{L}\cdot\text{s}) = k \times 0.0100 \text{ mol/L} \times (0.0100 \text{ mol/L})^3 \times (0.0200 \text{ mol/L})^2$$

$$k = \frac{0.002\ 42 \text{ mol}/(\text{L}\cdot\text{s})}{0.0100 \text{ mol/L} \times (0.0100 \text{ mol/L})^3 \times (0.0200 \text{ mol/L})^2}$$
$$= 6.05 \times 10^{-8} \text{ L}^5/(\text{mol}^5\cdot\text{s})$$

According to the evidence, the rate law is

$$r = 6.05 \times 10^{-8} \text{ L}^5/(\text{mol}^5\cdot\text{s}) [A_{(\text{aq})}] [B_{(\text{aq})}]^3 [C_{(\text{aq})}]^2$$

Reaction Kinetics, Extra Exercises

1. For each of the diagrams above, explain the *reasoning* in complete sentence statements, in terms of the changes in potential energy and kinetic energy over the duration of the reaction.

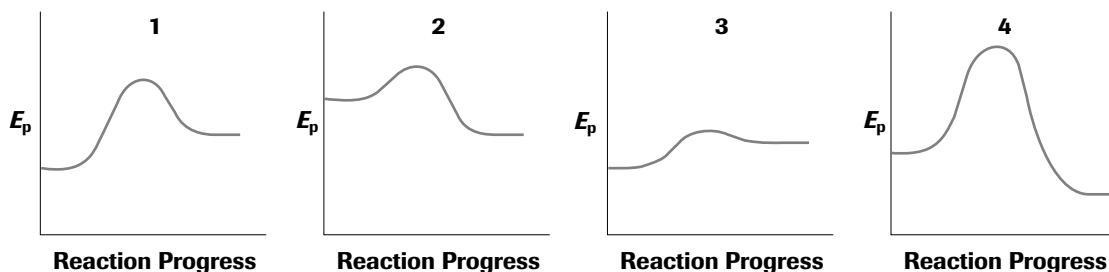
(a)

(b)

(c)

(d)

Reaction Kinetics, Extra Exercises, Solution



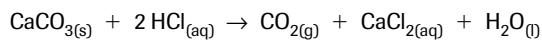
1. For each of the diagrams above, explain the *reasoning* in complete sentence statements, in terms of the changes in potential energy and kinetic energy over the duration of the reaction.
 - (a) In reaction 1, the activation energy is high, so the kinetic energy of the reactant molecules will likely not be sufficient to begin the reaction. Once underway, the increase in potential energy of this reaction results in kinetic (heat) energy being absorbed from the surroundings (in an endothermic reaction). Considerable energy would be required to keep this reaction going.
 - (b) In reaction 2, the activation energy is low, so the kinetic energy of the reactant molecules will likely be sufficient to begin the reaction. Once underway, the decrease in potential energy of this reaction results in kinetic (heat) energy being released to the surroundings (in an exothermic reaction) to help maintain or even increase the rate of the reaction.
 - (c) In reaction 3, the activation energy is low, so the kinetic energy of the reactant molecules will likely be sufficient to begin the reaction. Once underway, the increase in potential energy of this reaction results in kinetic (heat) energy being absorbed from the surroundings (in an endothermic reaction). An external energy source might be needed to keep the initially spontaneous reaction going.
 - (d) In reaction 4, the activation energy is high, so the kinetic energy of the reactant molecules will likely not be sufficient to begin the reaction. Once underway, the decrease in potential energy of this reaction results in kinetic (heat) energy being released to the surroundings (in an exothermic reaction) to maybe make the initially nonspontaneous reaction self-sustaining.

Lab Exercise 6.1.1: Determining a Rate of Reaction

INQUIRY SKILLS

- | | |
|---|--|
| <input type="radio"/> Questioning | <input type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input checked="" type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input type="radio"/> Conducting | |

The reaction between calcium carbonate (found naturally as chalk or limestone) and acid produces carbon dioxide gas.



This reaction is used by geologists to confirm the presence of limestone in a mineral sample (**Figure 1**). As you know, the rate of any reaction is determined by measuring the change in quantity or concentration of some reactant or product over a series of time intervals. In this experiment, the change in concentration of hydrochloric acid (in mol/L) and time (in min) are measured. The rates of reaction are, therefore, expressed as

$$-\frac{\Delta[\text{HCl}]}{\Delta t} = x \text{ mol}/(\text{L}\cdot\text{min})$$



Figure 1

Limestone is a mineral that can be detected by adding concentrated hydrochloric acid to a sample.

Question

What is the rate of reaction of calcium carbonate with acid over various time intervals, and at specific times in the reaction?

Prediction

- (a) Sketch a graph of concentration vs. time to show how you would expect the rate of reaction to change as the reaction proceeds.

(continued)

Experimental Design

The concentration of hydrochloric acid is determined (by calculation from measured solution pH values) at a series of times as the acid reacts with calcium carbonate solid. A graph of concentration vs. time is plotted and analyzed to determine rates of reaction at various times.

Materials

calcium carbonate chips
1.90 mol/L hydrochloric acid
pH meter
beaker

Evidence

Table 1 Concentration of Hydrochloric Acid Remaining

[HCl] (mol/L)	Time (min)
1.90	0.0
1.40	1.0
1.10	2.0
0.90	3.0
0.80	4.0
0.75	5.0
0.72	6.0

Analysis

(b) Plot a graph of [HCl] vs. time.

(c) Determine the average rate of consumption of HCl over the time interval from
(i) 0 to 2 min

(ii) 3 to 5 min

(continued)

(d) Determine the instantaneous rate of consumption of HCl at

(i) 1 min

(ii) 4 min

(e) Communicate the change in reaction rate in words.

Evaluation

(f) Can you detect any flaws in this Experimental Design?

(g) Compare your Prediction to your answer in (e). Account for any differences.

Synthesis

(h) State your experimental rate of consumption of hydrochloric acid at 1 min.

Use the appropriate notation to express the rate of consumption of CaCO_3 and the rate of production of CO_2 at the same time.

(i) How could conductivity and gas measurements be used to measure the rate of reaction?

Investigation 6.2.1: Chemical Kinetics and Factors Affecting Rate

INQUIRY SKILLS

- | | |
|--|--|
| <ul style="list-style-type: none">● Questioning○ Hypothesizing● Predicting● Planning● Conducting | <ul style="list-style-type: none">● Recording● Analyzing● Evaluating● Communicating |
|--|--|

You have learned about five factors that affect rate of reaction. In this investigation you will test as many of these factors as possible with a given chemical system. It is difficult to test all of these factors with just one chemical system, but many systems show clear changes in rates when nature of reactant, temperature, initial concentration, or surface area is varied. Catalysts (such as aqueous copper(II) sulfate and solid manganese dioxide) affect some chemical reactions, but not others. Some possible choices of reactions to test the five factors include the following:

- metals reacting with acids
e.g., $X_{(s)} + H_2SO_{4(aq)} \rightarrow XSO_{4(aq)} + H_{2(g)}$
(possible catalyst: aqueous $CuSO_4$)
- carbonate or bicarbonate salts reacting with acids
e.g., $XCO_{3(s)} + 2 HCl_{(aq)} \rightarrow XCl_{2(aq)} + H_2O_{(g)} + CO_{2(g)}$
- the decomposition of hydrogen peroxide
 $H_2O_{2(aq)} \rightarrow H_2O_{(l)} + O_{2(g)}$

Purpose

The purpose of this investigation is to design an experiment to test the factors that affect rate of reaction.

Question

- (a) Write an appropriate question that you will attempt to answer in this experiment.

Prediction

- (b) After designing your experiment and having it approved, but before carrying it out, predict an answer to your Question.

(continued)

Experimental Design

- (c) Choose a chemical system and design an experiment to investigate as many factors as practical that might affect the rate of reaction of the system. Clearly describe the set of conditions that will be used as the experimental control. Decide on the units you will use to calculate the rates of reaction (for example, amount of reactant/min or mol/L of product).

Materials

lab apron eye protection

- (d) List any necessary chemicals and equipment.

Procedure

- (e) Write a step-by-step procedure. Include any safety precautions that are needed, based on what you learn about the reactants and products from safety information sheets.

1. Have your procedure approved by your teacher before carrying it out.

Evidence

- (f) Design a table for recording observations, making sure that you can obtain and record both qualitative and quantitative observations.

(continued)

Analysis

- (g) Analyze your quantitative evidence both mathematically and graphically, and determine a rate of reaction for each of the experimental trials.
- (h) Answer your Question.

Evaluation

- (i) Evaluate your Experimental Design and Procedure. Suggest how they might be improved, if necessary.
- (j) Evaluate your Evidence and the Prediction that you made before starting the investigation.

Investigation 6.3.1: The Iodine Clock Reaction

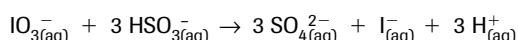
INQUIRY SKILLS

- | | |
|-----------------|-----------------|
| ○ Questioning | ● Recording |
| ○ Hypothesizing | ● Analyzing |
| ● Predicting | ● Evaluating |
| ○ Planning | ● Communicating |
| ● Conducting | |

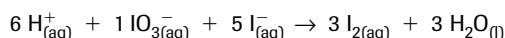
The rate of a chemical reaction depends on many physical and chemical factors, including temperature, the chemical nature of the reacting species, and the initial concentration of the reactants. In the following experiment, all factors except one are held constant.

The iodine clock reaction is the classic experiment used to investigate rates of reaction. It is called a clock reaction because at first, when two colourless solutions are mixed, no reaction appears to occur. Then, at a specific time, the mixture suddenly changes colour. The reaction under study combines aqueous iodate ($\text{IO}_{3(\text{aq})}^-$) and bisulfite ($\text{HSO}_{3(\text{aq})}^-$) ions and involves three reactions in sequence.

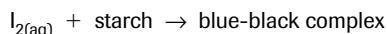
In the first reaction, aqueous iodate ions are reduced to iodide ions.



In the second reaction, the iodide ion is changed to molecular iodine.



In the final step, the iodine reacts with starch suspended in solution to form a blue-black complex.



The third reaction is extremely rapid in comparison to the first two, and serves to indicate the time of reaction for those two reactions. The shorter the time required, the greater is the rate of reaction. Because the times of reaction are related to the rates of reaction, we will be able to make rate-concentration comparisons.

Graphical analysis is an efficient way to determine order of reaction for iodate, a , where rate $\propto [\text{IO}_{3(\text{aq})}^-]^a$. Figure 3 in section 6.3 of the Student Text shows graphs that are characteristic of zeroth order ($n = 0$), first order ($n = 1$), and second order ($n = 2$) reactions.

Purpose

The purpose of this investigation is to gather and analyze experimental observations to determine the rate dependence of a reactant in a chemical system.

Question

What is the order of reaction with respect to the initial concentration of iodate ions in the iodine clock reaction?

Prediction

- Predict qualitatively what will happen to the time of reaction as the initial concentration of iodate ions is increased. What will happen to the rate of reaction?

(continued)

Experimental Design

A series of solutions will be prepared in which the only variable is the initial concentration of iodate ions. Equal amounts of starch, sodium bisulfite, and hydrochloric acid will be mixed with each of these solutions, so that the time from mixing to formation of a blue-black product can be timed. The evidence will be analyzed graphically.

- (b) Calculate the initial concentration of the iodate solution in each of the wells in microtray A, based on dilution from the concentration of the stock iodate solution.

Materials

lab apron
 eye protection
 0.020 mol/L potassium iodate solution (Solution A)
 0.00100 mol/L sodium bisulfite/hydrochloric acid/starch solution (Solution B)
 distilled water
 3 plastic micropipets, labelled A, B, and H₂O
 three 100-mL beakers, labelled A, B, and H₂O
 2 large-well microtrays or spot plates, labelled A and B
 stopwatch



Figure 2

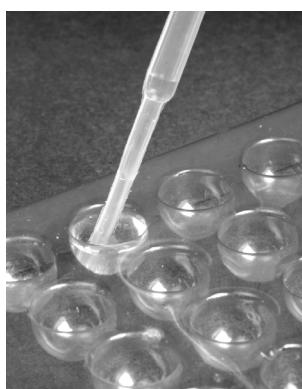


Figure 3

Procedure

1. Place the two microtrays on clean sheets of white paper. For microtray A, using the appropriate micropipets, place 1 drop of Solution A in Well 1, 2 drops of Solution A in Well 2, 3 drops in Well 3, and so on up to 10 drops of Solution A in Well 10.
2. Also in microtray A, place 9 drops of water in Well 1, 8 drops of water in Well 2, 7 drops in Well 3, and so on down to 1 drop of water in Well 9. There are now 10 drops of solution in each well (**Figure 2**).
3. For microtray B, using the appropriate micropipet, put 10 drops of Solution B in each of the first 10 wells.
4. With stopwatch ready, and using the water micropipet, transfer the contents of Plate A Well 10 to Plate B Well 10. In doing so, insert the tip of the micropipet below the surface of the liquid in Plate B to ensure that the solutions mix thoroughly and keep stirring (**Figure 3**). Start timing at the moment the micropipet is squeezed and stop when the colour first appears. Record your observations.
5. Rinse the water micropipet you just used at least twice with water, making sure that no water remains in the micropipet each time.
6. Repeat Steps 4 and 5 for each of the other pairs of wells, recording your observations.
7. Dispose of solutions down the drain with lots of running water.

(continued)

Analysis

- (c) Calculate the initial concentration of iodate solution in each of the wells at the instant of mixing with an equal volume of Solution B.
- (d) Make a general statement to summarize your qualitative observations.
- (e) Make appropriate evidence tables and plot graphs to identify the order of the reaction with respect to concentration of iodate ions.

(continued)

- (f) Write an expression to answer the Question.

Evaluation

- (g) What other variables, apart from initial bisulfite ion concentration, were controlled in this investigation?
- (h) Evaluate the Experimental Design, your lab skills, and the Evidence. Suggest any ways in which this experiment could be improved.
- (i) If you are confident in your Evidence, evaluate the Prediction that you made before starting the investigation.

Synthesis

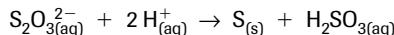
- (j) Why was it important to add specific volumes of water to the wells in micro-tray A?

Lab Exercise 6.4.1: The Sulfur Clock

INQUIRY SKILLS

- | | |
|-------------------------------------|---|
| <input type="radio"/> Questioning | <input type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input type="radio"/> Communicating |
| <input type="radio"/> Conducting | |

Like the iodine clock reaction in Investigation 6.3.1, the sulfur clock reaction occurs in stages, the final step of which causes a visible change. The overall reaction is between thiosulfate ions and acid to form elemental sulfur.



As particles of solid sulfur form, the solution becomes first cloudy and finally opaque. If a visible mark such as an “X” is written on a piece of filter paper and placed under the reaction beaker containing clear starting materials, the rates of reaction under different conditions can be compared by measuring the amount of time required for the “X” to become invisible in the increasingly cloudy solution.

Question

- (i) What is the order of reaction in the sulfur clock system with respect to thiosulfate ions?
- (ii) How many thiosulfate ions are involved in the rate-determining step?

Experimental Design

The rates of reaction under different conditions are tested by adding a fixed concentration of acid to a series of solutions of different concentrations of thiosulfate ion. The time is measured from the instant of mixing to the point at which an “X” marked under the reaction beaker becomes invisible.

Materials

laboratory apron	0.160 mol/L sodium thiosulfate
eye protection	pentahydrate solution
six 100-mL beakers	2.0 mol/L hydrochloric
2 volumetric pipettes	acid solution

Evidence

Table 2 Initial Concentration & Reaction Time Data for Sulfur Clock

Trial	Initial $[\text{S}_2\text{O}_3^{2-}]$ (mmol/L)	Initial $[\text{H}^+]$ (mol/L)	Time (s)
1	0.10	0.050	83
2	0.20	0.050	44
3	0.30	0.050	32
4	0.40	0.050	23
5	0.50	0.050	18

(continued)

Analysis

- (a) Make a general statement to summarize the qualitative observations in the investigation.

- (b) Make appropriate data tables and plot graphs to answer Question (i).

(continued)

- (c) Write a mathematical expression that shows how the rate of consumption of $\text{S}_2\text{O}_3^{2-}$ ions depends on $\text{S}_2\text{O}_3^{2-}$.
- (d) Answer Question (ii).

Evaluation

- (e) Why is it important to use the same “X” marked on the same piece of paper for all of the trials?
- (f) What measurement could be expected to cause the greatest experimental uncertainty?
- (g) Evaluate the Experimental Design. Suggest three ways in which it could be improved.

Synthesis

- (h) Create a possible mechanism for this reaction, given your experimental results.
- (i) Design an experiment to determine the order of this reaction with respect to hydrogen ions.

The Law of Chemical Equilibrium: Stoichiometry and Equilibrium, SIR Q

In the experiments that follow, you will use the SIR Q program to investigate chemical systems at equilibrium. Start the SIR Q program as directed by your teacher.

Experiment 1: Catalysis and Equilibrium

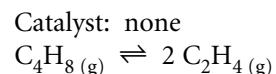
Purpose

To see the effect of a catalyst on a system approaching equilibrium.

Procedure

1. Settings:

SHOW Q
Temperature: 525°C
Select the equilibrium



2. Add 0.5 atm of $\text{C}_4\text{H}_8\text{(g)}$ into the reactor.
3. Run the reaction. Record your observations for each trial in Table 1.

Table 1 Catalysis and Equilibrium $\text{C}_4\text{H}_8\text{(g)} \rightleftharpoons 2 \text{C}_2\text{H}_4\text{(g)}$

Trial #	Amount of catalyst	$[\text{C}_4\text{H}_8\text{(g)}]$ at end of reaction (in atm)	$[\text{C}_2\text{H}_4\text{(g)}]$ at end of reaction (in atm)	Value of Q at end of reaction	Was equilibrium reached? If yes, at what time?
1.	none				
2.	1				
3.	2				
4.	3				

Questions

1. What effect, if any, did the amount of catalyst have on the time required to reach equilibrium?
2. Explain your answer to question 1 using rate theory.
3. Many students assume that equilibrium has been reached when the two graphs cross. Explain why this is incorrect.

