Chapter 1: Chemistry

Overview

Chapter 1 introduces the student to the concepts of science, the ability of science to solve problems as well as its limitations. The student is then introduced to the concept of matter and energy and the physical means by which they are measured.

Lecture Outline

1.1 Science and Technology: The Roots of Knowledge

Many students are confused about the differences between science and technology. Technology is the direct application of knowledge to solve problems. Science seeks an understanding of underlying principles.

1.2 Science: Reproducible, Testable, Tentative, Predictive, and Explanatory

The scientific method, from hypothesis to law, theory, or model, is best explained by example.

1.3 Science and Technology: Risks and Benefits

1.4 Solving Society's Problems: Scientific Research

Research can be basic or applied.

1.5 Chemistry: A Study of Matter and Its Changes

Matter is anything that has mass.
Mass is a measure of the amount of matter present.
Weight is mass times gravitational attraction.
Matter exhibits chemical and physical properties.
Chemical properties tell us how matter will combine to form new and different substances (a chemical change).
Physical properties are directly observable; color, state (solid, liquid, or gas), and texture are examples.
An excellent demonstration is to burn a candle (chemical change) and boil water (physical change).

1.6 Classification of Matter

A. States of Matter

- Solids, liquids, and gases.

B. Substances and Mixtures

- Substances have constant composition.
- The composition of a mixture is variable.
  - Homogeneous mixtures: appear the same throughout (milk, paint, and saltwater are examples).
  - Heterogeneous mixtures: appear different throughout (pizza, raisin bread, and chocolate chip cookies are examples).

C. Elements and Compounds

- Substances are either elements or compounds.
  - Elements: fundamental building blocks of all matter. (lead, silver, gold, carbon, and oxygen).
  - Compounds: two or more elements chemically combined in fixed ratios (water, ammonia, and propane).

D. Atoms and Molecules

1.7 The Measurement of Matter

Most students have been introduced to the metric system during their K-12 education; the SI system used in science is based on the metric system. The difference lies in the base units.

- Base units: kilograms (kg) for mass, meter (m) for length, and seconds (s) for time. The four other base units are shown in Table 1.4.

1.8 Density

Density is the mass-to-volume ratio of matter.

1.9 Energy: Heat and Temperature

The SI unit of temperature is the kelvin (K); however, the temperature scale used in the chemistry laboratory is normally the Celsius scale. The Celsius scale was developed with the freezing and boiling points of pure water at a pressure of 1 atmosphere as the reference frames. (Students find the history of the development of the Fahrenheit scale much more interesting!)

1.10 Critical Thinking

Demonstrations

1. Place samples of various elements (copper, sulfur, zinc, mercury, aluminum, carbon, etc.) in small stoppered bottles or flasks that can be passed around.

2. Compare a yardstick and a meter stick. A meter is slightly bigger than a yard. A sugar cube is approximately 1 mL. An ordinary paper clip weighs about 1 gram. A liter is a little bit (6%) larger than a quart.

3. Use Crispix and Raisin Bran to illustrate the difference between a compound and a mixture. Raisins and bran flakes are easily separated, and the ratio of raisins to flakes can vary. Crispix has a 1:1 ratio of corn and rice flakes, and they cannot be easily separated.

4. Ice floats on water because the density of ice is less than that of water. But alcohol has a much lower density. If you use alcohol instead of water, the ice will sink. Place a glass of water and a glass of alcohol side by side and add an ice cube to each one.

5. The “cartesian diver” is a popular demonstration about density. Completely fill a 2-liter plastic bottle with water and put just enough water into a medicine dropper so that it can barely float. Put the dropper into the bottle and cap the bottle. Squeeze and release the bottle to make the “diver” go up and down.

Review Questions

1. Science is testable, reproducible, explanatory, predictive, and tentative.

2. A hypothesis is a tentative explanation and must be verified or rejected through experiment.

3. The scientific method can only be used when all the variables in a system can be controlled.

4. Technology is the sum total of the processes by which humans modify the materials of nature; technology need not be based on science.

5. Risk-benefit analysis is an analysis of benefits versus risks, involving an attempt to calculate a desirability quotient (DQ).

6. (a) A benefit is anything that promotes well-being or has a positive effect. (b) A risk is any hazard that leads to loss or injury. A risk-benefit analysis is weighing both the risk and the benefit of a certain technology and determining if the hazard is worth the positive outcome. This tends to be a personal decision and can vary from group to group and person to person.

7. A DQ is benefits divided by risk, also known as a desirability quotient. A large DQ is the result of large benefits and small risks. While scientific investigation can help considerably in determining an accurate DQ, the calculation of benefits is almost entirely a social judgment and the risks and benefits are not always known. Benefits and risks are therefore judgments with a qualitative rather than quantitative aspect.

8. The common units used in the laboratory are (a) the gram (g) for mass and (b) the centimeter (cm) or millimeter (mm) for length.

9. The SI unit of volume is the cubic meter (m^3). The volume unit used in the laboratory is the cubic decimeter (dm^3) or the cubic centimeter (cm^3 or cc).

10.

Prefix	Symbol	Definition
Tera	T	10^{12}
Mega	M	10^{6}
Centi	c	10^{-2}
Micro	μ	10^{-6}
Milli	m	10^{-3}
Deci	d	10^{-1}
Kilo	k	10^{3}
Nano	n	10^{-9}

11. Research projects “a” and “b” are applied research. Research project “c” is basic research.

12. Research project “a” is applied research. Research projects “b” and “c” are basic research.

Problems

13. Penicillin has saved thousands of lives, causing harm to a very few. The use of penicillin for society as a whole has been very beneficial; it has a large DQ.

14. The risk of using sodium sulfite is greater in fruit juices than in wines as small children may be sensitive to sulfites and are more likely to consume fruit juice.

15. Hazards are greater than benefits for a person who is exposed daily to the paints than the hobbyist exposed once. The DQ would be increased for the professional painter and the hobbyist if both wore breathing masks capable of filtering out the *isocyanate*.

16. The risk is much higher for the dentist who stays in the room during every X-ray than for the patient receiving one X-ray per year.

17. The DQ for the use of antibiotics in treating common sore throats is low as there is a small risk to the untreated patient. The DQ for treating more serous influenza is higher as the risk to the untreated patient is greater.

18. There is no benefit for the man or the pregnant woman. There is a significant benefit to the unborn child whose mother has AIDS.

19. 100 g, 2 kg

20. A2 sized paper is larger. 1" = 25.4 mm thus:

$$594\text{ mm} \times \frac{1''}{25.4\text{ mm}} = 23.4''$$

$$420\text{ mm} \times \frac{1''}{25.4\text{ mm}} = 16.5''$$

21. Both a and b are correct. (c) would be more in the range of 100 kg and (d) would be in the range of 10 kg.

22. No. Weight is mass times the force of gravity! The sample weighed on the moon would have 1/6 th the weight of the sample weighed on Earth.

23. 250 mL

24. 40 cm^2

25. Use the following conversions: 1 km = 0.62 mi, $(1\text{ km})^3 = (0.62\text{ mi})^3 = 0.24\text{ mi}^3$

$$3.5\times10^8\text{ mi}^3 \times \frac{1\text{ km}^3}{0.24\text{ mi}^3} = 1.5\times10^9\text{ km}^3$$

26. Use the following conversions: 1 km = 0.62 mi, $(1\text{ km})^2 = (0.62\text{ mi})^2 = 0.38\text{ mi}^2$

$$1.4 \times 10^8\text{ mi}^2 \times \frac{1\text{ km}^2}{0.38\text{ mi}^2} = 3.7 \times 10^8\text{ km}^2$$

27. Use the following conversion: 1" = 25.4 mm, The aluminum tube will fit.

$$26.3\text{ mm} \times \frac{1''}{25.4\text{ mm}} = 1.03''$$

28. 18.7 mm

29. (a) Physical (b) Chemical (c) Physical (d) Chemical

30. (a) Physical (b) Physical (c) Chemical (d) Chemical

31. (a) Physical (b) Chemical (c) Physical

32. (a) Chemical (b) Physical (c) Physical

33. Substances: a, c, d; mixtures: b

34. Substances: a, b; mixtures: c, d, e, f

35. Homogeneous: a, b, c; heterogeneous: d

36. Homogeneous: a, d; heterogeneous: b, c

37. A substance has a definite, or fixed, composition that does not vary from one sample to another. Glucose samples, regardless of where they are collected, are 8 parts oxygen, 6 parts carbon, and 1 part hydrogen. Therefore, glucose is a substance.

38. Shampoo contains many substances such as water and soap; therefore it is a mixture. "Nothing artificial" reflects the sources of the compounds in the shampoo.

39. Elements: a, b; compounds: c, d

40. Elements: a, c; compounds: b, d

41. (a) Iron (b) Magnesium (c) Helium (d) Nitrogen

42. (a) O (b) P (c) K (d) Ar

43. f

44. d

45. (a) 1.0 dL (b) 8.5 pg (c) 1.05 mm

46. (a) 45 mg = 0.000045 kg = 4.5×10^{-5} kg
(b) 125 ns = 0.000000125 s = 1.25×10^{-7} s
(c) 10.7 µL = 0.0000107 L = 1.07×10^{-5} L
(d) 12.5346 kg

47. (a) 1000 mL = 1 L, $5.52 \times 10^4 \text{ mL} \times \dfrac{1 \text{ L}}{1000 \text{ mL}} = 552 \text{ L}$

(b) 1000 mg = 1 g $325 \text{ mg} \times \dfrac{1 \text{ g}}{1000 \text{ mg}} = 0.325 \text{ g}$

(c) 100 cm = 1 m, $27 \text{ cm} \times \dfrac{1 \text{ m}}{100 \text{ cm}} = 0.27 \text{ m}$

(d) 10 mm = 1 cm, $27 \text{ mm} \times \dfrac{1 \text{ cm}}{10 \text{ mm}} = 2.7 \text{ cm}$

(e) 1 ms = 1000 μs, $78 \,\mu\text{s} \times \dfrac{1 \text{ ms}}{1000 \,\mu\text{s}} = 0.078 \text{ ms}$

48. (a) 1000 mm = 1 m, $546 \text{ mm} \times \dfrac{1 \text{ m}}{1000 \text{ mm}} = 0.546 \text{ m}$

(b) 1000 ns = 1 μs , $65 \text{ ns} \times \dfrac{1 \,\mu\text{s}}{1000 \text{ ns}} = 0.065 \,\mu\text{s}$

(c) 1000 mg = 1 g, 1000 g = 1 k,

$87.6 \text{ mg} \times \dfrac{1 \text{ kg}}{1 \text{ x } 10^6 \text{ mg}} = 0.0000876 \text{ mg} = 8.76 \times 10^{-5} \text{kg}$

(d) $1 \text{ dm}^3 = 1\text{L}$, $46.3 \text{ dm}^3 \times \dfrac{1 \text{ L}}{1 \text{ dm}^3} = 46.3 \text{ L}$

(e) 1×10^6 pm = 1μm, $181 \text{ pm} \times \dfrac{1 \,\mu\text{m}}{1 \text{ x } 10^6 \text{ pm}} = 0.000181 \,\mu\text{m} = 1.81 \times 10^{-4} \mu\text{m}$

49. Larger units: (a) cm - centimeters (b) kg - kilograms (c) dL - deciliters

50. 7 x 10^{27} atoms

51. 31 cm = 310 mm = 0.31 m = 3.1 x 10^8 nm

52. (a) 352 mL = 0.352 L (b) 26 kL = 26,000 L

53. $D = \dfrac{m}{V}$; $D = \dfrac{500.0 \text{ g}}{53.6 \text{ cm}^3} = 9.33 \text{ g/cm}^3$

54. $D = \dfrac{m}{V}$; $D = \dfrac{22.4 \text{ g}}{25.0 \text{ mL}} = 0.896 \text{ g/mL}$

55. $D = \dfrac{m}{V}$; $D \times V = m$; $1.51 \text{ g/mL} \times 49.1 \text{ mL} = 74.1 \text{ g}$

56. $D = \dfrac{m}{V}$; $D \times V = m$; $1.43 \text{ g/cm}^3 \times 13.2 \text{ cm}^3 = 18.9 \text{ g}$

57. $D = \dfrac{m}{V}; V = \dfrac{m}{D}$

(a) $V = \dfrac{227\text{ g}}{0.660\ {}^{g}/_{mL}} = 343\text{ mL}$

(b) $V = \dfrac{454\text{ g}}{0.917\ {}^{g}/_{cm^3}} = 495\text{ cm}^3$

Note: Refer to Table 1.6 for the density data.

58. $D = \dfrac{m}{V},\ V = \dfrac{m}{D}$

(a) $V = \dfrac{475\text{g}}{8.94\ {}^{g}/_{cm^3}} = 53.1\text{ cm}^3$

(b) $V = \dfrac{253\text{ g}}{13.534\ {}^{g}/_{mL}} = 18.7\text{ mL}$

59. Three layers will form with mercury on the bottom (greatest density), then water, followed by hexane on the top.

60. The red maple and the balsa wood will float on the hexane, the ice will float on the water, the copper coin will float on the mercury, and the gold coin will sink to the bottom.

61. $D = \dfrac{m}{V}, \qquad D \times V = m$

$1.03\ {}^{g}/_{mL} \times 37900\text{ mL} = 39037\text{ g} \times \dfrac{1\text{ lb}}{453.6\text{ g}} = 86.1\text{ lbs}$

$86.1\text{ lbs} + 59.5\text{ lbs} = 145.6\text{ lbs}$

The answer is "yes."

62. $D = \dfrac{m}{V}, \qquad D \times V = m$

$14.0\text{ gal x } \dfrac{3.785\text{ L}}{1\text{ gal}} = 53.0\text{L x } \dfrac{1000\text{ mL}}{1\text{ L}} = 53{,}000\text{ mL}$

$0.758\ {}^{g}/_{mL} \times 53000\text{ mL} = 40200\text{ g} \times \dfrac{1\text{ kg}}{1000\text{ g}} = 40.2\text{ kg}$

63. Solve the equation for the volume of the crystal ball. The diameter of the crystal ball is 14.8 cm, thus the radius (r) is 7.4 cm.

$$V = \frac{4\pi r^3}{3} = \frac{4\pi \text{ x } (7.4 \text{ cm})^3}{3} = 1.70 \text{ x } 10^3 \text{ cm}^3$$

$$D \times V = m$$

$$3.18 \ {}^{g}\!/_{cm^3} \times 1.70 \text{ x } 10^3 \text{ cm}^3 = 5.41 \text{ x } 10^3 \text{ g}$$

64. Solve for the total volume required. $D = \frac{m}{V}$, $V = \frac{m}{D}$

$$V = \frac{500\text{g}}{1.031 \ {}^{g}\!/_{mL}} = 485 \text{ mL} = 485 \text{ cm}^3$$

Solve for height (h) of carton.

7.7 cm x 7.7 cm x h cm = 485 cm^3
59 cm^2 x h cm = 485 cm^3
h cm = 485 cm^3/ 59 cm^2 = 8.2 cm = 82 mm

65. ºC = K – 273, ºC = 77 – 273 = –196 ºC

66. K = ºC + 273, K = 37 + 273 = 310 K

67. 1 cal = 4.184 J, 1 kcal = 4.184 kJ, $161 \text{ kJ} \times \frac{1 \text{ kcal}}{4.184 \text{ kJ}} = 38.5 \text{ kcal}$

68. 1000 cal = 1 kcal, 584 cal = 0.584 kcal $0.584 \text{ kcal} \times \frac{4.184 \text{ kJ}}{1 \text{ kcal}} = 2.44 \text{ kJ}$

69. 1 microcentury = 1 x 10^{-6} centuries = 1x10^{-4} years x 365 days/year = 0.0365 days x 24 hours/day = 0.876 hours x 60 minutes/hour = 52.6 minutes.

