Exploratory Chemistry Experiments
BYU (Revised Edition 2013)

Lee D. Hansen, Bryon J. Wilson,
Francis R. Nordmeyer & John F. Cannon

Department of Chemistry and Biochemistry
Brigham Young University, Provo, UT 84602

What I hear, I forget
What I see, I remember
What I do, I understand

To the student:

The exploratory experiments are intended to be both fun and instructive, but to be successful as a learning aid they must be approached with the right attitude. If you approach the exploratory with a good-humored, thoughtful, and receptive attitude, we promise you will experience some of the fun, wonder, and excitement of doing science.
The following are acknowledged for assistance in developing experiments:

Justin T. Barratt
Matt C. Beard
Stephen R. Boden
Lora Bozarth
Michael A. Canady
Brenda Christensen
Lyneen R. Cluff
Dr. Coran L. Cluff
Lyneen R. Cluff
Robert H. Crawford
Brett S. Duersch
Angela R. Egbert
Judy L. Garner
Dr. Lu Giddings
Gurvais C. Grigg
Bret D. Heileson
Daniel J.B. Hemming
J. Brandon Hill
Steve Jensen
Michael D. Jones
Dave C. Jones

Angela R. Egbert Jones
David C. Jones
Angela R. Egbert Jones
Zenia Kim
L. Heath McDonald
Jennifer McGee
Dave Moulton
Jeremy B. Nicoll
Jim Pace
Hilary Porter Parry
Jennifer A. Rapking
Kristine Rasmussen
Julie Q. Reber
Michael W. Scheetz
Scott E. Sears
Cecily Vaughn
Teresa J. Warner
Rachel S. Wentz
Gary G. Willardson
Dr. Byron J. Wilson

We thank Jill Means for her drawing for Experiment 18.

Revisions and assistance in developing Brigham Young University Edition (2013) provided by:

Dr. Jeff Macedone
Kyli Bishop
Jackie Johnson
Jacob Hedelius

Kaylie Madsen
Dr. Scott Burt

Instructors may obtain an instructor’s manual by contacting the Brigham Young University Department of Chemistry and Biochemistry.

Copyright©2006 by Lee D. Hansen, Byron J. Wilson, Francis R. Nordmeyer & John F. Cannon

All rights reserved. No part of this publication may be reproduced, stored in a retrieval system, or transmitted in any form or by any means, electronic, mechanical, photocopying, recording, or otherwise, without the prior written permission of the copyright owner.

Printed in the United States of America

10 9 8 7 6 5 4 3 2
Preface to the BYU Revised Edition 2013:

It has been about 18 years since the Exploratory Lab was first integrated into the course curriculum for chemistry lecture classes at BYU, and six years since this manual has received a revision. We have found it necessary to update the most commonly used experiments to: increase clarity in procedures and questions, increase student success in experiments, increase reproducibility, utilize modern or custom equipment, and help experiments better fit the course curriculums.

The following have received updates: safety contract, Experiments 2, 8, 10, 17, 18, 24, 25, 41, 62, 68, 86, & 94

The following are in the process of being updated and updates will be added to this manual when complete: 15, 18, 21, 22, 24, 25, 62, 65, 68, 72, 83, & 95

In addition, the format in this edition has been updated to fit needs. We sincerely hope it will continue to be a positive resource for the Chemistry 101, 102, 105, and 285 courses.

Survey Results 2012:

A survey was conducted winter 2012 to access the effectiveness of the Exploratory Lab. Overall, results were positive. There were 85 Chemistry 101 and 20 Chemistry 102 students surveyed. Two-thirds of those surveyed said labs were helpful or they learned a great deal. In addition, two out of three students said that doing lab experiments helped them think more independently, and two in three said they gained something from labs that they could not have gained from class or personal reading. We expect that new changes to labs will continue to increase the positive response to the Exploratory Lab.
## Contents

Safety Contract ........................................................................................................................................... 1  
1. Absolute Zero .......................................................................................................................................... 2  
2. Atomic Spectra ........................................................................................................................................ 4  
3. Birds of a Feather .................................................................................................................................... 6  
4. Buffers ................................................................................................................................................... 8  
5. BYU Thermometer .................................................................................................................................. 10  
6. Case of the Unlabeled Bottles .................................................................................................................. 12  
7. Cloud Chamber ....................................................................................................................................... 14  
8. Collapsing Can ....................................................................................................................................... 16  
9. Crystalline Solids: Geometric Arrays of Atoms, Ions and Molecules ........................................... 18  
10. Dipping Bird ......................................................................................................................................... 20  
11. Electrolysis of Salt Solutions .................................................................................................................. 22  
12. Galvanic Lemon ..................................................................................................................................... 24  
13. Gaseous Molecules: Singles or Couples? ............................................................................................... 26  
15. How Big Is a Mole? ............................................................................................................................... 30  
17. Hungry Pyromaniac .............................................................................................................................. 34  
18. Hydrogen-Oxygen Rockets ....................................................................................................................... 36  
19. Kool-Aid Chromatography ..................................................................................................................... 38  
20. Lead Balloon .......................................................................................................................................... 40  
21. M&Ms—Moles & Molecules .................................................................................................................... 42  
22. Models of Molecules ............................................................................................................................ 44  
23. Oxidation States of Manganese ............................................................................................................ 46  
24. pH Indicators ......................................................................................................................................... 48
<table>
<thead>
<tr>
<th>Chapter</th>
<th>Title</th>
</tr>
</thead>
<tbody>
<tr>
<td>52.</td>
<td>Ring Around the Bathtub</td>
</tr>
<tr>
<td>53.</td>
<td>Signs of Chemical Change</td>
</tr>
<tr>
<td>54.</td>
<td>Spectronic 20® Rainbow</td>
</tr>
<tr>
<td>55.</td>
<td>Speed Limits: Determining a Rate Law</td>
</tr>
<tr>
<td>56.</td>
<td>Starch Chemistry</td>
</tr>
<tr>
<td>57.</td>
<td>Vanadis, Goddess of Love and Beauty</td>
</tr>
<tr>
<td>58.</td>
<td>Warm It Up If You Are in a Hurry</td>
</tr>
<tr>
<td>59.</td>
<td>Whiter than White</td>
</tr>
<tr>
<td>60.</td>
<td>The Zinc Iodine Cycle</td>
</tr>
<tr>
<td>61.</td>
<td>Acid-Base Properties of Salts</td>
</tr>
<tr>
<td>62.</td>
<td>Avogadro’s Hypothesis</td>
</tr>
<tr>
<td>63.</td>
<td>Boiling Water with Ice</td>
</tr>
<tr>
<td>64.</td>
<td>Breathalyzer Reaction</td>
</tr>
<tr>
<td>65.</td>
<td>Chemical Periodicity</td>
</tr>
<tr>
<td>66.</td>
<td>Chloride Ion in Gatorade®</td>
</tr>
<tr>
<td>67.</td>
<td>Chromatography of Ions</td>
</tr>
<tr>
<td>68.</td>
<td>Cobalt(II) Complexes</td>
</tr>
<tr>
<td>69.</td>
<td>Colors</td>
</tr>
<tr>
<td>70.</td>
<td>Common Ion Effect</td>
</tr>
<tr>
<td>71.</td>
<td>Competing Equilibria</td>
</tr>
<tr>
<td>72.</td>
<td>Dry Ice Under Duress</td>
</tr>
<tr>
<td>73.</td>
<td>Dynamic Equilibrium</td>
</tr>
<tr>
<td>74.</td>
<td>How Strong Is Your Bleach?</td>
</tr>
<tr>
<td>75.</td>
<td>Hydrogen and Oxygen Chemistry</td>
</tr>
<tr>
<td>76.</td>
<td>It’s A Colorful World</td>
</tr>
<tr>
<td>77.</td>
<td>Metal Alloy Analysis</td>
</tr>
<tr>
<td>78.</td>
<td>Molar Volume of Carbon Dioxide Gas</td>
</tr>
</tbody>
</table>
79. Organic Functional Groups................................................................. 158
80. Paramagnetism ............................................................................... 160
81. pH Rainbow .................................................................................. 162
82. Photoelectric Effect ....................................................................... 164
83. Polarity of Molecules ................................................................. 166
84. Popcorn Hypothesis ..................................................................... 168
85. Precipitation Pictures ................................................................. 170
86. Sink or Swim .................................................................................. 172
87. Slow Flows ...................................................................................... 174
88. Something Smells .......................................................................... 176
89. Thirsty Structures .......................................................................... 178
90. Viscous Vortex ............................................................................... 180
91. Classification Tests for Organic Functional Groups .................... 182
92. The Warmer, The Faster ............................................................. 184
93. Faster, Faster .................................................................................. 186
94. Measurement of Radioactivity ...................................................... 188
95. Molecular Models ........................................................................... 192
96. Molecular Shapes and Types of Bonds ........................................ 194
97. Thermodynamics of a Galvanic Cell ............................................ 196
98. Titration of a Weak Acid with a Strong Base ................................ 198
Future updates .................................................................................... 200
ALPHABETICAL LIST OF CONCEPTS
(by page number)

Absolute temperature .................. 3, 117, 185, 187
Acid-base buffers ................................ 9
Acid-base chemistry .......................... 63, 79, 87
Acids ............................................. 29, 59
Acids and bases .................. 123, 141, 143, 153, 199
Activation energy .................. 117, 185, 187
Amphotericism ............................... 63
Atmospheric pressure ......................... 17
Atomic spectra ................................ 5
Avogadro's number .................. 31, 65
Bases ............................................. 29, 59
Basic and amphoteric behavior of metal ions ................................ 53
Battery ......................................... 85
Boiling point .................................. 11, 127
Bond polarity .................................. 11, 167
Bond types ...................................... 195
Calorimetry .................................... 35
Catalysis ....................................... 113
Chemical change ............................ 75, 107, 121
Chemical equilibrium .......................... 103, 147
Chemical formulas .................. 37
Chemical properties .................. 13
Chemical reactions .......................... 51
Chirality ...................................... 97
Chromatography .............................. 39, 135
Classification of matter .................. 61
Close packing of spheres .................. 45
Colloids ........................................ 71
Color ........................................ 77, 109, 139
Color and energy .................. 5, 109
Colors of transition metal compounds ................................ 47
Colors of transition metal ions 75, 81, 115
Common ion effect .............................. 141
Complex ion formation .................. 53, 57
Complex ions ........... 63, 75, 77, 81, 137, 143
Concentration .................. 59, 149, 199
Concentration is proportional to observed color intensity .................. 57
Conductance by ions in solution .................. 79, 93
Counting by weighing .................. 43
Critical temperature .................. 145
Critical thinking .................. 169
Crystal structure ................................ 19
Decay rate .................................... 189
Decomposition reactions .................. 61
Dehydration .................................. 179
Density ....................................... 173
Dipole-dipole forces .................. 89
Dissociation of electrolytes 33, 79, 93
Electric dipole ................................ 7
Electrochemical series .................. 25
Electrochemistry .......................... 25, 85, 101, 121
Electrode reactions .................. 23
Electrolysis .................................. 155
Electromagnetic spectrum .................. 109
Electron configurations 77, 161, 193
Electrostatic interactions .................. 167
Energy content of foods .................. 35
Energy of reaction .................. 37
Enthalpy change .................. 197
Entropy change .................. 197
Enzymes .................................... 113
Equilibrium .................. 137, 141, 143
Experimental errors .................. 43
Fluorescence ................................ 119
Freezing point ................................ 11
Freezing point depression .................. 33
Functional groups .................. 159, 177, 183
Galvanic cell .................. 85, 197
Gas laws ...................................... 17
Gas properties .................. 125, 157
Gibbs Free Energy .................. 197
Graham's law of diffusion .................. 41
Group properties .................. 131
Half reactions .................. 85
Heat ........................................ 21
Heat capacity .................. 35, 149
Heat of reaction .................. 149
Hydrogen .................................. 37, 151
Hydrogen bonding 89, 179, 181
Hydrolysis .................................. 123
Hydrolysis of starch to sugar .................. 113
Hypertonic ................................ 83
Hypothesis .................. 169

—vii—
EXPLORATORY LAB C170 BNSN

Safety Contract

This must be submitted to BNSN C170 BEFORE beginning work in the lab.

STEP 1. COMPLETE THE FOLLOWING FIELDS

Your Full Name ____________________ Course ________________
Student ID# ________________________ Section ________________
Date ________________________________ Professor ______________

STEP 2. READ CAREFULLY:

While in the exploratory lab I commit to:
1. wear clothing that covers my legs and feet (no shorts or sandals).
2. wear goggles at ALL times while in the lab (eye-glasses do not count).
3. place my backpack in a designated area (hall lockers are also available).
4. read the procedure before going to lab.
5. only be in the lab with a TA.
6. clean up after myself, minimize waste, and dispose of waste properly.
7. ask questions if I am not sure how to do something.
8. not move work stations or use equipment not associated the weekly lab.

STEP 3. ACQUIRE THE FOLLOWING SIGNATURES

YOUR SIGNATURE

Have a TA sign on the line below after they have reviewed this contract with you.

__________________________________________
EXPLORATORY LAB TA SIGNATURE

STEP 4. Place in the appropriate slot of the blue box inside C170.

DONE!
1. Absolute Zero

**OBJECTIVE:** To observe how the behavior of gases defines the absolute scale of temperature.

**CONCEPT:** Absolute temperature

**INTRODUCTION:** The behavior of gases is described by mathematical equations called gas laws. These laws describe the relationships between the temperature, pressure and volume of a gas. This experiment is based on the ideal gas law, \( pV = nRT \), where \( p \) represents the pressure, \( V \) the volume, \( T \) the temperature and \( n \) the number of moles of gas. \( R \) is a fundamental constant, the value of which (0.082 L atm mol\(^{-1}\) K\(^{-1}\)) has been established previously in many quantitative experiments.

If the pressure and moles of gas are held constant, the equation can be modified to \( V = k_1T \) (Boyle’s law). Here \( k_1 \) replaces the constant term \( nR/p \). Similarly, by holding the volume of a closed system constant we can obtain \( p = k_2T \) (Amonton’s law). These forms of the ideal gas law predict that a gas has neither volume nor pressure when \( T = 0 \). The Kelvin temperature scale is established with this zero point (called absolute zero) and divisions (called Kelvins) which are the same size as centigrade or celsius degrees.

In this experiment, you will take several volume measurements at different temperatures and plot volume versus temperature. You will also make pressure measurements at various temperatures and plot these data in the same manner. The points for each set of data should be in a straight line. The intercept on the temperature axis is absolute zero on the temperature scale used in the experiment and is equal to 0 on the Kelvin scale.

**HAZARDS:** Syringe needles are very sharp. Use care not to stick yourself or somebody else. Don’t burn yourself with the hot water.

**PROCEDURE A:**
1. Four water baths ranging from 0 to 80°C and a 50 mL serum flask sealed with a septum and labeled with the exact volume, a thermometer, and a 20 mL ground glass syringe and needle will be needed. The syringe and serum flask must be dry. (Why? How does vapor pressure of a liquid change with temperature?) Crystals of dessicant are placed in the flask to remove water. These crystals indicate the presence of water by turning from blue to pink. If the crystals are pink, remove them, dry the flask with a tissue and add 10 to 15 crystals of fresh dessicant. Work the syringe plunger back and forth to eliminate any sticking.
2. Detach the syringe from the needle, leaving the needle in the stopper, pull 5 mL of air into the syringe and reattach to the needle. Inject 1 mL of air into the serum flask, and allow the plunger to rise and come to rest. This initial volume (flask + syringe) and the room temperature should be recorded in the table below as your first data point.
3. Immerse the flask in the lowest temperature water bath and allow the plunger to come to equilibrium. Gently tap the syringe to allow the plunger to reach an equilibrium position. Record the temperature and volume. Repeat the procedure in the remaining water baths.

**PROCEDURE B:**
1. A 250 mL filter flask with an attached pressure gauge will be needed. Make sure the dessicant in this flask is blue, not pink.
2. The gauge measures the pressure difference between the air in the flask and the atmosphere.
Atmospheric pressure can be read from a barometer or will be posted. The sum of barometric pressure and the gauge pressure will give you the total pressure in the flask. Make sure the units of pressure agree before adding the two pressures. Record the total pressure and room temperature.

3. Immerse the filter flask in an ice bath (0°C). Record the flask pressure and bath temperature after the gauge has stabilized, 1 to 2 minutes. If the gauge does not indicate a slightly positive pressure at this temperature, allow air to enter the flask by detaching the tubing from the needle, leaving the needle in the stopper, then reattach the tubing.

4. Transfer the flask to a 30°C water bath and repeat the procedure.

5. Repeat the procedure in the other two water baths at higher temperatures.

6. When you’ve finished your final measurement, remove the flask from the water bath and dry off the instruments and clean up any water that may have spilled on the lab bench.

<table>
<thead>
<tr>
<th></th>
<th>A. TEMPERATURE</th>
<th>TOTAL VOLUME OF FLASK AND SYRINGE (mL)</th>
<th>B. TEMPERATURE</th>
<th>TOTAL PRESSURE (Barometric + Gauge)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
<td>1</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
<td>2</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td>3</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
<td>4</td>
<td></td>
</tr>
</tbody>
</table>

Analyze your volume-temperature data: plot volume (y-axis) versus temperature (x-axis) for your measurements in Procedure A. Draw the best-fitting straight line through these points and determine the °C reading at the horizontal (x) intercept. Repeat this process to analyze your pressure-temperature data. Report and average your results below.