(continued)

Experiment 2: Stoichiometry and Equilibrium**Purpose**

To verify that equilibrium systems obey reaction stoichiometry.

Procedure: Reaction 1

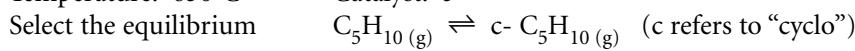
1. Settings:

SHOW Q

Temperature: 650°C

Catalyst: 3

Select the equilibrium



2. Run the reaction. Record your observations for each trial in Table 2.

Table 2 Stoichiometry and Equilibrium: Reaction 1 $C_5H_{10(g)} \rightleftharpoons c - C_5H_{10(g)}$

Trial #	[$C_5H_{10(g)}$] initially	[$c - C_5H_{10(g)}$] initially	[$C_5H_{10(g)}$] at eqm	[$c - C_5H_{10(g)}$] at eqm	[$C_5H_{10(g)}$] change during the reaction	[$c - C_5H_{10(g)}$] change during the reaction
1.	0.500	0.000				
2.	1.000	0.000				
3.	1.500	0.000				
4.	1.000	0.500				
5.	0.500	1.000				

Question

Use the reaction stoichiometry to explain the concentration changes observed.

Procedure: Reaction 2

1. Change the reaction to $C_4H_{8(g)} \rightleftharpoons 2 C_2H_{4(g)}$

2. Settings:

SHOW Q

Temperature: 525°C

Catalyst: 3

(continued)

3. Run the reaction. Record your observations for each trial in Table 3.

Table 3 Stoichiometry and Equilibrium: Reaction $2 C_4H_8(g) \rightleftharpoons 2 C_2H_4(g)$

Trial #	[$C_4H_8(g)$] initially	[$C_2H_4(g)$] initially	[$C_4H_8(g)$] at eqm	[$C_2H_4(g)$] at eqm	[$C_4H_8(g)$] change during the reaction	[$C_2H_4(g)$] change during the reaction
1.	0.600	0.000				
2.	1.000	0.400				
3.	1.200	0.400				
4.	1.000	0.500				
5.	0.500	1.000				

Question

Use the reaction stoichiometry to explain the concentration changes observed.

Procedure: Reaction 3

1. Change the reaction to $I_{2(g)} + Br_{2(g)} \rightleftharpoons 2 IBr_{(g)}$
2. Settings:

SHOW Q
Temperature: 450°C Catalyst: 3

3. Run trial 1 and record your observations in **Table 4**. Use your data from trial 1 to predict the equilibrium concentrations for trials 2–5.
4. Verify your predictions by running the trials. Record your observations for trials 2–5 in Table 4.

Table 4 Stoichiometry and Equilibrium: Reaction $3 I_{2(g)} + Br_{2(g)} \rightleftharpoons 2 IBr_{(g)}$

Trial #	[$I_{2(g)}$] initial	[$Br_{2(g)}$] initial	[$IBr_{(g)}$] initial	[$I_{2(g)}$] eqm	[$Br_{2(g)}$] eqm	[$IBr_{(g)}$] eqm
1	1.000	1.000	0.000			
2	1.000	1.500	0.000			
3	1.000	0.300	1.400			
4	0.000	0.000	1.400			
5	0.000	0.500	1.400			
6	0.200	0.400	1.400			

Question

Use the reaction stoichiometry to explain the concentration changes observed.

The Law of Chemical Equilibrium: Stoichiometry and Equilibrium, SIR Q, Solution

In the experiments that follow, you will use the SIR Q program to investigate chemical systems at equilibrium. Start the SIR Q program as directed by your teacher.

Experiment 1: Catalysis and Equilibrium

Purpose

To see the effect of a catalyst on a system approaching equilibrium.

Procedure

1. Settings:

SHOW Q

Temperature: 525°C Catalyst: none

Select the equilibrium $\text{C}_4\text{H}_8 \text{ (g)} \rightleftharpoons 2 \text{ C}_2\text{H}_4 \text{ (g)}$

2. Add 0.5 atm of $\text{C}_4\text{H}_8 \text{ (g)}$ into the reactor.

3. Run the reaction. Record your observations for each trial in Table 1.

Table 1 Catalysis and Equilibrium $\text{C}_4\text{H}_8 \text{ (g)} \rightleftharpoons 2 \text{ C}_2\text{H}_4 \text{ (g)}$

Trial #	Amount of catalyst	$[\text{C}_4\text{H}_8 \text{ (g)}]$ at end of reaction (in atm)	$[\text{C}_2\text{H}_4 \text{ (g)}]$ at end of reaction (in atm)	Value of Q at end of reaction	Was equilibrium reached? If yes, at what time?
1.	none	0.296	0.408	0.564	no
2.	1	0.290	0.421	0.612	no
3.	2	0.284	0.431	0.654	no
4.	3	0.284	0.433	0.660	yes, at approximately 10 min

Questions

1. What effect, if any, did the amount of catalyst have on the time required to reach equilibrium?

Equilibrium was reached sooner as more catalyst was added.

2. Explain your answer to question 1 using rate theory.

The catalyst lowers the activation energy for both the forward and reverse reactions, allowing equilibrium to be reached sooner.

(continued)

3. Many students assume that equilibrium has been reached when the two graphs cross. Explain why this is incorrect.

The two graphs cross when the concentrations (pressure) of both entities are the same. However, equilibrium is reached when the rates of the two reactions are equal, not the concentrations of reactants and products.

Experiment 2: Stoichiometry and Equilibrium

Purpose

To verify that equilibrium systems obey reaction stoichiometry.

Procedure: Reaction 1

1. Settings:

SHOW Q

Temperature: 650°C Catalyst: 3

Select the equilibrium $C_5H_{10(g)} \rightleftharpoons c - C_5H_{10(g)}$ (c refers to "cyclo")

2. Run the reaction. Record your observations for each trial in Table 2.

Table 2 Stoichiometry and Equilibrium: Reaction 1 $C_5H_{10(g)} \rightleftharpoons c - C_5H_{10(g)}$

Trial #	[$C_5H_{10(g)}$] initially	[$c - C_5H_{10(g)}$] initially	[$C_5H_{10(g)}$] at eqm	[$c - C_5H_{10(g)}$] at eqm	[$C_5H_{10(g)}$] change during the reaction	[$c - C_5H_{10(g)}$] change during the reaction
1.	0.500	0.000	0.238	0.262	0.262 less	0.262 more
2.	1.000	0.000	0.476	0.524	0.524 less	0.524 more
3.	1.500	0.000	0.715	0.785	0.785 less	0.785 more
4.	1.000	0.500	0.714	0.786	0.785 less	0.785 more
5.	0.500	1.000	0.714	0.786	0.785 less	0.785 more

Question

Use the reaction stoichiometry to explain the concentration changes observed.

The amount of $C_5H_{10(g)}$ and $c - C_5H_{10(g)}$ decreased and increased by the same amount. These changes are predictable since both compounds are in a 1:1 ratio in the chemical equation.

Procedure: Reaction 2

1. Change the reaction to $C_4H_{8(g)} \rightleftharpoons 2 C_2H_{4(g)}$
2. Settings:

SHOW Q

Temperature: 525°C Catalyst: 3

(continued)

3. Run the reaction. Record your observations for each trial in Table 3.

Table 3 Stoichiometry and Equilibrium: Reaction $2 \text{C}_4\text{H}_8\text{(g)} \rightleftharpoons 2 \text{C}_2\text{H}_4\text{(g)}$

Trial #	[$\text{C}_4\text{H}_8\text{(g)}$] initially	[$\text{C}_2\text{H}_4\text{(g)}$] initially	[$\text{C}_4\text{H}_8\text{(g)}$] at eqm	[$\text{C}_2\text{H}_4\text{(g)}$] at eqm	[$\text{C}_4\text{H}_8\text{(g)}$] change during the reaction	[$\text{C}_2\text{H}_4\text{(g)}$] change during the reaction
1.	0.600	0.000	0.357	0.485	0.243 less	0.485 more
2.	1.000	0.400	0.830	0.740	0.170 less	0.340 more
3.	1.200	0.400	0.995	0.810	0.205 less	0.410 more
4.	1.000	0.500	0.871	0.758	0.129 less	0.258 more
5.	0.500	1.000	0.668	0.664	0.332 less	0.664 more

Question

Use the reaction stoichiometry to explain the concentration changes observed.

The amounts of $\text{C}_4\text{H}_8\text{(g)}$ and $\text{C}_2\text{H}_4\text{(g)}$ decreased and increased in a 1:2 ratio as predicted by the ratio of the compounds in the balanced chemical equation.

Procedure: Reaction 3

1. Change the reaction to $\text{I}_{2\text{(g)}} + \text{Br}_{2\text{(g)}} \rightleftharpoons 2 \text{IBr}_{\text{(g)}}$
2. Settings:

SHOW Q

Temperature: 450°C Catalyst: 3

3. Run trial 1 and record your observations in Table 4. Use your data from trial 1 to predict the equilibrium concentrations for trials 2–5.
4. Verify your predictions by running the trials. Record your observations for trials 2–5 in Table 4.

Table 4 Stoichiometry and Equilibrium: Reaction $3 \text{I}_{2\text{(g)}} + \text{Br}_{2\text{(g)}} \rightleftharpoons 2 \text{IBr}_{\text{(g)}}$

Trial #	[$\text{I}_{2\text{(g)}}$]initial	[$\text{Br}_{2\text{(g)}}$]initial	[$\text{IBr}_{\text{(g)}}$]initial	[$\text{I}_{2\text{(g)}}$]eqm	[$\text{Br}_{2\text{(g)}}$]eqm	[$\text{IBr}_{\text{(g)}}$]eqm
1	1.000	1.000	0.000	0.278	0.278	1.444
2	1.000	1.500	0.000	0.1591	0.659	1.682
3	1.000	0.300	1.400	0.833	0.1335	1.735
4	0.000	0.000	1.400	0.1945	0.1945	1.011
5	0.000	0.500	1.400	0.923	0.0592	1.215
6	0.200	0.400	1.400	0.1907	0.391	1.419

Question

Use the reaction stoichiometry to explain the concentration changes observed.

In each trial, the concentrations of iodine and bromine both increased or decreased in a 1:1 ratio to each other as predicted by the balanced chemical equation. Furthermore, the concentration of iodine bromide increased or decreased in a 2:1 ratio as compared to iodine and bromine. This observation also agrees with the balanced chemical equation.

Investigation 7.1.1: Discovering the Extent of a Chemical Reaction

INQUIRY SKILLS

- | | |
|-----------------|-----------------|
| ○ Questioning | ● Recording |
| ○ Hypothesizing | ● Analyzing |
| ● Predicting | ● Evaluating |
| ● Planning | ● Communicating |
| ● Conducting | |

In a quantitative reaction—a reaction that goes to completion—the limiting reagent is completely consumed. We can test the final reaction mixture for the presence of the original reactants to determine whether a reaction has completely consumed at least one of the initial substances.

Purpose

The purpose of this investigation is to test the validity of the assumption that chemical reactions go to completion.

Question

What are the limiting and excess reagents in the chemical reaction of various quantities of aqueous sodium sulfate and aqueous calcium chloride?

Prediction

- Using your current state of knowledge, predict whether the reaction between aqueous sodium sulfate and aqueous calcium chloride is quantitative.



Figure 1
Filtering the final mixture

Experimental Design

Samples of sodium sulfate solution and calcium chloride solution are mixed in different proportions and the final mixture is filtered (**Figure 1**).

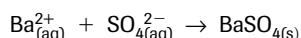
- Write the balanced equation for the overall reaction.

- Write the net ionic equation for the reaction that produces a precipitate.

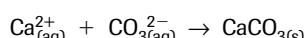
If the reaction proceeds to completion, at least one of the reagents will be completely consumed, and will not appear in the filtrate.

Samples of the filtrate are tested for the presence of excess reagents, using the following diagnostic tests.

- If a few drops of $\text{Ba}(\text{NO}_3)_{2(\text{aq})}$ are added to the filtrate and a precipitate forms, then excess sulfate ions are present.



- If a few drops of $\text{Na}_2\text{CO}_{3(\text{aq})}$ are added to the filtrate and a precipitate forms, then excess calcium ions are present.



(continued)



**Barium compounds are toxic.
Solutions containing barium
should be collected in a
marked disposal container at
the end of the lab.**

**Remember to wash your
hands before leaving the
laboratory.**

**Wear eye protection and a
laboratory apron.**

Materials

lab apron	2 50-mL or 100-mL beakers
eye protection	2 small test tubes
25 mL of 0.50 mol/L $\text{CaCl}_{2(\text{aq})}$	10-mL or 25-mL
25 mL of 0.50 mol/L $\text{Na}_2\text{SO}_{4(\text{aq})}$	graduated cylinder
1.0 mol/L $\text{Na}_2\text{CO}_{3(\text{aq})}$ in dropper bottle	filtration apparatus
saturated $\text{Ba}(\text{NO}_3)_{2(\text{aq})}$ in dropper bottle	filter paper
	wash bottle
	stirring rod

Procedure

- (d) Write a procedure for testing whether the reaction between aqueous sodium sulfate and aqueous calcium chloride is quantitative by reacting different proportions of the two reactants in a closed system.

1. Obtain your teacher's approval, then carry out the experiments.

Analysis

- (e) Write a statement describing what you observed, using the chemical names from the equation.
- (f) Referring to the net ionic equation, write a statement about the anomaly that you observed.

(continued)

Evaluation

(g) Evaluate your Prediction in (a), and the Experimental Design.

(h) Suggest improvements to the Experimental Design.

(i) Evaluate the wording of the Question.

Lab Exercise 7.2.1: Develop an Equilibrium Law

INQUIRY SKILLS

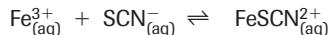
- Questioning
- Recording
- Hypothesizing
- Analyzing
- Predicting
- Evaluating
- Planning
- Communicating
- Conducting



Figure 2

The two reactants combine to form a dark-red equilibrium mixture. The red colour of the solution is the colour of the aqueous thiocyanate iron(II) product, $\text{FeSCN}^{2+}_{(\text{aq})}$.

The following chemical equation represents a chemical equilibrium.



This equilibrium is convenient to study because the reaction is clearly indicated by a colour change (**Figure 2**).

Question

What mathematical formula, using equilibrium concentrations of reactants and products, gives a constant for the iron(III) thiocyanate reaction system?

Experimental Design

Reactions are performed using various initial concentrations of iron(III) nitrate and potassium thiocyanate solutions. The equilibrium concentrations of the reactants and the product are determined from the measurement and analysis of the intensity of the colour. Possible mathematical relationships among the concentrations are tried, and then are analyzed to determine if the mathematical formula gives a constant value.

Observations

Table 1 Iron(III)-Thiocyanate Equilibrium at SATP

Trial	$[\text{Fe}^{3+}_{(\text{aq})}]$ (mol/L)	$[\text{SCN}^{-}_{(\text{aq})}]$ (mol/L)	$[\text{FeSCN}^{2+}_{(\text{aq})}]$ (mol/L)
1	3.91×10^{-2}	8.02×10^{-5}	9.22×10^{-4}
2	1.48×10^{-2}	1.91×10^{-4}	8.28×10^{-4}
3	6.27×10^{-3}	3.65×10^{-4}	6.58×10^{-4}
4	2.14×10^{-3}	5.41×10^{-4}	3.55×10^{-4}
5	1.78×10^{-3}	6.13×10^{-4}	3.23×10^{-4}

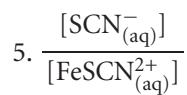
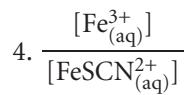
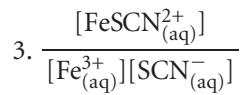
Analysis

- (a) Test the following mathematical relationships to see which gives a constant value:

1. $[\text{Fe}^{3+}_{(\text{aq})}][\text{SCN}^{-}_{(\text{aq})}][\text{FeSCN}^{2+}_{(\text{aq})}]$

2. $[\text{Fe}^{3+}_{(\text{aq})}] + [\text{SCN}^{-}_{(\text{aq})}] + [\text{FeSCN}^{2+}_{(\text{aq})}]$

(continued)



(b) Describe this relationship in words.

Student Worksheet

Investigation 7.3.1: Testing Le Châtelier's Principle

LSM 7.LA-3

INQUIRY SKILLS

- | | |
|--|--|
| <ul style="list-style-type: none">● Questioning● Hypothesizing● Predicting● Planning● Conducting | <ul style="list-style-type: none">● Recording● Analyzing● Evaluating● Communicating |
|--|--|

Le Châtelier's principle helps us to predict how reactions will respond under certain conditions, such as changes in temperature or pressure. In this investigation, you will make and test your own predictions on a variety of chemical systems.

Purpose

The purpose of this investigation is to test Le Châtelier's principle by applying stresses to seven different chemical equilibria.

Question

How does applying stresses to particular chemical equilibria affect the systems?

Prediction

- (a) Read the Experimental Design, Materials, and Procedure, and use Le Châtelier's principle to predict the change(s) that will occur when each equilibrium mixture is subjected to the stated stress.

(continued)

Experimental Design

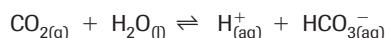
Stresses are applied to the following seven chemical equilibrium systems and evidence is gathered to test predictions made using Le Châtelier's principle. Control samples are used in all cases.