70. 1 millihelen = 1.0 x 10^{-3} or 0.001 helens. 1000 ships/helen x 0.001 hellens = 1 ship

71. law or an empirical law

72. A food calorie is 1000 cal or 1 kcal. Thus it takes 79 kcal to melt 1 kg of ice.

73. 200 kilowarhols = 200,000 warhols x 15 min/warhol = 3,000,000 min/60 min/hr = 50,000 hr/24 hrs/day = 2083.3 days/365 days/year = 6 years

74. (1) An experiment (2) A scientific law (3) A hypothesis (4) A theory

75. (1) An observation (2) A theory (3) A hypothesis
(4) An experiment (5) An observation

76. Subtract the weight of the empty glass container from the weight of the glass container plus the weight of the antifreeze to yield the weight of the antifreeze; then calculate the density of the antifreeze,

60.562 g – 48.462 g = 12.100 g antifreeze

$$D = \frac{m}{V} = \frac{12.100\,\text{g}}{8.00\,\text{mL}} = 1.51\,\text{g}/\text{mL}$$

77. Calculate the volume of the block, then calculate the mass.
 7.6 cm x 7.6 cm x 94 cm = 5400 cm^3

$$D = \frac{m}{V}, \qquad D \times V = m$$

$$0.11\,\text{g}/\text{cm}^3 \times 5400\,\text{cm}^3 = 590\,\text{g}$$

78. yardstick (36 in. = 0.91 m), rattlesnake (1.04 m), 1.21 m chain, 75-in. board (1.9 m)

79. 1.65 kg cabbage, 5 lb bag of potatoes (2.3 kg), 2500 g sugar (2.500 kg)

80. 1.53 m = 5’, 38.5 kg = 104 lb; the woman on the left

81. Find the mass of the chips only by subtracting the paper mass from the total mass.
 18.43 g – 1.21 g = 17.22 g

 Calculate density using the equation: $D = \frac{m}{V}$

$$D = \frac{17.22\,\text{g}}{3.29\,\text{cm}^3} = 5.23\,\text{g}/\text{cm}^3$$

82. $D = \frac{m}{V}$, solve for volume.

$$V = \frac{m}{D},\ V = \frac{3180\text{ g}}{7.9\text{ g/cm}^3} = 4.0 \text{ x } 10^2\text{ cm}^3$$

83. Use $D = \frac{m}{V}$, solve for volume.

$$V = \frac{m}{D},\ V = \frac{0.00579\,\text{g}}{19.3\,\text{g/cm}^3} = 0.00030\text{ cm}^3$$

 Use length (cm) x width (cm) x thickness (cm) = volume (cm^3).
 44.6 cm^2 x thickness (cm) = 0.0003 cm^3

$$\text{Thickness (cm)} = \frac{0.0003\,\text{cm}^3}{44.6\,\text{cm}^2} = 6.73 \text{ x } 10^{-6}\text{ cm}$$

84. Calculate the volume of each stainless steel rod:
 π x (1.27 cm)^2 x 610 cm = 3090 cm^3

Calculate the mass of each steel rod:

$$DV = m = 7.48\ {}^{g}\!/_{cm^3}\ \text{x}\ 3090\ {}^{cm^3}\!/_{\text{steel rod}} = 23100\ {}^{g}\!/_{\text{steel rod}}\ \text{x}\ \frac{1\,kg}{1000\,g} = 23.1\ {}^{kg}\!/_{\text{steel rod}}$$

$$\frac{1850\,kg}{23.1\ {}^{kg}\!/_{\text{steel rod}}} = 80\ \text{steel rods}$$

85. V = 36.1 cm x 36.1 cm x 36.1 cm; V = 47,000 cm^3
 m = 19.3 g/cm^3 x 47,000 cm^3; *m* = 907,000 g
 907,000 g x 1 kg/1000 kg = 907 kg
 907 kg x 1 metric ton/1000 g = 0.907 metric tons

86. $D = \frac{m}{V} = \frac{425\,g}{48.0\,cm^3} = 8.85\ {}^{g}\!/_{cm^3}$ this is silver

87. Calculate the volume of each planet.

$$V_{Jupiter} = \frac{4\pi(70{,}000\,km)^3}{3} = 1.44x10^{15}\ km^3$$

$$V_{Earth} = \frac{4\pi(6400\,km)^3}{3} = 1.10x10^{12}\ km^3$$

$$V_{Saturn} = \frac{4\pi(58{,}200\,km)^3}{3} = 8.26x10^{14}\ km^3$$

Calculate the density of each planet.

$$D_{Jupiter} = \frac{1.9x10^{27}\ kg}{1.44x10^{15}\ km^3} = 1.32x10^{12}\ {}^{kg}\!/_{km^3}\ x1000\ {}^{g}\!/_{kg}\ x\ {}^{km^3}\!/_{1x10^{15}\,cm^3} = 1.3\ {}^{g}\!/_{cm^3}$$

$$D_{Earth} = \frac{5.98x10^{24}\ kg}{1.10x10^{12}\ km^3} = 5.44x10^{12}\ {}^{kg}\!/_{km^3}\ x1000\ {}^{g}\!/_{kg}\ x\ {}^{km^3}\!/_{1x10^{15}\,cm^3} = 5.4\ {}^{g}\!/_{cm^3}$$

$$D_{Saturn} = \frac{5.68x10^{26}\ kg}{8.26x10^{14}\ km^3} = 6.87x10^{11}\ {}^{kg}\!/_{km^3}\ x1000\ {}^{g}\!/_{kg}\ x\ {}^{km^3}\!/_{1x10^{15}\,cm^3} = 0.69\ {}^{g}\!/_{cm^3}$$

Saturn would float on water!

88. $V_{HAT\text{-}P\text{-}1} = \frac{4\pi(96{,}600\,km)^3}{3} = 3.78x10^{15}\ km^3$

$$D_{HAT\text{-}P\text{-}1} = \frac{9.50x10^{26}\ kg}{3.78x10^{15}\ km^3} = 2.51x10^{11}\ {}^{kg}\!/_{km^3}\ x1000\ {}^{g}\!/_{kg}\ x\ {}^{km^3}\!/_{1x10^{15}\,cm^3} = 0.251\ {}^{g}\!/_{cm^3}$$

89. Green chemistry uses materials and processes that are intended to prevent or reduce pollution at its source and to meet the needs of the present generation without compromising the needs of future generations.

90**.** The Twelve Principles of Green Chemistry
1) Prevention 2) Atom Economy 3) Less Hazardous Chemical Syntheses 4) Designing Safer Chemicals 5) Safer Solvents and Auxiliaries 6) Design for Energy Efficiency 7) Use of Renewable Feedstocks 8) Reduce Derivatives 9) Catalysis 10) Design for Degradation 11) Real-time Analysis for Pollution Prevention 12) Inherently Safer Chemistry for Accident Prevention

91. a

92. d

Lab Notes

Note: This manual presents a page of notes to the instructor for each lab, followed by a sample Pre-lab Question Page and Sample Report Pages with sample answers to all of the questions and most of the data. It should be noted that if a process other than what is described in the investigation is followed, the data acquired may differ.

#1 Alchemy

Reagents

zinc, powdered	1 g/pair
6 M sodium hydroxide [NaOH]	25 mL/pair
6 M hydrochloric acid [HCl]	25 mL/lab

Common Materials

copper token, wire, or piece of plate	1/student
rulers, metric	1/pair
spray can of clear acrylic coating (optional)	1/lab

Laboratory Equipment

balance
caliper
evaporating dish
forceps
hotplate
stirring rod

Special Equipment

(none)

Notes

Pre-1982 pennies are pure copper but are difficult to clean thoroughly enough to plate out well and produce a satisfactory result.

The best tokens are new and very shiny pennies. Any material on the surface will hinder the adherence of the zinc to the copper. Even slightly darkened areas of fingerprints are enough to alter the quality of the finished product. Likewise, a dirty or corroded forceps tip can quickly reverse the plating on the coin where it is grasped.

The sodium hydroxide solution can be less concentrated, but it will slow the reaction.

Caution the students to refrain from heating the zinc/copper surface too long as the coin will melt. Constantly turning the coin while heating is a safer method. Use of a Bunsen burner instead of a hot plate will turn the token to a new color faster, but students run the risk of melting the token. If the students will lower the coin into the flame for only a second on each side and repeat continuously until the gold color appears, then heat each side one more time, they will ensure a complete reaction without misshaping the coin.

Disposal

Wet zinc dust exposed to air can burst into flames. **It is important that students not put wet zinc dust in the trash; it can cause fires in the trash.** The wet zinc solutions should be spread on a metal pan to dry. It must be in a metal pan for the chemical reaction to form zinc oxide. Dry zinc oxide will form, which can be buried in a landfill. The used hydrochloric acid should be flushed down the drain with plenty of water.

ALCHEMY PRE-LAB QUESTIONS

1) Legend holds that Archimedes was responsible for proving that a metallurgist had "cut," or diluted, the gold for a Ruler's crown with a lesser metal. Archimedes asked the Ruler for a mass of gold identical to what he had provided to the metallurgist. What measurements and calculations did Archimedes most likely make?
Assuming that the metallurgist was not stupid and the finished crown had the same mass as the originally supplied gold, Archimedes most likely took two volume measurements by liquid displacement—one of the identical mass of gold and one of the crown. Then finding that they were not the same, he deduced that since the smaller volume was the pure gold, the crown was only partially gold and contained a less dense metal. The metallurgist had indeed kept some gold for himself.

2) The alloy produced in this investigation is brass. There are several different types of brass. Use an Internet encyclopedia resource to find at least three different types of brass. Describe the differences in proportion and variety of metals used, and the effect on the resulting properties of the brass.
Answers will vary but will most likely include some of the following:

- ***Admiralty** brass contains 30% zinc and 1% tin that will inhibit dezincification.*
- ***Alpha brasses,** also known as Prince's metal, with less than 35% zinc, are malleable, can be worked cold, and are used in pressing, forging, or similar processes.*
- ***Alpha-beta brass**, also known as **Muntz metal** and called **duplex brass**, is 35-45 % zinc and is suited for working hot.*
- ***Aluminum brass** contains aluminum, which improves its corrosion resistance.*
- ***Arsenical brass** contains an addition of arsenic and frequently aluminum and is used for boiler fireboxes.*
- ***Beta brasses** with 45-50 % zinc content can only be worked hot, are harder, stronger, and suitable for casting.*
- ***Bronze** is an alloy of copper with tin and optionally zinc, silicon, nickel, and other metals.*
- ***Calamine brass** is a brass alloy manufacturing process from the first millennium BC that was not replaced in Europe until the late 18th century.*
- ***Cartridge brass** is a 30% zinc brass with good cold working properties.*
- ***Common brass**, or **rivet brass**, is a 37% zinc brass, is inexpensive, and is the standard for working cold.*
- ***Cupronickel** is an alloy of copper with nickel.*
- ***High brass** contains 65% copper and 35% zinc, has a high tensile strength, and is used for springs, screws, and rivets.*
- ***Leaded brass** contains addition of lead. It has excellent machinability.*
- ***Low brass** is a copper-zinc alloy containing 20% zinc with a light golden color, excellent ductility and is used for flexible metal hoses and metal bellows.*
- ***Naval brass**, similar to admiralty brass, is a 40% zinc brass and 1% tin.*
- ***Pinchbeck** is a brass that closely resembles gold in appearance.*
- ***Red brass** is an American term for Cu/Zn/Sn alloy known as gunmetal.*
- ***White brass** contains more than 50 % zinc and is too brittle for general use.*
- ***Yellow brass** is the American term for 33% zinc brass.*

[Information from Wikipedia]

3) Consider two different forms of brass. Brass A is 65% copper and 35% zinc, and Brass B is 75% copper and 25% zinc. Which form will be most dense? Explain your reasoning.
Brass B will have a greater density since there is more of the more dense copper in the alloy.

4) If the crown fit comfortably, what would be the difference between wearing a brass crown and wearing a gold crown of identical mass?
The pure gold crown would be much smaller in volume.

5) If one of the principles for green chemistry is prevention of waste, especially dangerous waste, would it be greener for several students to use the same steaming alkali solution or for each student to prepare his own?
Less solution that requires disposal is the greener choice. Therefore, as many students as possible using the same alkali solution would be preferred, provided it did not prohibitively lengthen the time required for the investigation.

ALCHEMY SAMPLE REPORT SHEET

I. MASS OF ORIGINAL TOKEN: _____*2.50 g*_____________

II. MASS OF DRY SILVER TOKEN: ___*2.51 g–2.53 g*_____
Instructor's Initials for silver token: ____________________

How does the silver token feel? (Is it smoother than the copper token? Are there rough spots?)
If the penny is very clean and bright to begin with, the coating will not change the texture.

Is the surface evenly coated with the zinc, or are spots of copper still visible?
Once again, if the penny is very clean and bright to begin with, and providing that it remains in the alkali solution sufficiently long, there should be no visible copper surface.

III. MASS OF COOLED GOLD TOKEN: ___*2.51 g–2.53 g*______
Instructor's Initials for gold token: _______________

How does the gold token feel?
If it is heated too long, the copper layer with its brass coating will bubble up and become rough. Otherwise, it should remain very smooth.

Is the gold color uniform over the entire coin?
If the silver coating was uniform, that is to say, there was no visible copper surface, then the gold color will also be uniform. Metal is an excellent conductor of heat, and the entire coin should reach the temperature required to form the alloy at the same time.

IV. DECISION
Do you tell the King the token has really turned to silver and then to gold? Do you tell him the silver-colored token is just zinc-coated copper and that the zinc has diffused into the copper to make brass, a solid solution of copper and zinc?

EXTRA
Don't lose your head because of a wrong decision. Consider the densities.

- [ ] Copper to Silver to Gold
- [x] Copper to Zinc coating to Brass

V. DENSITY
If you claim the token is not gold, can you prove it by a density calculation?

EXTRA
Density = mass/volume.

Yes. The density of gold is 19.3 g/cm^3, so the mass of a gold token would be greater than the experimental token's mass. Therefore, the token is not gold.

Many tokens are not flat, but an average thickness can be calculated from the measured mass and radius using the following equation:
$t = m/(\pi r^2 d)$, where r is radius (half the diameter), π is 3.14, t is thickness, d is density, and m is the measured (actual) mass.

Measurements of copper token:
diameter ____*1.9*____________ cm
radius ______*0.95*___________ cm
mass _______*2.50*_________ g

EXTRA
densities are:
copper 9.0 g/cm³
silver 10.5 g/cm³
gold 19.3 g/cm³
zinc 7.1 g/cm³

calculated average thickness _______*0.10*___________ cm

The volume of a token can be calculated from volume (v) = $\pi r^2 t$.
volume of token_____*0.27*_______ cm^3

You can calculate the mass of the token using the equation: mass (m) = density (d) × volume (v). The token thickness did not change significantly as it changed from copper to silver to gold, so you can calculate the expected masses of a solid silver and a solid gold token using the volume of the copper token and the densities of silver and of gold. You can then compare those masses with the actual masses.

	CALCULATED MASS	ACTUAL MASS
Original Token		*2.50 g*
Silver Token	*2.84 g*	*2.51–2.53 g*
Gold Token	*5.21 g*	*2.51–2.53 g*

VI. QUESTIONS

1. If the zinc adhered (stuck) to the copper instead of being bonded (chemically joined) to it, would the change to silver color be a chemical change?
   No, this is not a chemical change.

2. If the heating causes the zinc to bond with the copper, is the change to gold color a chemical change?
   Yes, the properties of the metal (color, density of the alloy, etc.) are changed.

3. What happened to the mass of a penny when the United States changed the penny's composition from pure copper to a copper-clad zinc coin? (Hint: Consider the densities given in the background material.)
   Provided the overall dimensions of the coin did not change, the mass of the penny decreased because the zinc has a smaller density than the copper.

4. What happened to the mass of a dime when the United States changed from pure silver to a silver-copper-silver sandwich?
   The mass of the dime decreased because the copper has a smaller density than the silver.

5. How long would it take a person who received an object "changed into gold" to realize he or she had been the victim of an early magic trick? What would most likely be the first indication?
As soon as something rubbed or scraped the thick alloy coating from the surface, the copper below it would become visible.

#2 Density Layers

Reagents

isopropyl alcohol	10 mL/pair

Common Materials

nontoxic antifreeze	1 mL/pair
brake fluid	1 mL/pair
30 wt motor oil	1 mL/pair
power steering fluid	1 mL/pair
transmission fluid	1 mL/pair
mineral oil	10 mL/pair
Ajax® laundry detergent (liquid)	10 mL/pair
Downy® fabric softener	10 mL/pair
Karo® syrup (both types: dark and clear)	10 mL/pair
molasses	10 mL/pair
vegetable oil	10 mL/pair
small cork	1/pair
paraffin	1 small piece/pair
thumbtack	1/pair
plastic paper clip	1/pair
aluminum paper clip	1/pair
a small stopper or pieces of rubber band	1/pair
food coloring—green	1/pair
ice cubes	1/pair

Laboratory Equipment

10-mL graduated cylinder
50-mL graduated cylinder
100-mL graduated cylinder
long stem funnel
small test tube
250-mL beaker

Special Equipment

(none)

Notes

The smallest rubber stopper issued to the students will suffice for the one called for in the experiment, but an even smaller one would be better.

Disposal

Most automobile fluids are oily and can contaminate water if they are not disposed of properly. Ethylene glycol is poisonous; for this reason we suggest using propylene glycol (nontoxidc antifreeze). Fortunately, service stations are part of our oil-recycling program. Pour the auto fluids into a marked waste container. The solutions of household products can be flushed down the drain.

DENSITY LAYERS PRE- LAB QUESTIONS

1) When you are walking through a parking lot and see a puddle of water with a rainbow effect on the surface, you are observing a thin-film optical phenomenon. Do you suspect the material in the film is miscible or immiscible with the water? Explain your reasoning.
   The substance must be immiscible in order for it to form a film on the surface. If it dissolved, there would be no top layer, or film.

2) What is the nature of film on the surface of the water in the first question, and what was its likely source?
   The particles are not miscible, so they are probably long-chain organic molecules. Since it is in a parking lot, it is most likely a petroleum product and could be oil or another auto fluid that has dripped from a car, or gasoline.