Volume-Temperature Absolute Zero = ___°C  
Pressure-Temperature Absolute Zero = ___°C  
Average Absolute Zero = ___°C  

**QUESTIONS:**

1. How do your results compare with the accepted value of absolute zero, –273.15°C?

2. What are the sources of error in your determination of absolute zero?

3. What relationship exists between pressure and volume at constant temperature?

4. Is a condition with T=V=p=0 actually achievable with a real gas? Why not?

5. What did you like about this experiment and how would you improve it?
2. Atomic Spectra

HAZARDS: Some of the chemicals used in this experiment are toxic. Avoid ingestion or skin contact. Be careful handling the Petri dishes; they do get hot. Please do not handle the discharge lamps.

CONCEPTS: Atomic spectra, Identifying elements, Color and energy

BACKGROUND:
When electrons in gas phase atoms are energized, either by an electric field or by heating, electrons can be excited to higher energy levels. When the source of the excitation is no longer present, the electrons “relax” back down to lower energy levels by releasing energy in the form of light. Because atoms have multiple energy levels in which electrons reside, photons of different colors can be emitted from the same element. The collection of lines observed from a particular element when it is emitting light is called its atomic spectrum.

PROCEDURE:
1. Find the yellow cardstock glasses. The gratings which make up the glasses lenses have regularly and closely spaced microscopic lines on their surfaces. These lines diffract light, a phenomenon which depends on wavelength. You can observe the wavelengths present in a light source by looking at the source through the lenses.

2. Locate the flashlight in the lab, and turn it on.
   a. What color is the flashlight without the glasses?
   b. Put on the diffraction glasses. Are there any breaks in the spectrum? That is, is it a “continuous” spectrum?

3. Next, find the gas discharge lamps (they are the tall black boxes with bulbs emitting light). Make the observations indicated below, and fill in the table.

<table>
<thead>
<tr>
<th>Substance in the lamp (find the label on the side)</th>
<th>Color seen with eye.</th>
<th>Number of lines observed using glasses</th>
<th>Colors of lines in spectrum</th>
<th>Draw a rough sketch of the spectrum (try to imitate brightness)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

4. Find the station where Petri dishes are filled with salts. Remove the lid from the empty Petri dish. Add a small amount of methanol from the squeeze bottle (between 1–3 mL, or 20–60 drops). Ignite the methanol with a match. Since methanol will be added to salts in other Petri dishes, this is a “control” experiment. What color is the flame?
5. For the next series of experiments, add enough methanol to wet the salt in the Petri dish (no more than 3 ml). Light it with a match and record the color of the flame.

<table>
<thead>
<tr>
<th>Substance in flame</th>
<th>Color</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium chloride</td>
<td></td>
</tr>
<tr>
<td>Sodium chloride</td>
<td></td>
</tr>
<tr>
<td>Potassium chloride</td>
<td></td>
</tr>
<tr>
<td>Calcium chloride</td>
<td></td>
</tr>
<tr>
<td>Copper(II) chloride</td>
<td></td>
</tr>
<tr>
<td>Strontium chloride</td>
<td></td>
</tr>
</tbody>
</table>

**QUESTIONS:**

1. What is different about the light given off by the flashlight and the discharge tubes? (Hint: a diffraction grating or prism is required to observe the difference.)

2. Since the lights are off, a station is set up where you can view a white fluorescent light through a narrow slit. Does the fluorescent bulb contain hot gaseous atoms, or is it a hot filament like a flashlight? What did you do to figure this out?

3. Thinking about your experience with the discharge lamps, and energy levels, is it possible that two elements could give the same emission spectrum as viewed through a diffraction grating?

4. Is it possible that two elements could produce the same color in a methanol flame? Explain.

5. Why is the spectrum of neon more complicated then hydrogen?

6. What evidence does this lab lend to our understanding of the structure of atoms?

7. Light speed, c, is $2.99792 \times 10^8$ m/s. All wavelengths of light travel at the same speed. One of the wavelengths emitted from hydrogen is 486.133 nm. Convert this wavelength to energy in J using $E=hc\nu$ where $E$ is energy, $h$ is Planck’s constant ($6.62607 \times 10^{-34}$ J·s) and $\nu$ is frequency ($c=\nu\lambda$). The energy represents the difference in energy between the two levels the electron is transitioning between.
3. Birds of a Feather

Name ___________________________________________ Course Number ________________
Identification Number _____________________________ Section Number ________________
Date ___________________________________________ Instructor ________________________

**OBJECTIVE:** To demonstrate intermolecular interactions in polar and nonpolar substances by observing solubilities in water and oil.

**CONCEPTS:** Like dissolves like, Electric dipole

**INTRODUCTION:** Polar molecules have large electric dipoles, i.e. charge separation on atoms in the molecule. Chemicals with similar molecular polarities tend to be mutually soluble, i.e. like dissolves like, and chemicals with dissimilar molecular polarities tend to be mutually insoluble, e.g. oil and water don’t mix.

In this experiment the solubilities of different compounds in two different liquids, a polar liquid (water) and a non-polar liquid (oil), will be tested. If the solute is polar, it will dissolve in water but not in oil. If it is nonpolar, the reverse will be true. To obtain good results, use only a tiny amount of solid, just barely enough to see. If you add too much, the solution may become saturated, and you will not be able to tell if the solid dissolves.

**HAZARDS:** Do not ingest or spill any of the materials in this lab. They may cause skin irritation or stain your clothing.

**PROCEDURE:**
1. Use one of the charts taped to the table. (Like the chart in these instructions.)

2. Tear off a piece of plastic wrap long enough to cover the chart and tape the corners down. Do not tape it on the paper itself or the table, but on the tape already on the table.

3. Begin with the NaCl on the left and work to the right (if you are left handed, do it the opposite way) to avoid getting anything on your hands or clothes.

4. Do the following for each compound and record all your observations in the chart:

   Place a crystal on the plastic in the proper square. Remember to put the lid back on the bottle when you are finished. To the crystals in each of the top squares, add one or two drops of water. Stir with a toothpick. Now add a drop of oil to each of the crystals in the bottom squares. Stir with a different toothpick. If the crystal does not dissolve in either liquid, you may have saturated the solution. Try adding one or two more drops of solvent. If that doesn't work, try again with a smaller crystal.

5. When you are finished, peel up the tape from the corners, fold the plastic wrap so that the chemicals are all inside, and put into the waste container labeled PLASTIC WRAP. Put all the
toothpicks into the waste container labeled TOOTHPICKS.

Solubility Chart

<table>
<thead>
<tr>
<th>Solvent</th>
<th>NaCl</th>
<th>NH₄Cl</th>
<th>C₁₀H₈ naphthalene</th>
<th>O=C(NH₂)₂ urea</th>
<th>CuSO₄</th>
<th>I₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>Water</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Oil</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Questions**:

1. If like dissolves like, which substances do you conclude are polar and which do you conclude are nonpolar? Were any surprising?

2. Suppose you have a mixture of two compounds. One is polar and one nonpolar. What could you do to separate them?

3. Describe the interactions between molecules of solute and solvent in the following cases.
   a. both polar
   b. both nonpolar
   c. polar solvent, nonpolar solute
   d. nonpolar solvent, polar solute

4. What did you like about this experiment and how could it be improved?
4. Buffers

Name __________________________________________ Course Number ___________________
Identification Number ___________________________ Section Number __________________
Date ___________________________ Instructor __________________________

OBJECTIVE: To demonstrate the pH buffering effect of a solution containing a weak acid and its conjugate base.

CONCEPTS: Acid-base buffers, pH, Indicators

INTRODUCTION: Acid-base buffers are used to moderate the pH change when an acid or a base is added to a solution. Added base is neutralized by the weak acid of the buffer and added acid is neutralized by the conjugate base of the buffer. Buffers play important roles in some everyday settings. For example, your blood as well as your favorite shampoo are buffered. Buffers function like cushions or shock absorbers for systems sensitive to acids and bases.

Buffers have a range of pH within which they function, i.e. the pKₐ of the buffer acid ±1 pH. Within the buffer range, buffers control fluctuations in pH when acids or bases are added, but outside this range they are not effective. In this experiment you will observe the effects of addition of strong acids and bases to buffered and unbuffered solutions. pH indicators will be used to measure the pH changes.

HAZARDS: Wear goggles. Strong acids and bases irritate the skin and damage clothes. Rinse thoroughly with water if any of the chemicals are spilled on your skin.

PROCEDURE:
Use a plate with at least 9 wells or use 9 test tubes in a rack. Number the wells or tubes 1 through 9.
1. Fill wells 1 through 3 one-fourth full of water.
2. Prepare acetic acid-acetate buffer in wells 4 through 6 by adding sufficient sodium acetate crystals (CH₃COONa) to just cover the bottom of wells 4 through 6, and then filling wells 4 through 6 one-fourth full with acetic acid solution (vinegar).
3. Prepare ammonia-ammonium ion buffer in wells 7 through 9 by adding sufficient ammonium chloride crystals to just cover the bottoms of wells 7 through 9, and then filling wells 7 through 9 one-fourth full with the ammonia solution (household ammonia).
4. Add 1 drop of universal pH indicator to each well.
5. Add 1 drop of 1 M HCl to wells 1, 4, and 7.
6. Add 1 drop of 1 M NaOH to wells 3, 6, and 9.
7. Observe the colors and use the chart provided in the lab to estimate the pH of each solution and fill in the table below.
8. Dispose of waste in the appropriate waste containers and clean up the lab area.

<table>
<thead>
<tr>
<th>Well</th>
<th>Contents</th>
<th>Color</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>water indicator HCl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>water indicator CH₃COONa</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>water indicator CH₃COOH Indicator HCl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>CH₃COONa</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5</td>
<td>CH₃COOH Indicator</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6</td>
<td>CH₃COONa</td>
<td></td>
<td></td>
</tr>
<tr>
<td>7</td>
<td>NH₄Cl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>8</td>
<td>NH₃ Indicator</td>
<td></td>
<td></td>
</tr>
<tr>
<td>9</td>
<td>NH₃ Indicator</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

---8---
**Questions:**

1. Explain what a buffer is and how it works.

2. Why did the pH in wells 1 and 3 change so much more than the pH in wells 4, 6, 7 and 9?

3. What happened to the sodium acetate (CH$_3$COONa) and ammonium chloride (NH$_4$Cl) on addition of the strong acid and base, respectively?

4. Write reactions for what occurred in wells 1, 3, 4, 6, 7, and 9 when HCl or NaOH was added.

5. Why is the initial pH different in the two buffer solutions? (Hint: The pK$_a$ value for CH$_3$COOH is 4.75 and for NH$_4^+$, 9.25.)

6. The pH of the water in well 2 should be slightly less than 7.0. Why? (Hint: The indicator is too dilute to significantly affect the pH. What common gas in air has acidic properties?)

7. Why are buffers so important?

8. Describe one buffer system you have encountered in your everyday life (other than those in this lab).

9. What did you like about this experiment and how would you improve it?
5. BYU Thermometer

OBJECTIVE: To demonstrate how temperature scales are defined and how temperatures are converted from one scale to another.

CONCEPTS: Temperature, Freezing point, Boiling point

INTRODUCTION: Temperature is not a substance, it is a measure of the average energy of the molecules in a material. Temperature is measured on arbitrary scales based on the expansion of liquids and gases and the change in electrical resistance of certain metals. Temperature scales are calibrated against agreed-upon, standard, easily reproducible temperatures.

HAZARDS: Be careful not to burn yourself or anyone else with the boiling water. If you break an alcohol-filled thermometer, be careful not to cut yourself on the broken glass when cleaning it up.

PROCEDURE: BYU has its own temperature scale. Calibrate the BYU thermometer provided against two fixed temperature points, i.e., the freezing (melting) (0°C) and boiling (100°C) points of water. Place the thermometer in an ice bath (a mixture of ice and water) and read the temperature. Then place the thermometer in boiling water and read the temperature.
Questions:
1. Calculate absolute zero in °BYU. (Absolute zero = −273.15°C)

2. Why are the melting point and boiling point called fixed (or constant) temperatures even though they are different on different temperature scales (e.g. °F and °C)?

3. Why not use just one fixed point to calibrate a thermometer?

4. Why is measurement of temperature important in chemistry?

5. What other fixed points are there?

6. What did you like about this experiment and how would you improve it?
6. Case of the Unlabeled Bottles

OBJECTIVE: To use solubility rules to identify several solutions of unknown ionic compounds.

CONCEPTS: Solubility rules, Chemical properties

INTRODUCTION: Professor Hansmeyer absentmindedly left some unlabeled bottles around. He knows what he put in them, but can't remember what went in which bottle. Hansmeyer is now somewhere in the tropics gathering data on intestinal parasites, but before he left he asked us to solve this mystery of the unlabeled bottles. The original bottles of known solutions are available. Your job is to label the unknowns. The known solutions are 0.1 M NaOH, KBr, KI, AgNO₃, BaCl₂, and Na₂SO₄. What ions are present in each of these solutions? (For example: BaCl₂(aq) ⇌ Ba²⁺(aq) + 2Cl⁻(aq).) Before you begin the lab, think about possible reactions that you can expect according to the solubility rules for ionic compounds. Which mixtures of solutions will yield precipitates and which will not? Each ion in the solutions will be tested with every other ion. Two 6 x 6 data tables are provided to record your observations when testing the known and unknown bottles.

HAZARDS: Be especially careful not to spill any chemicals on your clothes or skin. These solutions may stain clothes and skin, and the strong acid and base will irritate your skin and put holes in your clothes. If any chemicals are spilled on skin or clothing, immediately wash with large amounts of water.

PROCEDURE:
1. Obtain a 96-well microplate. Arrange the known bottles in the same order as the data table. In the left half of the plate, add four drops of each solution to each well in the horizontal row of the plate corresponding to the data table. Add four drops of each solution to each well in its vertical column, stir, and record your observations in the data table. Describe the precipitates. Record "NVR" where no visible reaction occurs.

<table>
<thead>
<tr>
<th></th>
<th>NaOH</th>
<th>KBr</th>
<th>KI</th>
<th>AgNO₃</th>
<th>BaCl₂</th>
<th>Na₂SO₄</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaOH</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>KBr</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>KI</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>AgNO₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>BaCl₂</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na₂SO₄</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2. In the right half of the plate, repeat the procedure for the unknowns. Compare observations.
3. Clean the microplate with running water and dispose of the waste as instructed. You may need to use a cotton swab to clean some of the wells.

**Questions:**

1. Write a quick note to Dr. Hansmeyer telling him which solution was in which bottle. (Number, name and formula.)

<p>| | | | | | |</p>
<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>#1</td>
<td>#2</td>
<td>#3</td>
<td>#4</td>
<td>#5</td>
<td>#6</td>
</tr>
<tr>
<td>#1</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>#2</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>#3</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>#4</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>#5</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>#6</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

2. If mixing two of your unknown solutions produces a brown precipitate, which solution combination is it?

3. If mixing two of your unknown solutions produces a yellow precipitate, which solution combination is it?

4. You are given solutions A, B, and C of the unknown solutions and the following results:
   - A and B produce a pale yellow precipitate,
   - B and C produce a brown precipitate,
   - A and C produce no visible reaction.
   Using your data table for known solutions, what are the possible identities for these three unknown solutions? Explain your answer.

5. What did you like about this experiment and how could it be improved?
7. Cloud Chamber

Name ____________________________ Course Number _______________________
Identification Number __________________ Section Number __________________
Date ______________________________ Instructor ____________________________

OBJECTIVE: To demonstrate emission of charged particles from a radioactive element.

CONCEPTS: Radioactivity, Isotopes

INTRODUCTION: Cloud chambers are used to detect particles from radioactive decay of nuclei. This experiment uses a vapor cooled below its condensation point (supercooled vapor) to detect the particles. The supercooled vapor is prepared by cooling the bottom of a chamber saturated with vapor. The warmer, upper part of the chamber acts as a reservoir for the vapor. When high energy particles pass through a supercooled gas, ions are created and condensation is initiated on these ions. The track of condensed liquid droplets shows the path of the particle. During the experiment, look for white condensation tracks (wisps of fog) coming from the tip of the needle where the radioactive source is located. Every once in a while a cosmic ray will pass through the chamber. It will appear as a streak that does not originate at the tip of the needle. You may be lucky and see one. It may take five to ten minutes for the chamber to begin working. BE PATIENT. If you are not able to see the results, talk to a teaching assistant. They may be able to help. Remember to have fun. Your set-up should look like this:

HAZARDS: Do not contact the radioactive source! Wash your hands after you complete the experiment. Avoid ingestion of radioactive materials. Contact with dry ice can cause frostbite. Always use the tongs or spoon provided when handling dry ice. Avoid breathing the alcohol vapors; they are hazardous if inhaled. The alcohol is also flammable; keep away from flames.