Part I: Dinitrogen Tetroxide–Nitrogen Dioxide Equilibrium (Demonstration)



Flasks containing an $\text{N}_2\text{O}_{4(\text{g})}$ – $\text{NO}_{2(\text{g})}$ equilibrium mixture are placed into cold- and hot-water baths.

Part II: Carbon Dioxide–Carbonic Acid Equilibrium



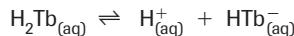
A $\text{CO}_{2(\text{g})}$ – $\text{HCO}_{3(\text{aq})}^-$ equilibrium mixture is placed in a syringe and subjected to increased pressure.

Part III: Cobalt(II) Complexes



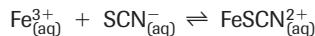
Water, a solution of silver nitrate, and heat are added to, and heat is removed from, samples of the equilibrium mixture, which are provided for you.

Part IV: Thymol Blue Indicator



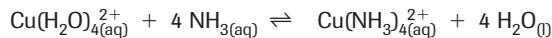
Hydrochloric acid and sodium hydroxide are added to different samples of the equilibrium mixture, which are provided for you.

Part V: Iron(III)–Thiocyanate Equilibrium



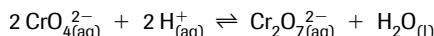
Iron(III) nitrate, potassium thiocyanate, and sodium hydroxide are added to samples of the equilibrium mixture, which are provided for you.

Part VI: Copper(II) Complexes



Aqueous ammonia and hydrochloric acid are added to samples of the equilibrium mixture, which are provided for you.

(continued)

Part VII: Chromate–Dichromate Equilibrium

Aqueous sodium hydroxide, hydrochloric acid, and then aqueous barium nitrate are all added to a sample of the equilibrium mixture, which is provided for you.

Materials

lab apron
 eye protection
 100-mL beaker
 large waste beaker
 6 to 12 small test tubes
 test-tube rack
 small syringe with needle removed (5 to 50 mL)
 solid rubber stopper to seal end of syringe
 distilled water
 crushed ice
 hot-water bath
 2 flasks or tubes containing an $\text{N}_2\text{O}_{4(\text{g})}$ – $\text{NO}_{2(\text{g})}$ mixture
 6.0 mol/L $\text{NaOH}_{(\text{aq})}$
 dropper bottles containing:
 carbon dioxide–bicarbonate equilibrium mixture ($\text{pH} = 7$)
 cobalt(II) chloride equilibrium mixture in ethanol
 0.2 mol/L $\text{AgNO}_3_{(\text{aq})}$
 thymol blue indicator
 bromothymol blue indicator
 0.1 mol/L $\text{HCl}_{(\text{aq})}$
 0.1 mol/L $\text{NaOH}_{(\text{aq})}$
 iron(III)–thiocyanate equilibrium mixture
 0.2 mol/L $\text{Fe}(\text{NO}_3)_3_{(\text{aq})}$
 0.2 mol/L $\text{KSCN}_{(\text{aq})}$
 0.1 mol/L $\text{CuSO}_4_{(\text{aq})}$
 1.0 mol/L $\text{NH}_3_{(\text{aq})}$
 1.0 mol/L $\text{HCl}_{(\text{aq})}$
 chromate–dichromate equilibrium mixture
 0.1 mol/L $\text{Ba}(\text{NO}_3)_2_{(\text{aq})}$



Be careful with the flasks containing nitrogen dioxide: This gas is highly toxic. Use in a fume hood in case of breakage.



The chemicals may be corrosive, irritating, and/or toxic. Exercise great care when using the chemicals and avoid skin and eye contact. Immediately rinse the skin with cold water if there is any contact and flush the eyes for a minimum of 15 min and inform the teacher.

Ethanol is flammable. Make sure there are no open flames in the laboratory when using the ethanol solution of cobalt(II) chloride.

Solutions containing heavy metal ions must not be flushed down the sink. Collect them in a marked container.

Wear a lab apron and eye protection.

Procedure

As you finish each part, dispose of the chemicals as directed by your teacher. Most may be washed down the drain with large amounts of water. Those containing heavy metals must be collected in a marked container for separate disposal.

(continued)



Be careful with the flasks containing nitrogen dioxide: This gas is highly toxic. Use in a fume hood.



Remember that the solute for this solution is alcohol, so keep it away from open flames.

Part I: Dinitrogen Tetroxide–Nitrogen Dioxide Equilibrium (Demonstration)

1. Place the sealed $\text{N}_2\text{O}_{4(\text{g})}$ – $\text{NO}_{2(\text{g})}$ flasks in hot- and cold-water baths and record your observations.

Part II: Carbon Dioxide–Carbonic Acid Equilibrium

2. Place two or three drops of bromothymol blue indicator in the carbon dioxide–bicarbonate equilibrium mixture.
3. Draw some of the carbon dioxide–bicarbonate equilibrium mixture into the syringe, then block the end with a rubber stopper.
4. Slowly depress the syringe plunger and record your observations.

Part III: Cobalt(II) Complexes

5. Obtain 25 mL of the equilibrium mixture with the cobalt(II) chloride complex ions.
6. Place a small amount of the mixture into each of five small test tubes. Use the fifth test tube as a control for comparison purposes.
7. Add drops of water to one test tube until a change is evident. Record your observations.
8. Add drops of 0.2 mol/L silver nitrate to another test tube and record your observations.
9. Heat a third sample of the equilibrium mixture in a water bath and record your observations.
10. Cool a fourth sample of the equilibrium mixture in an ice bath and record your observations.

Part IV: Thymol Blue Indicator

11. Add about 5 mL of distilled water to each of two small test tubes.
12. Add 1 to 3 drops of thymol blue indicator to the water in each test tube to obtain a noticeable colour.
13. Use one test tube of solution as a control.
14. Add drops of 0.1 mol/L $\text{HCl}_{(\text{aq})}$ to the experimental test tube to test for the predicted colour changes.
15. Add drops of 0.1 mol/L $\text{NaOH}_{(\text{aq})}$ to the same test tube to test for the predicted colour changes.

Part V: Iron(III)–Thiocyanate Equilibrium

16. Obtain about 20 mL of the iron(III)–thiocyanate equilibrium solution.
17. Place about 5 mL of the equilibrium solution in each of three test tubes.
18. Use one test tube as a control.
19. Add drops of $\text{Fe}(\text{NO}_3)_{3(\text{aq})}$ to an equilibrium mixture until a change is evident.
20. Add drops of 6.0 mol/L $\text{NaOH}_{(\text{aq})}$ to this new equilibrium mixture until a change occurs. (Remember that iron(III) hydroxide has a low solubility.)

(continued)

21. Add drops of $\text{KSCN}_{(\text{aq})}$ to another equilibrium mixture until a change is evident.

Part VI: Copper(II) Complexes

22. Obtain 2 mL of 0.1 mol/L $\text{CuSO}_4_{(\text{aq})}$ in a small test tube.
 23. Add three drops of 1.0 mol/L $\text{NH}_3_{(\text{aq})}$ to establish the equilibrium mixture.
 24. Add more 1.0 mol/L $\text{NH}_3_{(\text{aq})}$ to the above equilibrium mixture and record the results.
 25. Add 1.0 mol/L $\text{HCl}_{(\text{aq})}$ to the equilibrium mixture from step 24 and record the results.

Table 2 Diagnostic Test Colours

Ion or compound	Colour
$\text{CoCl}_{4(\text{aq})}^{2-}$	blue
$\text{Co}(\text{H}_2\text{O})_6^{2+}$	pink
$\text{H}_2\text{Tb}_{(\text{aq})}$	red
$\text{HTb}_{(\text{aq})}^-$	yellow
$\text{Tb}_{(\text{aq})}^{2-}$	blue
$\text{Fe}_{(\text{aq})}^{3+}$	yellow
$\text{SCN}_{(\text{aq})}^{2-}$	colourless
$\text{FeSCN}_{(\text{aq})}^{2+}$	red
$\text{Cu}(\text{H}_2\text{O})_4^{2+}$	pale blue
$\text{Cu}(\text{NH}_3)_4^{2+}$	deep blue
$\text{CrO}_{4(\text{aq})}^{2-}$	yellow
$\text{Cr}_2\text{O}_{7(\text{aq})}^{2-}$	orange
$\text{N}_2\text{O}_{4(\text{g})}$	colourless
$\text{NO}_{2(\text{g})}$	brown

Part VII: Chromate–Dichromate Equilibrium

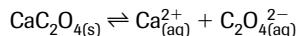
26. Obtain 15 mL of the chromate–dichromate equilibrium mixture.
 27. Place 5-mL samples of the equilibrium mixture into each of three small test tubes.
 28. Add 0.1 mol/L $\text{HCl}_{(\text{aq})}$ drop by drop to one sample or to 0.1 mol/L $\text{K}_2\text{Cr}_2\text{O}_7_{(\text{aq})}$ and record your observations.
 29. Add 0.1 mol/L $\text{NaOH}_{(\text{aq})}$ drop by drop to another sample (or, if you choose, the previous $\text{HCl}_{(\text{aq})}$ sample or 0.1 mol/L $\text{K}_2\text{Cr}_2\text{O}_7_{(\text{aq})}$) and record the results.
 30. Add 0.1 mol/L $\text{Ba}(\text{NO}_3)_2_{(\text{aq})}$ drop-by-drop to a third sample and record your observations. (Remember that barium chromate has a low solubility.) Compare to **Table 2**.
 31. Ensure that all equipment and surfaces are left clean and that your hands are washed thoroughly before leaving the laboratory.

Analysis

- (b) Answer the Question.
 (c) Summarize your observations in the form of a chart.

Investigation 7.6.1: Determining the K_{sp} of Calcium Oxalate

Calcium oxalate, $\text{CaC}_2\text{O}_{4(\text{s})}$, is a slightly soluble salt that dissolves according to the chemical reaction



Question

What is the K_{sp} of calcium oxalate?

Experimental Design

In this investigation, the solubility product constant of calcium oxalate, $\text{CaC}_2\text{O}_{4(\text{s})}$, is determined by mixing a fixed volume of 0.1 mol/L sodium oxalate with a serial dilution of aqueous calcium nitrate in a series of spotplate wells. The K_{sp} of calcium oxalate is determined by calculating the product of the concentrations of the ions in the well containing the highest concentration of ions with no visible precipitate.

Materials

lab apron	minimum 24-well spotplate (12 × 2)
eye protection	4 pipets
0.1 mol/L calcium nitrate, $\text{Ca}(\text{NO}_3)_{2(\text{aq})}$	distilled water
0.1 mol/L sodium oxalate, $\text{Na}_2\text{C}_2\text{O}_{4(\text{aq})}$	dark coloured paper

 **Sodium oxalate is toxic if ingested.**

Evidence

- Prepare a 12-row, 5-column table in which to record your observations and calculations (**Table 3**).

Table 3 Observations and Calculations for Investigation 7.6.1

Well #	A initial $[\text{Ca}_{(\text{aq})}^{2+}]$	B initial $[\text{C}_2\text{O}_{4(\text{aq})}^{2-}]$	C final $[\text{Ca}_{(\text{aq})}^{2+}]$	D final $[\text{C}_2\text{O}_{4(\text{aq})}^{2-}]$	E $(\text{C} \times \text{D})$ ion product
1					
2					
3					
4					
5					
6					
7					
8					
9					
10					
11					
12					

(continued)

Procedure

1. Place a spotplate on a dark sheet of paper.
2. Add 5 drops of distilled water to each of 11 consecutive wells in row A of the spotplate (wells A2 to A12), leaving the first well, A1, empty (**Figure 3**).

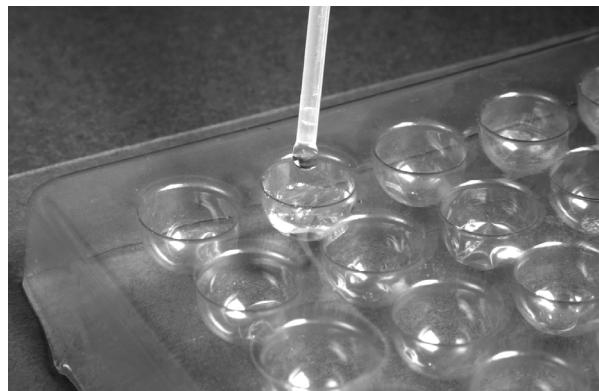


Figure 3

The dark paper should make it easier to detect the formation of a light-coloured precipitate.

3. Add 10 drops of 0.1 mol/L calcium nitrate to well A1.
4. Draw the solution from A1 into an empty pipet, and place 5 drops of the solution into A2. (Return any excess solution to the first well.)
5. Using the same pipet, transfer 5 drops of the solution in A2 to the water in A3.
6. Repeat step 5 for each of the remaining wells. Discard the 5 drops from well A12 into a sink with lots of running water.
7. Using a new pipet, place 5 drops of 0.1 mol/L sodium oxalate solution into wells B1 through B12.
8. Use a clean pipet to transfer the entire contents of well A1 into well B1. Mix well with the tip of the pipet. Continue this process for wells A2 into B2, A3 into B3, etc., until A12 and B12 have been mixed.
9. Examine all of the wells. Identify the first well that appears to have no precipitate.
10. Dispose of chemical wastes as instructed by your teacher.

Analysis

- (b) Calculate the initial $[Ca_{(aq)}^{2+}]$ in each well of row A. Record these in column A of your table.
- (c) Record the initial $[C_2O_{4(aq)}^{2-}]$ in each well of row B in column B of your table.
- (d) Determine the $[Ca_{(aq)}^{2+}]$ and $[C_2O_{4(aq)}^{2-}]$ in wells B1 to B12, and record these values in columns C and D, respectively.
- (e) Calculate the ion product for the contents of each well. Record these in column E of the table.

(continued)

- (f) Write the K_{sp} expression for calcium oxalate and the calculation that determined its value in this experiment.
- (g) Answer the Question by deciding which ion product in column E of the table corresponds to the K_{sp} of calcium oxalate. Give reasons for your decision.
- (h) Use your experimental value of K_{sp} to determine the solubility of calcium oxalate.

Evaluation

- (i) Identify sources of error and uncertainty in the Experimental Design. Provide suggestions for improvement.
- (j) Compare your experimentally derived value for the K_{sp} of calcium oxalate to the accepted value contained in a reliable source such as the *CRC Handbook of Chemistry and Physics*. Calculate the percentage difference. Comment on the validity of your experimental result.

Investigation 7.6.2: Determining K_{sp} for Calcium Hydroxide

INQUIRY SKILLS

- | | |
|--|--|
| <ul style="list-style-type: none">● Questioning○ Hypothesizing● Predicting● Planning● Conducting | <ul style="list-style-type: none">● Recording● Analyzing● Evaluating● Communicating |
|--|--|

Read Investigation 7.6.1 (p. 517 of the Student Text) and design your own investigation to determine the solubility product constant for calcium hydroxide.

Purpose

The purpose of this investigation is to use solubility equilibrium theory and experimental evidence to determine the solubility product constant for calcium hydroxide, and then to compare this value with the accepted reference value.

Question

- (a) Write a question that you will attempt to answer.

Prediction

- (b) Look up the accepted reference value for the solubility product constant for calcium hydroxide.

Experimental Design

- (c) Write your design.

Materials

- (d) List the materials you will need. Be sure to include any necessary safety precautions, particularly in the handling and disposal of calcium hydroxide.

(continued)

Procedure

(e) Write a detailed series of steps for your Procedure.

1. With your teacher's approval, carry out your Procedure.

Analysis

(f) Analyze your Evidence to answer the Question.

Evaluation

(g) Compare your experimental value to the reference value, and determine the percentage difference.

(h) How much confidence do you have in your experimental value? Account for any uncertainty.

Activity 8.1.1: Determining the pH of Common Substances

One reason for the wide acceptance of the pH scale is the availability of convenient, rapid, and precise methods for measuring pH. The purpose of this activity is to measure the pH of common household liquids and to determine the suitability of the various methods to measure the pH of the different substances.

Materials

lab apron
eye protection
red or blue litmus paper
wide-range pH paper (pH 1–14) and colour chart
pH meter and pH 7 buffer solution
wash bottle of distilled water
400-mL waste beaker
several 100-mL beakers and an equal number of watch glasses
several droppers
various household cleaning agents, such as “liquid” ammonia, glass cleaner, drain cleaner, and shampoo
various foods and beverages such as tap water, mineral water, juices, pop, vinegar, and milk



Many household cleaning products are corrosive or toxic. Handle them with care and do not mix them together.

Wash your hands upon completion of the investigation.

- Place a small amount of each substance to be measured in a different 100-mL beaker. The amount depends on the type of pH electrode supplied with the pH meter you are using. Your teacher will tell you what volume you should use.
 - Using a different dropper for each liquid, place a drop of the substance onto a small piece of litmus paper and lay the paper onto a watch glass. Record your observations.
 - Repeat your pH measurements using wide-range pH paper instead of litmus paper.
 - Measure the pH of each sample using a pH meter. For each sample, rinse the electrode of the pH meter with distilled water. Place the pH meter electrode in a standard pH 7 buffer solution and calibrate the instrument by adjusting the meter to read the pH of the buffer. Using a wash bottle and a 400-mL waste beaker, rinse the pH meter electrode with distilled water. Place the electrode in a beaker containing the sample and record the pH reading.
 - Dispose of samples in the sink with lots of running water. Discard the used pieces of pH paper and litmus paper in the wastepaper basket. Return the pH meter to storage, according to your teacher’s instructions.
 - Wash your hands with soap and water.
- (a) Classify the samples as acidic, neutral, or basic.

(continued)

LSM 8.LA-1

- (b) Provide a generalization regarding the acid–base properties of the various materials tested and their household uses.
 - (c) Provide advantages and disadvantages of using litmus paper, wide-range pH paper, and the pH meter in determining the pH of each of the samples tested.
 - (d) Why was the pH meter electrode calibrated with the buffer solution before use?