3) Assume you had to separate a series of layers of materials that have different densities. Suggest a possible method for separating each of the following:
   a) Solid objects of greater density than the liquid in which they are contained
      The less dense liquid could be poured from the top, leaving the more dense solid behind.

   b) Solid objects of less density than the liquid in which they are contained
      The solids could be filtered, or they could be skimmed from the surface.

   c) Two immiscible liquids of significantly different densities
      Most of the top layer could be poured from the top, and then the rest could be skimmed. Also, a burette, separatory funnel, or gravy cup could be used.

   d) Two miscible liquids of significantly different densities
      A burette or separatory funnel would be the best choice so that the interface is moved as little as possible.

4) We should be careful not to allow petroleum products to get into our waterways, and yet each time it rains we see evidence of those products in the environment. Can you suggest a creative method to trap these products as they run into street sewers?
   It would be very difficult to separate them as they run into the gutters, but a settling basin that held the liquids long enough for them to separate and then allowed the water to drain out of the bottom, holding the less dense petroleum liquids on the surface, would work well.

5) Motor oils are an environmental hazard. Are vegetable oils an environmental hazard?
   Vegetable oils are biodegradable materials whereas petroleum products remain in the environment for many years and can be poisonous.

6) If ethylene glycol is soluble in water, is it more dangerous to the environment? Is it more difficult to remove from the environment?
   Soluble compounds move throughout the environment with the water in which they are dissolved; therefore, ethylene glycol is likely to spread through soils and water systems once it gets into any waterway.

DENSITY LAYERS SAMPLE REPORT SHEET

I. AUTO LIQUIDS

List the fluids in order of decreasing density (most dense first).

antifreeze

brake fluid

transmission fluid

30 wt motor oil

manual power steering fluid Instructor's initials: __________

II. ICE ON LIQUIDS

At what position is the ice originally?

The ice will come to rest between the laundry detergent and the alcohol.

After Procedure 8 of the student manual, where is the liquid water that came from the melted ice

It will be in solution with the liquid of most similar density, the fabric softener.

III. HOUSEHOLD LIQUIDS *(The following are sample data only.)*

Mass of 10-mL graduated cylinder *28.07* g

Mineral Oil

Mass of cylinder and oil *36.66* g Volume *10* mL

minus mass of cylinder − *28.07* g Density $\frac{8.59\ \text{g}}{10\ \text{mL}}$ = *0.859* g/mL

Mass of oil *8.59* g

Isopropyl alcohol

Mass of cylinder and alcohol *35.64* g Volume *7.7* mL

minus mass of cylinder − *28.07* g Density $\frac{7.57\ \text{g}}{7.7\ \text{mL}}$ = *0.983* g/mL

Mass of alcohol *7.57* g

Ajax laundry detergent

Mass of cylinder and Ajax® *34.15* g Volume *9.4* mL

minus mass of cylinder − *28.07* g Density $\frac{6.08\ \text{g}}{9.4\ \text{mL}}$ = *0.647* g/mL

Mass of Ajax® *6.08* g

Downy fabric softener

Mass of cylinder and Downy®	*34.76* g	Volume	*6.7* mL
Minus mass of cylinder	− *28.07* g	Density	*6.69* g / *6.7* mL = *0.999* g/mL
Mass of Downy®	*6.69* g		

Karo

Mass of cylinder and Karo®	*41.07* g	Volume	*9.5* mL
minus mass of cylinder	− *28.07* g	Density	*13.0* g / *9.5* mL = *1.37* g/mL
Mass of Karo®	*13.0* g		

Molasses

Mass of cylinder and molasses	*39.13* g	Volume	*7.8* mL
minus mass of cylinder	− *28.07* g	Density	*11.06* g / *7.8* mL = *1.42* g/mL
Mass of molasses	*11.06* g		

OBJECT	ESTIMATED DENSITY	OBJECT	ESTIMATED DENSITY
Thumbtack	*1.4* g/mL	Aluminum paper clip	*0.75* g/mL
Rubber stopper/band	*1.3* g/mL	Plastic paper clip	*0.65* g/mL
Paraffin	*0.6* g/mL	Cork	*0.5* g/mL

III. QUESTIONS

1. Why shouldn't the auto liquids be poured down the drain?
   The liquids can contaminate water, and some are poisonous.

2. Of the automobile fluids, which ones mix well together? Which fluids form separate layers?
   The antifreeze and brake fluid mix well. The power steering fluid does not mix.

3. After observing the difference in the densities of the auto liquids, explain why it is important that the liquids used in an automobile be uncontaminated.
   The liquids do not mix entirely, and all have different densities. This will cause the improper functioning of an automobile.

4. What does it mean if an object sinks to the bottom of a cylinder filled with a liquid?
   This means that the object is greater than or equal in density of the liquid.

5. What does it mean if an object floats on top of a liquid?
   This means that the object is less dense than the liquid.

6. Karo® and molasses are mostly sugar in water. Why do they form separate layers?
   They form separate layers because they have different proportions of sugar and water, which makes their densities different.

7. Based on your calculations in Part III, which of the household liquids will initially float on water?
   The Ajax, Downy, alcohol, and mineral oil will float.

8. What can be inferred about the relative densities of ice and water from the ice-on-liquids procedure? Give an estimate of the value of the density of ice.
   Ice is less dense than water. Bodies of water are therefore insulated as they freeze on top first.

#3 Energy in Physical and Chemical Changes

Reagents

ammonium chloride [NH_4Cl], solid	6 g/pair
calcium chloride [$CaCl_2$], solid	9 g/pair
sodium hydroxide [NaOH], solid	4 g/pair
1.0 M sodium hydroxide [NaOH], solution	50.0 mL/pair
sulfuric acid [H_2SO_4], concentrated	5.0 mL/pair
1.0 M hydrochloric acid [HCl]	50,0 mL/pair
0.5 M hydrochloric acid [HCl]	100.0 mL/pair

Common Materials

Styrofoam® cup	1/pair
cardboard (for a lid)	1/pair
wire for stirrer	1/pair
rubberband	1/pair

Laboratory Equipment

thermometer	1/pair
100-mL graduated cylinder	1/pair
weighing paper or weighing boats	

Special Equipment

(none)

Notes

Disposal

All solutions can be flushed down the drain with plenty of water.

ENERGY IN PHYSICAL AND CHEMICAL CHANGES PRE-LAB QUESTIONS

1) In a certain ionic compound, more energy is required to break the lattice energy holding ions together in the crystal than is released when the ions dissolve in water. Will the overall process appear to be endothermic or exothermic? Explain your logic. Will the beaker feel warm or cool?
The energy required for a reaction is the endothermic process, while released energy is exothermic. Since more energy is required, the overall process will be endothermic. The beaker will feel cool as the energy leaves your hand.

2) Consider two solids that are stirred together in a glass beaker at room temperature and interact to produce a liquid while frost is seen to form on the outside of the beaker. Is the process endothermic or exothermic? Discuss the energy changes that can be assumed in the process.
The frost on the beaker would suggest that the process is endothermic. This means that the energy released to form the new substance is not nearly as great as the amount of energy required to break the bonds in the two original compounds.

3) Is the resulting liquid in Question 2 most likely the result of a physical or chemical change? Use energy changes to support your answer.
If the solid to liquid change was a physical process, one would expect the energy change to be endothermic so it would be hard to prove that the liquid was the result of a chemical change. However, the fact that the temperature decreased well below room temperature would suggest that another force is at work and that the process is indeed chemical.

4) Two substances in solution at room temperature are mixed in a test tube containing a thermometer to register temperature changes. There is no change on the thermometer, but the combined solution immediately becomes cloudy and a white powder eventually settles to the bottom of the test tube. Why is it incorrect to determine that no energy changes occurred?
Energy is required to separate dissolved particles from the solution, but it will be released when the new bonds form in the solid. The energy changes in the two opposite processes were simply equal in magnitude.

5) Ice and hot water are combined in the same container. Discuss the energy changes that will take place and any state changes that will result.
The energy from the hot water will be transferred to the ice to melt it. The ice absorbs heat (an endothermic process) as the hot water loses it (an exothermic process).

6) Considering the container discussed in Question 5, if ice remains after the temperature stabilizes, what can be assumed about the final temperature?
The final temperature would be the freezing point of the water.

ENERGY IN PHYSICAL AND CHEMICAL CHANGES SAMPLE REPORT SHEET

I. PART A: Heat of Solution and Dilution (Physical Changes)

	TEMPERATURE		
	INITIAL	FINAL	Δt (final – initial)
Ammonium chloride	________	________	________
Calcium chloride	________	________	________
Sodium hydroxide	________	________	________
Sulfuric acid	________	________	________

II. PART B: Heat of Reaction (Chemical Changes)

	TEMPERATURE		
	INITIAL	FINAL	Δt (final – initial)
HCl and NaOH solution	________	________	________
HCl and solid NaOH	________	________	________

III. QUESTIONS

1. Is there a difference in the final solution temperature of the HCl with NaOH solution and with solid NaOH? Can the difference by explained by the heat of solution of NaOH? (Hint—Compare the temperature changes for the HCl with solid NaOH to changes for both the water with solid NaOH and the HCl with NaOH solution.)
There is a difference because the solid must first be dissolved and then react.

2. Explain why acid is poured into water instead of water into acid, given the fact that if drops splatter during pouring, the drops are usually of the liquid poured into, not the liquid being poured.
Acid is poured into water so that the splattered drops will be water, not acid.

3. Could a solution of one of these compounds serve as a cold pack for an athletic injury?
An endothermic reaction could be used for a cold pack.

#4 Separation of a Mixture

Reagents

mixture of iron filings, cork, sugar, and sand

Common Materials

magnet

Laboratory Equipment

evaporating dishes (2)
glass rod
10-mL graduated cylinder
weighing paper

Notes

Disposal

All substances can be discarded into a trash receptacle.

SEPARATION OF A MIXTURE PRE-LAB QUESTIONS

1. Physical properties include color, density, hardness, mass, odor, temperature, structure, taste, boiling point, melting point, specific heat, conductivity, solubility, and magnetism. Which of these properties are being used in this experiment?
   Density, solubility, magnetism

2. From the list above, identify three physical properties that could not be used to separate components of a mixture.
   Odor, taste, specific heat, temperature

3. Chemists are moving toward greener chemistry. From that viewpoint, why is water a better solvent to use in this separation than a more toxic solvent, such as an ammonium solution or alcohol?
   Water, besides being one of the most versatile solvents, is naturally present in all parts of the environment and is therefore not an issue if it is allowed to escape the system. An ammonium solution, on the other hand, is reactive and is an irritant to many living systems. It is also capable of altering the pH of any water-based substances.

4. Why do you remove the iron filings before adding water?
   If the filings are wet, they will adhere to other components in the mixture, making it difficult to separate them.

5. Is separation of mixtures important in recycling? Explain your reasoning.
   The ability to use physical properties to separate substances in a mixture is critical because one substance may have a very different procedure for recycling than another.

SEPARATION OF A MIXTURE SAMPLE REPORT SHEET

I. RESULTS

Note: Data will vary from student to student and from lab to lab, depending on the ratio of substances in the mixtures.

Mass of evaporating dish #1	________g
Mass of evaporating dish #2	________g
Mass of weighing paper and iron filings	________g
− Mass of weighing paper	________g
Mass of the iron filings	________g
Mass of sugar and evaporating dish #2	________g
− Mass of evaporating dish #2	________g
Mass of dry sugar	________g
Mass of silicon dioxide and evaporating dish #1	________g
− Mass of evaporating dish #1	________g
Mass of dry silicon dioxide	________g
Mass of iron filings	________g
+ Mass of sugar	________g
+ Mass of silicon dioxide	________g
Total mass of iron, sugar, and silicon dioxide	________g
Initial mass of mixture	________g
− Mass of iron, sugar, and silicon dioxide	________g
Mass of cork	________g

II. QUESTIONS

1. What physical property is used to separate the iron filings from the mixture?
   magnetism

2. What physical property is used to separate the cork from the mixture?
   density

3. What physical property is used to separate the sugar from the mixture?
   solubility

4. What physical property is used to separate the silicon dioxide from the mixture?
   solubility and density

5. What would need to be done to the cork to be able to use it again?
   The cork would only need to be rinsed well to remove any remaining dissolved sugar and then dried.

6. What would be the effect on the mass of the cork if the sand and/or sugar were not dry when weighed?
   Any moisture remaining in the sugar or sand would have subsequently been subtracted from the total mass, resulting in an erroneously small mass for the cork.

#5 Conservation of Mass

Reagents
1 M magnesium chloride [$MgCl_2$]
1 M sodium hydroxide [NaOH]
solid sodium hydrogen carbonate, also called sodium bicarbonate [$NaHCO_3$]
1 M acetic acid [CH_3COOH]

Common Materials
large party balloons

Laboratory Equipment
centigram balance
10-mL graduated cylinder
500-mL Erlenmeyer flask with stopper to fit
250-mL Erlenmeyer flask
1 small test tube

Notes
Balloons need to fit tightly over the mouth of the flask. A rubber band to secure them may be in order.

Disposal
All materials are safe to flush down the drain after being diluted.

CONSERVATION OF MASS PRE-LAB QUESTIONS

1. An early researcher attempted to use a growing plant to support the idea of conservation of mass. Measuring the mass of all water added to the plant for the period of a year and accounting for the loss of mass from the soil, he was unable to prove that the mass gained by the plant was equal to the mass added by water plus the mass lost by the soil. Was the system open or closed? What processes and substances were not accounted for?
The system was an open system, and gases were allowed to enter and leave. The mass gained by the plant includes carbon from carbon dioxide absorbed from the atmosphere. The carbon mass was not included in the water or the nutrients lost from the soil.

2. Can you prove that the gas particles contained in the balloon have a measurable mass? If so, explain how you might proceed to experimentally verify the statement.
The mass of a balloon could be taken before it is inflated, and then again after it is full of gasses and tied. The mass for the inflated balloon will be greater.

3. Refer to Procedures 1 and 2 in the student manual. How could the sodium chloride (Na^+ and Cl^- ions) that remains dissolved in the solution be recovered?
The water could be evaporated, leaving the salt behind.

4. The dissociation of carbonic acid to form water and carbon dioxide gas in Procedure 4 is a major component of the soft drinks we consume. What happens to the acid content of a carbonated drink when the cap is removed and the carbon dioxide gas is able to escape?
As the carbon dioxide gas at the surface of the liquid escapes, more carbonic acid dissociates, replacing the gas. This lowers the acid content of the drink.

5. One of the principles of green chemistry is to design less hazardous chemical syntheses both to human health or the environment. How does this principle relate to the difference between using an open or a closed system during the syntheses?
An open system allows uncontrollable movement of gas particles. Any loss of material (that might otherwise be contained and used in another application) could be thought of as waste by-product from the process.

CONSERVATION OF MASS SAMPLE REPORT SHEET

I. REACTION OF MAGNESIUM CHLORIDE AND SODIUM HYDROXIDE
Note: Student measurements will vary.
Mass of the beaker, test tube, and contents

Before the reaction: ______________ g

After the reaction: − ______________ g

Difference: ______________ g

Description of substances
Magnesium chloride solution: ____
____

Sodium hydroxide solution: ____
____

Description of reaction: ____
____

Final substance(s): ____
____

II. REACTION OF SODIUM BICARBONATE AND ACETIC ACID
Trial 1
Mass of beaker, test tube, and contents
Before the reaction: ______________ g

After the reaction: − ______________ g

Difference: ______________ g

Trial 2
Mass of beaker, test tube, balloon, and contents
Before the reaction: ______________ g

After the reaction: − ______________ g

Difference: ______________ g

Descriptions of substances

Acetic acid solution: ______*clear, vinegar odor*______________________________

Sodium bicarbonate powder: ____*white, fine powder*_________________________

Description of reaction: _______*profuse bubbles, cooling flask*________________

Final substance(s): __________*clear liquid, possibly with powder on bottom*______

III. QUESTIONS

1. What indications were present to suggest that a reaction had taken place?
   Spontaneous temperature change of the system, production of a precipitate, as evidenced by the instant cloudiness of the mixture, gas production

2. Why could a glass container with a tightly fitting lid not be substituted for the balloon in the third reaction?
   The added gas molecules will increase the pressure in the container because it does not expand. The pressure could become great enough to break it.

3. How much did the gas weigh?
   Note: Student answers will vary.

4. Did you confirm the Law of Conservation of Mass? Give evidence from your experimental data to support your conclusion.
   Only the experiment in which the system is closed will have good enough data to support Conservation of Mass. Data cited will vary.

#6 Laws of Definite and Multiple Proportions

Reagents

plaster of Paris, calcium sulfate hemihydrate	distilled water
powdered gypsum, calcium sulfate dihydrate	3 M sulfuric acid

Laboratory Equipment

centigram balance	600-mL beaker
evaporating dish	ring stand
watch glass	2 test-tube clamps
2 25-mL graduated cylinders	glass stirring rod
desiccators	tongs
thermometer	

Common Materials

9-V battery	insulated copper wire

Notes

Powdering the substances as fine as possible will speed the dehydration process.