PROCEDURE:
1. The cloud chamber has been previously prepared for you. Do not try to open the cloud chamber. The sides and bottom of the sealed petri dish have been lined with absorbent paper that has been accordion folded, to increase surface area. There should be sufficient methanol present to completely saturate the folded paper and slightly cover the bottom of the petri dish. Note that the methanol is drawn up by the paper until the paper is saturated. The tip of the needle should be positioned near the center of the chamber and as close to the bottom of the chamber as possible without submerging it in the methanol. The radioactive material is on the tip of the needle. It should not be touching the folded paper. If the cloud chamber does not appear to be as described, ask a teaching assistant for help.
2. Fill a large, round dish with crushed dry ice. If the dry ice is in chunks, crush it into snow by placing several chunks in a folded towel and pulverizing them with a hammer. Remember to handle the dry ice with tongs or spoon. **Put the lid back on the dry ice container.**

3. Place the chamber on top of the dry ice and gently nestle it into the dry ice. This will cool the alcohol vapor at the bottom of the chamber. Do not pile the dry ice up around the sides. You only want to cool the bottom.

4. After several minutes, a fog of methanol vapor should begin to form in the chamber. By shining the flashlight on the vapor near the tip of the needle you should see the alpha particles leaving visible trails in the fog. Look for thin, white lines of condensation that originate near the radioactive source and travel outward in all directions.

5. If the chamber becomes too cold, the vapor will cloud and the best approach at this point is to start over. Remove the chamber from the dry ice and allow it to warm up. Begin cooling the chamber again.

6. After you have completed the experiment please put the dry ice back into its container. Clean up the lab area. **Remember to put the lid back on the dry ice.**

**Questions:**

1. Did you get the experiment to work? If so, what did you do to accomplish this, and if not, can you explain what went wrong?

2. Explain why charged particles leave visible tracks in the super-saturated vapor.

3. Why was the bottom cooled and the top not?

4. \(\alpha\) and \(\beta\) rays can both be seen with a cloud chamber, and are easily distinguished from the tracks they leave. Explain how this might be done.

5. How would the tracks of \(\alpha\) and \(\beta\) rays appear if the cloud chamber were in an electric field? In a magnetic field?

6. What did you like about this experiment and how would you improve it?
8. Collapsing Can

Name_________________________________________ Course Number_________________________________________
Student________________________________________ Section Number_________________________________________
Date_________________________________________ Instructor_________________________________________

**OBJECTIVE:** To observe the effects of atmospheric pressure.

**CONCEPT:** Atmospheric pressure. Gas laws.

**INTRODUCTION:**
Gas particles in the atmosphere exert a pressure on us that we refer to as atmospheric pressure. This pressure acts on our bodies and is approximately 15 lbs. per sq. inch (remember that pressure is force per area). Atmospheric pressure changes with altitude. One way that your body senses pressure is when your ears pop as you change altitude. Most of the time the inside and outside of our ears are at the same pressure. When atmospheric pressure changes, the pressure inside and outside our ears equilibrate; we hear and feel a “popping” sound. Generally, altitude changes are gradual enough to avoid physical damage. In this experiment, you will observe the effects of a system not balanced against atmospheric pressure.

**HAZARDS:** Use caution when using the hot plate and handling the hot soft-drink can. Please use tongs. Burns can cause severe skin damage. Remember to **TURN OFF** the hot plate.

**PROCEDURE:**
1. Determine the volume of your can by filling it up completely with water. Pour the water into 100ml graduated cylinder to the 100ml mark. Empty the cylinder and repeat as necessary. Volume initial __________ ml
2. Fill the can with just 15 ml of water.
3. Place the can on the hot plate, and turn the dial to the highest setting (careful with the dial!).
4. Once the water begins to boil rapidly, wait 60 seconds to allow the air originally in the can to be replaced by water vapor. (You can answer questions except 1B & 2B while you are waiting).
5. Using the tongs, lift and invert the hot can quickly, and submerge the top of the can into the water bath (see diagram). When the can is inverted into water, the opening is sealed by the liquid and the water vapor quickly condenses, reducing the pressure inside the can. The sudden difference in net force produces a dramatic effect!
6. Lift the can vertically out of the bath. Shake out water over the sink.
7. Turn OFF the hot plate.
8. Find the final volume of your can. Fill up the 100ml graduated cylinder to the 100ml mark with water from the sink. Pour water from the graduated cylinder into the collapsed can until water begins to overflow from the can. You may need to repeat this step.
9. Record the final volume of the can. Volume final ________________ ml.
10. Clean up the lab area. Dispose of the can in the recycling bucket.
The following information will help you answer the questions below. \( p \) is pressure, \( V \) is volume, \( T \) is absolute temperature, \( n \) is amount in moles of gas.

<table>
<thead>
<tr>
<th>Scientist to observe</th>
<th>Constants/fixed values</th>
<th>Variables</th>
<th>Relationship</th>
<th>Equation from Relationship</th>
</tr>
</thead>
<tbody>
<tr>
<td>Jacques Charles</td>
<td>( p, n )</td>
<td>( T, V )</td>
<td>( V \propto T )</td>
<td>( \frac{V_i}{T_i} = \frac{V_f}{T_f} )</td>
</tr>
<tr>
<td>Amedeo Avogadro</td>
<td>( T, p )</td>
<td>( n, V )</td>
<td>( V \propto n )</td>
<td>( \frac{V}{n} = c )</td>
</tr>
<tr>
<td>Robert Boyle</td>
<td>( T, n )</td>
<td>( p, V )</td>
<td>( p \propto \frac{1}{V} )</td>
<td>( p_iV_i = p_fV_f )</td>
</tr>
</tbody>
</table>

Atmospheric pressure in Provo is \( \sim 0.85 \) atm; because of this pressure, water boils at about \( 95^\circ C \) in Provo (this means \( c = 3.18 \) \( \frac{L}{mol} \) in Avogadro’s law). Room temperature in the lab is \( \sim 20.0^\circ C \). At \( 20.0^\circ C \), water has a vapor pressure of \( 0.0228 \) atm and liquid water has a density of \( 0.998 \) \( \frac{g}{ml} \).

**QUESTIONS:**

1. Part A. From Charles’s Law alone predict what the final volume of the collapsed can should have been in mL.

   Part B. Is your answer in Part A close to what you observed? Is it an overestimate, or just slightly off? If your calculated value is largely different in magnitude from your measured value, propose a reasonable idea to explain the difference.

2. Part A. From Avogadro’s law and the initial volume of your soda can, calculate the final volume of water vapor after it has condensed (in milliliters). Assume the can was first filled with only gaseous water molecules, and then all of the water condensed into a liquid (remember that water is not an “ideal gas”).

   Part B. Is your answer in Part A close to what you observed? Is it an overestimate, or just slightly off? If your calculated value is largely different in magnitude from your measured value, propose a reasonable idea to explain the difference.

3. Part A. Estimate the pressure inside the open can initially.
   Part B. Estimate the pressure inside the can while the water was boiling.
   Part C. Estimate the pressure in the can after it collapsed.

4. Use Boyle’s law and say initially the can is full of water vapor at \( 20.0^\circ C \). Calculate a predicted volume of the collapsed can.

5. How does question 3 relate to the accuracy of Boyle’s Law?
9. Crystalline Solids: Geometric Arrays of Atoms, Ions and Molecules

OBJECTIVE: To observe the properties of crystalline materials.

CONCEPT: Crystal structure

INTRODUCTION: There are a limited number of ways that ions can pack into crystals. These are most often based on cubes, tetrahedrons, and octahedrons. Because the shape of a macroscopic crystal is determined by how the ions are arranged, much can be learned by simply observing the angles between plane faces of a crystal. Crystals form and break along straight lines and planes of ions. Amorphous materials, on the other hand, form and break in conchoidal curves.

Because the crystal structure depends largely on the charge and size of the ions in the crystal and not so much on their identity, crystalline minerals often contain foreign ions as impurities. For example, Cu$^{2+}$ or Fe$^{3+}$ may substitute for Zn$^{2+}$.

HAZARDS: Be careful while traveling to the museum. Wear goggles while in the lab.

PROCEDURE: The first part of this activity is to observe the models of unit cells of crystals in the exploratory lab. The second part is to be done in a mineralogy museum. Examine the crystals on display there.

QUESTIONS: (Do 1-4 while observing the models, the remainder in the museum.)
1. Is there a relation between the shape of the unit cell for NaCl and the cubic shape of NaCl crystals? What angles would you expect to find between the cleavage planes?

2. Diamond is made of C atoms bonded together in a three-dimensional network of tetrahedral bonds. Can you find the cleavage planes?

3. Can you find any similarities between the diamond structure and the wurtzite structure?
4. Do you expect the angles between crystal planes for CaCO₃ and TiO₂ crystals to be different than those for NaCl? Why?

5. What are the macroscopic characteristics of a crystalline material, i.e. how do you know something is crystalline simply by looking at it?

6. Most materials are crystalline (e.g. rocks, metals, salt and sugar), not amorphous. Why?

7. Different samples of the same mineral may have different colors. List as many possible causes as you can for the different colors.

8. The same compound can occur in different crystal habits (same crystal structure, but different shaped crystals, e.g. flat rectangles instead of cubes) or polymorphs (same chemical formula, different crystal structures). Can you tell simply by looking at them if two crystals of the same material are polymorphs or just different crystal habits? (Hint: check the angles.)

9. Amorphous materials often have more useful properties than their crystalline analogs, e.g. glass, most plastics, and honey. How are amorphous materials made from crystalline materials? How are crystalline materials made from amorphous materials?

10. What did you like about this experiment and how would you improve it?
10. Dipping Bird

**OBJECTIVE:** To explain an observed phenomenon in terms of fundamental properties of materials.

**CONCEPTS:** Vapor pressure of a liquid, Heat, Work, Scientific method

**INTRODUCTION:**
How did humans learn what was edible and what was not? Perhaps some application of the scientific method was at work. Humans had the problem of hunger, they hypothesized that some foods were edible and some were not, experimented by trial and error, and based on results, made conclusions. In addition to using the scientific method, you will also be using concepts of pressure. A quick review of pressure concepts may help you figure out the puzzle we are about to pose to you.

Pressure can be measured using a mercury barometer. A barometer consists of a glass tube where one end is closed with a vacuum; the other end is open to air. Surrounding air pushes down on the reservoir, and hence the mercury is pushed up into the tube. The level rises until the weight of the mercury in the tube equals the force of the atmospheric pressure. Note that the vacuum does not create “suction,” but that a region of higher pressure will exert a force on a region of lower pressure. That is, the vacuum does not pull the mercury in; the atmospheric pressure pushes it up. Most liquids (and some solids) have a measureable pressure at room temperature. Not all atoms or molecules stay in the liquid; some have enough energy to leave as a gas. This is why water slowly evaporates. This is called vapor pressure.

In today’s lab you will apply these five steps of the scientific method (problem, hypothesis, experiments, results, and conclusion) to solve the mysterious workings of the “dipping bird”. You may have to repeat steps 2–4 several times until you get results that match your hypothesis. You will have a variety of materials available to try different experiments. Good luck!

**HAZARDS:** Be careful when handling the bird and the hand boiler because the glass may break. The liquid inside is methylene chloride with dye, and is toxic.

**MATERIALS:** 1 dipping bird, 3 cups filled with hot, cold and room temperature water, 1 hand boiler, bucket of water, timers, information sheet for methylene chloride, rubbing alcohol in a drip bottle

**PROCEDURE:**

1. Carefully immerse only the bird’s beak in the room temperature water cup. Do not allow the water to drip down to the lower parts of the bird. Stand the bird upright observe for about a minute. Record your observations:
2. To aid you in your quest for understanding of the dipping bird, a handboiler has been provided. Locate the handboiler, and hold one side of the hand boiler in your hand. Record your observations here:

3. Brainstorm with your group and formulate at least 2 other experiments that might help you in determining how the dipping bird and hand boiler work. There are a variety of materials available for you to use. The questions below may also be a helpful guide.

<table>
<thead>
<tr>
<th>HYPOTHESIS</th>
<th>EXPERIMENT</th>
<th>RESULTS</th>
</tr>
</thead>
</table>

**QUESTIONS:**
1. When you initially get out of a pool, or turn off the water in the shower, you have the sensation of feeling cold. Explain on the molecular scale why you feel cold even when the ambient air is room temperature.

2. Often, tires need to be filled in the late fall when it starts to get cooler. Describe the relationship between temperature and pressure.

3. What part of the device works like a straw? Explain how a straw works, while thinking about how the dipping bird works. (Hint: See introduction, do not use the word “suction”.)

4. Explain how the dipping works. As part of your explanation draw pictures showing the molecular view of what’s happening inside of the dipping bird.

| Before | Liquid Rising | During Dip |
11. Electrolysis of Salt Solutions

OBJECTIVE: To show the effects of an electrolytic current on salt solutions.

CONCEPTS: Oxidation-reduction reactions, Electrode reactions

INTRODUCTION: Electrolysis is carried out by immersing two electrodes in the solution to be electrolyzed and applying sufficient voltage to the electrodes to cause a chemical change. At one electrode, electrons are taken from the electrode by the substance being electrolyzed. This electrode, where the substance being electrolyzed is reduced, is the cathode. At the other electrode, electrons are given to the electrode by the substance being electrolyzed. This electrode, where the substance being electrolyzed is oxidized, is the anode.

Electrolysis can be used to decompose salt solutions through oxidation and reduction reactions. Reduction of the cation of the salt may occur at the cathode and oxidation of the anion of the salt may occur at the anode. Alternatively, water may be reduced at the cathode, producing H₂(g) and OH⁻ ions, and/or water may be oxidized at the anode, producing O₂(g) and H⁺ ions. The reaction that actually occurs can be deduced by observing what happens at each electrode.

HAZARDS: The chemicals used in this experiment can cause skin/eye irritations and health problems if ingested. If you come in contact with any of the solutions, thoroughly rinse the affected area immediately with water. If you come in contact with the 0.1 M AgNO₃, the affected area will turn black after exposure to light. This black stain is not harmful and will probably disappear after a few days. Whether you come into contact with the solutions or not, WASH YOUR HANDS AFTER THIS LAB. Wear your goggles!

PROCEDURE:
1. Rinse a small petri dish with distilled water and add 0.1 M KI (with phenolphthalein) to a depth of about 0.5 cm. Phenolphthalein is colorless in neutral and acidic solution and pink in basic solution.
2. Insert the tips of the electrolysis apparatus into the solution, making sure that both sharpened pencil tips are completely submerged. [The carbon pencil "leads" serve as electrodes. The apparatus is a 9 volt battery, 2 pencils, tape, and wire.] Note what happens at each of the electrodes, (+) and (−), and record your observations in the table. The desired results should be visible in 10-15 seconds.
3. After you have finished testing the solution, rinse the pencil tips in the sink with water, pour the used solution in the labeled waste container and thoroughly rinse the petri dish.
4. Repeat Procedures 1-3 using 0.1 M KCl containing blue food coloring.
5. Repeat Procedures 1-3 with 0.1 M AgNO₃ solution. [You should see silver plate out on the tip of one of the pencil electrodes.]
6. Repeat Procedures 1-3 with 0.05 M CuSO₄, except this time dip the pencil tips in the solution 3 or 4 times, leaving them in for just a few seconds each time. Then hold the tips in the light and look for a thin copper coat on one of them. If you leave the pencil tips in the solution too long, a black sludge will form and obscure the copper coat.
7. Do a final rinse of the pencil tips and the petri dishes.

<table>
<thead>
<tr>
<th>Solution</th>
<th>Observed at (+) electrode</th>
<th>Observed at (−) electrode</th>
</tr>
</thead>
</table>

<table>
<thead>
<tr>
<th>Solution</th>
<th>Observed at (+) electrode</th>
<th>Observed at (−) electrode</th>
</tr>
</thead>
</table>

---

Date ____________________________  Identification Number ____________________________  Name ____________________________
**Questions:**

1. You should have noted a color change at the negative electrode in the electrolysis of the KI solution.
   a. Write the chemical formula for the species produced at the negative electrode that is responsible for the color change:
   b. Based on your observations write the chemical formula for another species produced at the negative electrode: _____ What observation helped you to identify this species?
   c. The negative electrode is the (check one) anode ____, cathode ____. Explain how you decided.

2. Write the chemical formula for the species produced at the positive electrode in the KI solution: and in the KCl solution [Hint: What does chlorine bleach do to colors?]: ____. Explain how you decided in each case.

3. Based on your observations write the chemical formula for a species produced at the positive electrode in the AgNO₃ and CuSO₄ solutions: _____ What observation helped you to identify this species?

4. In the table below write balanced equations for each of the half-reactions that occurred.

<table>
<thead>
<tr>
<th>Solution</th>
<th>Half reaction at the (+) electrode</th>
<th>Half reaction at the (–) electrode</th>
</tr>
</thead>
<tbody>
<tr>
<td>KI</td>
<td></td>
<td></td>
</tr>
<tr>
<td>KCl</td>
<td></td>
<td></td>
</tr>
<tr>
<td>AgNO₃</td>
<td></td>
<td></td>
</tr>
<tr>
<td>CuSO₄</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
12. Galvanic Lemon

OBJECTIVE: To demonstrate electrochemical potentials.