Investigation 8.3.1: The pH of Salt Solutions

INQUIRY SKILLS

- | | |
|---|--|
| <input type="radio"/> Questioning | <input checked="" type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input checked="" type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input checked="" type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input checked="" type="radio"/> Conducting | |

When some salts are dissolved in water they form neutral solutions; others form acidic or basic solutions. The theory that a variety of ions may affect the acid–base characteristics of an aqueous solution can be tested in the laboratory.

Purpose

The purpose of this investigation is to test the theory that ions of soluble salts may affect the pH of aqueous solutions.

Question

Which salt solutions are acidic, which are basic, and which neutral?

Prediction

- (a) Predict the pH of a 0.1 mol/L aqueous solution of each of the salts in the Materials list.

Experimental Design

The pH of a variety of aqueous salt solutions is measured using a suitable pH measuring system.

Materials

- lab apron
eye protection
pure water
0.10 mol/L aqueous solutions of:
sodium carbonate
sodium phosphate
aluminum sulfate
aluminum chloride
sodium chloride
ammonium chloride
ammonium oxalate
ammonium acetate
ammonium carbonate
ammonium sulfate
potassium sulfate
copper(II) sulfate
iron(III) sulfate
iron(III) chloride
sodium hydrogen carbonate
sodium hydrogen sulfate
pH paper, pH meter, and/or universal indicator
containers (small beakers, test tubes, or spot plates)
waste beakers

(continued)

Procedure

- (b) Write a Procedure for your investigation. If you use a pH meter to measure pH, include a description of the setup procedure and any precautions you must consider.

1. With your teacher's approval, carry out your Procedure.

Analysis

- (c) Analyze your observations and use them to answer the Question. Display your answer in an appropriate table of evidence.

Evaluation

- (d) Compare your answers for (c) to your Predictions. Does the evidence support your predictions? Suggest an explanation for any discrepancies.

Activity 8.4.1: Quantitative Titration

In this activity, you will standardize a sodium hydroxide solution, then determine the concentration of an unknown acid by titration with the standardized base.

Experimental Design

This is a two-part activity. In the first part, you will standardize a sodium hydroxide solution by titrating it with the primary standard, potassium hydrogen phthalate, $\text{KHC}_8\text{H}_4\text{O}_{4(\text{aq})}$. The concentration of the unknown acid solution is then determined by titrating it with the standardized sodium hydroxide solution.

Materials

lab apron
eye protection
sheet of blank white paper
electronic balance
pH meter
distilled water
wash bottle with distilled water
125-mL Erlenmeyer flask
1000-mL glass or plastic bottle
rubber stoppers
 $\text{NaOH}_{(\text{s})}$
 $\text{KHC}_8\text{H}_4\text{O}_{4(\text{s})}$
1% phenolphthalein
vinegar, lemon juice, or other acid solution of unknown concentration
dropper
stirring rod
buret
buret stand
100-mL graduated cylinder
10-mL graduated cylinder
weighing paper



Acids are corrosive and may also be toxic. Solid sodium hydroxide is extremely corrosive, and especially dangerous to eyes and skin. Do not touch your eyes. If sodium hydroxide enters your eye, flush continuously with cold water for 10 min and get immediate medical attention. Wear a lab apron and eye protection.

Procedure

Part I: Standardization of $\text{NaOH}_{(\text{aq})}$

1. In a 125-mL Erlenmeyer flask, dissolve approximately 10 g of $\text{NaOH}_{(\text{s})}$ in 50 mL of previously boiled, distilled water.
2. Transfer 10.0 mL of the solution to a 1000-mL bottle and dilute with 500 mL previously boiled, distilled water. Stir the solution but do not shake it.
3. Weigh approximately 0.4 g $\text{KHC}_8\text{H}_4\text{O}_{4(\text{s})}$ and record the mass to three significant digits.
4. Place the $\text{KHC}_8\text{H}_4\text{O}_{4(\text{s})}$ into a clean, dry Erlenmeyer flask. Add 50.0 mL distilled water and two to three drops of phenolphthalein. Swirl to mix.

(continued)

5. Allow several millilitres of the $\text{NaOH}_{(\text{aq})}$ solution to flow through a buret, making sure that the solution wets all of the inside surfaces.
6. Fill the wetted buret with the $\text{NaOH}_{(\text{aq})}$ solution and record the volume in a suitable chart.
7. Place the Erlenmeyer flask containing $\text{KHC}_8\text{H}_4\text{O}_{4(\text{aq})}$ over a sheet of white paper and titrate with the $\text{NaOH}_{(\text{aq})}$ solution in the buret until the endpoint is reached.
8. Repeat steps 3 to 7 two more times and calculate the mean of the three volumes of $\text{NaOH}_{(\text{aq})}$ used to reach the endpoint. Rinse out the flask.

Part II: Determining the $[\text{H}^+_{(\text{aq})}]$ in a Solution of Unknown Concentration

9. Refill the buret you used in Part A with standardized $\text{NaOH}_{(\text{aq})}$ solution.
10. Place 25.00 mL of an unknown acidic solution into a clean, dry Erlenmeyer flask. Add 2 to 3 drops of phenolphthalein. Swirl to mix.
11. Place the Erlenmeyer flask containing the acid solution over a sheet of white paper and titrate with standardized $\text{NaOH}_{(\text{aq})}$ solution in the buret until the endpoint is reached.
12. Repeat steps 10 to 11 two more times, and calculate the mean of the three volumes of $\text{NaOH}_{(\text{aq})}$ used to reach the endpoint. Use this average value to calculate the concentration of $\text{H}^+_{(\text{aq})}$ in the acidic solution.
13. Calibrate a pH meter and use it to measure the pH of a sample of the unknown acid solution. Record the value.
14. Discard all solutions in the sink with lots of running water. Return materials and equipment to their proper location. Wash your hands with soap and water.

Analysis

Part 1

- (a) Use the average of your titration volumes to calculate the concentration of the $\text{NaOH}_{(\text{aq})}$ solution.
- (b) Why was $\text{KHC}_8\text{H}_4\text{O}_{4(\text{s})}$ used as the acid in the standardization titration?
- (c) Why should you not shake the $\text{KHC}_8\text{H}_4\text{O}_{4(\text{aq})}$ solution before titrating it with $\text{NaOH}_{(\text{aq})}$?
- (d) Why was boiled distilled water used to prepare the solutions?

(continued)

Part II

- (e) Use the evidence from your titration to calculate $[H_{(aq)}^+]$ of the unknown solution.

Evaluation

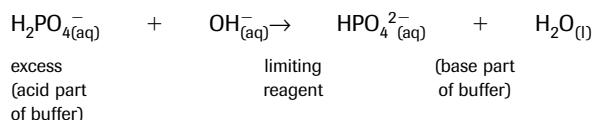
- (f) Evaluate the Procedure and suggest changes that might correct any sources of error.

Investigation 8.5.1: Buffer Action

INQUIRY SKILLS

- Questioning
 - Hypothesizing
 - Predicting
 - Planning
 - Conducting
 - Recording
 - Analyzing
 - Evaluating
 - Communicating

You will investigate the buffering capacity of a $\text{H}_2\text{PO}_4^{-} / \text{HPO}_4^{2-}$ buffer.



Question

How does the pH change when a strong acid and a strong base are slowly added to a $\text{H}_2\text{PO}_{4(\text{aq})}^-/\text{HPO}_{4(\text{aq})}^{2-}$ buffer?

Prediction

- (a) Predict how the pH of the $\text{H}_2\text{PO}_{4(\text{aq})}^-/\text{HPO}_{4(\text{aq})}^{2-}$ buffer will change when $\text{HCl}_{(\text{aq})}$ and $\text{NaOH}_{(\text{aq})}$ are added to different samples of the buffer, as described in the Procedure.

Experimental Design

- (b) Read the Procedure and describe the design in a brief paragraph.

Materials

- (c) Create a suitable list of materials. Have your list approved by your teacher before continuing. Include relevant safety precautions and disposal procedures beside each material in your list.

(continued)

Procedure

1. Obtain 50 mL of 0.10 mol/L $\text{KH}_2\text{PO}_{4(\text{aq})}$ and 29 mL of 0.10 mol/L $\text{NaOH}_{(\text{aq})}$ in separate graduated cylinders.
2. Pour the $\text{KH}_2\text{PO}_{4(\text{aq})}$ and then the $\text{NaOH}_{(\text{aq})}$ into a beaker to prepare a buffer with a pH of 7.
3. Pour an equal amount of the buffer into Test tubes 1 and 2.
4. Add 0.10 mol/L $\text{NaCl}_{(\text{aq})}$ as a control into Tubes 3 and 4.
5. Add two drops of bromocresol green to Tubes 1 and 3.
6. Add and count drops of 0.10 mol/L $\text{HCl}_{(\text{aq})}$ to Tubes 1 and 3 until the colour changes.
7. Repeat steps 5 and 6 with phenolphthalein and 0.10 mol/L $\text{NaOH}_{(\text{aq})}$ in Tubes 2 and 4.
8. Dispose of all solutions down the drain with lots of running water.

Evidence

- (d) After reading the Procedure, create a table in which to record observations.

Analysis

- (e) Answer the Question.

(continued)

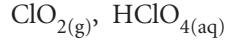
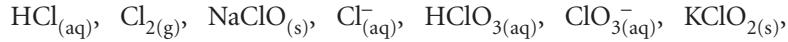
Evaluation

(f) Evaluate the Experimental Design.

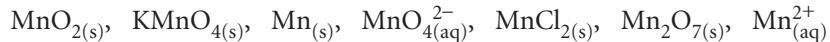
(g) Evaluate your Prediction based on the evidence gathered and your confidence in the design.

Oxidation Numbers, Extra Exercises

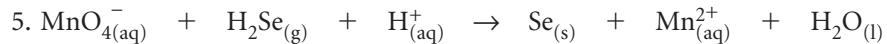
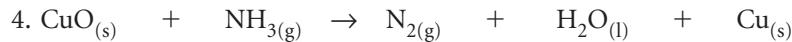
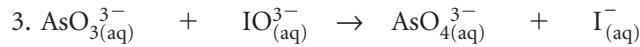
1. Assign oxidation numbers to chlorine in each of the following chemicals.



2. Assign oxidation numbers to manganese in each of the following chemicals.



For the following reaction equations, use oxidation numbers to identify the oxidation (the atom oxidized) and the reduction (the atom reduced).

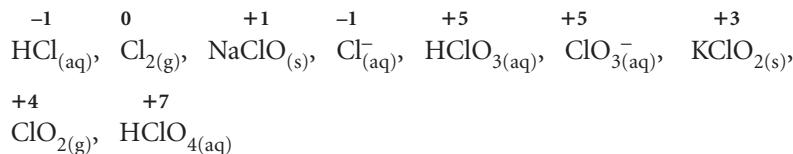


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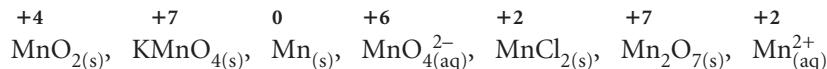
LSM 9.1-3

Oxidation Numbers, Extra Exercises, Solution

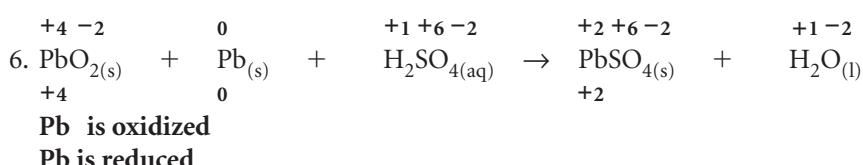
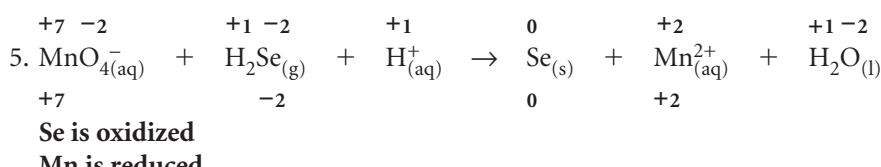
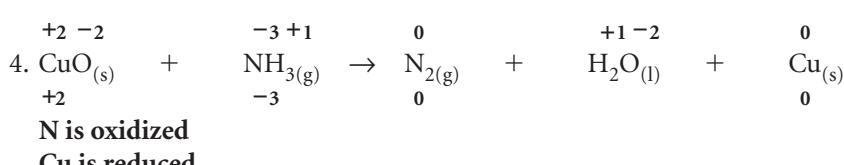
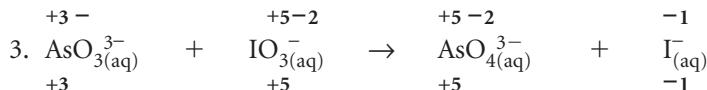
1. Assign oxidation numbers to chlorine in each of the following chemicals.



2. Assign oxidation numbers to manganese in each of the following chemicals.

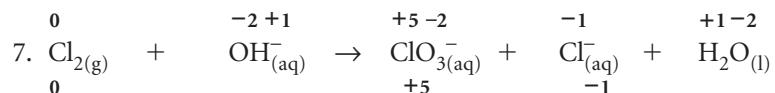


For the following reaction equations, use oxidation numbers to identify the oxidation (the atom oxidized) and the reduction (the atom reduced).



(Note that the lead(IV) ions in PbO_2 and the lead atoms in the metal element are both converted into lead(II) ions in the course of this reaction, which represents a car battery discharging.)

(continued)

LSM 9.1-4

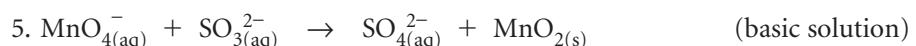
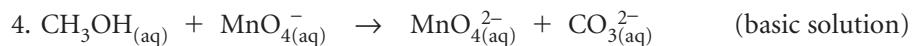
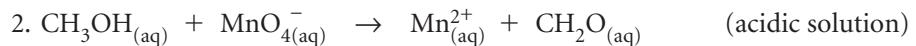
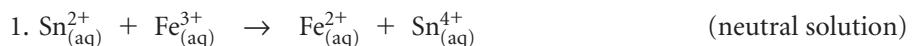
Cl is oxidized

Cl is reduced

(Note that some of the Cl atoms in the gaseous element oxidize while others reduce in this reaction. Such reactions are sometimes called auto-oxidation, or self-oxidation reduction.)

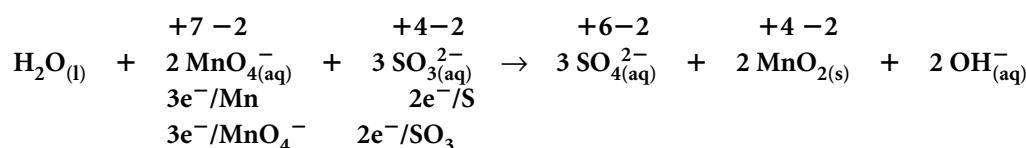
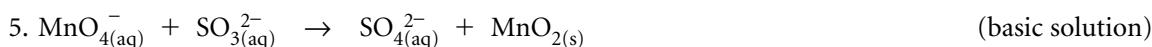
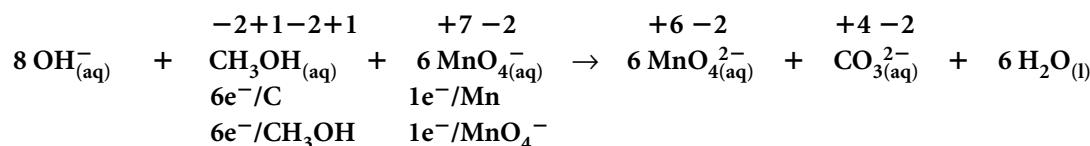
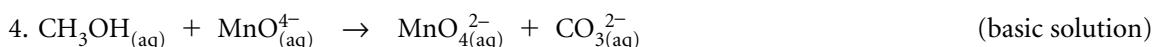
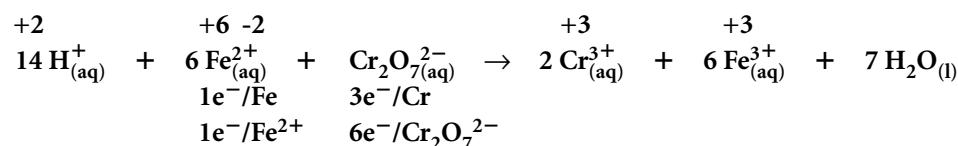
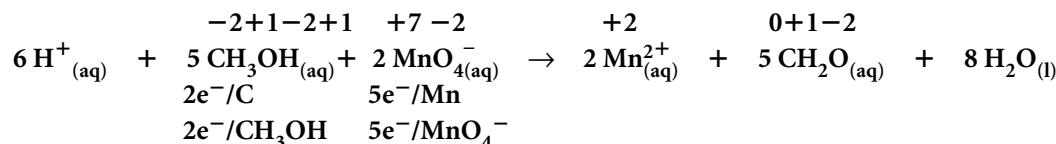
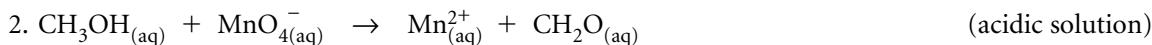
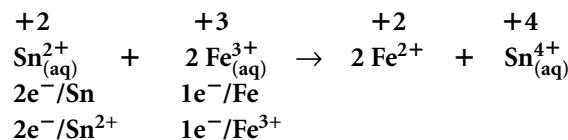
Balancing Redox Equations with Oxidation Numbers, Extra Exercises

Use the procedure summarized on page 667 of the Student Text to balance the following reaction equations:



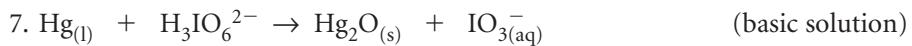
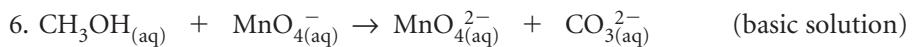
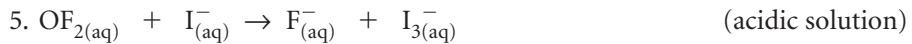
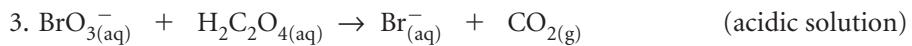
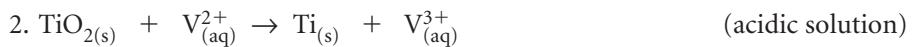
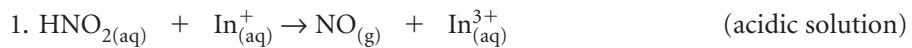
Balancing Redox Equations with Oxidation Numbers, Extra Exercises, Solution

Use the procedure summarized on page 667 of the Student Text to balance the following reaction equations:



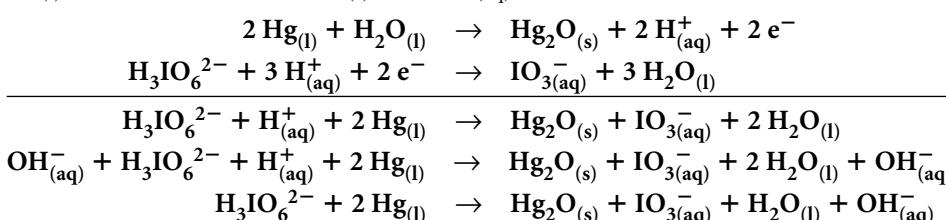
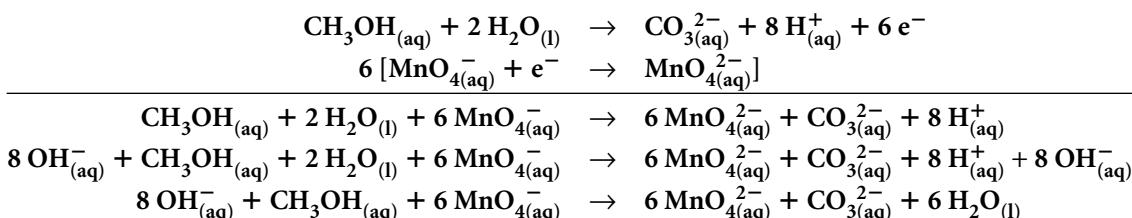
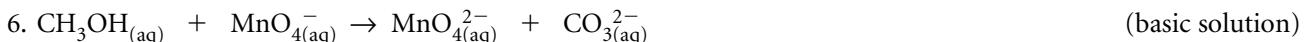
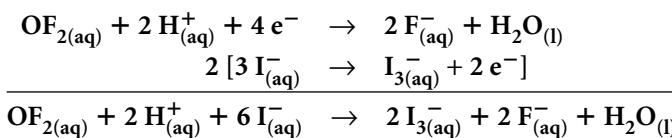
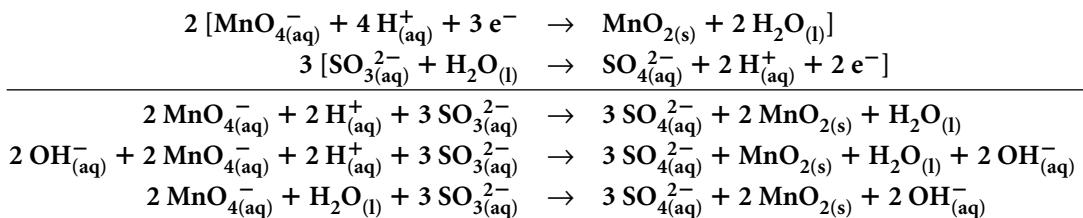
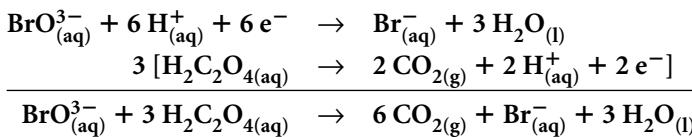
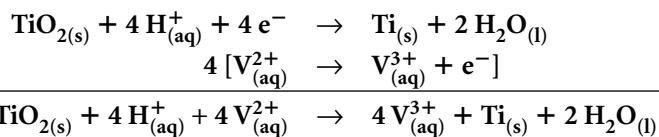
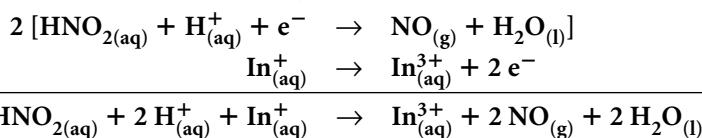
Balancing Redox Equations with Half-Reaction Equations, Extra Exercises

Use the procedure summarized on page 672 of the Student Text to balance the following reaction equations:



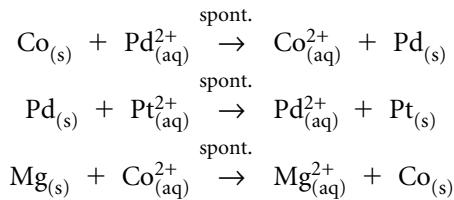
Balancing Redox Equations with Half-Reaction Equations, Extra Exercises, Solution

Use the procedure summarized on page 672 of the Student Text to balance the following reaction equations:

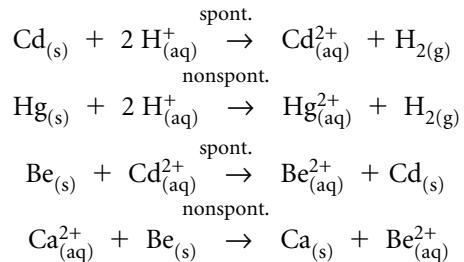


Creating a Redox Table, Extra Exercises

1. The following equations represent spontaneous reactions. From this evidence, set up a table of relative strengths of oxidizing and reducing agents. Write half-reaction equations and label the strongest oxidizing agent and reducing agent.

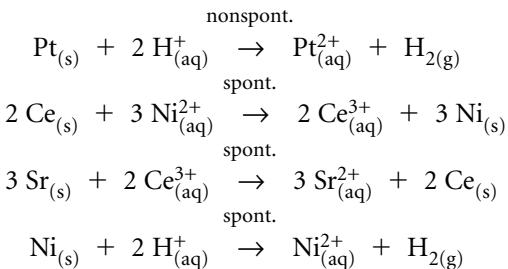


2. The following equations are interpretations of the evidence from the reactions of four metals with various cation solutions. Make a table of half-reactions and arrange the four metallic ions and the hydrogen ion in order of their decreasing tendency to react. Label the strongest oxidizing agent and reducing agent.



(continued)

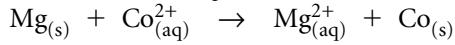
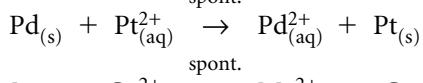
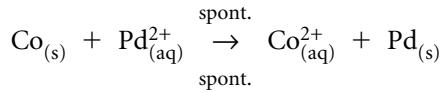
3. Four metals were placed into test tubes containing various ion solutions. Their resulting behaviour is communicated by the equations below. Create a half-reaction table and order the metallic ions and the hydrogen ion according to their tendency to react. Label the strongest oxidizing agent and reducing agent.



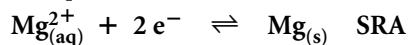
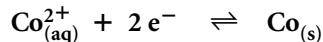
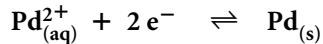
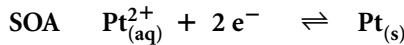
4. A student is required to store an aqueous solution of iron(III) nitrate. She has a choice of a copper, tin, iron, or silver container. Use the table of relative strengths of oxidizing and reducing agents in Appendix C11 of *Nelson Chemistry 12* and an appropriate generalization to predict which container would be most suitable for storing the solution.
5. An analytical chemist reacts an unknown metal X with a copper(II) sulfate solution, plating out copper metal. Metal X does not react with aqueous zinc nitrate. What is the order for these metal ions in decreasing tendency to react? What groups of metals are eliminated as a possible identity of the unknown metal? What other solutions might next be chosen to help identify the unknown metal?

Creating a Redox Table, Extra Exercises, Solution

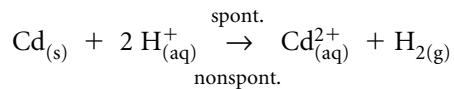
1. The following equations represent spontaneous reactions. From this evidence, set up a table of relative strengths of oxidizing and reducing agents. Write half-reaction equations and label the strongest oxidizing agent and reducing agent.



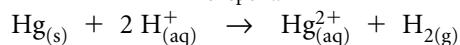
According to the evidence and the spontaneity rule, a table of oxidizing and reducing agents in order of strength is developed as follows:



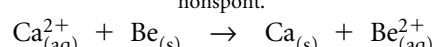
2. The following equations are interpretations of the evidence from the reactions of four metals with various cation solutions. Make a table of half-reactions and arrange the four metallic ions and the hydrogen ion in order of their decreasing tendency to react. Label the strongest oxidizing agent and reducing agent.



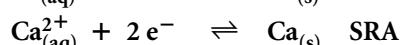
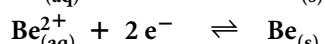
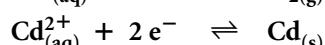
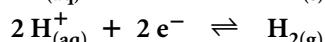
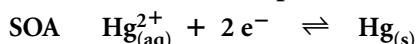
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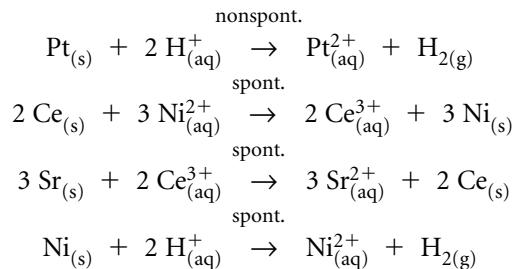


According to the evidence and the spontaneity rule, the following table of half-reactions is developed:

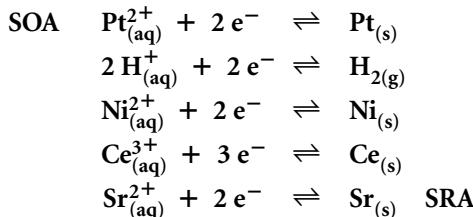


(continued)

3. Four metals were placed into test tubes containing various ion solutions. Their resulting behaviour is communicated by the equations below. Create a half-reaction table and order the metallic ions and the hydrogen ion according to their tendency to react. Label the strongest oxidizing agent and reducing agent.



According to the evidence and the spontaneity rule, the following table of half-reactions is created:

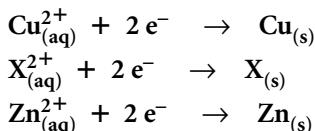


4. A student is required to store an aqueous solution of iron(III) nitrate. She has a choice of a copper, tin, iron, or silver container. Use the table of relative strengths of oxidizing and reducing agents in Appendix C11 of *Nelson Chemistry 12* and an appropriate generalization to predict which container would be most suitable for storing the solution.

Based upon the spontaneity rule, she should choose the silver container for storing the iron(III) nitrate solution. (This is, of course, from a scientific perspective. From an economic perspective, one would likely be concerned about the cost of using a silver container to store an inexpensive solution which could also be stored in a glass container.)

5. An analytical chemist reacts an unknown metal X with a copper(II) sulfate solution, plating out copper metal. Metal X does not react with aqueous zinc nitrate. What is the order for these metal ions in decreasing tendency to react? What groups of metals are eliminated as a possible identity of the unknown metal? What other solutions might next be chosen to help identify the unknown metal?

According to the evidence and the spontaneity rule, the metal ions in order of decreasing tendency to react are listed below:



Metal X must be between copper and zinc in a table of half-reactions. This eliminates any Group 1, 2, and 3 metals as a possible identity for Metal X. Groups 1, 2, and 3 metals are stronger reducing agents than zinc. The aqueous solutions of metal ions between copper and zinc ions in the table of relative strengths of oxidizing and reducing agents in Appendix C11 (page 805) of the Student Text may be used to help narrow down the possibility of what Metal X may be. Some possible solutions that can be used are solutions of lead(II), nickel(II), cobalt(II), cadmium(II), chromium(III), and iron(III) aqueous ions.

Predicting Redox Reactions, Extra Exercises

For each of the following questions, use the five-step method outlined in *Nelson Chemistry 12* on page 682 to predict and communicate the most likely redox reaction, and use the spontaneity rule to predict whether the reaction will occur spontaneously.

1. An aqueous solution of potassium permanganate was reacted with an acidic solution of sodium bromide and an orange-brown substance was formed.
 2. A strip of silver metal is placed in a solution of aqueous nickel(II) chloride.
 3. Liquid mercury is mixed with a paste of acidic manganese(IV) oxide.
 4. Hydrogen peroxide and silver nitrate are mixed.

(continued)

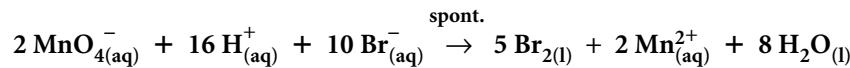
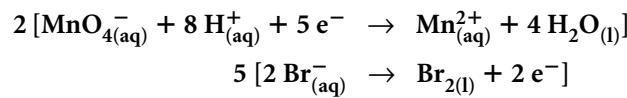
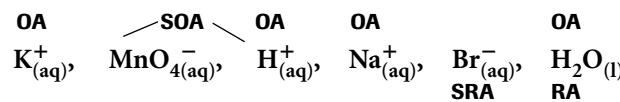
5. Potassium metal is placed into some water.

6. In a car battery lead and lead(IV) oxide electrodes are exposed to a sulfuric acid electrolyte. (Assume that the sulfuric acid ionizes to hydrogen and sulfate ions.)

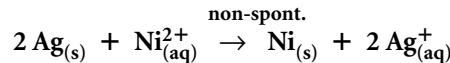
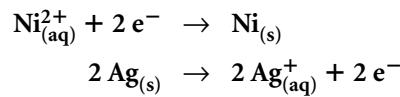
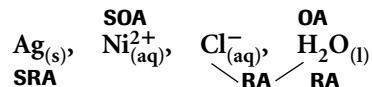
Predicting Redox Equations, Extra Exercises, Solution

For each of the following questions, use the five-step method outlined in *Nelson Chemistry 12* on page 682 to predict and communicate the most likely redox reaction, and use the spontaneity rule to predict whether the reaction will occur spontaneously.

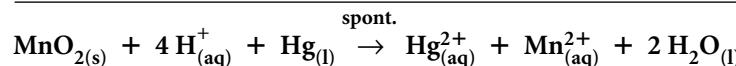
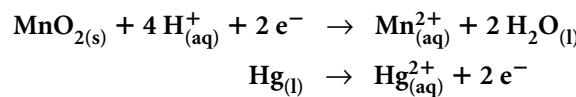
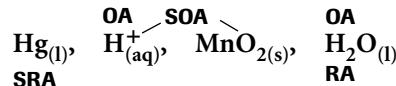
1. An aqueous solution of potassium permanganate was reacted with an acidic solution of sodium bromide and an orange-brown substance was formed.



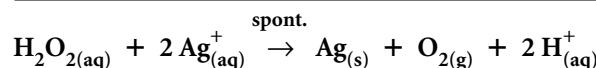
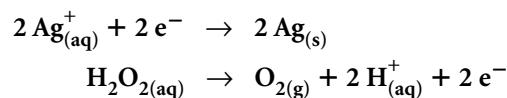
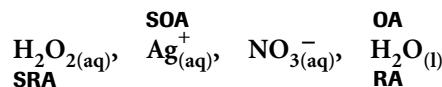
2. A strip of silver metal is placed in a solution of aqueous nickel(II) chloride.



3. Liquid mercury is mixed with a paste of acidic manganese(IV) oxide.

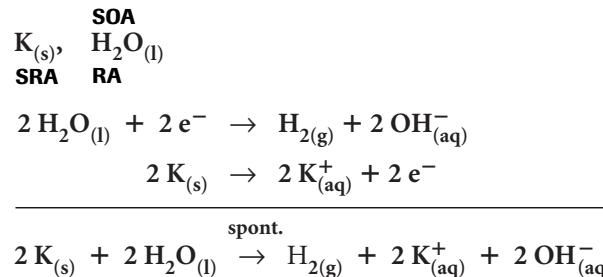


4. Hydrogen peroxide and silver nitrate are mixed.

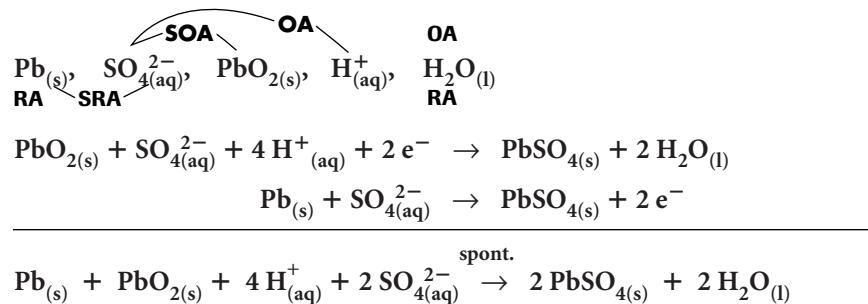


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5. Potassium metal is placed into some water.