Disposal

Solid materials may be disposed of in a landfill. Liquids may be flushed down the drain.

LAWS OF DEFINITE AND MULTIPLE PROPORTIONS PRELAB QUESTIONS

1. In our investigation, we will split one compound into its constituent elements and another compound into two simpler compounds. If in some of Proust's experiments, his decomposition reactions resulted in the formation of other compounds rather than elements, would the validity of his conclusion be affected? Explain your reasoning.
   Even if one or more of the substances formed in a decomposition reaction were compounds rather than elements, they would most likely be formed in consistent ratios from one experimental trial to another. Thus, his conclusions would not be affected.

2. If two or more compounds made of the same elements but in different ratios, were present in a sample analyzed by Proust, how might this affect his experimental results?
   The experimental ratios of multiple trials might have been inconsistent depending on the relative amounts of the two initial compounds. If the two compounds were always present in the same relative amount, the results of each trial would be consistent but would not reflect the actual elemental ratio of either compound.

3. Plaster is still intact in the Egyptian pyramid at Cheops. What does this indicate about the molecular stability and durability of the compound?
   Plaster is a very stable compound, resistant to decomposition and to combination with other substances in the environment. It is also extremely durable and remains intact for many centuries.

4. If water samples used in electrolysis experiments are not pure, how might this affect the data?
   Impurities that could also be broken down by electrolysis would produce additional substances. If those substances were gaseous, the products collected in the tubes would not be pure. However, if the products were solid or liquid, they would not affect the results.

5. One of the principles of green chemistry is to design safer chemicals. Why is it important to know which of two compounds is being made in a synthesis?
   Green chemistry principles discourage the production of needless waste in the form of leftover reagents or by-products, both of which require disposal. Therefore, if only one particular compound is needed from a synthesis, a chemist can design synthesis reactions to use the minimum of reagents and produce a minimum amount of by-product.

LAWS OF DEFINITE AND MULTIPLE PROPORTIONS SAMPLE REPORT SHEET

I. PART A: DATA

Trial	1	2
Volume of hydrogen	___________mL	___________mL
Volume of oxygen	___________mL	___________mL

Note: Answers will vary, but expect a 2:1 ratio of volumes.

II. PART A: CALCULATIONS *Answers will vary.*

1. The smaller volume is oxygen gas, and the larger volume is hydrogen. For each trial, divide the two volumes by the smaller one, so that the value for the oxygen volume becomes 1.00 and the value for the hydrogen volume is greater than one. Express the hydrogen value to the nearest hundredth.

2. Express the oxygen to hydrogen ratio as 1.00: _____ where both values are expressed to the hundredths place.

Trial	1	2
Oxygen volume/oxygen volume	____________	__________
Hydrogen volume/oxygen volume	____________	__________
Oxygen to hydrogen ratio	1.00:________	1.00:______

Final ratio should approach 1.00:2:00.

III. PART B: DATA *Answers will vary.*

	Gypsum	Plaster of Paris
Appearance before heating	__________________________	_______________________
Appearance after heating	__________________________	_______________________
Observations during heating	__________________________	_______________________
Mass of evaporating dish and watch glass	_________________g	_________________g

Mass of dish, glass, and powder before heating _________________g _________________g

Mass of dish, glass, and powder after heating _________________g _________________g

IV. PART B CALCULATIONS *Answers will vary.*

1. For each powder, calculate the mass of the *hydrated compound* by subtracting the mass of the dish and glass from the total mass before heating. Record these values.

   Gypsum _______________g Plaster of Paris _______________g

2. For each powder, calculate the mass of the *anhydrous compound* by subtracting the mass of the dish and glass from the final mass after heating. Record these values.

   Gypsum _______________g Plaster of Paris _______________g

3. For each powder, calculate the mass of water removed from the compound by finding the difference in the masses of the hydrated and anhydrous forms of the compound. Record these values.

   Gypsum _______________g Plaster of Paris _______________g

4. For each powder, find the moles of anhydrous compound by dividing the mass of anhydrous compound by its molar mass. Record these values. (The molar mass of $CaSO_4$ is 136.14 g/mol.)

   Gypsum _______________mol Plaster of Paris _______________mol

5. For each powder, find the moles of water that were removed by dividing the mass of the water removed by its molar mass (18.0152 g/mol). Record these values.

   Gypsum _______________mol Plaster of Paris _______________mol

6. For each powder, find the ratio of moles anhydrous $CaSO_4$ to moles water in the hydrate by dividing the moles water by the moles of anhydrous $CaSO_4$. Record these values expressed to the nearest 0.01 and in the form "1.00: __."

   Gypsum 1.00:_______________ Plaster of Paris 1.00:_______________

7. Find the percent error for your experimental values by using the following formula:

   $$\% \text{ Error} = |A-E| / A \times 100\%$$

   where A is the accepted value and E is the experimental value for the hydrate number. The vertical lines on each side of A-E mean take the absolute value. The accepted hydrate values for gypsum and plaster of Paris are 2.00 and 0.50, respectively. Record the percents error in your calculations table.

   Gypsum ______________%   Plaster of Paris ______________%

8. Gather experimental ratios from several classmates as comparison to your work.

   Gypsum 1.00:______________   Plaster of Paris 1.00:______________

   Gypsum 1.00:______________   Plaster of Paris 1.00:______________

   Gypsum 1.00:______________   Plaster of Paris 1.00:______________

   Gypsum 1.00:______________   Plaster of Paris 1.00:______________

   Gypsum 1.00:______________   Plaster of Paris 1.00:______________

V. QUESTIONS

1. The dehydration of gypsum can be observed to occur as a two-step process, with a definite increase in temperature between the steps. From the introductory discussion of the relationship between gypsum and plaster of Paris, can you explain this phenomenon?
   During the dehydration process, the heat absorbed by the gypsum is equal to the heat of hydration for the substance. Heat is absorbed until all the moisture is removed. Once all the moisture is driven off, the plaster of Paris that remains has a different heat of hydration. The temperature must reach a new, higher temperature before water is driven from the plaster.

2. Why is it important to place the dehydrated substance into a sealed dessicator while it cools?
   The moisture in the air will be reabsorbed by the dehydrated substances as they cool, and the results will be erroneous.

3. Use your observations and data to confirm the existence of more than one ratio for the combination of calcium sulfate and water?
   The ratio of calcium sulfate to water in gypsum is higher than the ratio of calcium sulfate to water in plaster of Paris. In addition, there are two steps to the dehydration of gypsum.

4. Many geological studies have been performed to determine if the natural processes that are responsible for the dehydration of gypsum in Earth's crust result in any instability of rock formations. Why is this knowledge important, and how might laboratory simulation of these processes help scientists understand the potential effects of what is happening deep in Earth's crust?
   If heat and/or the passage of time might result in a loss of hydration of any components in the crust of the earth, stability could potentially be affected. For areas high in gypsum content, this might be a clue to the future stability of the rock structure. Laboratory studies of related processes could shed light on the likelihood of shifting crust and help in the prediction of earthquakes.

#7 Periodicity

Common Materials

scissors clear tape

Notes

1. Name several ways in which Mendeleev's arrangement of elements was different from earlier representations.
   Earlier arrangements of elements were in list form and relied primarily on increasing mass. Mendeleev incorporated several other characteristics, including density and compounds formed by the elements. Using both mass and other properties, Mendeleev arranged the elements in a chart format so that similar properties lined up when they were arranged by increasing mass.

2. What is the advantage in having an arrangement of elements that identifies groups of properties as well as the orderly increase in mass?
   The arrangement allows for a prediction of properties by the position of the element on the chart. Any undiscovered elements may even be located by knowledge of likely compounds in which they could be found.

3. Early criticisms of Mendeleev's table centered on the fact that, when determining the placement of elements, in a few places he chose to follow the trend in properties over the trend of increasing mass. For example, the mass of iodine is actually less than the mass of tellurium, but the remaining properties fall correctly if tellurium is placed before iodine in the table. If you were Mendeleev, how would you defend the placement of the two elements?
   Properties are more indicative of the nature of elements than mass. Some elements have very similar masses but quite different properties.

4. With the eventual understanding of basic atomic structure and the use of the number of protons (atomic number) as the identifying characteristic for elements, the perceived discrepancies in Mendeleev's table were shown to be inconsequential and his arrangement was affirmed. What does this say about the relationships among the number of protons, mass, and elemental properties?
   Elemental properties are dependent on the atomic number (number of protons) rather than on the mass.

5. How does this investigation relate to the green principle of prevention of waste if it only uses paper and no chemicals?
   By using paper, we are producing only waste that can be easily recyclable.

PERIODICITY SAMPLE REPORT SHEET

I. PART A

Name the five characteristics that can each be found in a group of the aliens and describe the trend:

Characteristic	Trend
1. *hair*	*length decreases across each row*
2. *eyes*	*number decreases across each row*
3. *stripes*	*number increases across each row*
4. *oval*	*size increases across each row*
5. *stars*	*number increases across each row*

Name the two general characteristics that are shared by all aliens but that vary:

Characteristic	Trend
1. *diamonds on belt*	*number increases down each column*
2. *size*	*decreases across each row from left to right*

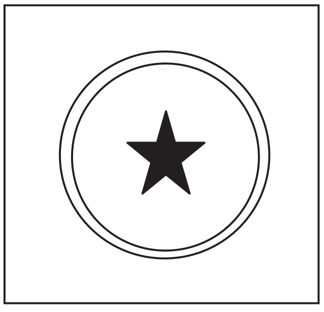
Instructor's initial for accurate arrangement of the pictures: ____________________

Note: Without any numerical data, there is more than one possible arrangement for the aliens, depending on whether the student arranged the size and number of diamonds on the belt in increasing or decreasing order. One possibility has been described above.

II. PART B

In the box at right, draw the design that is missing from the collection.

Note: Regardless of the direction in which the properties are arranged in this exercise, the missing element will be the same.



III. PART C

In the table below, tape the blocks of elemental properties in table format so that they are arranged in proper order and are aligned by trend.

1 R_2O 6.9 +1	2 RO 9.0 +2	3 R_2O_3 10.8 +3	4 RO_2 RH_4 12.0 −4 or +4	3 R_2O_5 RH_3 14.0 −3	2 RH_2 16.0 −2	1 R_2O_7 RH 19.0 −1
1 R_2O 23.0 +1	2 RO 24.3 +2	3 R_2O_3 27.0 +3	4 RO_2 RH_4 28.1 +4	3 R_2O_5 RH_3 31.0 −3	2 RO_3 RH_2 32.1 −2	1 R_2O_7 RH 35.5 −1
1 R_2O 39.1 +1	2 RO 40.1 +2	3 R_2O_3 69.7 +3	4 RO_2 RH_4 72.6 +4	3 R_2O_5 RH_3 74.9 −3	2 RO_3 RH_2 79.0 −2	1 R_2O_7 RH 79.9 −1

IV. QUESTIONS

1. From the information used in Part C, identify the trends for each property provided for the elements.

Property	Trend
a. Valence	________*increases and then decreases across each row*____________
b. Oxide	________*ratio of R:O increases across each row*_________________
c. Hydride	________*ratio of R:H decreases across each row*_________________
d. Mass	________*increases across and down*____________________________
e. Ion Charge	________*increases across each row to group 4 then then the negative charge decreases*____

2. Mendeleev kept data about measurable properties of elements. Explain how this enabled him to predict specific properties of missing elements.
Numeric data that follows a definite trend is predictable. Mendeleev would have been able to fill in gaps that formed when he lined up elements by similar properties.

#8 Atoms and Light

Reagents

magnesium ribbon	2 in./lab
0.1M chloride salts of Li, Na, Ba, Sr, K, and Cu	5 mL
europium oxide [Eu_2O_3]	enough for demonstration
concentrated HCl (for cleansing flame test wires)	10 mL/lab

Common Materials

candle	1/lab
incandescent bulb in socket	1/lab

Laboratory Equipment

laboratory burner	1/lab

Special Equipment

flame test wires	1/solution
diffraction gratings	1/student
hydrogen discharge tube	1/lab
helium discharge tube	1/lab
neon discharge tube	1/lab
mercury discharge tube	1/lab
ultraviolet light	1/lab
discharge apparatus	1/lab
earth light (optional)	1/lab

Notes

The room must be totally dark in order for the experiment to work well. Sometimes the students hold the diffraction grating turned 90 degrees from the correct orientation. Check each student's view, but this is best done as a demonstration experiment. There may be a hazard from UV light, when burning magnesium or using the mercury discharge tube. The teacher must have adequate light sources and equipment; perhaps the Physics Department can supply any materials that are not available through the Chemistry stockrooms.

The HCl is only necessary for cleaning the flame loops. It is not necessary to have it for each pair of students.

Nitrate salts can also be used for the flame tests.

Diffraction gratings are available from Flinn Scientific, Batavia, IL and from Frey Scientific, 905 Hickory Lane, Box 8101, Mansfield, OH 44901-8101.

An additional light source is the Earth Light by Phillips Petroleum. It gives white light, rather warm, from three phosphors, red, green, and violet, as shown by the diffraction grating. An Earth Light may be purchased from Phillips Lighting, P.O. Box 156, Orangeburg, NY 10962-0156.

Disposal

There should not be any disposal to deal with if the test solutions are capped tightly and saved from semester to semester.

ATOMS AND LIGHT PRE-LAB QUESTIONS

1) Name several ways in which atoms can be excited.
Heating, electricity, energy from chemical reactions, and absorbing electromagnetic radiation are all ways in which electrons can be excited.

2) When a fuel such as wood is burning in a fireplace, visible light is one of the products of the reaction. What is the source of this light we call the "flame?" (Give your answer in the form of a sequence of events.)
Heat released in the reaction excites the electrons. It also warms the gaseous products so that they rise from the fuel. As the electrons in the atoms reemit the energy that first excited them, they do so as wavelengths of electromagnetic radiation that we perceive as colors of light.

3) How might the normal color of a flame mask the true color of a salt during a flame test?
If the fuel is a source that gives color to the flame, the color will be added to the spectrum resulting from the substance being tested. It would be best to use a fuel that burns with a colorless flame.

4) When the electric element on a home cooktop is turned on, it first becomes light orange and then turns red hot. If more current is supplied, it can become hotter. If there was no restriction on the amount of current, at some point the temperature would be sufficient for the burner to glow white. What has happened to the number of individual colors of light being emitted if the burner appears white to the naked eye?
As the temperature increases, more and more different electrons in the atom are excited. Each electron gives its own color, or colors, of light as the energy is reemitted. When enough different colors of the visible spectrum are present, the overall effect will appear to be white light

5) Video cameras measure colors in Kelvin temperatures. Why is this appropriate?
Colors of light are the key to the temperatures of stars. Since all electromagnetic radiation originates in stars, this is a true and universally accepted standard to colors of light.

ATOMS AND LIGHT SAMPLE REPORT SHEET

I. Describe the view of each of the following through a diffraction grating:
1. Candle—*complete spectrum*
2. Burning magnesium—*complete spectrum*
3. Incandescent bulb (tungsten)—*complete spectrum*
4. Hydrogen—*three lines, one blue, one violet, and one red. A green line can also be seen.*
5. Helium—*2 green lines, 3 yellow-green, 6 yellow, 7 orange, and 3 red lines*
6. Neon—*several lines in red, orange, yellow and green*
7. Mercury—*2 blue lines, 2 yellow-green lines, 3 yellow, 2 orange, and 1 red line*

II. Describe the color of salts of these metals in a flame:
1. Lithium—*red*
2. Sodium—*orange*
3. Barium—*lime green*
4. Strontium—*scarlet*
5. Potassium—*violet (requires cobalt glass plate to see)*
6. Copper—*blue-green*

III. QUESTIONS

1. Are the results the same for candle light, magnesium light, and tungsten?
   Yes.

2. Could the colors of the flames be used to identify the different metal ions?
   Yes.

3. Why is a yellow shirt yellow?
   The pigment in a yellow shirt reflects yellow light, so the shirt appears yellow.

4. Large mercury vapor lamps are often used as night-time security lights. What color would you expect these to appear to the naked eye?
   The color of the lamp is the same as the color of the mercury discharge tube—light aqua/blue.

5. What applications can you think of for europium compounds?
   Student responses will vary. Possibilities include TV monitors, computer monitors, and red indicator lights.

6. What can we learn from a study of light from different stars?
   By studying the light given off by stars, the elements produced can be determined and then from these elements, the temperature can be estimated. We can also determine how fast the stars are receding from the Earth.