CONCEPTS: Electrochemistry, Electrochemical series

INTRODUCTION: A voltaic (after Alessandro Volta, 1800) or galvanic (after Luigi Galvani, 1780) cell uses a chemical reaction to produce electrical energy. When different metals, such as copper and zinc, are placed in a solution of ions, a potential difference (a voltage) develops between them. If the metals are connected externally by a wire, an electrical current (electrons) flows in the wire and a current of ions flows in the solution. The voltage and current result from oxidation-reduction reactions at the electrode surfaces.

Using a lemon as the solution of ions and a variety of metal strips, you will construct several voltaic cells. The lemon juice serves as the conductive solution and the metal strips or wires as the electrodes. Use a voltmeter to determine the potential between several different combinations of metal electrodes. Compare your results with potentials calculated from a table of E\text{°} values. Lemons are very acidic, so use the E\text{°} values for acid solutions.

HAZARDS: Please handle the voltmeter carefully. **DO NOT EAT OR SUCK ON THE LEMON!** Lead is toxic and it is impossible to know what other students may have put on or in the lemons. NEVER eat anything in the laboratory area.

PROCEDURE:

1. Before beginning, wash the metal strips with soap and water. Rinse thoroughly and dry.

2. Insert the copper halfway into one of the pre-cut slits on the lemon and the zinc metal strip in the other. Turn on the voltmeter. Touch one voltmeter lead to the copper strip and the other to the zinc. Record the voltage in the corresponding box of the table provided. You may see the voltage reading jumping around. Wait a moment for it to stabilize before recording the voltage. The stabilizing process is dissolution of metal oxides or metal in the lemon juice. Repeat the measurement if you are unsure. Reverse the placement of electrodes and observe what happens.

3. Repeat the above steps for each remaining combination of metal pairs by replacing the zinc strip with another metal while leaving the copper in place. Be sure to keep the same voltmeter lead attached to the Cu electrode for all the measurements and record the polarity, plus or minus. In the appropriate box of the table, record the cell potential for each metal tested.

4. Using the table of standard electrode potentials in your textbook, calculate E\text{°} for each electrode combination. Record these calculated values in the appropriate box of the table.

5. Remember to **TURN OFF** the voltmeter and return the metal strips to the table.
**CELL POTENTIALS**

<table>
<thead>
<tr>
<th></th>
<th>ZINC</th>
<th>IRON</th>
<th>MAGNESIUM</th>
<th>LEAD</th>
<th>TIN</th>
</tr>
</thead>
<tbody>
<tr>
<td>COPPER (measured V)</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>COPPER (calculated $E^\circ$)</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**QUESTIONS:**

1. Write the redox reaction for the Cu-Zn couple. Suggest where the Cu$^{2+}$ ions come from.

2. What does a negative voltage mean?

3. Why might the measured voltages not agree with the $E^\circ$ values?

4. Arrange the metals in order in an activity series (most active to least active) based on your data.

5. Does this order agree with that given in the text? If not, can you suggest a reason for the disagreement?

6. What purpose does the lemon serve in this experiment? Could another vegetable or fruit be used? If so, suggest some and explain why they would work.

7. Why shouldn't the metal strips be left in the lemon for an extended time?

8. If you have metallic fillings in your teeth, chewing strips of aluminum foil may produce an odd sensation. Explain what is causing the odd sensation.

9. What did you like about this experiment and how could it be improved?
13. Gaseous Molecules: Singles or Couples?

**OBJECTIVE:** To demonstrate that some gases are monatomic and some are diatomic.

**CONCEPT:** Molecular mass is proportional to gas density

**INTRODUCTION:** Gaseous elements have one or two atoms per molecule. How do we know which are monatomic and which are diatomic? The answer is found by measuring the density of the gas. Equal volumes of gases at the same temperature, pressure and volume have equal numbers of molecules. (Whose principle is this?) For a given atomic mass, a monatomic gas has half the density of a diatomic gas.

You will weigh some gaseous elements in a fixed volume at a constant temperature and pressure. The ideal gas law allows the calculation of the molar mass for each gas. The molecular formulas can then be calculated from the molar mass of the element.

**HAZARDS:** Follow the directions on the gas tanks explicitly. **DO NOT** turn knobs without understanding what you are doing. **DO NOT GUESS!** You may ruin one of the valves, or worse, cause an explosion. Ask a TA for help if you have any questions. Take care that you do not knock the cylinders over and create an unguided missile.

**PROCEDURE:**
1. Connect the filter flask to the vacuum pump via the rubber tubing and stopcock. Pump out most of the air (about 2 minutes). Close the stopcock, disconnect the flask from the pump, weigh the evacuated flask, and record the weight.
2. Open the stopcock to fill the flask with air. Weigh the flask filled with air. Remove the stopper and place the hose from the oxygen tank in the flask, turn on the gas and let it run for about 30 seconds to displace the air. Quickly remove the hose and stopper the flask. Weigh the flask filled with oxygen.
3. Repeat this procedure with nitrogen, argon, and helium as available.
4. Record the atmospheric temperature and pressure.
   - Oxygen
   - Nitrogen
   - Argon
   - Helium
   - Air

<table>
<thead>
<tr>
<th>Weight of flask + gas</th>
<th>Oxygen</th>
<th>Nitrogen</th>
<th>Argon</th>
<th>Helium</th>
<th>Air</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Weight of empty flask</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Weight of gas</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Pressure</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Temperature</td>
<td></td>
</tr>
<tr>
<td>Volume of flask</td>
<td></td>
</tr>
</tbody>
</table>

5. Use the ideal gas law, pV=nRT, to determine whether the gas is monatomic or diatomic.
Number of moles \( n = \frac{pV}{RT} \)

<table>
<thead>
<tr>
<th></th>
<th>oxygen</th>
<th>nitrogen</th>
<th>argon</th>
<th>helium</th>
<th>air</th>
</tr>
</thead>
</table>

(Remember to use the Kelvin temperature scale, \( 0^\circ C = 273.15 K \), and \( R = 0.082 \text{ L atm mol}^{-1} \text{ K}^{-1} \).)

Molar mass of gas = \( \frac{\text{weight of gas}}{\text{number of moles}} \)

<table>
<thead>
<tr>
<th></th>
<th>oxygen</th>
<th>nitrogen</th>
<th>argon</th>
<th>helium</th>
<th>air</th>
</tr>
</thead>
</table>

Number of atoms per molecule = \( \frac{\text{molar mass of gas}}{\text{molar mass of element}} \)

<table>
<thead>
<tr>
<th></th>
<th>oxygen</th>
<th>nitrogen</th>
<th>argon</th>
<th>helium</th>
<th>XXXX</th>
</tr>
</thead>
</table>

What is the molecular formula of each of the gases?

- Oxygen
- Nitrogen
- Argon
- Helium

6. Clean up the lab area.

**QUESTIONS:**

1. You have calculated the molar mass of air. Air is a mixture. What is the meaning of your number?

2. Which would lift more if used in a balloon, \( \text{H}_2 \) or He? How much more?

3. Why is He safer than \( \text{H}_2 \) to use in balloons?

4. What did you like about this experiment and how would you improve it?
### 14. Household Materials: Acids or Bases?

**OBJECTIVE:** To become familiar with common acids and bases.

**CONCEPTS:** Acids, Bases, pH, Indicators

**INTRODUCTION:** You will test various household items to determine whether they are acids or bases. Acidic and basic character of a solution can be determined with indicator paper which indicates pH by color. Universal indicator paper contains a mixture of indicators that changes to red with strongly acid solutions and blue with strongly basic solutions with a range of colors in between. The color scale to indicate the pH is posted in the lab. The pH is equal to the negative log$_{10}$ of the hydrogen ion concentration in the solution. Pure water has a pH = 7; therefore, pH 7 is neutral, neither basic nor acidic. Acidic solutions have a pH less than 7 and basic solutions a pH greater than 7. The actual pH depends both on the strength and concentration of the acid or base in the solution.

**HAZARDS:** None of the substances in this lab should be ingested. Avoid contact with skin when handling these materials. If any is spilled on your skin, flush with water. To protect other students, carefully clean up all spills and the work area when finished.

**PROCEDURE:**

<table>
<thead>
<tr>
<th>Bottle</th>
<th>Substance</th>
<th>pH</th>
<th>Acid/Base/Neutral</th>
<th>Bottle</th>
<th>Substance</th>
<th>pH</th>
<th>Acid/Base/Neutral</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Salt</td>
<td>6</td>
<td></td>
<td>6</td>
<td>Ammonia</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>Carbonated drink</td>
<td>7</td>
<td></td>
<td>7</td>
<td>Scouring powder</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>Hand soap</td>
<td>8</td>
<td></td>
<td>8</td>
<td>Sugar</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>Dishwasher detergent</td>
<td>9</td>
<td></td>
<td>9</td>
<td>Distilled water</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5</td>
<td>Chlorine</td>
<td>10</td>
<td></td>
<td></td>
<td>Vinegar</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

1. Place ten small pieces of indicator paper on a piece of plastic wrap. Put 1 or 2 drops of each of the ten solutions on a piece of indicator paper. Observe and record the pH in the table above.

2. Discard the used indicator paper and plastic wrap in the appropriate waste container.

3. Return all pipets to the proper bottles and replace the lids on all the liquids. Clean the lab table.
**QUESTIONS:**

1. What do the terms acidic and basic mean?

2. Which materials appear to be neutral (pH between 6 and 8)? Note that distilled water should be slightly acidic because it contains CO$_2$ from the air.

3. Which materials were acidic? Did any appear to be more acidic than the others and if so how could you tell?

4. Which materials were basic? Did any appear to be more basic than the others, and if so, how could you tell?

5. Is there a problem in using indicator paper to determine the pH of some of the materials? Explain.

6. What other acids and bases do you come in contact with every day?

7. What did you like about this experiment and how would you improve it?
15. How Big Is a Mole?

OBJECTIVE: To measure and observe 1 mole each of several substances.

CONCEPTS: Mole, Avogadro's number, Molecular mass

INTRODUCTION: The SI unit for the amount of something is mole, abbreviated mol. The word "mole" comes from the Latin word "moles" which means "a mass." The word "molecule" means "a small mass." A mole is defined as the amount of substance that contains as many entities (atoms, molecules or other particles) as there are atoms in 12 grams of pure carbon-12 atoms. Therefore 1 mole = 6.02 x 10²³ particles. The number 6.02 x 10²³ is called Avogadro's number in honor of Amedeo Avogadro (1776-1856). Thus, a mole of apples is 6.02 x 10²³ apples just as a dozen apples is 12 apples. A mole of cherries would also contain 6.02 x 10²³ cherries, but would weigh much less and occupy much less volume than a mole of apples. In the same way, a mole of sulfur and a mole of carbon each contain 6.02 x 10²³ atoms, but they have different masses and different volumes. The molar mass of an element is equal to the number of grams of the element containing 6.02 x 10²³ atoms. Thus, one mole of sulfur has a mass of 32.066 grams and one mole of carbon 12.011 grams. The molar mass of a molecular compound is equal to the number of grams of the compound containing 6.02 x 10²³ molecules. In an older terminology, which is still used informally, the molar mass of an element is called its atomic weight and the molar mass of a compound is called its molecular weight.

In this experiment you will weigh out a mole of several different substances and observe the differences in mass and volume.

HAZARDS: Do not consume any of the substances provided. REUSE all substances to reduce waste.

PROCEDURE:

1. Use a periodic table to determine the molar mass of each substance listed in the table below.

2. Weigh out a mole (a molar mass, grams) of each of the six substances provided. Estimate the volume of the sample from the marks on the beaker. Do NOT attempt to determine the volume by displacement of water. Keep the solid samples clean and dry - do NOT put them in a wet beaker.

3. Carefully return all substances to the proper containers and clean up any spills.
<table>
<thead>
<tr>
<th>SUBSTANCE</th>
<th>Molar Mass, (grams)</th>
<th>Mass of the mole as you measured it (grams)</th>
<th>Approx. volume of sample (mL)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sugar, C_{12}H_{22}O_{11}</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Water, H_{2}O</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Salt, NaCl</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Aluminum, Al</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Copper, Cu</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Lead, Pb</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**QUESTIONS:**

1. What is a mole?

2. Order the molar masses of the substances from the lowest to the highest. Do the molar volumes follow the same order? Why or why not?

3. Why is the mole so useful in chemistry?

4. Why does a mole of water weigh less than a mole of sugar?

5. Would it be practical to sell eggs by the mole? Explain.

6. What did you like about this experiment and how would you improve it?
16. How Low Can You Get?

Name ____________________________ Course Number ____________________________
Student ID ____________________________ Section Number ____________________________
Date ____________________________ Instructor ____________________________

**OBJECTIVE:** To demonstrate the effects of solute concentration on the freezing point of water.

**CONCEPTS:** Freezing point depression, Dissociation of electrolytes

**INTRODUCTION:** A pure substance has a characteristic freezing or melting point. Any solute will lower the freezing point. The magnitude of the freezing point depression for a given *solvent* depends only on the concentration of *solute particles*.

Freezing point depression is a "colligative" property; that is, it depends on the total collection of particles present, not their identity. In the case of a nonelectrolyte like sucrose there is only one kind of particle present in aqueous solution. On the other hand, electrolytes dissociate into separate positive and negative ions; with NaCl, both Na\(^+\) and Cl\(^-\) particles are present in aqueous solution and the freezing point depression is twice what it would be for a nonelectrolyte with the molar mass of NaCl.

**HAZARDS:** Beware of frostbite.

**PROCEDURE I: Salt in Water**

1. Fill a 6 oz styrofoam cup half full with ice.

2. Weigh out 10 g of NaCl in a plastic weighing boat.

3. Fill the cup with water to the ice level and record the temperature. __________

4. Add the 10 g portion of NaCl, stir, and record the new equilibrium temperature (freezing point) after the temperature stabilizes. __________

5. Stir and add 10 g portions of NaCl and ice as needed and record the lowest temperature you are able to obtain. __________

**QUESTIONS I:**

1. Which substance is the solvent? Which is the solute?

2. How low did you get? If a solution is saturated, can you depress its freezing point any further?
3. Would salt melt the ice on a road at -20°F? Why or why not? (You may need to convert Celsius to Fahrenheit to figure this out.)

**PROCEDURE II: Sugar in Water**

1. Repeat the above procedure using sucrose (table sugar) instead of salt.

2. Record the temperature after 10 g of sucrose is mixed with the ice water. _______

3. Record the lowest temperature attainable with sucrose. ________________

**QUESTIONS II:**

1. Which produces the lower freezing point in the half cup of water, 10 g NaCl or 10 g C_{12}H_{22}O_{11}? Why?

2. Which solution gives a lower freezing point, 1 m NaCl or 1 m sucrose?

3. Why do all solutes lower the freezing point?

4. What did you like about this experiment and how would you improve it?
17. Hungry Pyromaniac

**Objective:** To demonstrate the energy stored in the chemical bonds of ordinary food items.

**Concepts:** Calorimetry, Energy content of foods, Heat capacity

**Introduction:** As we eat, respiration releases the energy from food and makes it available for movement, growth, and thinking. Food is the fuel for the human body; gasoline is a common fuel for vehicles. The engine burns the fuel which in turn gives off energy. That energy is then directed towards propelling the vehicle forward. The energy used to move a car is similar to the energy that we get from our meals. In this lab we use calorimetry, the measurement of heat, to calculate how much energy is contained in certain foods.

**Background:** We will measure the amount of energy in a food by combusting (burning) it. Using known values of temperature change (ΔT), the mass (m) of the sample, and its specific heat (c), we will calculate the amount of heat given off (Q) using the equation, Q = mcΔT. Keep in mind that a food Calorie (note the upper case C) is equal to 1000 small calories or 1 kcal. (1C = 1000c or 1C = 1 kcal) A small calorie (cal) is the energy required to raise the temperature of one gram of water one degree Celsius, and is equal to 4.184 Joules.

**Hazards:** Care should be taken when burning items. This experiment generates a lot of smoke; do all burning inside a fume hood. Do not eat any of the food samples provided as they may be contaminated by other chemicals and could cause injury, illness or death.

**Procedure:**
1. Pour 150 ml of water into the empty soda can using the graduated cylinder provided.
2. Choose three food samples and weigh out an equal mass of each onto a piece of aluminum foil (about 1 gram, it doesn’t need to be exact but you do need to write down all of the digits displayed on the balance). Record your values in the table on the back page.
3. Also record the initial temperature of the water in the can before burning the samples.
4. Place the food sample with the aluminum foil on the wire mesh stand.
5. Warm the sample by moving the lighter all over the sample, then hold the lighter in one place to ignite it. Try to minimize heat loss to the atmosphere by placing the burning sample close to the can. The winged screw on the back of the ring clamp will allow you to move the can up and down. You may also want to fold up the foil in back to reduce the hot air sucked in by the vent.
6. Periodically stir the water in the can with the stirrer.
7. When the food stops burning, record the final water temperature.
8. Record the mass of the sample after burning. Subtract this mass from the initial sample mass to get the mass of the food burned. Record the results in the table.
9. Repeat steps 1–6 for the two other food samples.
10. Calculate the heat of combustion using a heat capacity of 0.150 kcal/°C (this is an estimate for the aluminum can and 150 mL of water).
11. To calculate the heat of combustion per gram of food divide the heat of combustion in step 10 by the change in mass of the food burned.
12. Clean up the area and throw away all waste.