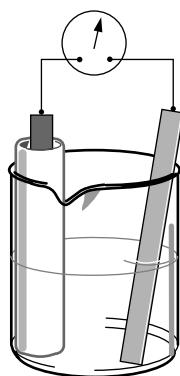
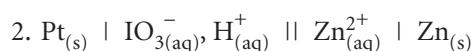
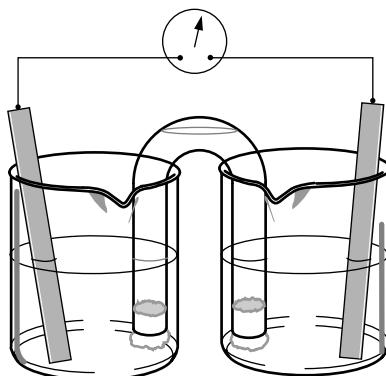
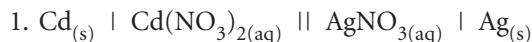


6. In a car battery lead and lead(IV) oxide electrodes are exposed to a sulfuric acid electrolyte. (Assume that the sulfuric acid ionizes to hydrogen and sulfate ions.)



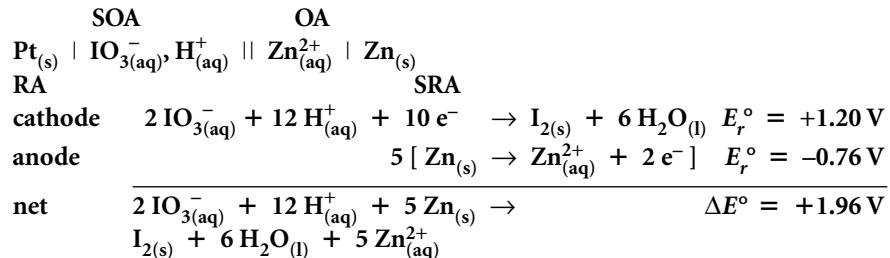
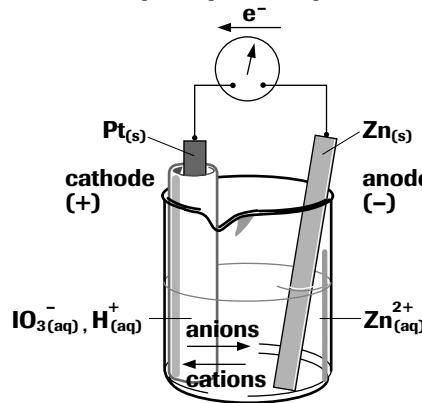
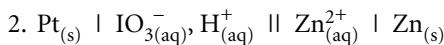
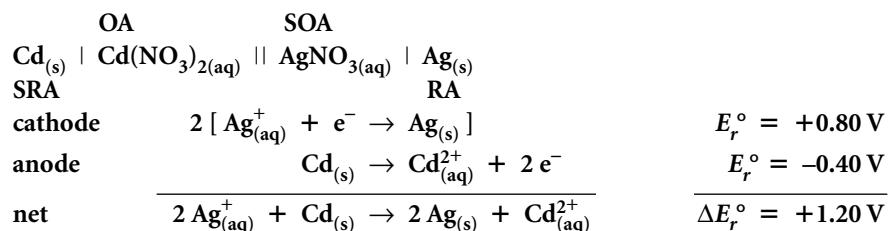
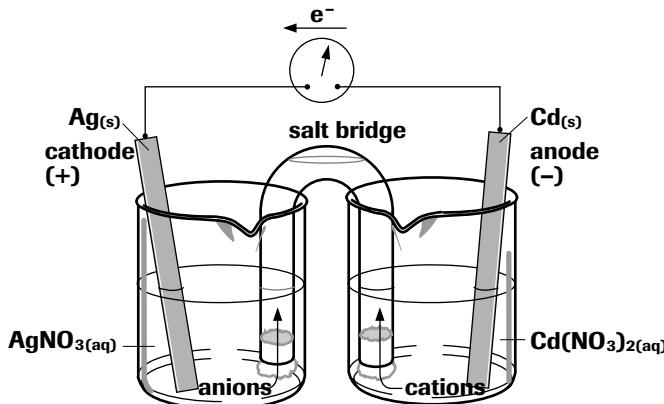
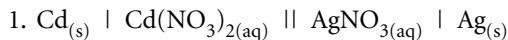
Galvanic Cells, Extra Exercises

For each of the following cells, use the given cell notation to identify the strongest oxidizing and reducing agents. Write chemical equations to represent the cathode, anode, and net cell reactions. Label electrodes, electrolytes, electron flow, and ion movement. Predict the cell potential.



Galvanic Cells, Extra Exercises, Solution

For each of the following cells, use the given cell notation to identify the strongest oxidizing and reducing agents. Write chemical equations to represent the cathode, anode, and net cell reactions. Label electrodes, electrolytes, electron flow, and ion movement. Predict the cell potential.



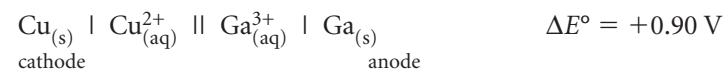
Cell Potentials, Extra Exercises

1. Assume that the reference half-cell is changed to a standard mercury-mercury(II) half-cell.
 - (a) What would be the reduction potential of a standard chlorine half-cell?
 - (b) What would be the reduction potential of a standard nickel half-cell?
 - (c) What would be the net cell potential of a standard chlorine-nickel cell?
 - (d) Why is the answer to part c the same as the answer obtained using the standard hydrogen half-cell as the reference?
2. For each of the following standard cells, write the cell notation, label electrodes, and determine the cell potential.
 - (a) cobalt-hydrogen standard cell
 - (b) zinc-aluminum standard cell
 - (c) tin(IV)-zinc standard cell
3. For each of the following cell notations, write the cathode, anode, and net cell reaction equations and calculate the cell potential.
 - (a) $\text{Pb}_{(\text{s})} \mid \text{Pb}_{(\text{aq})}^{2+} \parallel \text{Ni}_{(\text{aq})}^{2+} \mid \text{Ni}_{(\text{s})}$
 - (b) $\text{Pt}_{(\text{s})} \mid \text{SO}_{4(\text{aq})}^{2-}, \text{H}_{(\text{aq})}^+, \text{H}_2\text{SO}_{3(\text{aq})} \parallel \text{Ag}_{(\text{aq})}^+ \mid \text{Ag}_{(\text{s})}$
 - (c) $\text{Cd}_{(\text{s})} \mid \text{Cd}_{(\text{aq})}^{2+} \parallel \text{ClO}_{4(\text{aq})}^-, \text{H}_{(\text{aq})}^+, \text{Cl}_{(\text{aq})}^- \mid \text{C}_{(\text{s})}$

(continued)

LSM 9.5-6

4. Use the standard cell described below to determine the standard reduction potential of the gallium half-cell.

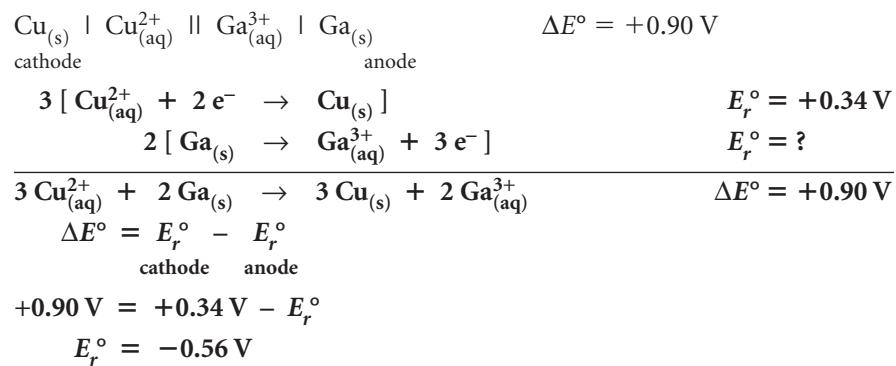


Cell Potentials, Extra Exercises, Solution

(continued)

LSM 9.5-7

4. Use the standard cell described below to determine the standard reduction potential of the gallium half-cell.



Chemical Bonding

INQUIRY SKILLS

- | | |
|-----------------|-----------------|
| ○ Questioning | ● Recording |
| ○ Hypothesizing | ● Analyzing |
| ● Predicting | ● Evaluating |
| ● Planning | ● Communicating |
| ● Conducting | |

In this investigation you will observe some common single displacement reactions and then interpret the changes in terms of electron transfer.

Purpose

The purpose of this investigation is to gather some evidence about single displacement reactions.

Question

What are the products of the single displacement reactions for the following sets of reactants?

- copper and aqueous silver nitrate
- aqueous chlorine and aqueous sodium bromide
- magnesium and hydrochloric acid
- zinc and aqueous copper(II) sulfate
- aqueous chlorine and aqueous potassium iodide

Prediction

- (a) According to the single displacement generalization, predict the balanced chemical equations for each set of reactants listed above.

Experimental Design

Small quantities of reactants are mixed and diagnostic tests (Appendix A6) are used to determine the products of each reaction.

- (b) For each chemical equation listed in your prediction, record diagnostic tests that you will use. (Some diagnostic tests will be very specific and some will be a general observation you expect.)

Materials

lab apron	dropper bottles of 0.1 mol/L
eye protection	—aqueous silver nitrate
five small test tubes	—aqueous sodium bromide
two test-tube stoppers	—aqueous copper(II) sulfate
test-tube rack	—aqueous potassium iodide
emery paper or steel wool	—hydrochloric acid
wash bottle	chlorine water (or bleach)

(continued)



The substances used in this experiment are toxic and corrosive. Avoid skin contact. Wash any splashes on the skin or clothing with plenty of water. If any chemical is splashed in your eyes, rinse for at least 15 min and inform your teacher.



Magnesium and cyclohexane are highly flammable. Keep away from open flame.



Keep the cyclohexane sealed to avoid evaporation and inhalation of the vapours. Dispose of the hydrocarbon mixtures as directed by your teacher.

Make sure matches are extinguished by dipping in water.

matches
copper wire or strip
zinc strip

cyclohexane
magnesium metal ribbon

Procedure

- Set up five test tubes, each filled to a depth of 2–3 cm with one of the five aqueous solutions.
- Add the appropriate element to each solution (see Question) in each test tube.
- Perform diagnostic tests on each of the five mixtures. Record your observations.
- Dispose of the solutions as directed by your teacher.

Evidence

- (c) Design a table to record your observations.

Analysis

- (d) Interpret your evidence and record the products that you can reasonably conclude you obtained in each reaction.

Evaluation

- (e) Evaluate the experimental design, materials, and procedure by considering any possible flaws and improvements.
- (f) Use your answer to (a). What is your judgment of the quality and quantity of evidence obtained?
- (g) How confident are you in the answer obtained? Provide your reasons.
- (h) Evaluate the Prediction and provide your reasons.

Lab Exercise 9.1.1: Oxidation States of Vanadium

Vanadium is a transition metal that forms many different ions with different oxidation states (**Table 1**). Vanadium and its compounds have many different uses, including colouring for glass, ceramics, and plastics.

Table 1 Colours of Vanadium Ions

Ion name	Ion formula	Colour
vanadate(V)	$\text{VO}_{3(\text{aq})}^-$	yellow
vanadate (IV)	$\text{VO}_{(\text{aq})}^{2+}$	blue
vanadium(III)	$\text{V}_{(\text{aq})}^{3+}$	green
vanadium(II)	$\text{V}_{(\text{aq})}^{2+}$	purple

Purpose

The purpose of this lab exercise is to investigate some redox chemistry of vanadium compounds.

Question

What are the oxidation states and changes in oxidation number for vanadium ions?

Evidence

Table 2 Reactions of Vanadium Ions

Procedure	Final solution colours
(1) ammonium vanadate(V) dissolved in sulfuric acid	yellow
(2) yellow solution with three subsequent additions of small quantities of zinc dust	yellow turned blue, then green, then violet
(3) violet solution left sitting in an open container	slowly turned green
(4) yellow solution mixed with potassium iodide solution	very dark, almost black
(5) blue solution mixed with potassium iodide solution	stayed blue; no change
(6) violet solution slowly mixed with acidic potassium permanganate	violet to green to blue to yellow

(continued)

Analysis

- (a) Using **Table 1**, identify the vanadium ions in the sequence of reactions in **Table 2**.
- (b) In each case, is the vanadium being oxidized or reduced? Justify your answer, using oxidation numbers.
- (c) Explain the observations made in (3) to (6) in **Table 2** above. Suggest what is causing these changes.
- (d) Vanadium ion chemistry can be quite complicated. For example, the initial yellow solution is likely an equilibrium between $\text{VO}_{3(\text{aq})}^-$ and $\text{VO}_{2(\text{aq})}^+$. Does this alter your analysis in the previous questions? Explain briefly.

Investigation 9.3.1: Spontaneity of Redox Reactions

INQUIRY SKILLS

- | | |
|-----------------|-----------------|
| ○ Questioning | ● Recording |
| ○ Hypothesizing | ● Analyzing |
| ● Predicting | ● Evaluating |
| ● Planning | ● Communicating |
| ● Conducting | |

In the past, we have usually assumed that all chemical reactions are spontaneous; that is, they occur of their own accord once reactants are placed in contact, without a continuous addition of energy to the system. Spontaneous redox reactions in solution generally provide visible evidence of a reaction within a few minutes.

Purpose

The scientific purpose of this investigation is to test the assumption that all single displacement reactions are spontaneous.

Question

Which combinations of copper, lead, silver, and zinc metals and their aqueous metal ion solutions produce spontaneous reactions?

Prediction

- (a) State and justify your answer to the question.



The solutions used are toxic—especially the lead solution—and irritants. Avoid skin contact. Remember to wash your hands before leaving the laboratory.

Dispose of reaction products according to your teacher's instructions.

Rinse all of the metal strips thoroughly and return them so they can be used again.

LEARNING TIP

Distinguishing Lead and Zinc

When cleaned, lead and zinc metals look very similar. However, lead is much softer. You can distinguish between them because lead strips bend and scratch much more easily than zinc strips.

Experimental Design

A drop of each metal ion solution is placed in separate locations on a clean area of each of the four metal strips.

Materials

lab apron
eye protection
reusable strips of copper, lead, silver, and zinc metals
0.10 mol/L solutions of copper(II) nitrate, lead(II) nitrate, silver nitrate, and zinc nitrate in dropper bottles
steel wool or emery paper

Procedure

- (b) Write a brief procedure for this investigation. Have your teacher approve your procedure before you start.

(continued)

Evidence

- (c) Design a table to record your observations.

Analysis

- (d) Based on your evidence, which combinations of reactants produced spontaneous reactions?

Evaluation

- (e) Were the experimental design, materials, and procedure sufficient to answer the question? Justify your answer.

- (f) Suggest an improvement to increase the certainty of the evidence.

- (g) Evaluate the prediction, including your reasons.

- (h) What does your answer to (g) tell you about the assumption on which the prediction was based? What should be done with this assumption?

Lab Exercise 9.3.1: Building a Redox Table

INQUIRY SKILLS

- Questioning
- Recording
- Hypothesizing
- Analyzing
- Predicting
- Evaluating
- Planning
- Communicating
- Conducting

Suppose that a research team is developing a table of relative strengths of oxidizing and reducing agents. One team member had completed an investigation summarized in Table 3, Section 9.3, of the Student Text, and another had completed the investigation reported in Practice question 9, Section 9.3. A third member used the combination of metals, nonmetals, and solutions shown below. By completing this exercise, you will see how scientists have developed more extensive tables of relative strengths of oxidizing and reducing agents.

Purpose

The purpose of this lab exercise is to construct a table of relative strengths of oxidizing and reducing agents.

Question

What is the table of relative strengths of oxidizing and reducing agents for copper, silver, bromine, and iodine?

Evidence

Table 3 Reactions of Metals and Nonmetals with Solutions of Ions

	I _(aq) ⁻	Cu ²⁺ _(aq)	Ag ⁺ _(aq)	Br _(aq) ⁻
I _(aq) ⁻	X	X	✓	✓
Cu _(s)	✓	X	✓	✓
Ag _(s)	X	X	X	✓
Br _(aq) ⁻	X	X	X	X

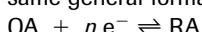
X no evidence of a redox reaction

✓ evidence redox reaction occurred

LEARNING TIP

Tables of Oxidizing and Reducing Agents

All tables of oxidizing and reducing agents follow the same general format.



Analysis

- Using the results from **Table 3**, prepare a table of relative strengths of oxidizing and reducing agents.

(continued)

Synthesis

- (b) Compare Table 3 in Section 9.3, your analysis table from Practice question 9 in section 9.3, and **Table 3** from (a) above. Note that there are several substances that appear in two of these tables. Combine all three tables in one larger table showing the order of oxidizing and reducing agents.

Investigation 9.3.2: The Reaction of Sodium with Water

INQUIRY SKILLS

- | | |
|---|--|
| <input type="radio"/> Questioning | <input checked="" type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input checked="" type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input checked="" type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input type="radio"/> Conducting | |

This demonstration will provide practice in both predicting and testing a chemical reaction.

Purpose

The purpose of this demonstration is to test the five-step method for predicting redox reactions.

Question

What are the products of the reaction of sodium metal with water?

Prediction

- (a) Answer the question and provide your reasoning.

Experimental Design

- (b) Write a general plan for this reaction.

- (c) Suggest one diagnostic test for each predicted product, using the “If [procedure], and [evidence], then [analysis]” format for every product predicted. (This format is described in Appendix A6.)

- (d) What control(s) should be used with these tests?



This reaction of sodium metal with water must be demonstrated with great care, because a great deal of heat is produced.

Evidence

- (e) Record all observations, including the controls, in a suitable table.

(continued)

Analysis

(f) According to the evidence collected, what products were obtained?

Evaluation

(g) Assuming the evidence is of suitable quality, evaluate your prediction.

(h) Evaluate the method of writing redox reactions used to make your prediction.

(i) This experiment is not sufficient to provide a judgment of the method of predicting redox reactions. What should be done next?

Activity 9.4.1: Developing an Electric Cell

In this activity, an aluminum soft-drink can is both the container and one of the electrodes (**Figure 1**). The other electrode is a solid conductor such as graphite from a pencil, an iron nail, or a piece of copper wire or pipe. The electrolyte may be a salt solution or an acidic or basic solution. Although the overall performance of a cell depends on many factors, only the voltage is investigated here.