#9 Flame Tests and Analysis

Reagents

1M chloride salt solutions of Na, K, Li, and Cu (II)	10 mL/pair
1M iron (II) sulfate [$FeSO_4$]	0.5 mL/pair
18M sulfuric acid [H_2SO_4]	0.5 mL/pair
1M calcium nitrate [$Ca(NO_3)_2$]	5 mL/pair
1M silver nitrate [$AgNO_3$]	5 mL/pair
1M hydrochloric acid [HCl]	5 mL/pair
conc. hydrochloric acid [HCl]	10 mL/pair
unknown salts	
6M nitric acid (HNO_3)	(for silver reclamation)
1M sodium hydroxide (NaOH)	(for silver reclamation)
pH test papers	(for silver reclamation)

Common Materials

notebook paper	1 sheet/pair
scissors	

Special Equipment

flame test wires	1/pair
filter paper	(for silver reclamation)

Laboratory Equipment

laboratory burner
2 50-mL beakers
10-mL graduated cylinders
small test tube
1 400-mL beaker
1 pair forceps

Notes

An alternative to having each pair of students test each solution at their own lab area is to set up an area with Bunsen burners and the solutions. In this arrangement, each test solution should be supplied with its own flame loop. This also reduces contamination of the loops and preserves the integrity of the tests for the students.

The iron sulfate should be made fresh the day of use.

Caution: Store nitrates away from acids.

Disposal

The calcium sulfate precipitates can be placed into the trash. Pour the solutions down the drain with plenty of water. The silver compounds should be collected and disposed of as hazardous waste.

An alternate method for silver disposal is to reclaim it as metal. Dissolve the silver salt in 6 M nitric acid (HNO_3). Neutralize with 1 M sodium hydroxide (NaOH) to pH 7. Place a clean copper strip in the solution. Set for 20 minutes. The silver metal will wipe off the copper easily. Wipe the silver into a filter cone (see Investigation #5). Filter the silver metal. Allow it to dry.

FLAME TEST AND ANALYSIS PRELAB QUESTIONS

1) Would a flame test work to identify the specific substances in a solution if it contained multiple cations? Explain your answer.
Not with the naked eye because the different colors would be seen simultaneously and each would mask the other. However, if viewed through a spectroscope, the individual lines in the spectra could be used to make a proper identification of the elements present.

2) What is the advantage of a qual scheme that not only identifies, but simultaneously precipitates, a particular ion?
If the ions are precipitated as they are identified, they can be further analyzed, recycled, or simply discarded as impurities that have been removed from the sample.

3) Why would it be important for a technician at a water treatment plant to check the incoming water often for the presence of barium, lead, mercury, or other toxic metallic cations?
These metals should not be present in any water system, either public municipal water supplies, or waterways in nature. Because they are persistent in the environment and are toxic to most all living things, they should not be allowed to enter the food chain at any level.

4) How might the plant operators deal with the presence of such ions?
When undesirable cations are detected in a water system, the technician could add an anion that is insoluble in combination with the ion in question, thus precipitating it.

5) The chloride ion concentration in ocean water is quite high; nearly 55% of the ions present are chloride. What would you expect to be the fate of any ions of silver that found their way into the ocean? Explain your reasoning.
Silver is quite insoluble as silver chloride and would be precipitated immediately upon entering the salt water.

6) Which is more “green”: burying the silver chloride precipitate or reclaiming the silver as metal?
Reclaiming the silver will allow it to be used in another application. Burying it risks introducing the silver ion to the environment where it could enter the food chain. Reclaiming is more green.

FLAME TEST AND ANALYSIS SAMPLE REPORT SHEET

I. FLAME TEST COLORS

Sodium	*orange*	
Potassium	*violet*	*(requires a cobalt glass to view)*
Lithium	*red*	
Copper	*blue-green*	
Unknown	__________	
Unknown	__________	

	POSITIVE	NEGATIVE
II.* Nitrate test	__________	__________
Sulfate	__________	__________
Chloride	__________	__________
Acetate	__________	__________

III.* SALT CATION ANION

SALT	CATION	ANION
	__________	__________

**These answers will depend on the unknown salt used.*

IV. QUESTIONS

1. Describe some situations in which an analysis of material is necessary.
   Student responses will vary. Possibilities include determination of contaminants in water, input into manufacturing processes, quality control, and environmental studies.

2. Is a scheme necessary in those situations?
   Yes

3. Why is the $Ca(NO_3)_2$ added before $AgNO_3$ in the anion scheme?
   $AgNO_3$ will cause a precipitate while $Ca(NO_3)_2$ will not.

4. What would happen if $AgNO_3$ were added before $Ca(NO_3)_2$? (Hint: Ag_2SO_4 is an insoluble compound.)
   The Ag_2SO_4 *and AgCl would precipitate, and the test for sulfate would not work.*

5. Why is the brown ring test done on the original solution instead of the solution left after the sulfate and chloride ions have been removed?
   The nitrate is added to the solution that would cause it to test positive for that ion.

#10 Ionic versus Covalent

Reagents

ethanol [CH_eCH_2OH]	18 mL/pair
covalent (pick three or choices of your own)	2 g of each/pair
table sugar	
Styrofoam	
salicylic acid	
paraffin	
Crisco	
ionic (pick 3 or choices of your own)	2 g of each/pair
sodium chloride [$NaCl$]	
potassium iodide [KI]	
sodium sulfate [Na_2SO_4]	
potassium chloride [KCl]	
calcium chloride [$CaCl_2$]	
sodium hydrogen carbonate [$NaHCO_3$] (sodium bicarbonate)	
Epsom salts [$MgSO_4 \cdot 7\ H_2O$]	

Common Materials

aluminum foil	4”×4”/pair
can lids (optional)	

Laboratory Equipment

hot plate
6 small test tubes
microscale conductivity tester
10-mL graduated cylinder
evaporating dish

Special Equipment

Notes

Covering tin can lids with the foil forms a stable and flat surface and prevents the substances from mixing together as they are heated.

There are other choices for typical ionic and covalent compounds besides the ones that are listed. Choose compounds that are easily available but test them beforehand.

If the students put too much of a substance into the water or ethanol, they may have difficulty “seeing” whether or not the compounds dissolve because the solution becomes saturated. Be sure to emphasize the importance of dropping only one or two grains into the solvent.

Heating the compounds too long will result in burning of the sugar and decomposition of some ionic compounds. It might be wise to instruct the students to remove the heat source as soon as the first three have melted.

Disposal

All of the listed substances except thioacetamide can be thrown into the trash. All of the solutions can be flushed down the drain with lots of water. Thioacetamide should be mixed with a sodium hypochlorite solution and allowed to cool before being flushed down the drain with plenty of water.

IONIC Versus COVALENT PRELAB QUESTIONS

1) Table sugar melts at a much lower temperature than does table salt. Which of these two would you most expect to form an electrolytic solution?
   Table salt because high melting point is a characteristic of ionic compounds and ionic compounds generally dissociate to form electrolytic solutions.

2) The labels have fallen off of two containers on a lab counter. The label from one container reads "potassium iodide," an ionic compound. The other container label reads "dextrose," a covalent compound and simple sugar. Describe two tests you could perform to determine which container should receive each label.
   - *Heat small samples of each compound and compare melting points.*
   - *Dissolve a small amount of each compound in water and check for electrolytic properties.*

3) Ocean water contains a lot of different ions, but the majority of them are of only seven varieties. Sodium and chloride ions together make up nearly 82% of the ions present. What property of ionic compounds is responsible for this presence?
   Ionic compounds are highly soluble in polar substances such as water. Sodium and chlorine are common in the Earth's crust and enter the runoff water from melting glaciers, rain, and other precipitation, eventually entering streams and rivers, and ultimately the oceans.

4) Why would we not find a large amount of silicon dioxide (generally accepted as sand) dissolved in ocean water?
   Silicon dioxide is a covalently bonded compound, and although the bonds are polar, the molecules are not polar; thus it does not dissolve in water.

IONIC Versus COVALENT SAMPLE REPORT SHEET

I. PART A: MELTING POINT

Test Tube	Description of Substance	Melting Time (seconds)
sugar	*white, granular*	*5*
styrofoam	*large chunks*	6
salicylic acid	*white powder*	*10*
thioacetamide	*yellow, granular*	*5*
paraffin	*translucent chunks*	*3*
Crisco	*white gel*	*3*
sodium chloride	*white, granular*	*30+*
potassium iodide	*white, granular*	*20+*
sodium sulfate	*white, granular*	*30+*
potassium chloride	*white, granular*	*30+*
calcium chloride	*white to pink, granular*	*30+*
sodium hydrogen carbonate	*white powder*	*30+*
Epsom salts	*white, granular*	20+

II. PART B: SOLUBILITY AND CONDUCTIVITY

Test Tube	Water Soluble	Ethanol Soluble	Conductivity
Covalent			
sugar	*Y*	*SLIGHTLY*	*N*
styrofoam	*N*	*N*	*N*
salicylic acid	*N*	*Y*	*Y*
thioacetamide	*Y*	*Y*	Y
paraffin	*N*	*SLIGHTLY*	*N*
Crisco	*N*	*Y*	*N*
Ionic			
sodium chloride	*Y*	*N*	*Y*
potassium iodide	*Y*	*N*	*Y*
sodium sulfate	*Y*	*N*	*Y*
potassium chloride	*Y*	*N*	*Y*
calcium chloride	*Y*	*N*	*Y*
sodium hydrogen carbonate	*Y*	*N*	*Y*
Epsom salts	*Y*	*N*	*Y*

III. PART C: CLASSIFYING THE COMPOUNDS

Test Tube	Ionic or Covalent	Compound Name

This depends on which substances are chosen.

IV. QUESTIONS

1. What syllable in the name of some compounds is a very good indicator of the class to which it belongs?
   All the -ide *endings denote that they are ionic compounds.*

2. According to the data gathered in this lab, what is the most definitive property for use in classifying ionic and covalent compounds?
   The melting times and conductivity gave the most accurate results with the exception of salicylic acid and thioacetamide for the conductivity test. (Organic compounds are covalent, but the acids will conduct because of the hydrogen ions that they lose.)

3. Consider the solubility properties of the compounds and apply your findings to hair care. When hair is rolled and then dried, the curl leaves the hair as soon as it gets moist again. However, if the hair is treated with a permanent wave, the curl does not wash out. Which type of bond is involved in a "wet set," and which type is involved in a permanent set?
   Ionic bonds are involved with the "wet set" because they dissolve in water. However, covalent bonds are involved with a permanent set because they do not dissolve in water.

4. Is the statement that pure compounds that are liquid are covalent true? If so, give three examples.
   Mercury and bromine are elements and are liquid. Almost no ionic compound is liquid. Many covalent compounds are liquid, including water, several alcohols, hexane, ether, and other organic compounds.

#11 Iron(II) and Iron(III) Ions

Reagents

0.4 M cobalt(II) nitrate [$Co(NO_3)_2$]	5 mL/pair
12 M hydrochloric acid [HCl]	15 mL/pair
0.1 M iron(III) chloride [$FeCl_3$]	5 mL/pair
0.2 M iron(II) ammonium sulfate [$Fe(NH_4)_2(SO_4)_2$]	5 mL/pair
0.1 M potassium thiocyanate [KSCN]	5 mL/pair
0.1 M potassium ferrocyanide [$K_3Fe(CN)_6$]	5 mL/pair
0.1 M potassium ferricyanide [$K_4Fe(CN)_6$]	5 mL/pair

Common Materials (for microscale version)

clear plastic sleeve, letter size	1 sheet/pair
a sheet of plain white paper	1 sheet/pair

Laboratory Equipment

large test tube	1/pair
small test tubes	6/pair
test-tube rack	1/pair
watch glass	1/pair
10-mL graduated cylinder	1/pair
stirring rod	1/pair

Notes

This laboratory exercise works well in conjunction with the Double Replacement Reactions investigation.

It is VERY important that all solutions be freshly prepared.

The procedure can be performed in a microscale fashion, with the use of white paper in a clear plastic sleeve as the reaction surface for Parts B and C, and a microtitration plate for Part A. Replace all volumes in the procedure with "drop" in place of "mL." It is easiest to fill and label beral pipettes with the solutions and supply them for each pair of students.

Disposal

The small amount of solids generated in this experiment can be absorbed by a paper towel and thrown into the trash. They should not be washed down the drain. Solutions from this investigation are collected in a marked container and then precipitated with sulfuric acid, filtered, and disposed of in a landfill. The remaining filtrate is not toxic and can be poured down the drain with plenty of water.

IRON(II) AND IRON(III) IONS PRELAB QUESTIONS

1) According to the information in this investigation, could a chemist ever be certain of the oxidation state of a cobalt ion from its color alone? Explain.
No. Since some cobalt ions vary in color depending on the nature of the surrounding environment, color alone is not enough to determine the oxidation number of a cobalt ion.

2) In your own terms, briefly explain the differences in monatomic ions, polyatomic ions, and complex ions.
Monatomic ions are composed of a single atom with a charge. Polyatomic ions are a group of covalently bonded atoms with a charge that is assigned to the group and not a single atom. Complex ions are composed of individual ions and/or small molecules that are grouped together with a single assigned charge.

3) How might it be an advantage to a chemist if the color of a substance was a definite indication of its oxidation state?
If color alone was a definitive indication of the oxidation state of an ion, many analytical processes could be simplified.

4) Which of the following reagents contain polyatomic ions? $Co(NO_3)_2$, HCl, $FeCl_3$, $K_3Fe(CN)_6$, $Fe(NH_4)_2(SO_4)_2$, $K_4Fe(CN)_6$, KSCN. List the specific ions.
$Co(NO_3)_2$, $Fe(NH_4)_2(SO_4)_2$, KSCN, $K_3Fe(CN)_6$, and $K_4Fe(CN)_6$.

5) Which of the 12 principles of green chemistry is followed because these tests are at room temperature?
No energy is used to raise or lower temperatures.

6) What is the environmental impact of the use of energy in an experiment or synthesis?
Any use of energy, whether as a flame from natural gas or as electricity, requires that the natural resources for the energy be obtained, and in some cases that the energy be generated. This has an environmental impact. Only electricity produced by geothermal, wind, or solar-powered generators results in no net gain in carbon dioxide or pollutants such as particulate matter. Additionally, the use of ovens, hot plates, and burners may ultimately cause an increase in the air-conditioning use for the room.

IRON(II) AND IRON(III) IONS SAMPLE REPORT SHEET

I. EXPERIMENTAL DATA
PART A: Cobalt(II) Ions

mL 12 M HCl	Color
1	*peachy-pink*
2	
3	*pink*
4	*pinkish-violet*
5	
6	*violet*
7	
8	*blue-violet*
9	
10	*very blue*
11	
12	

PART B: Iron(II) and Iron(III) Thiocyanide Compounds

Compound	Color
Fe^{3+} + SCN^-	*teal-green*
Fe^{2+} + SCN^-	*yellowish-green*

PART C: Tests for Iron(II) and Iron(III) Ions

Iron Ion	Ferricyanide (Fe^{3+}) $K_3Fe(CN)_6$	Ferrocyanide (Fe^{2+}) $K_4Fe(CN)_6$
Fe^{3+} $FeCl_3$	*deep blue*	*blood red*
Fe^{2+} $Fe(NH_3)_2(SO_4)_2$	*orange red*	*teal blue*

II. QUESTIONS

1. At what point in the addition of HCl to the cobalt solution did the concentration of $[Co(H_2O)_6]^{2+}$ roughly equal the concentration of $[CoCl_4]^{2-}$? How could you tell?
   At about 6 mL, the color became halfway between pink and blue—it was violet. This would suggest half pink ion and half blue ion.

2. What visual result would you expect to see from an increase in the relative amount of water in a solution made of ethanol (alcohol) and water, plus the $[CoCl_4]^{2-}$ and $[Co(H_2O)_6]^{2+}$ ion combination?
   An increase in the percentage of water molecules surrounding the cobalt complexes would probably encourage more of the water-containing complexes to form and that would make it turn blue.

3. Why is the thiocyanate ion test only useful for a solution that contains iron ions in only one oxidation state?
   If both the 3+ and 2+ ions are present, the darker color of the 3+ ion would mask the lighter hue of the 2+ ion.

4. What test(s) could you perform to prove beyond doubt that a solution contained both Iron(II) and Iron(III) ions?
   In each of two samples of the solution to be tested, add one of the two (ferroocyanate and ferricyanate) ions. If both ions turn the solution blue, then both ions are present. If one is red, only one ion is present.

5. If a chemist is presented with a solution that is known to contain ionized iron atoms, how many steps are necessary to determine if it contains only one, or both, iron ions? What are they, and how would you proceed?
   Since the thiocyanate ion is inconclusive for both ions at once, it would be best to use only the ferro- and ferricyanate ion tests. The same procedure as for Question 4 will work.

#12 Chemical Reactions

Reagents

copper (II) carbonate [$CuCO_3$]	0.1 g/pair
sodium bisulfate [$NaHSO_4$]	0.4 g/pair
potassium ferrocyanide [$K_4Fe(CN)_6$] (or sodium ferrocyanide)	0.2 g/pair
1 M copper (II) sulfate [$CuSO_4$]	20 mL/pair
1 M potassium iodide [KI]	1.0 mL/pair
6 M hydrochloric acid [HCl]	10 mL/pair
6 M sodium hydroxide [NaOH]	27 mL/pair
pH test paper	
copper turnings	.05 g/pair
6 M nitric acid	10 mL/pair
ammonium sulfide [$(NH_4)_2S$]	1.0 g/pair

Common Materials

steel wool, fine	0.4 g/pair

Special Equipment

filter paper	1 sheet/pair

Laboratory Equipment

2 small test tubes
2 large test tubes
50-mL beaker
10-mL graduated cylinder
eyedropper
stirring rod
hot plate

Notes

Caution: Store nitric acid away from copper and ammonium sulfide.