<table>
<thead>
<tr>
<th>Name of food type 1</th>
<th>Name of food type 2</th>
<th>Name of food type 3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Initial mass of food sample with aluminum foil, g (step 2)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Final Mass (mass of food after combustion with aluminum foil) g (step 8)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Mass of food burned, g</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Initial temperature of water in can, °C (step 3)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Final temperature of water in can, °C (step 7)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Change in temperature, ΔT in °C</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Calories (kcal) of heat = (0.150 kcal/°C) × ΔT (step 10)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Heat of combustion per gram of food, kcal/g (step 11)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Posted heat of combustion per gram of food, kcal/g</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Questions:**
1. Do your experimental results seem close to the accepted values posted in the lab?
2. What is the main source of error in this experiment?
3. Consider a food company with a new recipe for potato chips. Comment on how valuable the technique from this lab would be to determine the number of Calories to list on the potato chip package? Include ideas on how this experiment could be improved to make the experiment more accurate.
18. Hydrogen-Oxygen Rockets

Name __________________________ Course Number ________________________
Student ID __________________________ Section Number ________________________
Date __________________________ Instructor __________________________

**OBJECTIVE:** To demonstrate the importance of proportions of reactants in a reaction.

**CONCEPTS:** Chemical formulas, Stoichiometry, Energy of reaction, Law of combining volumes, Hydrogen, Oxygen

**INTRODUCTION:** To understand chemistry, one must first understand the proportions required for a reaction to be complete, or understand “stoichiometry” as it is referred to in chemistry. As a practical example, you will optimize the proportions in a mixture of H\textsubscript{2} and O\textsubscript{2} gas to maximize flight distance of a plastic bulb. When hydrogen and oxygen are present in the appropriate amounts for a complete reaction, we maximize the amount of energy that can be released from the fuel in the form of hot water molecules. At these conditions, the reactants are present in “stoichiometric” amounts. It is important to note that for a given temperature and pressure equal volumes of different gases have an equivalent number of molecules. This means that one liter of H\textsubscript{2} gas has the same number of molecules as one liter of O\textsubscript{2} gas.

**HAZARDS:** Do not adjust any of the gas flow devices, if you are having trouble with any part of the gas delivery device, ask the TA to take care of the problem. An additional hazard in this lab, is the Tesla coil which generates high voltage (~50,000 Volts ) and a high frequency spark that can be painful and cause serious burns. **Be extremely careful with the Tesla coil!**

**PROCEDURE:**

1. Fill the “rocket” (the top of a plastic pipet) by submerging it in the water tank, and squeeze out all of the air. Note that the water does not run out of the rocket when held vertically with the stem down.

2. Place the rocket on the tip of the hydrogen supply, hydrogen will bubble in to the bulb. The amount of gas you allow into the bulb will depend on the experimental run (see step 6).

3. Transfer the rocket to the oxygen supply and fill the rocket with oxygen until only the stem (about a quarter of an inch) has water left in it. The distance the rocket travels depends on the amount of water remaining in the stem, so be consistent with how much water you leave in the bulb.

4. Place the rocket on the launcher. Tilt the launcher to a 45 degree angle.

5. Pick up the Tesla coil by the brown portion, using one hand, and then plug it in. To ignite the rocket, the tip of the Tesla coil will make a spark when brought close to the very tip of the nail, as in the figures at right. Record the distance the rocket travels, starting at the launcher and ending where the rocket first hits the ground.

6. Repeat the above procedure and fill in the chart below.
Questions:
1. Which contains more molecules, a bulb full of hydrogen or a bulb filled with oxygen?

2. Why wouldn’t a rocket launch when it is filled with hydrogen gas only?

3. Which volume ratio went the longest distance?

4. Why did the rocket go different distances with different ratios?

5. Write a balanced equation for the reaction of $H_2$ and $O_2$. (hint: refer to intro for products)

6. Are your results consistent with the balanced reaction in question 5? If not, propose some reasons why.
OBJECTIVE: To demonstrate one method for distinguishing a mixture from a pure compound by separating the mixture.

CONCEPTS: Separations, Chromatography, Pure substances, Mixtures

INTRODUCTION: Chromatography is probably the most widely used and most powerful of all the techniques of chemical separation. Chromatography was invented in the late nineteenth century by the Russian botanist Mikhail Tswett, who used a column of a powdered solid such as calcium carbonate to separate naturally occurring plant pigments. Bands of color separated along the length of the column when a solution of the pigments was passed through the column. Tswett claimed that the bands represented different compounds. Chromatography is now commonly used to investigate the chemistry of complex natural mixtures.

Chromatographic separations are achieved by passing a sample (generally a mixture) through a system of two phases. One phase is stationary and one is mobile (moving) and moves the sample. Each component of the sample will distribute differently between the two phases. Therefore the components of the sample move at different rates. Components that have a high affinity for the stationary phase will move slower than components that tend to stay mostly in the mobile phase. Thus, individual components of a sample can be selectively separated. If a substance cannot be separated with any of several combinations of stationary and mobile phases, it can probably be assumed to be a pure material, i.e. a compound or an element.

HAZARDS: Do not consume any of the substances in this lab. Sodium hydroxide (NaOH) can burn your skin, so be sure to wash your hands if you get it on you. Clean up spills and be sure to wear your goggles.

PROCEDURE:
1. Place a small petri dish within a larger one and fill the small petri dish approximately half way with 0.1 M NaOH solution (the mobile phase).

2. Using a pencil (no ink!), place a mark at the center of a circular filter paper (the stationary phase) that will cover the large petri dish. Draw a circle about 2.5 cm (1 inch) in diameter around this mark.

3. From the selection of concentrated Kool-Aid solutions choose four samples, preferably of different colors, to use in your experiment. Place one small drop of each of the Kool-Aid solutions on the circle about 90 degrees apart and label each dot (in pencil!) according to its flavor.

4. Punch a small hole through the center mark with your pencil. Roll one of the pieces of filter paper into a wick and insert it through the hole until it is a little more than halfway in. Place the circular filter paper with samples on it on the petri dishes so that only the wick touches the NaOH.
5. Allow time for the liquid to soak into the paper and flow outward until the liquid almost reaches the edge of the filter paper. The colors will move along with the liquid, some faster, some slower. Remove the paper and allow it to air dry on a paper towel.

6. You are welcome to test other materials such as ink to determine if they are mixtures or pure substances.

7. Dispose of all waste in the appropriately labeled containers. Please clean up the lab area when you are finished.

**QUESTIONS:**

1. Describe your results. Which samples separated into multiple components? From your results, which samples are pure compounds and which are mixtures?

2. What is the stationary phase in your experiment? The mobile phase? What causes the sample components to move?

3. Which color of dye in the samples has the greatest affinity for the stationary phase? The mobile phase?

4. In what ways could the separation of the dyes be improved? (There are at least three ways!)

5. Why can’t ink be used to mark the paper? Why is it all right to use a pencil?

6. What did you like about this experiment and how would you improve it?


20. Lead Balloon

<table>
<thead>
<tr>
<th>Name</th>
<th>Course Number</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Student ID</th>
<th>Section Number</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Date</th>
<th>Instructor</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**OBJECTIVE:** To observe some consequences of Graham's law of diffusion.

**CONCEPTS:** Graham's law of diffusion, Kinetic molecular theory of gases

**INTRODUCTION:** Diffusion, the ability of a substance to mix spontaneously with and spread throughout another, is a property we experience daily in many ways. Diffusion of gases is demonstrated by the smell from a passing skunk, the aroma of mom's homemade bread greeting you at the door, and the fragrance encircling the homecoming dance. Another example of diffusion is the movement of a gas through a permeable membrane. In the latter case the gas dissolves in, and moves through, the molecules making up the membrane.

Gases of low molecular mass have greater average molecular speeds and therefore greater diffusion rates than gases of high molecular mass. Under identical conditions of temperature and partial pressure, the ratio of the diffusion rates of two gases is inversely proportional to the square root of the ratio of their molecular masses.

\[
\frac{\text{Rate of diffusion of gas 1}}{\text{Rate of diffusion of gas 2}} = \frac{\sqrt{\text{Molecular mass of gas 2}}}{\sqrt{\text{Molecular mass of gas 1}}}
\]

This relation is commonly called Graham's Law after Thomas Graham, a Scottish chemist who studied gases in the early 1800s.

**HAZARDS:** SF₆ is nontoxic, but can cause asphyxiation.

**PROCEDURE:** Three balloons filled respectively with air, helium and sulfur hexafluoride in approximately equal volumes are available. Observe the different masses of the balloons resulting from the different densities of the gases. Observe how the balloons displayed in the lab change over 24 hours. Record your observations.
**Questions:**

1. Why is the $\text{SF}_6$-filled balloon heavier than the air-filled balloon, which is in turn heavier than the $\text{He}$-filled balloon?

2. How are the weights of the gases in the balloons related to the molecular masses?

3. From a table of atomic masses, determine the molar masses of $\text{He}$, $\text{SF}_6$ and the average molar mass of air (assume 20% $\text{O}_2$, 80% $\text{N}_2$).

4. What differences did you observe in the way the balloons changed over a 24 hour period? Explain these changes in terms of Graham's law.

5. Which gas do you think would be the hardest to get out of your lungs and why? Suggest the best posture to assume to remove this gas from the lungs.

6. What did you like about this experiment and how would you improve it?
21. M&Ms—Moles & Molecules

<table>
<thead>
<tr>
<th>Name</th>
<th>Course Number</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Identification Number</th>
<th>Section Number</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Date</th>
<th>Instructor</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**OBJECTIVE:** To demonstrate the concept of "counting by weighing" and relate it to the concept of a mole.

**CONCEPTS:** Counting by weighing, Mole, Experimental errors

**INTRODUCTION:** This experiment illustrates the concept of counting by weighing. Any kind of small candy that comes in uniform sized packages will work. The number of pieces of candy in the package is determined by weighing all of the pieces together and then weighing one piece. Atoms and molecules are counted the same way. The weight of one atom is difficult to measure. Avogadro’s number was defined to make calculations and measurements easy, so that 1amu = \(\frac{6}{mol}\). This way chemists can look at a periodic table and relate the atomic weight to either amu or to \(\frac{g}{mol}\). For example, looking at the periodic table lithium has an atomic mass of 6.941. Because of Avogadro’s number when units are included the value can be either \(6.941 \frac{amu}{atom}\) or \(6.941 \frac{g}{mol}\) depending on which works best for the situation. A mole is just a number so

\[
\text{Avogadro’s number} = 1 \text{ mole} = 6.02 \times 10^{23} \text{ items} = 6.02 \times 10^{23} \frac{amu}{g} = 6.02 \times 10^{23} \frac{atoms}{mol}
\]

**HAZARDS:** Do not consume any of the food in this lab!

**PROCEDURE:**
1. Write down the claimed mass from the packaging (A)
2. Place a cup on a balance and zero the reading. Pour the M&Ms into the cup and record the mass (B).
3. Weigh and record the mass of just one M&M (C).
4. Calculate the total number of M&Ms (B/C = D).
5. Count the actual number of M&Ms (E).

**DATA:**

<table>
<thead>
<tr>
<th>Units</th>
<th>Meaning of Number</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>Claimed Mass</td>
</tr>
<tr>
<td>A</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Units</th>
<th>Meaning of Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>B</td>
<td>Mass M&amp;Ms</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Units</th>
<th>Meaning of Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Units</th>
<th>Meaning of Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>D</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Units</th>
<th>Meaning of Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>E</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Units</th>
<th>Meaning of Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>C</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Units</th>
<th>Meaning of Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>D</td>
<td></td>
</tr>
<tr>
<td>E</td>
<td></td>
</tr>
</tbody>
</table>
**Questions:**

1. Was there an error in your calculated count? If your mass measurement of a single M&M is high by 0.02 g, what difference in calculated count would result?

2. Determine the average mass of one M&M from the mass of five or ten M&Ms (C₂). Recalculate D (D₂). Do you come closer? Explain.

3. What are some reasons that the original calculated count (D) could have been in error? (Hint: Do all M&Ms weigh the same?)

4. What is the most efficient way for M&M/Mars Co. to measure the candy into the package? Is it accurate?

5. What would a mole of M&Ms weigh?
22. Models of Molecules

<table>
<thead>
<tr>
<th>Name</th>
<th>Course Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Identification Number</td>
<td>Section Number</td>
</tr>
<tr>
<td>Date</td>
<td>Instructor</td>
</tr>
</tbody>
</table>

**OBJECTIVE:** To make and observe models of molecular shapes.

**CONCEPT:** Close packing of spheres, VSPER theory, Molecular shapes

**INTRODUCTION:** Shapes of molecules are of considerable use in predicting both physical properties (polarity, boiling point, etc) and chemical properties (lack of reactivity, as with SF$_6$; or reactivity, as with lone pairs or pair vacancies for Lewis acid-base reactions). In biochemical reactions, enzymes will only interact with molecules of specific shapes. The shapes of molecules are described in terms of the geometric object defined by the positions of the nuclei around a central nucleus. Most molecular models use different colored balls to represent different atomic nuclei and connecting rods to indicate pairs of electrons.

In this experiment, model sets with linear, trigonal planar, tetrahedral, trigonal bipyramidal, and octahedral center shapes are available. The connecting tubes represent pairs of electrons, both in bonds and as lone pairs. Small spheres represent atoms bonded to the central atom. These models fairly accurately represent bond angles, but not relative sizes of atoms.

**PROCEDURE:** Models of CO$_2$ (linear), BF$_3$ (triangular planar), CH$_4$ (tetrahedral), PF$_5$ (trigonal bipyramidal), and SF$_6$ (octahedral) are displayed to show you how they are made.

1. Make models of the following molecules. A linear molecule (example BeCl$_2$(g)). A trigonal planar molecule (example BCl$_3$). A tetrahedral molecule (example SiCl$_4$). A trigonal bi-pyramidal molecule (example PCl$_3$). An octahedral molecule (example SF$_6$). Sketch these models. Label your sketches.

2. Make models of molecules with one lone pair of electrons by removing one ball and its corresponding tube from each of the models constructed in 1 above. You will then have models of bent (SO$_2$), trigonal pyramidal (NH$_3$), see-saw (SF$_4$), and square pyramidal (BrF$_5$) molecules. Note there are two ways to remove a ball from the trigonal bi-pyramid since the axial and equatorial positions are not equivalent. Only molecules with a vacant equatorial position, i.e. a distorted tetrahedron (known as a see-saw), are ever observed. Sketch these models. Label these sketches. Do not draw the lone pairs, only the nuclei and bonds. (Lone pairs are not considered in naming the shape of a molecule.)

3. Removing another ball from the trigonal pyramid, distorted tetrahedron, and square pyramid gives models of bent (H$_2$O), T-shaped (ClF$_3$), and square planar (XeF$_4$) molecules. Sketch the shapes of these molecules and label your sketches. Note the angle of the bent molecule is 109.5° instead of 120°.

4. Isomers are molecules having the same chemical formula, but different properties. Two
important types of isomers are geometric and optical.

A. Geometric isomers.
Make a model of the octahedral ion CoF₄Cl₂³⁻. Use two different colors of balls to represent the F⁻ and Cl⁻ ligands. There are geometric isomers of this ion. A cis-isomer has the Cl⁻ ligands next to each other. A trans-isomer has the Cl⁻ ligands on opposite sides of the molecule. How many geometric isomers are there? Sketch the cis and trans isomers of CoF₄Cl₂³⁻ and label your sketches.

B. Optical isomers.
Optical isomers are nonsuperimposable mirror images. Make models of the optical isomers of CHFClBr. Sketch these isomers and label your sketches.

5. Close packing of spheres. For atomic solids composed of spherical atoms, such as metals, one of the two close packing arrangements (cubic close pack or face centered cubic close pack) is usually found in natural crystals.
   a. Place 15 or 20 tennis balls on the table top in close-pack arrangement with six balls around each ball. Use barriers to keep them from rolling away. Notice that the greatest density of balls (i.e. most balls/unit area) is achieved when each ball has six nearest neighbors. Can you maintain this "close packing" of the balls in three dimensions?
   b. Place a second layer of balls on top of the first layer. The "close packing" in the second layer is maintained because the balls in the second layer fit into the trigonal depressions formed by three balls in the lower layer. Show there are two ways to place the second layer. There are twice as many depressions as can be used, i.e. two sets of depressions. (Hint: the two ways differ by moving the entire second layer in one direction by approximately one ball radius.)
   c. Place a third layer of balls on top of the second layer. The same two choices of depressions over which to place the third layer exists, but now it will make a difference. Note that one set of depressions has balls from the first layer directly under it (either look down through the holes or drop in a pencil in to verify this). If you put the third layer in these depressions, and continue the same pattern (i.e., balls over balls), we call this ABABAB stacking or hexagonal close pack. If you put the third layer in the depressions over empty spaces in the first layer, and continue the same pattern (i.e., balls over holes), we call this ABCABC stacking or cubic close pack.
   d. Stack enough ABABAB and ABCABC layers so that you can see what the pattern is. Identify the hexagonal repeating unit in ABABAB stacking. Identify the face-centered cubic repeating unit in ABCABC stacking. (Hint: the face-centered unit cell is not horizontal; it is balanced on one point and may be rotated).