The purpose of this activity is to use a technological problem-solving (trial-and-error) approach to construct a working electric cell with the highest possible voltage. You will need to control variables and work in a systematic way. Be prepared to alter your materials if your results are not promising and to maximize your results.

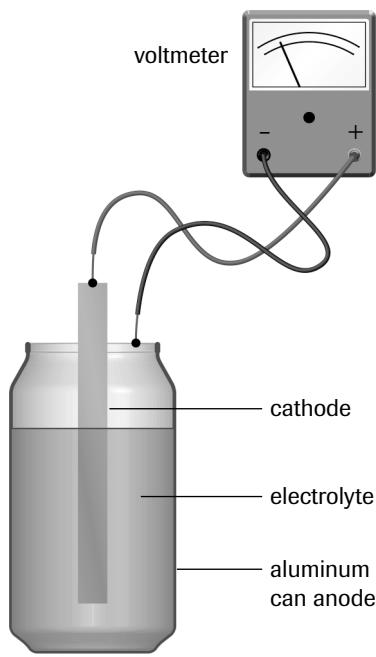


Figure 1

An aluminum can cell is an efficient design, since one of the electrodes also serves as the container.

Materials

- lab apron
- eye protection
- various electrodes
- acidic, basic, and neutral ionic solutions
- bottle of distilled water
- steel wool or emery paper
- voltmeter
- 2 plug-and-clip wires
- an aluminum can with the top removed

Design

- Using the same electrolyte and the aluminum can as the control variables, test two or three different materials as the second electrode. Measure the voltage of each cell. (Scrape the paint from the can where the wire is attached.)
- Using the same two electrodes as the control variables, test two or three possible electrolytes. Measure the voltage of each cell.
- Test additional combinations, based on the analysis of the initial trials.

Evidence

- (a) Keep careful records of all observations, including what worked and did not work.



Be careful when handling acidic and basic solutions because they are corrosive. Wear eye protection and work near a source of water. Electrolytes may be toxic or irritants; follow all safety precautions. Avoid eye and skin contact.

Dispose of solutions according to your teacher's instructions.

(continued)

- (b) Set up a table or organized list to record your observations and analysis of each trial in your problem-solving cycle.

Analysis

- (c) What is the best result you obtained? Report the details of the design and voltage.

Evaluation

- (d) Evaluate the quality of your evidence, including any sources of experimental error or uncertainty.

- (e) If you were to continue this process, what changes or improvements would you make?

- (f) Evaluate the suitability of your final electric cell for potential commercial development. Consider a variety of factors.

Lab Exercise 9.4.1: Characteristics of a Hydrogen Fuel Cell

INQUIRY SKILLS

- | | |
|-------------------------------------|---|
| <input type="radio"/> Questioning | <input type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input type="radio"/> Communicating |
| <input type="radio"/> Conducting | |

A fuel cell supplies electrical energy in the same way as any electric cell. Ions transfer the charge within the cell, and electrons flow through an external circuit between the electrodes. In a hydrogen fuel cell, the electrodes are not consumed, and there is no need to reverse the cell to recharge it. The reactants, hydrogen and oxygen, are continuously supplied and consumed to produce water and electricity.

Commercial hydrogen fuel cells, such as the Ballard cell, utilize a solid electrolyte called a proton exchange membrane (PEM). This polymer is bonded to two porous carbon cloth electrodes (**Figure 2**).

Purpose

The purpose of this lab exercise is to compare the electrical characteristics of a hydrogen fuel cell and a dry cell.

Question

How do the trends of the voltage-current and power-current graphs compare for a hydrogen fuel cell and a typical dry cell?

Experimental Design

Hydrogen and oxygen gases, produced from the electrolysis of water, are passed into a hydrogen fuel cell. The voltage of the cell is measured with an open circuit. Various resistances are added to the circuit, and the voltage and current are measured for each resistance (load). The fuel cell is disconnected and replaced by a dry cell. The same procedure is followed to collect voltage and current measurements for the same set of resistances in the circuit.

Evidence

Fuel cell		Dry cell	
Voltage (V)	Current (mA)	Voltage (V)	Current (mA)
0.81	0	1.44	0
0.80	5	1.42	5
0.79	6	1.41	10
0.78	16	1.40	22
0.73	66	1.34	90
0.70	115	1.30	146
0.67	175	1.25	212
0.61	315	1.14	335

(continued)

Analysis

- (a) Construct and label a graph with voltage on the vertical axis and current on the horizontal axis. Plot and label the line for the fuel cell and another line for the dry cell.

(b) Describe and compare the trend shown by the voltage-current line for the fuel cell with that of the dry cell. Does the hydrogen fuel cell behave like a regular (dry) cell?

(c) The power output of a cell is a measure of the quantity of energy per second delivered by the cell. For each voltage and current, calculate the power supplied by using the formula $P = VI$. (Recall that power is measured in watts (W) and that $1 \text{ V} \cdot \text{A} = 1 \text{ W}$.) Summarize your results in a table for both the fuel cell and the dry cell.

(continued)

- (d) Repeat (a) and (b) to construct a power curve for each cell. Plot power on the vertical axis and current on the horizontal axis.

- (e) Use the results of your analysis to answer the Question.

Synthesis

- (f) Compare the designs of fuel cells and dry cells.

- (g) What makes fuel cells more practical for powering electric cars compared to other types of cells?

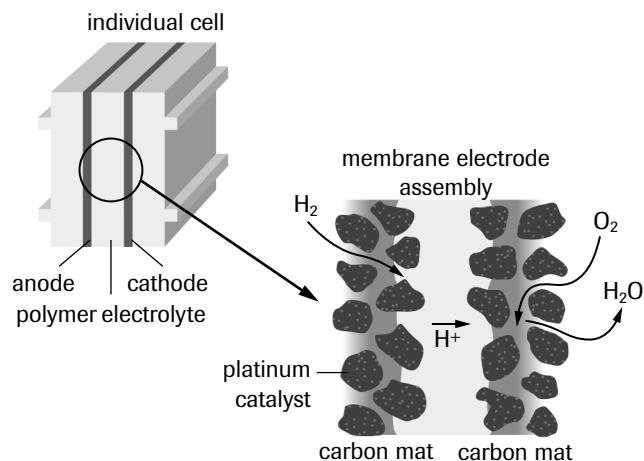


Figure 2

The membrane electrode assembly includes the electrodes (carbon mats) and the proton exchange polymer. Each carbon particle has tiny platinum particles on its surface.

Activity 9.5.1: Galvanic Cell Design

The purpose of this activity is to demonstrate the design and operation of a galvanic cell used in scientific research. A cell with only one electrolyte is compared with similar cells containing the same electrodes but two electrolytes (Figure 3).

Materials

- lab apron
- eye protection
- silver and copper electrodes
- steel wool (for cleaning electrodes)
- voltmeter with leads
- 4 medium-sized beakers
- U-tube
- cotton plugs
- porous cup
- bottle of distilled water
- solutions of sodium nitrate, silver nitrate, copper(II) nitrate
- Construct the three cells shown in Figure 3.
- (a) Which design is most similar to Volta's invention? Compare the three cell designs.

 **Solutions used are irritants and are toxic if ingested. Avoid contact with skin and eyes. Silver nitrate will temporarily blacken your skin.**

Dispose of solutions according to your teacher's instructions.

- Use a voltmeter to determine which electrode is positive and which is negative, and measure the electric potential difference (voltage) of each cell.

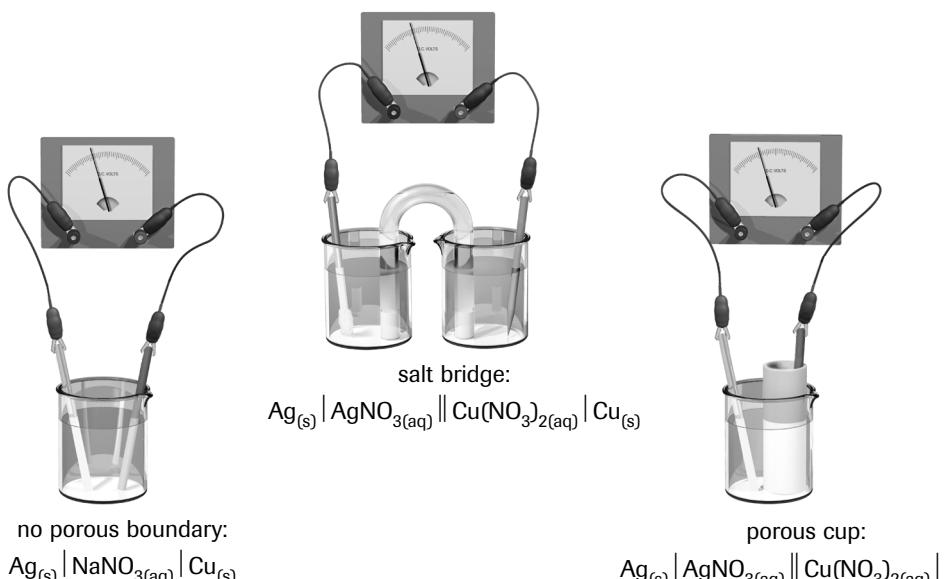


Figure 3
Three different cell designs

(continued)

- (b) According to the voltmeter test, which electrode is the cathode and which is the anode?
- (c) Why is your answer to (b) the same for all three cells?
- (d) Suggest a reason why two of the voltages measured are very similar and the third is very different.
- With the voltmeter connected, remove and then replace the various parts of each cell.
- (e) Why does the voltmeter reading go to zero when one of the parts of the cell is removed?
- (f) What common device in your home and school also “breaks” the circuit?
- For each cell, connect the two electrodes directly with a wire. Record any evidence of a reaction after several minutes, and after one or two days. Measure the electric potential difference after several days.
- (g) What is the design of a control that can be used to compare changes with each of the three cells?
- (h) State some diagnostic tests that could be done to obtain more specific evidence for the operation of each cell.
- (i) Suggest a reason why all solutions were nitrates.

Investigation 9.5.1: Investigating Galvanic Cells

INQUIRY SKILLS

- | | |
|---|--|
| <input type="radio"/> Questioning | <input checked="" type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input checked="" type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input checked="" type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input checked="" type="radio"/> Conducting | |

In this investigation, you are given the opportunity to construct galvanic cells and compare your observations with the rules and concepts you have learned.

Purpose

The purpose of this investigation is to compare the predictions of cell potentials and electrodes of various cells with those measured in the lab.

Question

In cells constructed from various combinations of copper, lead, silver, and zinc half-cells, what are the standard cell potentials, and which is the anode and cathode in each case?

Prediction

- (a) According to redox concepts and a redox table (Appendix C11), prepare a table of all possible combinations of half-cells and answer the question.

Experimental Design

- (b) Write a brief general plan to answer the question, including the identification of variables.

Materials

- lab apron
safety glasses
voltmeter and connecting wires
U-tube with cotton plugs, porous cups, or filter-paper strips
four 100-mL beakers or well plate
distilled water
steel wool or emery paper
 $\text{Cu}_{(s)}$, $\text{Pb}_{(s)}$, $\text{Ag}_{(s)}$, and $\text{Zn}_{(s)}$ strips
1.0 mol/L $\text{CuSO}_4_{(\text{aq})}$, $\text{Pb}(\text{NO}_3)_2_{(\text{aq})}$, $\text{AgNO}_3_{(\text{aq})}$, $\text{NaNO}_3_{(\text{aq})}$, and $\text{ZnSO}_4_{(\text{aq})}$



Some of the solutions used are toxic and/or irritants. Avoid skin and eye contact.

(continued)

Procedure

- (c) Based on the equipment supplied, write a specific procedure to collect the evidence to answer the question. Be sure to include safety and disposal instructions. Have your teacher approve your procedure before starting.

Evidence/Analysis

- (d) Prepare a table to record your observations and the predicted cell potentials. Include a column for expressing the accuracy of each result (in terms of a percentage difference).
- (e) Note and record any unexpected observations.

Evaluation

- (f) List all sources of experimental error or uncertainty. Considering this list, state your judgment of the overall quality of the evidence obtained.
- (g) For each cell, compare the electrodes you predicted to be the cathodes and the anodes with your evidence. How well do these agree?
- (h) Limitations of the equipment and materials mean that some experimental uncertainties are unavoidable. Assuming that about 5% difference is unavoidable, is the agreement between your predicted and measured cell potentials acceptable? Justify your answer.

(continued)

LSM 9.LA-9

- (i) Is there any pattern to the accuracy of your measured cell potentials? Suggest some reasons to explain this.

 - (j) Evaluate the design, procedure, and materials. Note any flaws or possible improvements.

Investigation 9.6.1: The Corrosion of Iron

INQUIRY SKILLS

- | | |
|---|--|
| <ul style="list-style-type: none"><input type="radio"/> Questioning<input type="radio"/> Hypothesizing<input checked="" type="radio"/> Predicting<input checked="" type="radio"/> Planning<input checked="" type="radio"/> Conducting | <ul style="list-style-type: none"><input checked="" type="radio"/> Recording<input checked="" type="radio"/> Analyzing<input checked="" type="radio"/> Evaluating<input checked="" type="radio"/> Communicating |
|---|--|

The knowledge gained from this experiment is used to help explain corrosion and to develop methods of corrosion prevention.

Purpose

The purpose of this investigation is to test your predictions of factors affecting the rate of corrosion of iron.

Question

What factors, chemical and electrical, affect the rate of corrosion of iron?

Prediction

- (a) Based on your experience and knowledge of electrochemistry, predict some factors that may affect the rate of corrosion of iron. Provide your reasoning.

Experimental Design

Several iron nails or pieces of iron wire are thoroughly cleaned with steel wool, rinsed with water, and dried with a solvent (alcohol or acetone). In part 1, the iron is exposed to different conditions in separate test tubes. A clean piece of iron in a dry empty test tube is the control. All test tubes are observed immediately and after one day. In part 2, two iron-carbon cells are connected to 9-V batteries, with the electrodes attached oppositely for the two cells.

Materials

- (b) Prepare a list of materials. Check with your teacher to make sure your materials are suitable and available.

(continued)



The solvent used is very flammable and may be toxic.

Procedure

- (c) Using the experimental design and your list of materials, write a procedure to answer the question. Include safety and disposal instructions. Obtain approval from your teacher before proceeding.

Evidence

- (d) Prepare a table to record your observations.

(continued)

Analysis

- (e) Compare the evidence from the factors you tested with the control. Which factors appear to accelerate the rate of corrosion of iron? To the extent possible, arrange your factors from least to greatest apparent effect.

- (f) List any factors that did not seem to affect the rate of corrosion of iron compared to the control.

Evaluation

- (g) Evaluate the design, noting any flaws or improvements that could be made. Check the variables and control to see if these are all appropriate. Suggest changes if necessary.

- (h) Evaluate the procedure and materials. Were you able to gather sufficient and suitable evidence? Suggest improvements where required.

- (i) What are the main sources of experimental error or uncertainty? How serious would you judge these to be?

- (j) Using your answers to (g) to (i), state your judgment of the quality of the evidence. How confident are you about the evidence?

- (k) Assuming your evidence is of reasonable quality, state your judgment of the prediction.

- (l) Judge the reasoning you used to make your prediction. How useful was it? Why is there a need for more empirical and theoretical knowledge about corrosion?

Stoichiometry of Cell Reactions, Extra Exercises

1. A student wishes to set up an electrolytic cell to plate copper onto a belt buckle. Predict the length of time it will take to plate out 2.5 g of copper from a copper(II) nitrate solution using a 2.5-A current. At which electrode should the buckle be attached?
 2. Determine the mass of chlorine produced when a 200-A current flows for 24.0 h through a cell containing molten sodium chloride (a Downs cell). At which electrode is the chlorine produced?

(continued)

LSM 10.3-3

3. How long would it take a 500-A current to produce 1.00 kg of aluminum from aluminum oxide dissolved in molten cryolite (Hall-Héroult cell)?
 4. A trophy company is setting up a nickel-plating cell using an electrolyte containing nickel(II) ions. Predict the current required to produce nickel metal at the rate of 5.00 g/min.

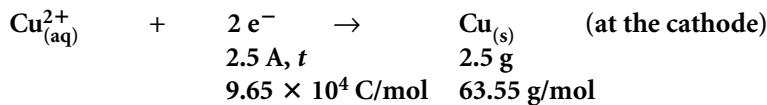
(continued)

LSM 10.3-3

5. In an electrolysis experiment, a student passed 1.57 A of current through an aqueous solution of lead(II) nitrate containing a 7.63-g piece of lead foil as the cathode. What is the final mass of the lead electrode after 17.0 min?

Stoichiometry of Cell Reactions, Extra Exercises, Solution

1. A student wishes to set up an electrolytic cell to plate copper onto a belt buckle. Predict the length of time it will take to plate out 2.5 g of copper from a copper(II) nitrate solution using a 2.5-A current. At which electrode should the buckle be attached?



$$n_{\text{Cu}} = 2.5 \text{ g} \times \frac{1 \text{ mol}}{63.55 \text{ g}}$$

$$n_{\text{Cu}} = 0.039 \text{ mol}$$

$$n_{\text{e}^-} = 0.039 \text{ mol Cu} \times \frac{2 \text{ mol e}^-}{1 \text{ mol Cu}}$$

$$n_{\text{e}^-} = 0.079 \text{ mol}$$

$$n_{\text{e}^-} = \frac{It}{F}$$

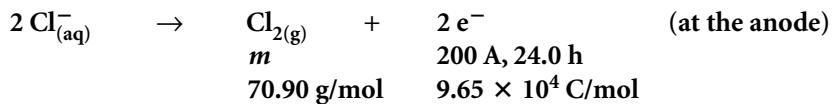
$$t = \frac{nF}{I}$$

$$= \frac{0.079 \text{ mol} \times 9.65 \times 10^4 \frac{\text{C}}{\text{mol}}}{2.5 \frac{\text{C}}{\text{s}}}$$

$$t = 3.0 \times 10^3 \text{ s} \quad (\text{or } 51 \text{ min})$$

The time required is 3.0 ks (or 51 min). The buckle should be attached to the cathode.