Disposal

The solutions can be flushed down the drain with plenty of water, and the solids can be placed into a waste container to be buried in a toxic waste dump.

CHEMICAL REACTIONS PRELAB QUESTIONS

1) A chemical reaction is defined as a process in which the atomic and electronic arrangements of substances are changed. If a substance does not change during a process, should it be represented as a reactant or as a product in the equation?
No, unchanged substances are not represented in a proper chemical equation. Sometimes only an ion appears on one side of the equation, while it is part of a compound on the other side. The missing part of the compound from which the ion came was obviously left unchanged in the reaction, so it does not appear on either side of the equation.

2) Many times a substance such as a catalyst is required for reaction but is not changed in the course of the reaction. Where would be an appropriate place to indicate the presence of the substance in the equation?
A catalyst is a necessary "condition" for a reaction and is placed over the yield sign with other conditions such as temperature or pressure.

3) When testing the pH of a substance, a drop of the substance is added to a piece of pH paper. A color change indicates the pH. Could the change in the pH paper itself be a reaction? Support your answer.
Since the color of the substance changes with a change in pH, it is probably a reaction. The absorption or reflection of specific wavelengths of light is a property that has changed.

4) Some reactions are very subtle and give almost no indicators. Others are so over-the-top that there is no doubt a reaction has occurred. Consider the chemical processes involved in a Fourth of July fireworks display. List all of the reaction indicators that you can think are evident in the explosion of fireworks.
Energy changes are evident as heat, light, and sound. There are color changes in some displays. When the firework is initially launched, the rocket portion produces expanding gases. New particulate matter is seen as smoke.

5) Give an example of each of the following changes that would NOT be a chemical change:
   a) temperature change
      heating a cup of coffee in a microwave
   b) formation of a gas
      boiling water to produce steam
   c) formation of a solid
      freezing water to form ice
   d) color change
      adding food coloring to cookie dough

6) When considering the green principles outlined in Hill's text, would it be more "green" if a reaction used chemicals other than a strong acid and a strong base to generate heat?
There are many alternates to acid–base reactions as a process for the generation of heat, depending on the circumstances. Most any other method other than the use of a chemical reaction would be preferable.

CHEMICAL REACTIONS SAMPLE REPORT SHEET

I. COLOR CHANGE

1. Complete the equation:
   Cu^{2+} + Fe → *Fe^{2+}* + *Cu (s)*

2. Where did the iron (II) ions come from?
   The iron ions came from the steel wool.

II. DISAPPEARANCE OF METAL

1. What does the color of the solution indicate?
   The blue color of the solution indicates the presence of copper ions.

2. Where did the copper go?
   The copper goes into solution as Cu^{2+} ions.

III. NEUTRALIZATION

1. Initial pH _____*2*_______

   Final pH _____*10*______

IV. PRECIPITATION

1. Write the reaction.
   Cu^{2+} + ZnS → Zn^{2+} + CuS

2. What is the color of the final liquid?
   The final liquid should be colorless. The precipitate will be some form of blue (milky, chalky, dark, etc.).

V. HEAT PRODUCTION

1.

			pH
6 M HCl			*2*
6 M HCl	+	2 mL NaOH	*2*
		4 mL NaOH	*2*
		6 mL NaOH	*2*
		8 mL NaOH	*2*
		10 mL NaOH	*2*
		12 mL NaOH	*10*
		14 mL NaOH	*10*
6 M NaOH			*10*

2. Write the reaction.
   HCl + NaOH → H_2O + NaCl

3. Why were you instructed to place the beaker on the lab bench?
   The beaker was placed on the lab bench because it gets very hot.

VI. GAS PRODUCTION

1. Write the equation for the reaction.
   $2NaHSO_4 + Fe \rightarrow 2Na^{+} + (SO_4)^{2-} + H_2 + Fe^{2+}$

2. Where did the hydrogen come from?
   The hydrogen came from the sodium hydrogen sulfate reacting with the iron. This formed $H_2(g)$.

3. Where did the iron(II) ions come from?
   The iron ions came from the steel wool.

#13 Double Replacement Reactions

Reagents

1 M copper nitrate [$Cu(NO_3)_2$]	2 mL/pair
1.5 M sodium hydroxide [NaOH]	3 mL/pair
0.5 M hydrochloric acid [HCl]	27 mL/pair
sodium bicarbonate [$NaHCO_3$]	0.5 g/pair
1 M calcium nitrate [$Ca(NO_3)_2$]	2 drops/pair
1 M sodium carbonate [Na_2CO_3]	6 drops/pair
1 M potassium iodide [KI]	10 drops/pair
1 M lead (II) nitrate [$Pb(NO_3)_2$]	2 drops/pair
1 M copper (II) sulfate [$CuSO_4$]	6 drops/pair
1 M sodium chloride [NaCl]	2 drops/pair
phenolphthalein	1 drop/pair

Common Materials

clear plastic sleeve, letter size	1 sheet/pair
a sheet of plain white paper	1 sheet/pair

Laboratory Equipment

10-mL graduated cylinder	droppers for solutions
100-mL graduated cylinder	50-mL beaker
3 small test tubes	wash bottle
Thermometer	

Notes

This laboratory exercise works well in conjunction with the Single Replacement Reactions exercise. It can be started as soon as the "5 minute" check has been made in the other lab.

Disposal

Except for the lead solutions, the solutions from this investigation are not extremely toxic and can be poured down the drain with plenty of water. The small amount of lead solution used in this experiment can be absorbed by a paper towel and thrown into the trash. It should not be washed down the drain.

DOUBLE REPLACEMENT REACTIONS PRELAB QUESTIONS

1. When solutions of two dissolved salts are combined and result in the formation of an insoluble salt in the form of a precipitate, where are the remaining two ions?
   The remaining two ions are not changed in the reaction and are referred to as spectators.

2. Suppose the precipitate in Question 1 is filtered from the solution that originally contained four ions. How could the salt that remains in solution be recovered?
   The liquid in the solution, that is, the solvent that is probably water, could be evaporated. The remaining two ions would then be free to form an ionic solid.

3. If there are ions of several soluble salts in one solution, could you devise a way to separate only one of the ions from the solution?
   If there is another ion that forms an insoluble salt with the desired ion but not with any of the others, the ion in question could be precipitated as a solid and filtered to remove it.

4. Why is the volume of reactants used not important to the results in Part A of this lab?
The procedure is not intended to be quantitative, which is a measurement of how much ion is present, but rather qualitative. The production of a precipitate is not dependent on how much of the two ions are introduced into a common vessel.

5. From the procedure in Parts B and C, why do you think there are cautionary statements on any household products that are moderately or strongly acidic or basic?
The heat that can result from diluting or neutralizing some acids and bases can cause splattering and spewing. Gases are also produced from the combination of specific acids and bases that are harmful. (Especially anything containing hydrochloric or muriatic acid, HCl, and the base ammonia as the result is a caustic gas, ammonium chloride.)

6. Are cleaning supplies strongly acidic or strongly basic following the green principle of designing safer chemicals?
No. Production of heat, possible by-products, pH changes, or damage to other substances are all potential results when using strongly acidic or basic compounds.

DOUBLE REPLACEMENT REACTIONS SAMPLE REPORT SHEET

I. PART A: Predictions and Results Based on Solubility Rules

SOLUTIONS	**PREDICTION**	**OBSERVATION**	**PRECIPITATE**
calcium nitrate & sodium carbonate	*Y*	*cloudy white*	$CaCO_3$
potassium iodide & lead (II) nitrate	*Y*	*mustard yellow*	PbI_2
copper (II) sulfate & sodium hydroxide	*Y*	*cloudy and blue*	$Cu(OH)_2$
sodium chloride & potassium iodide	*N*	*nothing*	*none*
sodium carbonate & potassium iodide	*N*	*nothing*	*none*
copper sulfate & sodium carbonate	*Y*	*cloudy and gel like*	$CuCO_3$
potassium iodide & sodium hydroxide	*N*	*nothing*	*none*
copper sulfate & potassium iodide	*Y*	*brown*	CuI_2

II. PART B: Gas-Forming Reactions

	Temperature		Temperature
HCl Solution	__________	$NaHCO_3$ and HCl mixture	__________

Observations

Bubbles form rapidly, temperature drops.

Na_2SO_3 and HCl mixture __________

Observations

Bubbles form rapidly, temperature drops.

III. PART C: Acid–Base Reactions

	Indicator Color
HCl	*colorless*____
HCl and NaOH mixture	*pink*________

Observations

Slight temperature rise

IV. QUESTIONS

1. Out of the eight precipitation reactions, how many did you correctly predict? _______
2. In how many of those reactions were you able to correctly determine the precipitate even before observing the reaction? ________
3. Describe the similarities and differences of the precipitates that were formed. Did you find any characteristics consistent with compounds containing a particular ion?
   The copper carbonate was gel-like. The precipitates are various colors. Sodium and potassium are always soluble.
4. Explain why such tools as the Periodic Table, Activity Series, and solubility charts are important to the work of a research chemist.
   These tools let chemists know what reactions will occur without having to try them out.
5. Write a balanced equation for each reaction that occurred in Part A. (If no precipitate formed, give the reactants and write "no reaction" after the yield sign.)

   a. $Ca^{2+} + CO_3^{2-} \rightarrow \mathit{CaCO_3}$
   b. $Pb^{2+} + I^{-} \rightarrow \mathit{PbI_2}$
   c. $Cu^{2+} + OH^{-} \rightarrow \mathit{Cu(OH)_2}$
   d. $Na^{+} + Cl^{-} + K^{+} + I^{-} \rightarrow$ *no reaction*
   e. $Na^{+} + CO_3^{2-} + K^{+} + I^{-} \rightarrow$ *no reaction*
   f. $Cu^{2+} + CO_3^{2-} \rightarrow \mathit{CuCO_3}$
   g. $Na^{+} + OH^{-} + K^{+} + I^{-} \rightarrow$ *no reaction*
   h. $Cu^{2+} + I^{-} \rightarrow \mathit{CuI_2}$

#14 Mole Relationships

Reagents

1M calcium chloride [$CaCl_2$]	30 mL/pair
1M sodium sulfate [Na_2SO_4]	40 mL/pair
1M calcium nitrate [$Ca(NO_3)_2$]	20 mL/pair
1M copper (II) sulfate [$CuSO_4$]	20 mL/pair

Special Equipment

filter paper	4 pieces/pair
Buchner funnel	
filter flasks	2 flasks/pair
hoses for filter system	1 set/pair

Laboratory Equipment

centigram balance	
10-mL graduated cylinder	1/pair
small beakers	2 beakers/pair
watch glass	1/pair

Disposal

Pour together all the solutions of calcium and sulfate. Filter the calcium sulfate precipitate. Discard all precipitates in a container to be buried in a landfill. Pour the solutions down the drain with lots of water.

MOLE RELATIONSHIPS PRELAB QUESTIONS

1) What remains in solution at the end of step 1?
Na^+ and Cl^- ions remain in solution.

2) If sodium ions and chloride ions were left in solution, how could they be recovered?
The solvent (water) could be evaporated to leave the ionic sodium chloride crystals behind.

3) If the reaction was 1M chloride ion with 1 M calcium ion, would you still use 10 mL of each solution?
No, because there are two chloride ions per molecule in the calcium chloride solution. You would only need a 0.5 M solution to have 1.0 M of chloride ions.

4) If it was determined that the water supply for a city had a minute level of lead ion contamination, the water treatment plant would likely attempt to drop the lead out by precipitating it. Would they want the negative ion they add to the water or the existing lead ion to be the limiting reactant? Explain your reasoning.
It would be more efficient to make the lead the limiting reactant by adding more of the negative ion than is needed. This will also function to help ensure that the maximum amount of lead is removed from the system.

MOLE RELATIONSHIPS SAMPLE REPORT SHEET

I. LIMITING REACTANT *Answers will vary.*

Record the mass of $CaSO_4$ + Filter paper for each run:

Run #1	5 m mol Na_2SO_4	+	5 m mol $CaCl_2$	*8.84* g
Run #2	10 m mol Na_2SO_4	+	10 m mol $CaCl_2$	*9.69* g
Run #3	20 m mol Na_2SO_4	+	10 m mol $CaCl_2$	*9.37* g
Run #4	10 m mol Na_2SO_4	+	20 m mol $CaCl_2$	*10.51* g

Precipitation in filtered solution

Mark each solution that produces a large amount of precipitate.

				$CuSO_4$	$Ca(NO_3)_2$
Run #1	5 m mol Na_2SO_4	+	5 m mol $CaCl_2$		
Run #2	10 m mol Na_2SO_4	+	10 m mol $CaCl_2$		
Run #3	20 m mol Na_2SO_4	+	10 m mol $CaCl_2$		*X*
Run #4	10 m mol Na_2SO_4	+	20 m mol $CaCl_2$	*X*	

IV. QUESTIONS

1. The last three $CaSO_4$ masses should be the same except for a small water–mass difference. Explain this statement.
   The mass of the filtered $CaSO_4$ product is the same in all three reactions because the limiting reactant in each reaction determines the amount of product that can be formed. In the third run the Ca^{2+} is the limit, while in the fourth run the SO_4^{2-} is the limit.

2. What should be the difference in $CaSO_4$ mass between run 1 and run 2?
   The product mass in the first run should be one-half that of the second run because the moles of each reactant used were half.

3. Which compound was in excess in run 3?
   sodium sulfate

4. Which compound was the limiting reagent in run 3?
   calcium chloride

5. Which compound was in excess in run 4?
   calcium chloride

6. Which compound was the limiting reagent in run 4?
   sodium sulfate

#15 Molar Volume of a Gas

Reagents

magnesium ribbon	about 15 cm/pair
6.0 M hydrochloric acid [HCl]	about 35 mL/pair

Common Materials

sandpaper	1 piece/pair
thread	3 12" lengths/pair
ruler (cm)	1 or 2/lab

Special Equipment

50 mL eudiometer	1/pair
1-hole rubber stopper to fit eudiometer	1/pair
1000-mL graduated cylinder	1/pair
wash bottle with distilled water	1/pair
barometer	1/lab

Laboratory Equipment
- 10-mL graduated cylinder
- 600-mL beaker
- ring stand
- utility (or burette) clamp
- thermometer
- centigram balance

Notes

The thickness of magnesium ribbon varies slightly from supplier to supplier. It would be wise to do a trial run on the lab to ensure that the length of magnesium used will not produce more than about 48 mL of hydrogen gas.

Caution students about putting too much pressure on the stopper when inserting it into the eudiometer so as not to split the opening of the glass tube.

Any containers deep enough to hold water to reach the level of the gas in the eudiometers can be substituted for the 1000-mL graduated cylinders.

Disposal

There are no harmful substances at the end of the lab. The fluid from the eudiometer tube can be flushed down the drain with plenty of water.

MOLAR VOLUME OF A GAS PRELAB QUESTIONS

1) If gases expand to fill any container, what invisible barrier keeps an atmosphere around the Earth?
   Gravity holds the atmosphere close to the Earth. The atmosphere is denser closest to the Earth's surface because that is where gravity is the strongest. There is no barrier.

2) If the gas collected dissolves in water, as does oxygen or carbon dioxide, what effect will that property have on the final experimental value for the volume of gas produced?
   The measured volume will be smaller than the actual volume of gas produced because some will remain dissolved in the water.

3) A eudiometer is marked in graduations of 0.1 mL. If a water meniscus sits on the third mark above the 46-mL mark, what is the volume reading?
   45.7 mL because the eudiometer is marked to show the amount of gas collected with the 0.0 mark at the top.

4) If bubbles of the gas are allowed to remain on the inner surface of the tube and on the thread, what effect will that have on the measured volume of gas in the tube?
   The effect will be the same as for dissolved gases; they will not be included in the measured volume.

5) In Procedure 4, the mass of the magnesium is limited to less than 0.040 g. Would it be against atom economy of the green principles listed in Hill's text to have more than the amount of magnesium that would react?
   Any waste of material is not in keeping with green chemistry.

6) Give an explanation for the fact that some hydrogen bubbles could actually escape *down* through the hole in the stopper.
   If the reacting magnesium is quite low in the tube, the tiny bubbles could get caught in the flow of water that is exiting through the hole in the stopper.

7) List as many sources of error as possible in the data for your experimental value for the moles hydrogen produced and for the determination of the standard molar volume of a gas.
   Error could be introduced with bubbles stuck to the side of the eudiometer, on the stopper, and on the thread that are not measured above the meniscus; bubbles that escaped through the hole in the stopper; impurities remaining on the magnesium ribbon, not allowing the water in the tube to reach the same temperature as the water outside the tube so that the temperature measurement is incorrect; not having the inner and outer surfaces of water at the same level; difficulty reading the meniscus inside the eudiometer; an inaccurate or uncalibrated barometer; human error in all measurements.