6. What did you like about this experiment and how could it be improved?
23. Oxidation States of Manganese

| Name __________________________ | Course Number __________________________ |
| Identification Number __________________________ | Section Number __________________________ |
| Date __________________________ | Instructor __________________________ |

OBJECTIVE: To produce compounds or ions containing manganese in different oxidation states and observe their colors.

CONCEPT: Colors of transition metal compounds

INTRODUCTION: Manganese is typical of d-transition elements in that its ions are colored and are different colors in different oxidation states. In this experiment you will be looking at compounds or ions with Mn in +2, +4, +6, and +7 oxidation states.

The reactions that produce the different oxidation states of manganese are:

(+2): \[2\text{MnO}_4^- + 6\text{H}^+ + 5\text{HSO}_3^- = 2\text{Mn}^{2+} + 5\text{SO}_4^{2-} + 3\text{H}_2\text{O}\] (manganese(II) ion)

(+4): \[2\text{MnO}_4^- + 3\text{HSO}_3^- = 2\text{MnO}_2(\text{s}) + 2\text{SO}_4^{2-} + \text{HSO}_4^- + \text{H}_2\text{O}\] (manganese dioxide)

(+6): \[2\text{MnO}_4^- + 3\text{OH}^- + \text{HSO}_3^- = 2\text{MnO}_4^{2-} + \text{SO}_4^{2-} + 2\text{H}_2\text{O}\] (manganate ion)

(+7): \[\text{MnO}_4^-\] (the experiment begins with this ion of manganese, permanganate ion)

HAZARDS: Rinse with lots of water if you spill any of these solutions on your skin or clothes. Potassium permanganate will stain clothes and skin. The strong acid (H$_2$SO$_4$) and strong base (NaOH) will irritate your skin and put holes in your clothes. To protect others, carefully clean up any spills.

PROCEDURE:

1. Designate four wells in a plate (or four test tubes) as +2, +4, +6, and +7. This will help you to keep track of the different reactions.

2. Put 2 drops of 0.01 M KMnO$_4$ solution in each of the 4 wells. Add 2 drops of 3 M H$_2$SO$_4$ solution to the well labeled +2. Add 2 drops of 2 M NaOH solution to the well labeled +6.

3. While stirring, add 0.01 M NaHSO$_3$ solution dropwise to the well labeled +2 until there is no further color change (10–15 drops).

4. While stirring, add 0.01 M NaHSO$_3$ solution dropwise to the well labeled +4 until a brown precipitate appears (10–15 drops).

5. While stirring, add 0.01 M NaHSO$_3$ solution dropwise to the well labeled +6 until the bright green color of MnO$_4^{2-}$ appears (10–15 drops).

6. Clean up any spills.
**Questions:**

1. Give some reasons for the different colors of the different compounds and ions of manganese.

2. Why isn't the Mg$^{2+}$ ion colored while the Mn$^{2+}$ ion is pale pink?

3. For each of the reactions on the previous page, what is being oxidized and what is reduced?

4. Why are different products formed from the same MnO$_4^-$ and HSO$_3^-$ reactants in the different wells?

5. Why doesn't manganese in the +7 state exist as the Mn$^{7+}$(aq) ion?

6. What purpose does well +7 serve?

7. What did you like about this experiment and how would you improve it?
24. pH Indicators

Name __________________________ Course Number __________________
Identification Number ________________________ Section Number _________________
Date ________________________________ Instructor _________________________

OBJECTIVE: To demonstrate how indicators can be used to determine the pH of a solution.

CONCEPTS: pH, Ionization of weak acids and bases, Indicators

INTRODUCTION: Indicators are used to give a visual representation of the conditions within a solution. A pH indicator is a weak acid that changes color upon losing a H⁺ ion (proton). Whether the indicator will be present in its protonated or unprotonated form depends on the concentration of hydrogen ions ([H⁺]) and its dissociation constant, Kᵢ. Mathematically, pH is related to hydrogen ion concentration by the following relationship: pH = −log₁₀([H⁺]). When half the indicator has reacted, i.e. at the midpoint of the color change, the Kᵢ = [H⁺]. In this lab you will experiment with how indicators change color with changes in pH.

HAZARDS: Wear your goggles. Wash your hands or affected area with copious amounts of water if you have any skin contact with laboratory chemicals, especially if you have a localized burning feeling on your skin.

PROCEDURE:
1. Obtain a 96-well plate. In the first well, add 9 drops of water with a pipette. Use this volume as a gauge to fill the rest of the wells. Fill up the rest of the wells with distilled water (“DW”- single knob faucet) at the sink. Tip the plate slightly (over the sink) to drain and shake off excess water from the plate. It is important that the wells are almost completely filled, about 90% full. Do not overfill wells! This step should be quick, the 90% is an approximation. Many students have been successful by first overfilling the 96-well plate and then slowly turning the plate upside-down over the sink. This drains excess water, but water sticks inside the wells leaving them about 90% full.
2. The wells are numbered. Place the 96-well plate on a white paper towel to better see the numbers. It is labeled 1 through 12 across the top, and going down the left side the letters A through H are used.
3. Using a glass Pasteur pipette, draw out the water from all the wells in the first and last columns.
4. Place 8–10 drops of 0.1M HCl in each well in the first column (i.e. in wells A1–H1) or until full. Place 8–10 drops of 0.01M NaOH in each well in the last column (i.e. wells A12–H12) or until full.
5. Your top row is your “stock solution” row, or a place to store chemicals you will need later. With a clean pipette, transfer one drop of acid from well A1 to each well in column 2. Return excess acid to well A1.
6. Transfer one drop from A2 to each well in column 3. Continue this procedure up to column 6. (Column 7 should be left unchanged). This will establish a series of wells with differing degrees of acidity.
7. Next, you will need to prepare a set of wells of differing degrees of basicity. So, continue with the dilutions working backwards from A12 to column 8.
8. Note that you have not added anything but water to the wells in column 7.
9. Add 1 drop of an indicator solution to each of the 12 wells in row B. Choose 6 more different indicators to add to each row C–H.
10. Determine the pH of the solution in each well. (Hint: The first well is at pH 1, well 7 is at pH 7 and
well 12 is at pH 12.) Write these values for pH in the chart below. Also determine the hydrogen ion concentration at each pH.

11. Next, record a color for each pH in the table for each indicator you are using, you may use extended arrows if the color remains about the same.

12. When you finish, use the wash bottle to rinse the solutions from the 96 well plate into the plastic waste receptacle provided.

+----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+
|                | Column          | 1               | 2               | 3               | 4               | 5               | 6               | 7               | 8               | 9               | 10              | 11              | 12              |                |
+----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+
|                | pH              | [H^+]           | [H^+]           | [H^+]           | [H^+]           | [H^+]           | [H^+]           | [H^+]           | [H^+]           | [H^+]           | [H^+]           | [H^+]           | [H^+]           |                |
| Row            | Indicator       | Row             | Row             | Row             | Row             | Row             | Row             | Row             | Row             | Row             | Row             | Row             | Row             |                |
+----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+
| B               |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                |
+----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+
| C               |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                |
+----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+
| D               |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                |
+----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+
| E               |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                |
+----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+
| F               |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                |
+----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+
| G               |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                |
+----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+
| H               |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                 |                |
+----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+-----------------+

**QUESTIONS:**

1. For each indicator, find the pH(s) at which it changes color, and put a small asterisk in that box.
   What is the [H^+] at the color change (do this for three of your indicators)?
   - Indicator Name: [ ]
   - pH(s): [ ]
   - [H^+]: [ ]

2. A particular experiment will be performed where a pH change occurs around pH 4. Which of the indicators you used will best fit this experiment?

*Optional:* What did you like about this experiment and how would you improve it?
25. Precipitation Stoichiometry

CONCEPTS: Chemical reactions, Stoichiometry, Precipitation, Limiting reactants

INTRODUCTION: When chemicals react, there is an optimum amount of each reactant required to maximize the yield of the product. In this lab, you will maximize the amount of precipitate formed for three insoluble compounds, make observations, and predict the balanced chemical reaction.

HAZARDS: Be careful not to get any of the chemicals on you or your clothing. You will dislike the stains or holes that result. Please wear your goggles. Pb is toxic; avoid contact or ingestion.

PROCEDURE:
Please work in groups to minimize the amount of waste generated.

1. Using a permanent marker, number and label 15 test tubes. The 15 tubes will be divided into three groups, A, B, and C. Group A will consist of numbers 1 to 3, group B will be numbers 4 to 8, and group C, numbers 9 to 15.
2. Add reagents to each test tube as shown in the tables and stir each mixture by capping the test tube and shaking it.
3. Centrifuge or allow the tubes to stand so the precipitates will settle. If using a centrifuge to separate the precipitate from the liquid, be sure to counter-balance the centrifuge by using a second tube filled with water to the same level as the solution in the first tube and set them opposite each other in the centrifuge head.
4. Rate each test tube for the amount of precipitate, with the largest number representing the most precipitate and the smallest the least. Use the same integer for test tubes that appear to contain the same amount of precipitate.

<table>
<thead>
<tr>
<th>Group A</th>
<th>Tube number</th>
<th>1</th>
<th>2</th>
<th>3</th>
</tr>
</thead>
<tbody>
<tr>
<td>Drops of 1 M CaCl₂</td>
<td>5</td>
<td>10</td>
<td>15</td>
<td></td>
</tr>
<tr>
<td>Drops of 1 M Na₂CO₃</td>
<td>15</td>
<td>10</td>
<td>5</td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Group A</th>
<th>Amount of Precipitate</th>
<th></th>
</tr>
</thead>
</table>

<table>
<thead>
<tr>
<th>Group B</th>
<th>Tube number</th>
<th>4</th>
<th>5</th>
<th>6</th>
<th>7</th>
<th>8</th>
</tr>
</thead>
<tbody>
<tr>
<td>Drops of 1 M KI</td>
<td>4</td>
<td>8</td>
<td>12</td>
<td>16</td>
<td>20</td>
<td></td>
</tr>
<tr>
<td>Drops of 1 M Pb(NO₃)₂</td>
<td>20</td>
<td>16</td>
<td>12</td>
<td>8</td>
<td>4</td>
<td></td>
</tr>
</tbody>
</table>

| Group B | Amount of Precipitate | |
|---------|----------------------|---|---|---|---|---|
5. Dispose of all waste in the proper waste container. **RINSE THE TEST TUBES WITH DISTILLED WATER OVER THE CHEMICAL WASTE CONTAINER AND THEN PLACE THEM IN THE SOLID WASTE CONTAINER.**

**REACTIONS:**
- Group A: Calcium and carbonate ions react.
- Group B: Lead(II) ions and iodide ions react.
- Group C: Iron(III) ions and hydroxide ions react.

**QUESTIONS:**
1. **Based on the results of your experiments,** write the formula for the precipitate that formed in each of the tubes.
   - A. 
   - B. 
   - C. 

2. What is the limiting reagent in each test tube? For each tube number, make a declaration by placing an x next to the number of drops in the boxes in the tables.

3. Ignoring, if necessary, the results of your experiments, write a balanced equation for the reaction you expected to occur in each group of test tubes. Write (s), for solid, after the formula of the precipitate you expected to form in each case.
   - A. 
   - B. 
   - C. 

4. Do the results for the balanced equations in Question 3 agree with the results of your experiments in each case? Yes ___ No ___ If you answered “No,” suggest reasons why your results don’t agree with what you expected.
26. Reactions of Some Metal Ions

OBJECTIVE: To observe precipitation and complexation reactions of some metal ions.

CONCEPTS: Visual detection of chemical reaction, Basic and amphoteric behavior of metal ions, Complex ion formation

INTRODUCTION: The simple, monatomic metal cations may all be classed as either amphoteric or basic depending on their reaction with OH\(^-\) (aq) ion. Basic metal ions either do not react with OH\(^-\), e.g. Na\(^+\), K\(^+\), or form insoluble hydroxides in excess OH\(^-\). Amphoteric metal ions form insoluble hydroxides in the presence of a near stoichiometric amount of OH\(^-\), and redissolve in excess OH\(^-\). The metal ions in the latter half of the d-transition series often form complexes with NH\(_3\), while most other metal ions react only with the OH\(^-\) formed by hydrolysis of NH\(_3\).

In this experiment, 1 M solutions of NaOH and ammonia will be added to 0.1 M solutions of Mg\(^{2+}\), Al\(^{3+}\) and Cu\(^{2+}\). These cations are respectively representative of basic, amphoteric and ammonia complexing cations. Observe the reaction of each cation with stoichiometric and excess amounts of OH\(^-\) and stoichiometric and excess amounts of NH\(_3\).

HAZARDS: Wear goggles. Acids and bases can irritate the skin and damage your eyes and clothing, so use caution. Rinse your skin or clothing if any of these chemicals are spilled on you.

PROCEDURE:
1. In a 24 well plate, place about 10 drops of each of the three metal ion solutions in four wells, twelve wells altogether. Note that these solutions are only one-tenth as concentrated as the reagent solutions added in the next steps.

2. Add 1 M NH\(_3\) solution dropwise, with stirring, to the first of the wells containing each metal ion. Stop the addition as soon as you see a precipitate. You have produced the insoluble hydroxides of each ion. To the second well containing each metal ion, add excess ammonia solution (up to 20 drops) and note any further reaction.

3. To the third well for each metal ion, add 1 M NaOH solution dropwise, stirring after each drop. Stop the addition as soon as you see a precipitate. You have produced the insoluble hydroxides of each ion. To the last set of three wells, add excess NaOH solution (up to 15 drops) to each and note any further reaction.

4. Complete the table, giving evidences of any chemical reactions such as precipitates or colors produced.

5. These reactions can be reversed by adding 1 M acetic acid dropwise. Try it. Record your observations.

6. Please clean up the work area and dispose of solutions in the proper waste containers. Clean up any spills to protect the next person from contact with the acids or bases.
<table>
<thead>
<tr>
<th></th>
<th>Ammonia Solution</th>
<th>NaOH Solution</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>first addition</td>
<td>excess</td>
</tr>
<tr>
<td><strong>Mg^{2+}</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Al^{3+}</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Cu^{2+}</strong></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**QUESTIONS:**

1. Which ions precipitate as hydroxides?

2. Which ions are amphoteric?

3. Which ions form complexes with NH$_3$?

4. What colored species were noted? (Indicate their colors.)

5. Indicate the colors of the following precipitates and complex ions:
   
   \[
   \text{Mg(OH)}_2 \quad \text{Al(OH)}_3 \\
   \text{Cu(OH)}_2 \quad \text{Al(OH)}_4^- \\
   \text{Cu(NH}_3\text{)}_4^{2+}
   \]

6. What happened when you added the acetic acid to wells with precipitate? To wells with NH$_3$ complexes?

7. Explain what you observed by writing equations for the reactions producing the precipitates and complex ions listed in question 5.

8. What did you like about this experiment and how could it be improved?
27. Temperature Effect on Solubility

OBJECTIVE: To demonstrate supersaturation and the effect of temperature on solubility.

CONCEPTS: Saturated solutions, Le Chatelier’s principle

INTRODUCTION: In this experiment, you will observe two water-soluble ionic compounds showing different solubility behavior as temperature is varied. Both experiments can run concurrently in the same sand bath.

You will observe that the solubility of sodium acetate increases rapidly as temperature increases. This phenomenon permits the preparation of a supersaturated solution. You will also observe the endothermic and exothermic nature of two processes: (1) sodium acetate dissolving in water and (2) the reverse of (1), sodium acetate crystallizing from a supersaturated solution. The effect of temperature on these reactions is a simple demonstration of LeChatelier’s principle.

Not all ionic compounds show the same temperature effects on solubility. The solubility of calcium acetate decreases as temperature increases.

HAZARDS: Calcium and sodium acetate will irritate eyes and mucous membranes. Use extreme care in heating test tubes. Do not overheat. ALWAYS LOOSEN THE CAP ON THE TEST TUBES BEFORE HEATING. WEAR YOUR GOGGLES!

PROCEDURE:
1. Obtain a 16 x 125 mm test tube containing 12 mL of saturated Ca(C₂H₃O₂)₂ solution. If the tube contains less than 12 mL, add water, shake, and allow the tube to come to chemical equilibrium. Note if solid is present and record the temperature of the tube.

2. LOOSEN THE CAP about 1/4 turn to prevent pressure build-up and place the test tube in a hot (70–80°C) sand bath. Be careful not to let the tube touch the metal sides or bottom. Push the tube far enough into the bath so the sand covers the liquid. Heat the tube for about five minutes. Remove the tube from the sand and record your observations of amount of solid and the temperature. Allow the tube to cool. Record your observations. Return it to the test tube rack.

3. Weigh out 10 g of Na₂C₂H₃O₂·3H₂O in a weighing tray. Transfer the crystals to a clean 16 x 125 mm test tube, and while holding the bottom of the tube in the palm of your hand, add 2 to 3 mL of distilled water with a graduated cylinder and stir. The water will not dissolve all of the crystals. Record the observations of temperature change you detect with your hand.