2. Determine the mass of chlorine produced when a 200-A current flows for 24.0 h through a cell containing molten sodium chloride (a Downs cell). At which electrode is the chlorine produced?



$$n_{\text{e}^-} = \frac{It}{F}$$

$$= \frac{200 \frac{\text{C}}{\text{s}} \times 24.0 \text{ h} \times 3600 \frac{\text{s}}{\text{h}}}{9.65 \times 10^4 \frac{\text{C}}{\text{mol}}}$$

$$n_{\text{e}^-} = 179 \text{ mol}$$

$$n_{\text{Cl}_2} = 179 \text{ mol e}^- \times \frac{1 \text{ mol Cl}_2}{2 \text{ mol e}^-}$$

$$n_{\text{Cl}_2} = 89.5 \text{ mol}$$

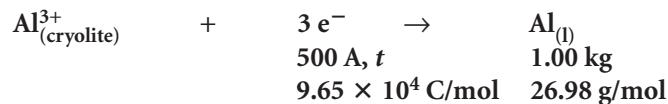
$$m_{\text{Cl}_2} = 89.5 \text{ mol} \times \frac{70.90 \text{ g}}{1 \text{ mol}}$$

$$m_{\text{Cl}_2} = 6.35 \text{ kg}$$

The mass of chlorine produced is 6.35 kg at the anode of the cell.

(continued)

3. How long would it take a 500-A current to produce 1.00 kg of aluminum from aluminum oxide dissolved in molten cryolite (Hall-Héroult cell)?



$$n_{\text{Al}} = 1.00 \text{ kg} \times \frac{1 \text{ mol}}{26.98 \text{ g}}$$

$$n_{\text{Al}} = 0.0371 \text{ kmol}$$

$$n_{\text{e}^-} = 0.03371 \text{ kmol Al} \times \frac{3 \text{ mol e}^-}{1 \text{ mol Al}}$$

$$n_{\text{e}^-} = 0.111 \text{ kmol}$$

$$n_{\text{e}^-} = \frac{It}{F}$$

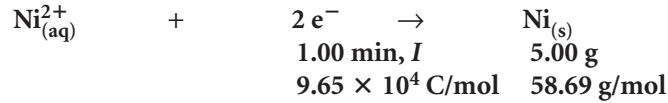
$$t = \frac{nF}{I}$$

$$= \frac{0.111 \text{ kmol} \times 9.65 \times 10^4 \frac{\text{C}}{\text{mol}}}{500 \frac{\text{C}}{\text{s}}}$$

$$t = 21.5 \text{ ks} \quad (\text{or } 358 \text{ min})$$

The time required to produce the aluminum is 21.5 ks (or 358 min).

4. A trophy company is setting up a nickel-plating cell using an electrolyte containing nickel(II) ions. Predict the current required to produce nickel metal at the rate of 5.00 g/min.



$$n_{\text{Ni}} = 5.00 \text{ g} \times \frac{1 \text{ mol}}{58.69 \text{ g}}$$

$$n_{\text{Ni}} = 0.0852 \text{ mol}$$

$$n_{\text{e}^-} = 0.0852 \text{ mol Ni} \times \frac{2 \text{ mol e}^-}{1 \text{ mol Ni}}$$

$$n_{\text{e}^-} = 0.170 \text{ mol}$$

$$n_{\text{e}^-} = \frac{It}{F}$$

$$t = \frac{nF}{I}$$

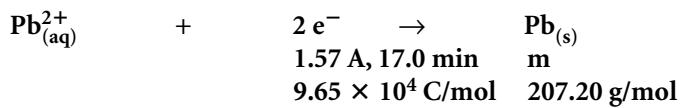
$$= \frac{0.170 \text{ mol} \times 9.65 \times 10^4 \frac{\text{C}}{\text{mol}}}{1.00 \text{ min} \times 60 \frac{\text{s}}{\text{min}}}$$

$$I = 274 \text{ A}$$

The current required is 274 A.

(continued)

5. In an electrolysis experiment, a student passed 1.57 A of current through an aqueous solution of lead(II) nitrate containing a 7.63-g piece of lead foil as the cathode. What is the final mass of the lead electrode after 17.0 min?



$$\begin{aligned} n_{e^-} &= \frac{It}{F} \\ &= \frac{1.57 \frac{\text{C}}{\text{s}} \times 17.0 \text{ min} \times 60 \frac{\text{s}}{\text{min}}}{9.65 \times 10^4 \frac{\text{C}}{\text{mol}}} \end{aligned}$$

$$n_{e^-} = 0.0166 \text{ mol}$$

$$n_{\text{Pb}} = 0.0166 \cancel{\text{mol}} e^- \times \frac{1 \text{ mol Pb}}{2 \cancel{\text{mol}} e^-}$$

$$n_{\text{Pb}} = 0.00830 \text{ mol}$$

$$m_{\text{Pb}} = 0.00830 \cancel{\text{mol}} \times \frac{207.20 \text{ g}}{1 \cancel{\text{mol}}}$$

$$m_{\text{Pb}} = 1.72 \text{ g}$$

$$\text{final } m_{\text{Pb}} = 7.63 \text{ g} + 1.72 \text{ g}$$

$$m_{\text{Pb}} = 9.35 \text{ g}$$

The final mass of the lead electrode is 9.35 g.

Investigation 10.1.1: A Potassium Iodide Electrolytic Cell

INQUIRY SKILLS

- | | |
|-----------------|-----------------|
| ○ Questioning | ● Recording |
| ○ Hypothesizing | ● Analyzing |
| ○ Predicting | ● Evaluating |
| ○ Planning | ● Communicating |
| ● Conducting | |

In this investigation, you will first observe a simple cell without an external battery or power supply and then compare this with the observations when a battery is connected. In this way, you will see firsthand the operation of an electrolytic cell.

Purpose

The purpose of this investigation is to observe the operation of an electrolytic cell and to determine its reaction products.

Question

What are the products of the reaction during the operation of an aqueous potassium iodide electrolytic cell?

Experimental Design

Inert electrodes are placed in a 0.50 mol/L solution of potassium iodide and connected directly with a wire. Then a battery or power supply is added to the circuit to provide a direct current of electricity to the cell. The litmus and halogen diagnostic tests (Appendix A6) are conducted to test the solution near each electrode. The litmus and halogen tests before the battery or power supply is added serve as a control for the same tests done after electric power is supplied.

Materials

- lab apron
- eye protection
- U-tube
- 2 clean carbon electrodes
- two connecting wires
- 3-V to 9-V battery or power supply
- red and blue litmus paper
- ring stand and utility clamp
- small test tube with stopper
- plastic pipet with long tip
- dropper bottle of cyclohexane
- bottle of distilled water
- 0.50 mol/L KI_(aq)



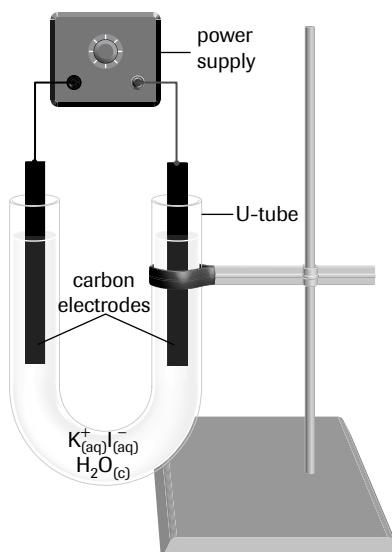
Cyclohexane is highly flammable. Do not use near an open flame. Avoid inhaling fumes of cyclohexane.

Dispose of the solutions as directed by your teacher.

Procedure

1. Set up the KI_(aq) cell as shown in **Figure 1**, but with a single wire connecting the electrodes (i.e., no power supply).
2. Record observations of the cell.
3. Use the medicine dropper to remove some solution near each electrode in the cell.
4. Test each solution using the litmus and halogen tests.

(continued)

**Figure 1**

A U-tube is a convenient container for the aqueous potassium iodide solution because the inert, carbon electrodes can be separated by a reasonable distance.

5. Using two connecting wires, hook up the power supply (**Figure 1**).
6. Turn on the power supply and record observations made while the cell is operating.
7. Perform both diagnostic tests at each electrode by repeating steps 3 and 4.
8. Deposit any hexane mixtures into the labelled disposal container.

Evidence

- (a) On a separate piece of paper, create a convenient table to compare diagnostic test results. In your table, identify the electrodes according to the terminal (red—positive, black—negative) on the power supply. Be sure to record any other general observations of the cell.

Analysis

- (b) Interpret your evidence and answer the question as well as you can with the evidence collected.

Evaluation

- (c) Evaluate the design. For example, were the control tests sufficient to show clearly whether changes occurred?
- (d) Which product were you not able to identify? Why not?
- (e) Suggest some improvements to the materials and to the procedure used to collect and identify the gas produced.
- (f) How certain are you about the other two products? Justify your answer.
- (g) Why is it necessary to set up the apparatus so that the electrodes are not touching?
- (h) Explain the observations at the bottom of the U-tube.
- (i) Overall, how would you judge the quality of the evidence obtained? Provide reasons.

Investigation 10.1.2: Investigating Several Electrolytic Cells

INQUIRY SKILLS

- | | |
|---|--|
| <input type="radio"/> Questioning | <input checked="" type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input checked="" type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input type="radio"/> Conducting | |

The purpose of this demonstration is to evaluate the method of predicting the products of a reaction occurring in an electrolytic cell.

Questions

What are the products obtained when an electrolytic cell is made by immersing inert electrodes in

- aqueous copper(II) sulfate?
- aqueous sodium sulfate?
- aqueous sodium chloride?

Prediction

- (a) According to redox concepts and the table of reduction potentials, predict the products of the reaction occurring in each electrolytic cell. For your reasoning, show the equations for the cathode and anode half-reactions and the net cell reaction. Calculate the minimum potential difference that must be applied in each case.

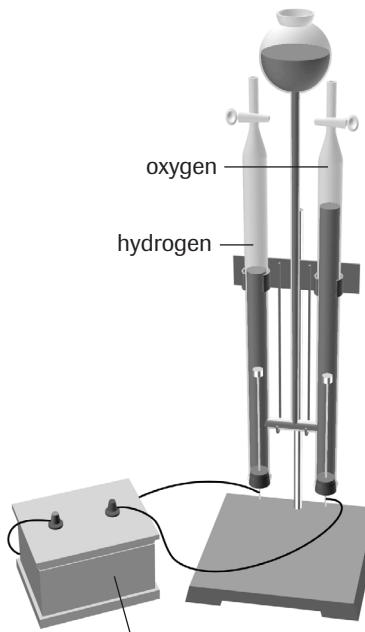


Figure 2

A Hoffman apparatus is an electrolytic cell that is very useful when doing electrolysis reactions that produce gases. The gases rise in a graduated tube (like a burette) and displace the solution back into the bulb (in the middle). The gases can easily be removed by opening the stopcock at the top of each tube.

Experimental Design

Electrolysis of aqueous copper(II) sulfate is carried out in a U-tube with carbon electrodes. Electrolysis of aqueous sodium sulfate and sodium chloride is carried out with platinum electrodes in a Hoffman apparatus (**Figure 2**) so the gases can be collected. Diagnostic tests, with any necessary control tests, are conducted to establish whether any of the predicted products are obtained.

(continued)



Copper(II) sulfate is toxic and and an irritant. Avoid skin and eye contact. If you spill copper(II) sulfate solution on your skin, wash the affected area with lots of cool water.



During electrolysis, corrosive substances are produced; avoid skin and eye contact.

Remember to wash your hands when you have finished this investigation.

Evidence

- (b) Set up a table that includes the titles: Cell, Cathode, Anode. Record the cell notation and applied voltage under Cell. Record the observations for cathode and anode half-cells, including observations made during the control tests.

Analysis

- (c) Based on your interpretation of the evidence collected, prepare a table or list of the products at each electrode for each of the three cells.

Evaluation

- (d) Evaluate the experimental design. Consider whether the question was answered, any flaws were present, and the controls were adequate. Suggest some improvements.
- (e) What is your judgment of the overall quality of the evidence? How certain are you about this? Provide reasons.

(continued)

- (f) Evaluate each of the three predictions, considering both cathode and anode products.
- (g) Overall, how would you judge the redox concepts and the table of reduction potentials used to make these predictions? Provide reasons.

Synthesis

- (h) Which one of the half-reactions that was expected to occur did not occur? Using the table of reduction potentials, write the equations for the cathode and anode half-reactions and the net reaction to obtain the observed products. Calculate the minimum potential difference required to force this reaction to occur.
- (i) Compare the minimum potential differences for oxygen as a product (see Prediction) and chlorine as a product.
- (j) Suggest one reason why it was possible to produce chlorine.
- (k) Considering that one half-reaction out of six did not agree with the prediction, should our rules be restricted, revised, or replaced? Discuss briefly.

Investigation 10.3.1: Investigating an Electrolytic Cell

INQUIRY SKILLS

- | | |
|---|--|
| <input type="radio"/> Questioning | <input checked="" type="radio"/> Recording |
| <input type="radio"/> Hypothesizing | <input checked="" type="radio"/> Analyzing |
| <input type="radio"/> Predicting | <input checked="" type="radio"/> Evaluating |
| <input type="radio"/> Planning | <input checked="" type="radio"/> Communicating |
| <input checked="" type="radio"/> Conducting | |

Suppose we design an experiment to determine the value of a scientific constant such as the Faraday. Since this constant is well known and reliable, we can use the accuracy of our experimentally determined constant to evaluate the design of our experiment or our understanding of the processes occurring.

Purpose

To test the reactions in an electrolytic cell by determining the value of a known constant.

Question

What is the value of the Faraday in the electrolysis of copper(II) sulfate solution?

Prediction

- (a) According to a modern reference, what is the value of the Faraday?

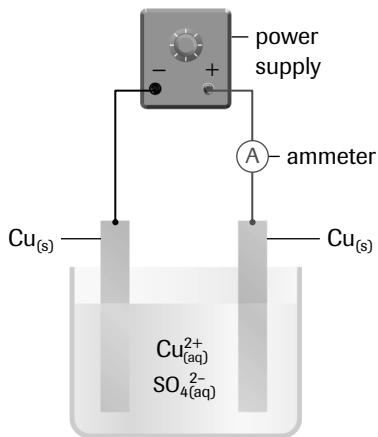


Figure 3



**Copper(II) sulfate is toxic and
an irritant. Avoid skin and
eye contact. If you spill
copper(II) sulfate solution on
your skin, wash the affected
area with lots of cool water.**

Experimental Design

An electrolytic cell with copper electrodes and copper(II) sulfate solution is set up as shown in **Figure 3**. Measurements of the change in mass of the electrodes, current, and time are used to determine the Faraday constant. An important controlled variable is the electric current.

Materials

- | | |
|--|------------------------------|
| lab apron | variable power supply |
| eye protection | stopwatch |
| two 250-mL beakers | cardboard holder |
| 2 strips of copper foil (8 cm by 3 cm) | centigram balance |
| steel wool or emery paper | bottle of distilled water |
| connecting wires | 0.5 mol/L copper(II) sulfate |
| ammeter (0–5 A) | |

(continued)

Procedure

1. Carefully clean the copper strips with steel wool, rinse with water, and dry. Once they are cleaned and dry, do not touch the surface of the strips with your fingers, except at the top.
2. Label each copper strip and then record the mass of each strip.
3. Fill a beaker close to the top with copper(II) sulfate solution.
4. Place the copper strips into the cardboard holder and set this into the beaker of solution.
5. Assemble the rest of the apparatus as shown in **Figure 3**. Note which copper strip is the cathode and which is the anode.
6. Set the variable power supply voltage to about the halfway position.
7. Plug in the power supply and reset the stopwatch to zero.
8. Turn on the power supply and start the stopwatch at the same time.
9. Immediately after the power is on, check the ammeter and adjust the power supply voltage to set the current between one half and two amperes. If necessary, adjust the power supply voltage to keep the current constant.
10. Record the current used.
11. Record observations of the contents of the cell.
12. Let the cell run for at least 25 min, longer if possible.
13. Turn off the power supply and stopwatch at the same time. Record the time.
14. Fill the other beaker with distilled water.
15. Carefully remove the cardboard holder with the copper electrodes and dip the electrodes into the distilled water to remove any $\text{CuSO}_4\text{(aq)}$. Be careful not to lose any deposit.
16. Let the electrodes dry completely. (An alcohol or acetone rinse may be available.)
17. Measure the mass of each copper electrode. Wait 5 min and measure the mass again.
18. Recycle the copper(II) sulfate solution to the appropriate container.

Analysis

- (b) Calculate the change in mass of each electrode.

(continued)

- (c) Using the average of the two masses determined in (b), determine the amount, in moles, of electrons.

- (d) Calculate the charge in coulombs that passed through the cell, using the current and the time.

- (e) Using your results from (c) and (d), answer the Question.

Evaluation

- (f) Evaluate the evidence collected in this experiment. Consider the design, materials, procedure, and skills. Note any flaws and improvements.

- (g) What is your judgment of the quality of the evidence? How certain are you about this?

- (h) Calculate the accuracy of your experimental result.

- (i) Assume that the evidence was of reasonable quality. What does the accuracy tell you about the processes occurring within the cell? Justify your answer, using equations for the half-reactions and any other observations you have.

(continued)

Synthesis

- (j) Describe a technological application of the cell used in this investigation.