MOLAR VOLUME OF A GAS SAMPLE REPORT SHEET

I. DATA TABLE *(Student data will vary.)*

	Trial 1	Trial 2	Trial 3
Mass of magnesium (g)	*0.04343 g*		
Gas volume (mL)	*38.72 mL*		
Temperature (^oC)	*19.2 ^{o}*		
Barometric pressure (inHg)	*28.8 inHg*		

II. CALCULATIONS

1) Find the moles magnesium consumed by dividing the mass of magnesium by its molar mass (MM_{Mg} = 24.31 g/mol).
Trial 1 0*.0017869 mol*

2) Convert your volume in mL to Liters by dividing by 1000 mL/L.
Trial 1 *0.03872 L*

3) Your thermometer is probably calibrated in degrees Celsius. Convert the experimental temperatures to Kelvin by adding 273.15.
Trial 1 *292.4 K*

4) If your barometer is calibrated in inches of mercury, convert the measurement to centimeters of mercury by multiplying by 2.54 cm/in. Then rewrite the measurement in millimeters of mercury, "mmHg", by multiplying by 10.
Trial 1 *731.4 mmHg*

5) Find the experimental temperatures on the table of equilibrium vapor pressures for water and subtract the partial pressure of the water to get the "dry" pressure of the hydrogen gas for each trial.
Trial 1 *714.7 mm Hg*

6) Convert the dry gas pressures from torr (mmHg) to atmospheres (atm) by dividing the value by 760 mmHg/atm.
Trial 1 *0.940 atm*

7) Use the relationship n = (PV)/(RT) where P is the corrected pressure in atm, V is the experimental volume in liters, T is the Kelvin temperature, and R is 0.08206 L·atm/mol·K to find the moles of hydrogen gas collected.
Trial 1 *0.0015178 mol*

8) Examine the balanced equation for the reaction between magnesium and hydrochloric acid in the experimental background material. There is a 1:1 mole ratio of magnesium consumed to hydrogen gas produced. Using your determination of the moles magnesium consumed in the reaction as the theoretical, or expected, yield T, and the moles hydrogen gas actually produced as A, find a percent yield for each trial using A/T x100% = % Yield.
Trial 1 *84.9 %*

9) Use the Combined Gas Law to find the equivalent volume of hydrogen gas at standard conditions. Use the relationship $V' = (PVT')/(TP')$ where P, V, and T are your experimental values and T′ is 273 K, while P′ is 1.00 atm. This is the volume of a gas at standard conditions (Standard Volume).
Trial 1 *0.03402 L*

10) Using a ratio, find the experimental volume of one mole of hydrogen gas at standard conditions (Exp. SMV), where EV is your experimental standard volume of hydrogen and Emol is the experimental moles of hydrogen as follows:

$$EV/Emol = SMV/1.00mol \text{ or } (EV/Emol)/1.00 \text{ mol} = SMV$$

Trial 1 *22.41 L*

11) Using 22.4 L and the accepted value A, and your Exp. SMV as the experimental value E, find a percent error for your value of the Standard Molar Volume of a gas using the relationship:
$\% \text{ Error} = |A-E|/A \times 100\%$ where $|A-E|$ is the absolute value of the difference in the accepted and experimental values.
Trial 1 <u>*0.04464 %*</u>

III. POST-LAB QUESTIONS

1) Why was it necessary to tie the magnesium and to pin it against the mouth of the eudiometer with the thread? What does this say about the relative densities of the metal and the water?
Although magnesium is more dense than water, it might still float to the top ot the reaction tube once the reaction has started because of the buoyancy provided by the hydrogen gas bubbles that cling to the magnesium. Tying the magnesium ribbon will help to minimize the amount of magnesium that floats to the top.

2) When the eudiometer tube was being raised in the large cylinder to match the water levels inside and outside, what happened to the inner water level? Use gas pressures to explain this observation.
As the eudiometer tube is raised to make the water in the tube match the level of the water outside the tube, it is the higher pressure of the gas on the surface of the water in the open container that supports the water in the tube. As the water levels get closer, less pressure is exerted through the water in the tube and the pressure on the gas is therefore reduced so that it expands, lowering the water level in the tube.

3) Examine the table of equilibrium vapor pressures for water supplied by your lab instructor. Assume all measurements for the hydrogen gas were kept constant except for the room temperature. Would an increase of 5 degrees have meant the presence of more or less hydrogen gas?
At a higher temperature, the amount of water vapor would increase. If the pressure and volume are constant, this means less of the gas mixture is actually hydrogen.

#16 Hydrated Compounds

Reagents

hydrated copper (II) sulfate crystals [$Cu(SO_4)\cdot 5H_2O$]	3 g/pair
hydrated magnesium sulfate crystals [$Mg(SO_4\cdot 7H_2O$]	3 g/pair

Common Materials

Laboratory Equipment

evaporating dish	2/pair
watch glass	2/pair
tongs	1/pair
eyedropper	1/pair
hot plate	1/pair
centigram balance	
desiccator	

Notes

Disposal

HYDRATED COMPOUNDS PRELAB QUESTIONS

1) If heat is required to remove the water molecules from hydrated crystals, what would you expect to be the result of dropping water onto a dehydrated crystal?
If dehydration is an endothermic process, then re-hydration would be expected to be exothermic.

2) Salt allowed to sit undisturbed in a shaker can "cake" or become stuck together. This happens because most salt contains a bit of calcium chloride and magnesium chloride as minor impurities. Explain what is happening.
The calcium chloride and magnesium chloride are hygroscopic and function as desiccants.

3) Once the heating of a mass of hydrated crystal is completed, it will be placed into a desiccator to cool. From the name of the device, what do you think its purpose is?
A desiccant is a substance that absorbs moisture from the surrounding atmosphere. The desiccator contains a desiccant that will absorb moisture from the air that is sealed within the container as a second material dries and thereby keep the humidity of the air low enough to continue removing moisture from the second material.

4) In the bottom of the desiccator is a solid substance. Sometimes this substance is a commercial product called Dririte^{TM}, but sometimes it is anhydrous calcium chloride. What is the function of the substance in the bottom of the desiccator? Name another substance that could be used.
The substance functions to absorb moisture from the surrounding air so that it facilitates the drying of any substance sealed in the desiccator. Magnesium chloride could also be used.

5) Does the substance in the bottom of the desiccator necessarily need to be discarded once it has fully hydrated? What is an alternative?
The desiccant could be heated to drive out the moisture it has absorbed and thus would be "regenerated."

6) If the substance used in the desiccator can be reused, is that in agreement with the principle of green chemistry that indicates that a renewable material is better than using new material as listed in Hill's text?
Regenerating the desiccant prevents the need to use more chemical. This is definitely a use of renewable material. Yes, it is in keeping with the green principle.

7) The manager of a laboratory stockroom might prefer to have the hydrated form of a crystal rather than the anhydrous form. What might be a practical reason for that preference?
A hygroscopic or deliquescent substance will absorb any moisture that comes into contact with it and thereby change its formula weight and properties. It is much more efficient to store a fully hydrated crystal whose constitution and mass will be constant.

HYDRATED COMPOUNDS SAMPLE REPORT SHEET

Parts A & B Data Table

	Magnesium Sulfate	Copper Sulfate
Appearance before heating	*white*	*blue*
Appearance after heating	*white*	*light gray*
Observations during heating	*condensation*	*condensation*
Mass of dish and glass	________ g	________ g
Mass of dish, glass, and salt before heating	________ g	________ g
Mass of dish, glass, and salt after first heating	________ g	*(N/A)*
Mass of dish, glass, and salt after final heating	________ g	________ g

Parts A & B Calculations

1) For each salt, calculate the mass of the *hydrated salt* by subtracting the mass of the evaporating dish and watch glass from the total mass before heating.
Magnesium sulfate ____________g Copper(II) sulfate ____________g

2) For each salt, calculate the mass of the *anhydrous salt* by subtracting the mass of the dish and glass from the final mass after heating.
Magnesium sulfate ____________g   Copper(II) sulfate ____________g

3) For each salt, calculate the mass of water removed from the salt by finding the difference in the masses of the hydrated and anhydrous forms of the salt.
Magnesium sulfate ____________g   Copper(II) sulfate ____________g

4) For each salt, find the moles of anhydrous salt by dividing the mass of anhydrous salt by its molar mass. Record these values. (The molar mass of $CuSO_4$ is 159.60 g/mol, and the molar mass of $MgSO_4$ is 120.38 g/mol.)
Magnesium sulfate ____________mol   Copper(II) sulfate ____________mol

5) For each salt, find the moles water that was removed by dividing the mass of the water removed by its molar mass (18.0152 g/mol.)
Magnesium sulfate ____________mol   Copper(II) sulfate ____________mol

6) For each salt, find the ratio of moles salt to moles water in the hydrate by dividing the moles water by the moles of anhydrous salt. Record these values expressed and "1: __".
Magnesium sulfate 1.00:___7_____   Copper(II) sulfate 1.00:___5_____

7) Find the percent error for your experimental values by the following formula:
% Error = |A-E| / A x 100%
where A is the accepted value and E is the experimental value for the hydrate number. The accepted hydrate values for magnesium chloride and copper sulfate are 7.00 and 5.00, respectively.
Magnesium sulfate ______<1_____%   Copper(II) sulfate ____<1______%

8) Round the mole ratio to the nearest whole number to get a hydrate number and report your experimental empirical formula for the hydrated salt in the form of XY·nH20 where n is the experimental hydrate number.
Magnesium sulfate ___*$MgSO\cdot 7H_2O$*___   Copper(II) sulfate ___*$CuSO_4\cdot 5H_2O$*___

9) Gather hydrate numbers from several classmates as comparison to your work.
Magnesium sulfate ____________   Copper(II) sulfate ____________
Magnesium sulfate ____________   Copper(II) sulfate ____________
Magnesium sulfate ____________   Copper(II) sulfate ____________
Magnesium sulfate ____________   Copper(II) sulfate ____________
Magnesium sulfate ____________   Copper(II) sulfate ____________

Post-Lab Questions

1) Once the mass of the clean and dry evaporating dish and watch glass have been recorded, they are handled only with clean tongs. Why is this important?
Any oils or foreign particles on the dish will alter the mass and change the final results.

2) What is the purpose of the watch glass?
The watch glass prevents loss of material from spattering.

3) If the final mass is taken while the dish is still hot, what would be the likely effect on the experimental ratio of moles water to moles salt? Would it be too high or too low? Explain the logic of your reasoning.
Taking a mass on a substance while it is hot will cause the balance to read less than the actual mass. This would cause the results to appear as if more water was removed than was actually removed. Therefore, the experimental ratio of water to salt would be too high.

4) Examine data from some classmates. Does there seem to be consistency in the hydrate numbers you found for the two salts? Comment. What law does this data support?
There should be good consistency in all numbers, and this is an example of the Law of Definite Proportions.

5) Why was the magnesium sulfate heated until the mass was constant?
If there is no indication of dryness by the color of the substance, the only method of determining if it is totally dehydrated is to heat it repeatedly until there is no more water to drive off.

6) A student runs an experiment and heats white magnesium sulfate ($MgSO_4$) crystals and measures the masses before and after heating. During the process, she notices condensation on the watch glass covering the substance. The masses she records are 18.49 and 9.03 g.
i. What two evidences are there that the substance was a hydrated crystal?
Condensation and reduced mass
ii. Calculate the hydrate number for the substance.

7) From your observations when adding water to the copper sulfate, explain the energy changes taking place.
The substance returned to its original blue color, popped and hissed, and released steam. From the sound and steam, it can be deduced that heat was released when the water was reabsorbed into the crystal.

8) Cobalt chloride has an interesting property. A $CoCl_2 \cdot 4H_2O$ is blue, while $CoCl_2 \cdot 6H_2O$ is pink. Can you think of a good use for this substance?
It is often used as an indicator for the moisture content of a desiccant.

#17 Solubility

Reagents

alcohol (ethyl or isopropyl)	5 mL/pair

Common Materials

syrup	10 mL/pair
cooking oil	10 mL/pair
liquid soap	1 mL/pair
chewing gum	1 stick/pair
peanut butter	1 small scoop/pair
cloth	8 in. × 8 in./pair

Special Equipment

Laboratory Equipment

balance
50-mL beaker
eyedropper
stirring rod
10-mL graduated cylinder

Notes

Disposal

The solutions in this investigation are not toxic and can be poured down the drain with plenty of water.

SOLUBILITY PRE-LAB QUESTIONS

1) If the liquid layers in the procedure were added to the cylinder in the reverse order, that is, with the densest first and the least dense last, what would you expect to happen at the interface of each liquid?
   The substances would naturally move to position the most dense substance on the bottom and the least dense on top. The movement would cause the water and syrup to mix so that there would only be two layers after the substances rearranged themselves.

2) What properties would you expect the molecules of two miscible liquids to have in common?
   They would have similar properties of polarity. They would either both be polar or both be nonpolar.

3) Suppose a garment has been stained by a substance that is made of nonpolar molecules. What kind of molecule would you expect to be most effective at removing the stain?
   A nonpolar molecule would be most efficient in removing a nonpolar stain. Dry cleaning agents are nonpolar solvents.

4) If a liquid with a density of 1.2 g/mL is mixed in equal proportions with a liquid that has a density of 0.9 g/mL, what would you expect of the combined density if
   a) the two liquids are completely miscible with additive volumes?
      The new density would be an average of the two densities.
   b) the two liquids are completely miscible without additive volumes?
      The new density would be between the two original densities but not necessarily the average.

c) the two liquids are completely immiscible?
Two distinct layers would form, with the more dense substance on the bottom.

d) the two liquids are miscible but of greatly varying viscosities?
The two substances would tend to form two layers with the more dense substance on the bottom, but the interface between the two layers might not be well defined because of some mixing.

SOLUBILITY SAMPLE REPORT SHEET

I. DENSITIES OF LIQUIDS

	SYRUP	WATER	OIL
mass of cylinder and liquid	________g	________g	________g
minus mass of cylinder	−________g	−________g	−________g
mass of liquid	________g	________g	________g

$$\text{Density of liquids} = \frac{\text{mass of liquid}}{10\ \text{mL}}$$

	SYRUP	WATER	OIL
=	*1.382* g/mL	*1.034* g/mL	*0.959* g/mL

Which liquid is the most dense? *syrup*

Which liquid is on the bottom of the beaker? *syrup*

Which two layers mix in Procedure 4? *water and syrup*

	Top Layer	Bottom Layer
mass of cylinder and liquid	________g	________g
minus mass of cylinder	−________g	−________g
mass of liquid	________g	________g

$$\text{Density of liquids} = \frac{\text{mass of liquid}}{5\ \text{mL}}$$

	Top Layer	Bottom Layer
=	*1.106* g/mL	*1.094* g/mL

Is the density of the mixture an average of the original two layers?
Yes

Does soap act as an emulsifying agent?
Yes

Why would forming an emulsion be important when washing greasy dishes?
The grease must be held to the water by the soap in order to remove it from the dishes.

II. MIXTURE VOLUMES

Volume water 5 mL
Volume alcohol 5 mL
10 mL
Measured volume *9.5* mL

If the volumes are different, explain why?
The molecules will be held tightly together and will fit together closely enough to need less space. It is as if the H_2O molecules fit into holes between the alcohol molecules.

III. QUESTIONS

1. Which ingredient in peanut butter allows it to remove gum? Ingredients are listed on the peanut butter jar.
   Oil

2. Why would using peanut butter rather than scissors be better for removing gum from hair?
   Scissors would cut the gum out while peanut butter will remove the gum and leave the hair undamaged.

3. Why would using peanut butter rather than water be better for removing gum from hair?
   The water will not dissolve the gum, but the peanut butter will dissolve the gum.

4. Would vegetable oil work in a similar manner to peanut butter to remove gum?
   Yes, oil would work, but it is a liquid. The solid peanut butter is easier to work with.

5. A substance called lecithin allows the vinegar and oil in some dressings to remain mixed. What qualities would you expect lecithin to have in common with soap?
   It is also an emulsifier. In this way, lecithin can attract the oil with one end of the molecule and attract the polar molecules of the vinegar with the other end, allowing them to mix.

#18 Viscosity

Reagents

rubbing alcohol (isopropyl)	30 mL/pair

Common Materials

soluble starch	30 g/pair
vegetable oil	30 mL/pair
honey	30 mL/pair
dishwashing liquid	30 mL/pair
ketchup	30 mL/pair
glass marbles	3/pair

Special Equipment

stopwatch	1/pair
test-tube rack	1/pair
identical test tubes, large	11/pair
stopper to fit large test tube	1/pair
hot plate	

Notes

The soluble starch solution is made by stirring the powdered starch into cold water. Heat the solution with occasional stirring to a slow boil until the solution is clear. When the solution cools and thickens, it will appear white and it will be ready to use. It cannot be made up more than one or two days ahead because it will mold easily.