4. Place the test tube of Na₂C₂H₃O₂·3H₂O in the hot sand bath and stir periodically. When all of the crystals are dissolved, remove the test tube and let it cool in a test-tube rack (after it is sufficiently cool you may run cold tap water over it to rapidly bring it to room temperature). If crystals begin to form before it is cooled, heat to redissolve in the sand bath. The goal is to obtain a supersaturated solution. When the solution has cooled, drop a small sodium acetate crystal into the test tube while holding the tube in the palm of your hand. Record your observations.

5. When you are done, use the spray bottle to clean out the test tube of sodium acetate into the waste container provided, and place the cleaned test tube into the soap solution. Clean up the area.
**Questions:**

1. What happened to the temperature when the sodium acetate dissolved? Is dissolution of sodium acetate in water endothermic or exothermic?

2. What happened to the temperature when the sodium acetate crystallized from supersaturated solution? Is the crystallization of sodium acetate endothermic or exothermic?

3. Do your observations of sodium acetate solubility with increasing temperature agree with the predictions of LeChatelier's principle? Explain.

4. Why didn't the sodium acetate solution re-form crystals until you added a "seed" crystal?

5. How does the solubility of Ca(C₂H₃O₂)₂ change with temperature?

6. Is dissolution of calcium acetate endothermic or exothermic?

7. What could supersaturated sodium acetate solution be used for commercially?

8. What did you like about this experiment and how would you improve it?
28. Testing for Proteins with the Biuret Test

OBJECTIVE: To determine the relative concentration of protein in several foods.

CONCEPTS: Complex ion formation, Concentration is proportional to observed color intensity

INTRODUCTION: The biuret test for proteins consists of dissolving the protein in a basic solution (NaOH or other base) and adding copper sulfate solution. The Cu$^{2+}$ ions form a purple complex with the amide link between the amino acids of the protein. This purple complex indicates the presence of a protein, and the intensity of the color formed indicates the relative amount of protein in the sample. Eleven different food samples are provided. Test each for the presence of protein. Be aware that the native color of the sample can affect the results of the biuret test.

![Amino Acid and Amide structures](image)

HAZARDS: Sodium hydroxide (NaOH) is a strong base. Wear safety goggles when handling the chemicals and avoid contact with skin or clothing. If NaOH solution comes in contact with your skin, rinse the affected area thoroughly with water. **DO NOT EAT** any of the food samples; they may be contaminated or spoiled.

PROCEDURE:
1. Perform the tests on a 12-well, ceramic, spot plate. Use the table to record your results. Before testing, guess which of the food samples will contain protein and record it in the appropriate box. Let the samples sit for about 5 minutes after adding the copper sulfate solution to allow the color to fully develop.

2. Run a blank or zero control. Add 6 drops of water to a well, and while stirring, add 3 drops of 6 M NaOH solution and 6 drops of 0.05 M CuSO$_4$ solution. Observe the blue color.

3. Next run a positive control. Place 3 drops of egg white (high in protein) in a second well. Add 3 drops of water and gently stir with the toothpick. Add 3 drops of 6 M NaOH to the mixture and gently stir. Add 6 drops of 0.05 M CuSO$_4$ and note the color change. A positive test produces a blue-violet (purple) color.

4. Repeat this procedure for the other food samples, i.e. add 3 drops water, 3 drops 6 M NaOH, and 6 drops 0.05 M CuSO$_4$. Use an amount of sample equivalent in volume to the 3 drops of egg white. If the test color is obscured by the food, draw some of the solution into a strip of dry filter paper. Check the blank and egg white the same way.
5. When finished, clean up the area and properly dispose of the tested samples.

<table>
<thead>
<tr>
<th>Food Sample</th>
<th>Guess Protein (Y/N)</th>
<th>Tested Protein (Y/N)</th>
<th>Results (color change, intensity etc.)</th>
</tr>
</thead>
<tbody>
<tr>
<td>egg white</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>egg yolk</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>yeast</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>corn flakes</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>refried beans</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>sugar</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>applesauce</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>peas</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>peanut butter</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>white bread</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>beef</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**QUESTIONS:**

1. Why did you add just NaOH and CuSO₄ (with no food sample) to the first well?

2. What is the purpose of adding 3 drops of water to the sample?

3. If you had to choose several of the above foods for their high protein content, which ones would you choose and why?

4. Of the substances above that contain protein, which do you think has the least? The most? Explain how you decided.

5. What did you like about this experiment and how would you improve it?
29. Titration of a Strong Acid with a Strong Base

OBJECTIVE: To learn how the concentration of a strong acid can be determined by titration with a strong base.

CONCEPTS: Acids, Bases, Concentration, Titration

INTRODUCTION: Titration is a technique used to determine the concentrations of solutions. The acid concentration of a known volume of solution can be determined from the volume and base concentration of a second solution required to exactly react with the first, if the stoichiometry (an acid-to-base mole ratio) of the reaction is known. The acid (HCl) and base (NaOH) used in this experiment react in a one to one mole ratio.

\[ \text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{H}_2\text{O} + \text{NaCl(aq)} \]

HAZARDS: Wear your safety goggles. Make sure you have no skin contact with either the acid or the base. Both are strong skin irritants. If contact does occur, wash immediately with a large amount of water.

PROCEDURE:

1. Obtain a 24-well plate and rinse thoroughly with distilled water. To each of three wells transfer 25 drops of 0.50 M HCl. Add one or two drops of phenolphthalein indicator solution to each well. (If you are red-green colorblind, use thymol blue or meta cresol purple indicator instead of phenolphthalein.)

2. To one of the wells add NaOH drop-wise, while stirring with a toothpick, until a faint pink color remains for 30 seconds after stirring. Record how many drops it takes to reach this endpoint.

3. Repeat the procedure for the other two wells.

4. Compare your data. All three titrations should agree within one drop. If they do not, you may want to repeat the titration.

5. Dispose of waste in the proper waste container.

6. Clean up the lab area.
**QUESTIONS:**
1. What is the ratio of volume of acid to volume of base? What is the ratio of the concentrations of the acid and base? (Hint: The ratios are related but not equal.) What is the concentration of the NaOH solution?

2. What errors may result from using different dropper bottles for the acid and base?

3. Will it matter if the wells contain a few drops of pure water when you begin the experiment? Why or why not?

4. What is the purpose of phenolphthalein?

5. How would this experiment be different if sulfuric acid, H₂SO₄, was used instead of hydrochloric acid?

6. What substances have you recently used outside this lab that are either acids or bases?

7. What did you like about this experiment and how would you improve it?
30. Water, a Compound or an Element?

**OBJECTIVE:** To demonstrate how a compound may be decomposed into elements.

**CONCEPTS:** Classification of matter, Decomposition reactions

**INTRODUCTION:** Compounds are substances made up of simpler forms of matter called elements. They may be distinguished from elements by causing them to decompose into different parts. Decomposition can often be caused by heating a material or subjecting it to light or electricity. In this experiment a common substance, water, is tested by subjecting it to the electric current from a 6-volt source. Decomposition can be observed directly. Bubbles of gas are produced where the surface of each of the electrodes contacts the water. One of these gases forms explosive mixtures with air and is thus identified as hydrogen. The other gas is oxygen.

Pure water is a very poor conductor of electricity, so a sodium sulfate solution is used instead of pure water in the experiment. The sodium sulfate serves to speed up the reaction but is not decomposed itself.

**HAZARDS:** Clean up any spills before leaving and remember to remove the electrodes from solution when you have finished. Wash your hands before leaving.

**PROCEDURE:**
1. Make sure the container has enough 1 M Na₂SO₄ solution to cover the inverted funnels by 1 to 2 cm. Turn on the power supply to the electrodes under the funnels. Gas will begin to bubble off the electrodes and up through the inverted funnels.

2. Collect a sample of gas from the (–) electrode. First, fill a plastic bulb with solution by submerging it, open side up, in the solution and squeezing it 3 or 4 times, then invert it over the funnel at the (–) electrode to collect the gas. Hold the bulb over the funnel until it is 1/2 filled with gas. Remove the bulb from the funnel, but continue to hold it with the open end down. The liquid prevents the gas from escaping.

3. To test the collected gas, squeeze the bulb carefully to remove the solution and then release it to draw in air. Hold a burning match about 1 cm from the bulb’s mouth and then squeeze the gas out of the bulb onto the flame. Work quickly after expelling the water so that the gas will not escape. If nothing occurs, repeat the experiment, but reduce the time between expulsion of the water and the ignition with the match.

4. Repeat the test with gas from the (+) electrode. Try it with pure gas (no air) in the bulb.

5. Observe which electrode evolves gas more rapidly.
**Questions:**

1. How do your observations indicate that water is a compound and not an element?

2. How do your observations indicate that the gas collected at the (−) electrode was hydrogen gas and not oxygen gas?

3. How do your observations indicate that oxygen gas was produced at the (+) electrode?

4. From the relative rates of gas evolved, what is the chemical formula for water?

5. When iron combines with oxygen, rust is formed. Would you classify rust as an element or a compound? Why?

6. Name three elements and three compounds you have encountered within the past week outside of this lab.

7. What did you like about this experiment and how would you improve it?
31. Ammonia—Basicity and Complexation

**OBJECTIVE:** To observe reactions of ammonia and hydroxide ion in aqueous solutions.

**CONCEPTS:** Complex ions, Amphoterism, Acid-base chemistry

**INTRODUCTION:** Ammonia acts as a weak base in the Brønsted sense (i.e. \( \text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^{+} + \text{OH}^{-} \)) and as a complexing agent with some metal ions as in the reaction \( \text{Cu}^{2+} + 4\text{NH}_3 \rightleftharpoons \text{Cu} (\text{NH}_3)_4^{2+} \). In this latter reaction, ammonia is also acting as a base, this time in the Lewis sense, because it donates its non-bonding electron pair to \( \text{Cu}^{2+} \), which is acting as a Lewis acid by accepting the electron pair on ammonia.

Many metal ions precipitate in the presence of hydroxide ion because hydroxides are generally insoluble. These precipitates appear as flocculent solids which in most cases settle rapidly to the bottom of the container, but in some cases settle only slowly. When these slowly-settling solids are white they are sometimes difficult to see precisely because they don’t settle rapidly. In cases like this the precipitate manifests itself as a slight cloudiness in the solution.

At higher concentrations hydroxide ion acts as a complexing agent for amphoteric cations. Since \( \text{Al}^{3+} \) and \( \text{Zn}^{2+} \) are each amphoteric, they form the complex ions \( \text{Al(OH)}_4^{1-} \) and \( \text{Zn(OH)}_4^{2-} \) respectively in the presence of excess hydroxide ion.

This experiment demonstrates the action (or lack of action) of \( \text{OH}^{-} \) and \( \text{NH}_3 \) on \( \text{Al}^{3+}(\text{aq}) \) and \( \text{Zn}^{2+}(\text{aq}) \).

**HAZARDS:** Avoid contact with the solutions, especially the \( \text{NaOH} \) and \( \text{NH}_3 \) solutions. If any solution comes in contact with your skin, thoroughly rinse the affected area immediately with water. Wear safety goggles!

**PROCEDURE:**

1. Place 10 drops of 0.10 M \( \text{Zn(NO}_3)_2 \) in each of two 10x75 mm test tubes. Label these test tubes 1 and 2. Likewise, place 10 drops of 0.10 M \( \text{Al(NO}_3)_3 \) in each of two different 10x75 mm test tubes. Label these test tubes 3 and 4.

2. Add 6 M \( \text{NH}_3 \) dropwise with vigorous mixing after each drop to each of the four test tubes until a precipitate is visible in each. One drop may be sufficient. Avoid an excess of ammonia.

3. To test tubes 1 and 3, add 6 M \( \text{NaOH} \) dropwise with vigorous mixing after each drop until all the precipitate dissolves in each test tube.

4. Add 10 drops of 6 M \( \text{NH}_3 \) to test tube 2. Mix vigorously.

5. Add 10 drops of 6 M \( \text{NH}_3 \) to test tube 4. Mix vigorously.
**QUESTIONS:** (Remember that the default for balanced equations is net ionic equations. Also keep in mind that balanced equations should be written in terms of the predominant species present.)

1. Write balanced equations for the reactions that occurred in test tubes 1 and 3 in procedure 2.

2. Write balanced equations for the reactions that occurred in test tubes 1 and 3 in procedure 3.

3. How do the reactions in procedure 3 prove that \( \text{Al}^{3+} \) and \( \text{Zn}^{2+} \) are amphoteric?

4. Did the precipitate dissolve in procedure 4? Yes ___ No ___ If you answered “Yes,” write a balanced equation for the reaction that caused the precipitate to dissolve. If you answered “No,” explain why the precipitate didn’t dissolve.

5. Did the precipitate dissolve in procedure 5? Yes ___ No ___ If you answered “Yes,” write a balanced equation for the reaction that caused the precipitate to dissolve. If you answered “No,” explain why the precipitate didn’t dissolve.
32. Avogadro’s Number

OBJECTIVE: To experimentally determine Avogadro’s Number.

CONCEPTS: Avogadro’s number, Mole

INTRODUCTION: Masses of various elements in proportion to their atomic masses contain equal numbers of atoms. If the mass in grams of a sample of an element is numerically equal to its atomic mass, we have a gram atomic mass of that element. The number of atoms in a gram atomic mass of any element is the same and defines Avogadro’s number. By a similar reasoning process we can say that the number of molecules in one gram molecular mass (mole) is also equal to Avogadro’s number.

In this experiment a known mass of oleic acid is spread over water in a film one molecule thick and the area of this monolayer measured. From the mass of the oleic acid and its gram molecular (or molar) mass, the fraction of a mole of oleic acid in the film may be calculated. Then, knowing the density of oleic acid allows the number of molecules in the layer to be calculated, and therefore the number of molecules per mole (Avogadro’s number).

HAZARDS: Pentane is volatile and flammable. Lycopodium powder is combustible. Keep away from ignition sources.

PROCEDURE:

1. Fill an aluminum pie tin halfway with water. Dust the surface with enough lycopodium powder to completely cover the surface.

2. With a small (0.25 to 1 mL) syringe, transfer 0.20 mL of the oleic acid in pentane solution to the center of the prepared surface. Be sure to recork the bottle. Wait 3-5 minutes for the pentane to evaporate. The oleic acid repels the powder on the surface so the area covered by the oleic acid molecules is easily observed.

3. From the concentration (on the container) and volume of the oleic acid solution, calculate the grams of oleic acid added to the surface. Record this and all other answers in the table.

4. Oleic acid (C_{18}H_{34}O_2) has a molar mass of 282 g mol^{-1}. Calculate the fraction of a mole of oleic acid on the surface.

5. Place a clear plastic sheet with a 1 cm grid copied on it over the aluminum pie tin, and, with a water-soluble marker, trace the outline of the area covered by the oleic acid. Determine the area occupied by the oleic acid by counting the number of squares. (Each square on the plastic sheet represents 1 cm^{2}.)

6. The density of oleic acid is 0.89 g/cm^{3}. From the number of grams added to the surface, determine the volume, i.e. volume=mass/density of the oleic acid.
7. Knowing the area and volume of the oleic acid, calculate the height of the monolayer, i.e.
   height=volume/area. Note that the measured height is equal to the length of an oleic acid molecule.

8. Now calculate the area occupied by each molecule and determine the number of molecules on the
   surface. Assume the molecule is a cylinder with height equal 4 times the diameter of the base. The
   area of one molecule in cm² then equals (height)² (sin 60°)/16 or (height)² (√3)/32. (This formula is
   readily derived by observing a model of the oleic acid molecule and by assuming hexagonal packing
   on the surface, i.e. a circular object occupies a hexagonal area.) Find the number of molecules on
   the surface by dividing the area of the entire monolayer by the area of one molecule.

<table>
<thead>
<tr>
<th>grams of oleic acid</th>
<th>moles of oleic acid</th>
<th>area of oleic acid</th>
<th>volume of oleic acid</th>
<th>height of oleic acid film</th>
<th>area per molecule</th>
<th>number of molecules in film</th>
</tr>
</thead>
</table>

9. **NOW YOU HAVE THE POWER TO CALCULATE AVOGADRO'S NUMBER!!!!** (molecules/mole)

**QUESTIONS:**

1. What value did you obtain for Avogadro's number? Is your answer close to the expected value of
   6.02 x 10²³?

2. What are the possible sources of error?

3. How would you prove that the oleic acid formed a monolayer?

4. What is the conversion factor between the atomic mass unit (amu) and the gram? (Hint: one ¹²C
   atom has a mass of 12 amu, and there are 6.02 x 10²³ ¹²C atoms in 12 g of ¹²C.) How is this
   conversion factor related to Avogadro's number?

5. What did you like about this experiment and how would you improve it?
33. Bouncing Putty Polymer

OBJECTIVE: To prepare a simple polymer and observe its properties.

CONCEPTS: Macromolecules, Polymerization

INTRODUCTION: A polymer is a giant molecule made by linking many repeating units together. The physical properties of the polymer depend on the chemical structure of repeating units and the extent of crosslinking. The stiffness (viscosity) of the polymer depends on the amount of crosslinking because the crosslinks between polymer molecules greatly increase molecular size, e.g. one crosslink per molecule doubles the average molecular mass. The difficulty of disentangling polymer molecules increases exponentially with molecular size and branching. The polymer used in this experiment is similar to "Silly Putty®," a commercial polymer available in toy stores.

HAZARDS: Don't eat the polymer or get it in your eyes.