Other common materials that work well are shampoo, conditioner, mustard, and liquid fabric softener. .

Disposal

All the solutions in this experiment can be flushed down the drain with plenty of water.

VISCOSITY PRE-LAB QUESTIONS

1) Starch is a linear protein molecule. Knowing this, what might you expect to be the effect on the viscosity of a starch solution if it was allowed to sit for a while?
The linear molecules will settle easily and because of the amount of contact between the molecules, they may be able to form numerous and strong attractions to each other. The result would be that the solution would become more and more viscous as it sits.

2) If two objects of the same material, density, and mass, but of different shape, are each allowed to fall through the same viscous fluid, how would shape affect the apparent measurement of the fluid's viscosity?
It would be expected that the more aerodynamic shape would meet with less resistance and friction from the fluid and would therefore have a slightly increased speed.

3) Considering the cohesive forces between particles in a highly volatile fluid—that is, one that evaporates easily—what would you suspect concerning the viscosity of the fluid? Explain your reasoning.
If the fluid evaporates easily, the cohesive forces must be small. It would then be expected that the viscosity is very low because cohesive forces are the major factor in viscosity.

4) Air is not considered to be a very viscous fluid. However, automobiles and planes are designed with aerodynamics in mind. How does the speed of an object in a fluid affect the force of resistance it experiences from the viscosity?
As an object moves at increasing speeds through any fluid, the number of fluid particles impacting and then moving over the surface is increased. This increases the total frictional force that is then measured as resistance.

5) This experiment was designed to use materials that if mixed with plenty of water would not contaminate the environment. How does this relate to the 12 principles of green chemistry listed in Hill's text? The principles of green chemistry are designed to protect the environment and people from harm.
There are no toxic substances, or substances that would require a long period of time to decompose. Water is already present in the environment, and starch will be readily broken down by bacteria and other organisms.

VISCOSITY SAMPLE REPORT SHEET

I. PART A: PRODUCING A VISCOSITY SCALE

Data Table

Mass of one test tube: ____________ g Volume in each tube: 30.0 mL

Test Tube	Water/starch ratio	Mass	Time for Marble to Reach Bottom
1	0:30	______ g	*>5 hours* seconds
2	3:27	______ g	*hours 32 min* seconds
3	6:24	______ g	*28 min* seconds
4	9:21	______ g	*3 min* seconds
5	12:18	______ g	*47 sec* seconds
6	15:15	______ g	*6.5 sec* seconds
7	18:12	______ g	*2.1 sec* seconds
8	21:9	______ g	*1.2 sec* seconds
9	24:6	______ g	*0.8 sec* seconds
10	27:3	______ g	*0.7 sec* seconds
11	30:0	______ g	*0.8 sec* seconds

Calculations Table
Density Calculation

Test Tube	Total mass	–	test tube mass	=	mass of solution	Density = mass/volume
1	_______ g	–	_______ g	=	________ g	__________ g/mL
2	_______ g	–	_______ g	=	________ g	__________ g/mL
3	_______ g	–	_______ g	=	________ g	__________ g/mL
4	_______ g	–	_______ g	=	________ g	__________ g/mL
5	_______ g	–	_______ g	=	________ g	__________ g/mL
6	_______ g	–	_______ g	=	________ g	__________ g/mL
7	_______ g	–	_______ g	=	________ g	__________ g/mL
8	_______ g	–	_______ g	=	________ g	__________ g/mL
9	_______ g	–	_______ g	=	________ g	__________ g/mL
10	_______ g	–	_______ g	=	________ g	__________ g/mL
11	_______ g	–	_______ g	=	________ g	__________ g/mL

Viscosity Scale
Divide each time measurement by the smallest one to get a ratio. Give the ratio to three significant digits.

Test Tube	Viscosity Scale
1	___________
2	___________
3	___________
4	___________
5	___________
6	___________
7	___________
8	___________
9	___________
10	___________
11	___________

II. Part B: DETERMINATION OF RELATIVE VISCOSITIES OF COMMON SUBSTANCES

Data Table

Substance	Time to Bottom of Tube	Mass Measurement	Relative Viscosity from Scale
Vegetable oil	*14 s* seconds	____ g	____
Rubbing alcohol	*0.6 s* seconds	____ g	____
Honey	*3 h 12 m* seconds	____ g	____
Dishwashing liquid	*29 s* seconds	____ g	____
Ketchup	*> 5 hrs* seconds	____ g	____
Mustard	*> 5 hrs* seconds	____ g	____
Molasses	*> 5 hrs* seconds	____ g	____
Hair conditioner	*47 s* seconds	____ g	____

Calculations Table

Substance	Mass –	Test Tube Mass =	Mass of Substance	Density = Mass/Volume
Vegetable oil	____ g –	____ g =	____ g	____ g/mL
Rubbing alcohol	____ g –	____ g =	____ g	____ g/mL
Honey	____ g –	____ g =	____ g	____ g/mL
Dishwashing liquid	____ g –	____ g =	____ g	____ g/mL
Ketchup	____ g –	____ g =	____ g	____ g/mL

III. QUESTIONS

1. Do the densities of the starch solutions vary directly with the viscosity ratings?
   Yes

2. Do the densities of the other substances vary directly with the viscosity ratings?
   Not all. The ketchup has the highest rating but not the highest density.

3. If you answered differently for the first two questions, give an explanation for the apparent discrepancy.
   The molecular structure of ketchup allows the proteins to form a matrix out of the chains that have formed strong attractions to one another.

4. What factors, other than the viscosity of the fluid, affect how fast the marble falls down the test tube? Does this in any way affect the results as they have been gathered? Why or why not?
   Density differences in the fluids also affect the rate at which an object falls through them. No, this does not affect the results because the marble is more dense than any the fluids measured.

5. How would the data be affected by using a pointed object to fall through the fluids rather than a spherical one?
The shape might allow it to move through the fluids more easily because of the streamlined aerodynamics.

6. Relate your answer in Question 5 to the general shape of fish and submarines. How is this applicable to automobiles that also travel through a gaseous fluid – air?
The shape of fish and submarines allows them to move through the fluids with less drag caused by the viscosity. Automobiles can also have less drag if they are shaped aerodynamically, thus saving energy.

7. Honey and molasses are both sugar solutions. Knowing this fact and that molasses is more dense than honey, can you make a prediction as to the viscosity of molasses? Why is this generalization okay to use when comparing these two substances?
Since molasses is more dense, it must have a higher percentage of sugar and will then be more viscous. Since both solutions are made of mostly sugar and water, a direct comparison can be made because no other major factor affects the viscosity. This is similar to our experimental plan in that the viscosity difference is a function of the ratio of sugar to water, rather than starch to water as we have examined here.

#19 Gas Laws

Reagents

sodium hydrogen carbonate	0.20 g/pair
1M hydrochloric acid	3 mL/pair

Common Materials

drinking straw	1/student
ruler (cm)	1/pair

Special Equipment

burette	1/pair
very small (75-mm-long) test tube	1/pair
clear tubing	4 ft./pair

Laboratory Equipment

- #1 1-hole rubber stopper with short glass tube
- 10-mL graduated cylinder
- 100-mL graduated cylinder & 1-hole stopper to fit
- 250-mL Erlenmeyer flask & 1-hole stopper to fit
- 1000-mL beaker
- thermometer
- barometer

Notes

In the first procedure, it is very important that there are no bubbles in the tubing. This is why clear tubing has been suggested.

In Procedure 14, be sure the stopper does not block the pour spout so that the water cannot escape the cylinder during the gas collection.

Disposal

The solutions produced in this investigation can be poured down the drain.

GAS LAWS PRELAB QUESTIONS

1) In Procedure 1, air is trapped inside the small cylinder above the water level. The air in the top of the burette is open to the atmosphere. As the burette is raised and lowered, the water exerts varying pressures on the trapped gas in the small cylinder. How is the gas pressure at the top of the water in the burette affected? Explain.
When the water levels are even, the pressure of the trapped air is the same as the pressure of the open air at the surface of the burette. When the water level in the closed cylinder is lower than the water in the burette, the water exerts pressure on the trapped air as it attempts to reach the same elevation as the water in the burette. Therefore, the air in the cylinder has greater pressure. However, the air pressure on the water in the burette is a direct result of current atmospheric pressure, and so it does not change.

2) Gas volumes change with temperature when all other factors are held constant. What would you expect to see happen to a fully inflated Mylar® balloon if it was taken outside into freezing weather? Why?
The cooler temperatures would allow the gas molecules to move with less energy and thereby exert less pressure on the inside of the balloon. The balloon would have a smaller volume of air than the volume of the balloon at room temperature, so the balloon would appear to be wrinkled.

3) Basketballs inflated at one temperature and then used at another may not behave as expected. Why do balls often seem "flat" in the wintertime?
Most balls are inflated indoors. When they are taken outside to play, the reduced temperatures cause the internal pressure to be lowered. This results in a smaller coefficient of elasticity for the trapped air and a ball that does not bounce as well.

4) If a gas is collected in a rigid container so that volume, temperature, and pressure are known, what relationship or law will allow the experimenter to find the amount of gas collected?
The Ideal Gas Law relates these three values to the amount of gas in moles.

5) The early experiments on gases often used mercury where this experiment used water. Mercury vapors are toxic. Does changing from mercury to water in this experiment follow the green principle of inherently safer chemistry for accident prevention?
Yes. Should the apparatus come apart during the experiment, mercury and its vapors would not be introduced into the atmosphere.

GAS LAWS SAMPLE REPORT SHEET

I. Boyle's law	Air volume	Pressure	PV
a. even levels	*5.0 mL*	*721.4 torr*	*3600*
b. first lifted burette	*4.6 mL*	*757.4 torr*	*3500*
c. second lifted burette			
d. first lowered burette	*5.4 mL*	*643.4 torr*	*3500*
e. second lowered burette			

II. Charles's law	Temperature (K)	Pressure	$\frac{P}{T}$
Room temp	*296 K*	*674 torr*	*2.28*
Lower temp	*279 K*	*417 torr*	*1.50*
Higher temp	*335 K*	*1150 torr*	*3.40*
Highest temp			

III. Gay-Lussac's law	Volume	Temperature	$\frac{V}{T}$
Room temp	*65.0 mL*	*320 K*	*.20*
Lower temp	*79.8 mL*	*280 K*	*.29*
Higher temp	*91.0 mL*	*348 K*	*.26*
Highest temp			

IV. Ideal Gas Law

Gas volume: Final volume – Initial volume = Gas volume

<u>*68.4*</u> mL – <u>*0*</u> mL = <u>*68.4*</u> mL

<u>*68.4*</u> mL $\times \dfrac{10^{-3}\ \text{L}}{\text{mL}}$ = <u>*0.0684*</u> **L (*V*)**

Temperature: <u>*21.0*</u> °C +<u>273</u>° = <u>*294*</u> **K (*T*)**

Pressure: Measured pressure – Vapor pressure = Dry gas pressure

<u>*732.6*</u> mm Hg – <u>*18.7*</u> mm Hg = <u>*713.8*</u> mm Hg

<u>*713.8* mm Hg</u> $\times \dfrac{1\ \text{atm}}{760\ \text{mmHg}}$ = <u>*0.939*</u> **atm (*P*)**

Amount: <u>*0.25*</u> g $NaHCO_3 \times \dfrac{1\ \text{mol}}{84\ \text{g}}$ = <u>*0.0684*</u> mol (*n*) $NaHCO_3$ = <u>*0.0684*</u> **mol (*n*)** CO_2

$R = \dfrac{PV}{nT}$ = <u>*0.0733*</u> $\dfrac{\text{L}\cdot\text{atm}}{\text{K}\cdot\text{mol}}$

V. QUESTIONS

1. Are the values for PV all the same?
   They should be very close.

2. Are the values for $\dfrac{V}{T}$ all the same?
   They should be very close.

3. Are the values for $\dfrac{P}{T}$ all the same?
   They should be very close.

4. What was the percent error of your value for R?
   $\dfrac{R_{actual} - R_{calculated}}{R_{actual}} \times 100\% = \%\ \text{error}$ *(10.4 %)*

5. If you plotted your temperature–volume data, at what temperature would the volume be zero?
   From this data it would be at –42 K. It should occur at 0K.

6. If you check the tire pressure after the car has been driven long enough for the tires to gain heat, what would you find?
   The air in the tires heats up with the friction of driving and causes more pressure in the tire. If the pressure were checked when the tire was cool, it would be too high after the tire had been driven a while. However, tires are designed to handle the extra pressure.

#20 Diffusion and Graham's Law

Reagents

hydrochloric acid [HCl] concentrated	5 mL/pair
ammonia [NH_3] concentrated	5 mL/pair
distilled water	

Common Materials

Johnson's® safety swabs or q-tips	10 swabs/pair
ice cubes	2 cubes/pair
food coloring (blue or green)	2 drops/pair
ruler (cm)	1/pair

Special Equipment

1" glass tube	1/pair
1-hole stoppers to fit glass tube	2/pair
stopwatch	1/pair
grease pencil	1/pair
cork borer (if stoppers are solid)	1/pair

Laboratory Equipment

4 250-mL beakers
thermometer
2 watch glasses or evaporating dishes (for transporting moistened swabs)
barometer

Notes

As the students get more adept at spotting the ammonium chloride ring, their times get shorter. It might be wise to have them do one or two practice runs before they begin taking data.

The vapors from the two concentrated solutions can become uncomfortable. If individual containers of solutions are used at several locations in the lab, students will transport the moistened swabs shorter distances. If one container is used, it would be best to place it under the hood since it will remain open much of the lab period as many students will be using it. Either way, dropper bottles will produce less vapor than an open beaker and a separate dropper. Suggesting that students transport the moistened swabs in a evaporating dishes or watch glasses might be a good idea.

Disposal

All solutions can be flushed down the drain with plenty of water. The cotton swabs can be squeezed and thrown into the trash.

DIFFUSION PRE-LAB QUESTIONS

1) Give several evidences of diffusion from your own experience.
Answers will vary but should include things like: knowing someone is painting their nails without being able to see them, walking into a house and knowing what is for dinner, knowing there is a fire somewhere without seeing smoke, looking at last night's leftover iced tea and seeing the gradual fading of color near the top where the ice was instead of a separate layer of water from the melted ice, and so on.

2) If gas molecules are moving at a tremendous speed, why does it require several minutes for a scent to diffuse across a room?
Molecules don't travel in a straight line but rather in a very random path depending on collisions with other molecules.

3) How would an increase in atmospheric pressure be expected to affect the rate of diffusion in a gas?
The rate of diffusion would be slower as there would be more molecules in the pathway and therefore an increased number of collisions and more random directions of motion.

4) Explain how an increase in temperature would be expected to affect the rate of diffusion in a gas.
The rate of diffusion would increase because molecules are moving faster.

5) How does the molar mass of gas particles affect the relative rates of diffusion?
The smaller the molar mass, the faster the rate of diffusion.

DIFFUSION SAMPLE REPORT SHEET

I. DATA AND CALCULATIONS
PART A: The Effect of Temperature on Diffusion in a Liquid

Beginning time: ____________________

	Beaker #1	Beaker #2	Beaker #3
Temperature			
Ending time			
Elapsed Time	*depends on temperature*		

PART B: The Effect of Particle Mass on Diffusion in Gases

	Trial #1	Trial #2	Trial #3	Trial #4
Time (in sec)	*148*	*120*	*103*	
D $_{(HCl\ to\ ring)}$ (cm)	*8.0*	*10.1*	*12.3*	
D $_{(NH_3\ to\ ring)}$ (cm)	*22.0*	*20.2*	*17.7*	
R $_{(HCl)}$ (cm/sec)	*0.054*	*0.084*	*0.119*	
R $_{(NH_3)}$ (cm/sec)	*0.13*	*0.16*	*0.17*	

	Molar Mass	Average Rate
HCl	*36.46*	*0.083*
NH_3	*17.03*	*0.15*

Rate Ratio:
Experimental: *1.8:1*

Expected: *1.5:1*

II. QUESTIONS

1. What was the effect of the different temperatures on the rates of diffusion in the beakers of water?
   The colder the temperature, the slower mixing occurred.

2. Did you notice any "currents," or convection, as the color moved through the water? If so, account for the cause of the currents.
   The water is not homogeneous. The currents occur because of the mixing of the waters of different temperatures. Warm water rises and the cooler water sinks.

3. Write the balanced equation for the formation of ammonium chloride.
   $HCl(g) + NH_3(g) \rightarrow NH_4Cl(s)$

4. Use the experimental ratio and the accepted (expected) ratio to calculate percent error.
   $$\frac{|E - A|}{A} \times 100\%$$
   $$\textit{Percent error} = \frac{1.8 - 1.5}{1.5} \times 100\% = 20\%$$

5. List possible sources of error that account for the percent error in your experimental data.
   An error in timing as one learns to recognize the ring, width of grease mark, swabs not put in at the same time or put in differently are examples of types of errors.