PROCEDURE:

1. Place enough Elmer's Glue-All® in a 100 mL beaker to just cover the bottom. Add an equal volume of water. Stir well. Add one or two drops of food coloring. Mix to make your own colors.

2. Add an amount of saturated borax solution equal to about 1/2 the volume of the glue and water mixture. Stir well. The borax acts as a crosslinking agent in this mixture by forming labile bonds between polymer chains.

3. Remove the sticky glob of polymer and work it into a ball with your hands. The stickiness disappears after a few minutes. (If the polymer didn't form into a glob, discard it and start over, using less borax solution.)

4. Try bouncing the polymer ball. Observations:

5. Pull the ball apart, then squeeze the pieces back together. Observations:

6. Lay the ball on the table and leave it undisturbed for several minutes. Observations:
7. Clean the beaker and stirring rod with a brush and water. Wash your hands.

8. If you want to take the polymer with you, wrap it in plastic wrap to prevent drying and to protect your clothing. If it does become dry, wet it with water and knead for several minutes.

QUESTIONS:

1. The polymer in Elmer's Glue-All® is made from vinyl acetate, \( \text{H}_3\text{CC} \text{O} \text{C} \text{CH}_2 \) \( \text{H} \). Write the equation for the polymerization reaction. (Hint: Think about polymerization of ethylene, \( \text{H}_2\text{C} \text{CH}_2 \).)

\[
\begin{align*}
\text{O} & \quad \text{H} \\
\text{H}_3\text{CC} & \text{O} \quad \text{C} \text{CH}_2
\end{align*}
\]

2. The polymer resists deformation when you attempt to deform it rapidly, but deforms easily under slow, gentle pressure. How can this be explained in terms of the molecular structure of the polymer?

3. In the absence of borax, \( \text{Na}_2\text{B}_4\text{O}_7 \cdot \text{A} \text{10H}_2\text{O} \), Elmer's Glue-All® slowly dries to a hard material. Explain how the borax functions to cause the glue to (a) set rapidly and (b) stay soft.

4. What did you like about this experiment and how would you improve it?
# 34. Chemical Clock

**Name**

**Student ID**

**Date**

**Course Number**

**Section Number**

**Instructor**

**Objective**: To determine the effect of reactant concentration on the rate of a reaction.

**Concepts**: Reaction kinetics, Rate laws

**Introduction**: Clock reactions proceed with no observable change until one reactant is entirely consumed, then an abrupt, visible change occurs. The clock reaction in this experiment involves the following sequence of reactions:

\[(a) \quad IO_3^{1-} + 3HSO_3^{1-} = I^- + 3HSO_4^{1-}\]

\[(b) \quad 5I^{1-} + IO_3^{1-} + 6H^+ = 3I_2(aq) + 3H_2O\]

I\(_2\)(aq) from reaction (b) will not appear in the solution until all the HSO\(_3\)^{1-} is oxidized. After the HSO\(_3\)^{1-} is gone, I\(_2\)(aq) from step (b) reacts with starch to form a dark blue complex. So the appearance of a dark blue color indicates that all HSO\(_3\)^{1-} has been consumed. In this experiment the time required for HSO\(_3\)^{1-} to be entirely consumed is measured for different initial concentrations of IO\(_3\)^{1-}. From these times the rate of the reaction at the different initial IO\(_3\)^{1-} concentrations can be calculated.

For a chemical reaction to occur the reactant molecules or ions must collide. The reaction rate thus depends on the concentrations of the reactants. In this experiment only the initial concentration of IO\(_3\)^{1-} will be varied. The simplified rate equation, rate = k[IO\(_3\)^{1-}]\(^n\), thus applies, where k is the rate constant, [IO\(_3\)^{1-}] is the molar concentration of IO\(_3\)^{1-}, and \(n\) is the order of the reaction with respect to IO\(_3\)^{1-}. If \(n=0\), rate = k[IO\(_3\)^{1-}]\(^0\) = k, so the rate will remain constant even though [IO\(_3\)^{1-}] is varied. If \(n=1\), rate = k[IO\(_3\)^{1-}], so the rate will be directly proportional to [IO\(_3\)^{1-}]. If \(n=2\), rate = k[IO\(_3\)^{1-}]\(^2\), so the rate will quadruple when [IO\(_3\)^{1-}] is doubled. Therefore by examining the rate data obtained in this experiment to see which relationship holds as [IO\(_3\)^{1-}] is varied, the order of reaction with respect to IO\(_3\)^{1-} can be determined.

**Hazards**: Substances used in this experiment can be irritating to your skin, so be sure to wash your hands after doing the experiment. Wear safety goggles!

**Procedure**: [Note: at least 2 students are required to carry out Procedure 2 properly.]

1. Place two 24-well plates together so two rows of 6 wells are adjacent. In each of the wells in the bottom row of the upper plate place 4 drops of 0.10 M NaHSO\(_3\) solution and 5 drops of starch solution. In the first well of the lower plate place 12 drops of 0.05 M KIO\(_3\) solution. In the next wells place decreasing amounts of iodate solution and increasing amounts of water as shown in the table.

<table>
<thead>
<tr>
<th>Plate 1</th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>4 d HSO(_3)^{1-}</td>
<td>4 d HSO(_3)^{1-}</td>
<td>4 d HSO(_3)^{1-}</td>
<td>4 d HSO(_3)^{1-}</td>
<td>4 d HSO(_3)^{1-}</td>
</tr>
<tr>
<td>5 d starch</td>
<td>5 d starch</td>
<td>5 d starch</td>
<td>5 d starch</td>
<td>5 d starch</td>
</tr>
<tr>
<td>12 d IO(_3)^{1-}</td>
<td>10 d IO(_3)^{1-}</td>
<td>8 d IO(_3)^{1-}</td>
<td>6 d IO(_3)^{1-}</td>
<td>4 d IO(_3)^{1-}</td>
</tr>
<tr>
<td>0 d water</td>
<td>2 d water</td>
<td>4 d water</td>
<td>6 d water</td>
<td>8 d water</td>
</tr>
</tbody>
</table>

| Plate 2 | | | | |
|---|---|---|---|
| | | | |
2. Now, noting the time to the nearest second when the solutions run together, tilt plate 1 up over plate 2 so the contents in plate 1 run into the corresponding wells in plate 2. Set plate 1 aside and quickly begin gently to swirl plate 2 to mix the contents. Record (in the table on the next page) the elapsed time when the color change occurs in each well. Repeat this experiment at least once.

<table>
<thead>
<tr>
<th>Well #</th>
<th>Drops IO$_3^{1-}$</th>
<th>Time to color change (seconds)</th>
<th>Rate = 1/Time</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td>Trial 1</td>
<td>Trial 2</td>
</tr>
<tr>
<td>1</td>
<td>12</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>10</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td>8</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td>6</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5</td>
<td>4</td>
<td></td>
<td></td>
</tr>
<tr>
<td>6</td>
<td>0</td>
<td>Control (shouldn't change)</td>
<td>XXXXXXXXXXXXX</td>
</tr>
</tbody>
</table>

3. Calculate the average time to color change for each of the wells and record it in the “Ave Time” column. Since the rate of a reaction is inversely proportional to the time, a relative rate can be determined by taking the inverse of the average time to color change in each well. Calculate the relative rate for each well and record it in the “Rate = 1/Time” column.

QUESTIONS:

1. Note that one of the products of reaction (a) is I$^{1-}$. This means that I$^{1-}$ is produced as soon as the reaction begins and is therefore present before all the HSO$_3^{1-}$ is consumed. Since I$^{1-}$ and IO$_3^{1-}$ are both present before all the HSO$_3^{1-}$ is consumed, what prevents I$_2$(aq) from forming and reacting with the starch to produce a dark blue color before all the HSO$_3^{1-}$ is consumed? [Hint: it has nothing to do with the availability of H$^{1+}$ for reaction (b). Adequate acid for reaction (b) is present from dissociation of the HSO$_4^{1-}$ produced in reaction (a).] You may find some of the following data helpful in formulating an answer.

\[
\begin{align*}
\text{HSO}_4^{1-} + 2H^{1+} + 2e^{1-} & = \text{HSO}_3^{1-} + H_2O & E^0 & = 0.172 \text{ V} \\
I_2(aq) + 2e^{1-} & = 2I^{1-} & E^0 & = 0.5355 \text{ V} \\
\text{IO}_3^{1-} + 6H^{1+} + 6e^{1-} & = I^{1-} + 3H_2O & E^0 & = 1.085 \text{ V}
\end{align*}
\]

2. The order of reaction with respect to IO$_3^{1-}$ is (check the correct order) 0 1 2. Show how you used your data to determine your answer.
35. Chemical Sunset

OBJECTIVE: To prepare a colloidal suspension and observe the Tyndall effect, i.e. light scattering by a colloid.

CONCEPTS: Colloids, Light scattering

INTRODUCTION: A light beam passing through a homogeneous solution is invisible unless the light is scattered by particles. Colloidal-sized (100–10,000 nm) particles scatter light in all directions. Shorter wavelengths such as violet (350 nm) are scattered more effectively than longer wavelengths such as red (750 nm). Colloidal precipitates are often produced by reactions in homogeneous solutions. For example, colloidal-sized particles are produced in the atmosphere by condensation of water (fog and ice crystals).

Adding acid to a solution of sodium thiosulfate initiates a homogeneous reaction

\[ 2 \text{H}^+(\text{aq}) + \text{S}_2\text{O}_3^{2-}(\text{aq}) \rightarrow \text{S(s)} + \text{H}_2\text{SO}_3(\text{aq}) \]

in which the solid sulfur, S(s), is produced gradually in colloidal-sized particles.

HAZARDS: Attend to chemical spills immediately. Rinse any spills from skin and clothing with generous amounts of water. Wear goggles. Some people may be sensitive to the SO₂(g) evident near the end of the reaction. Use adequate ventilation.

PROCEDURE:

1. Place about 100 mL of 0.1 M sodium thiosulfate (Na₂S₂O₃) solution in a clean 150 mL beaker. Place the beaker in a box to exclude as much room light as possible.

2. Shine a flashlight beam horizontally through the solution and onto a white paper taped inside the box. Observe the flashlight beam as it passes through air, through the solution, and as it strikes the white paper. Record what you see. (Ignore reflections from any large particles that may be present.)

3. With the flashlight on, add about 10 mL of 0.1 M HCl to the thiosulfate solution and stir for a few seconds. Observe the flashlight beam in the solution and record any changes you observe. As time passes, also observe and record any color changes in the beam exiting the solution.

4. After several minutes, when the production of colloid is complete, dispose of the solution in the container provided. Turn off the flashlight and wash the beaker.
**Questions:**

1. Is it easier to see where the beam of light passes through the solution before or just after the colloid begins to form? Why?

2. Explain why sunsets appear reddish.

3. Explain why the sky is blue.

4. What does the sky look like to astronauts beyond the earth's atmosphere? Why?

5. What did you like about this experiment and how would you improve it?
36. Chiral Molecules

OBJECTIVE: To observe the rotation of plane polarized light by solutions of chiral molecules in a simple polarimeter.

CONCEPTS: Optical isomers, Polarized light

INTRODUCTION: When light is passed through a "polarizer," only those waves in planes parallel to an invisible "grid" in the polarizer pass through. Out-of-plane waves are absorbed. Light passed by such a polarizer is called plane polarized light. If two polarizers are oriented 90° to each other, no light is transmitted.

Chiral molecules are said to be "optically active" because they have the ability to rotate the plane of polarization of polarized light. Models of chiral molecules show them to be nonsuperimposable on their mirror images (as are your hands). Such molecules exhibit chiral (or optical) isomerism, one isomer rotating plane polarized light in one direction, and the other rotating it in the other. A 1:1 mixture of the chiral isomers will not rotate the plane because their effects cancel.

HAZARDS: Don't drop the glass bottles.

PROCEDURE:

1. Place one gray polarizing filter on the light source. Place a second over the first so the orientation lines are parallel. What do you observe?

2. Rotate the top polarizer slowly until it is 90° relative to the first. What do you observe?

3. Lay a square bottle of water on its side between the two polarizers and repeat steps 1 and 2. Do you notice any difference? Are water molecules chiral?

4. Place a square bottle of 100% corn syrup (Karo White Syrup®) between the two polarizers and repeat steps 1 and 2. Describe what you observe. Are all colors rotated equally? Choose a color and record the approximate angle of rotation in the table. Note that it is much easier to determine the angle of rotation by referencing to the positions of the polarizer where no light is transmitted than by referencing to the angle of maximum transmission.

5. Measure and record the angle of rotation of the other 3 syrup solutions. Use the angle of minimum transmission of the color of the light you chose to measure the rotations.
### Questions:

1. Why is it necessary to use monochromatic light in an accurate polarimeter?

2. How can you tell if an organic molecule is chiral?

3. The structure of α-D-glucose, a sugar in corn syrup, is shown below. Is this molecule superimposable on its mirror image?

4. What effect on the angle of rotation would you expect if a thicker sample cell with the same concentration was used?

5. What did you like about this experiment and how would you improve it?
37. Chromatic Chromium

OBJECTIVE: To produce compounds and complex ions of chromium in different oxidation states and observe their colors.

CONCEPTS: Oxidation-reduction reactions, Colors of transition metal ions, Complex ions; Chemical change

INTRODUCTION: The colors of transition metal ions vary with oxidation state and coordinated ligands. Chromium can exist in a variety of oxidation states, most of which can be produced in aqueous solution. Several compounds and complexes of both Cr(III) and Cr(VI) are made in this experiment using only Cr(III)(aq) and an acid, a base, and hydrogen peroxide as reagents. In these experiments, hydrogen peroxide, H₂O₂, functions as an oxidizing agent in basic solutions and a reducing agent in acidic solutions. Under some conditions the peroxide ion, O₂²⁻, also functions as a ligand.

Some of the reactions of chromium you will observe in this experiment are:

<table>
<thead>
<tr>
<th>Oxid. State, Soln., Kinetics</th>
<th>Reaction</th>
<th>Well</th>
</tr>
</thead>
<tbody>
<tr>
<td>(+3, acid)</td>
<td>Cr(III)(aq)(the experiment begins with the violet-colored ion, Cr(H₂O)₆³⁺)</td>
<td>1</td>
</tr>
<tr>
<td>(+3, neutral, fast)</td>
<td>Cr(H₂O)₆³⁺ + 3OH⁻(aq) → Cr(OH)₃(s) + 6H₂O(l) (gray-green ppt)</td>
<td>2</td>
</tr>
<tr>
<td>(+3, basic, fast)</td>
<td>Cr(OH)₃(s) + OH⁻(aq) → Cr(OH)₄⁻(aq) (green)</td>
<td>3</td>
</tr>
<tr>
<td>(+6, basic, slow)</td>
<td>2Cr(OH)₄⁻(aq) + 3H₂O₂ + 2OH⁻ → 2CrO₄²⁻ + 8H₂O (yellow)</td>
<td>4</td>
</tr>
<tr>
<td></td>
<td>transient peroxo complexes of Cr(IV, V, and VI)</td>
<td>5</td>
</tr>
<tr>
<td>(+6, acid, fast)</td>
<td>2CrO₄²⁻ + 2H⁺ → Cr₂O₇²⁻ + H₂O (orange)</td>
<td>6</td>
</tr>
<tr>
<td>(+6, acid, transient)</td>
<td>2H⁺ + Cr₂O₇²⁻ + 4H₂O₂ → 2CrO(O₂)₂ + 5H₂O (blue)</td>
<td>7</td>
</tr>
<tr>
<td>(+6, acid, slow)</td>
<td>Cr₂O₇²⁻ + 3H₂O₂ + 8H⁺ → 2Cr(III) + 3O₂(g) + 7H₂O (probably green)</td>
<td>7</td>
</tr>
</tbody>
</table>

HAZARDS: Safety goggles are required. Strong acids and bases can damage your eyes, skin, and clothes. Avoid exposure to H₂O₂ (hydrogen peroxide) and the chromium compounds. Be sure to wash your hands at the conclusion of the experiment.

PROCEDURE:  

1. Put a white sheet of paper under a 24-well plate and label it as shown in the diagram.
2. Add 10 drops of 0.1 M Cr³⁺ to each of wells 1–7.
3. Add 10 drops of 0.05 M CrO₄²⁻ to well 9 and 10 drops of 0.01 M Cr₂O₇²⁻ to well 8.
4. Add 2 drops of 1.5 M NaOH to each of wells 2–7. Stir.
5. Add 6 drops of 1.5 M NaOH to each of wells 3–7. Stir.
6. Add 1 drop of 10\% \text{H}_2\text{O}_2 to each of wells 4-7 and stir for at least 3 minutes. Wait for the clear, lemon-yellow color of \text{CrO}_4^{2-} to fully develop. Compare to well 9.

7. Add 3 drops, stirring after addition of each drop, of 2 M HNO\textsubscript{3} to each of wells 5-7. Watch for the transient blue color of Cr(VI)-peroxo complexes and the red-brown color of peroxy complexes of Cr(IV) and Cr(V). The latter two may precipitate.

8. Add 5 more drops, stirring after addition of each drop, of 2 M HNO\textsubscript{3} to wells 6-7. Compare to well 8.

9. Add 1 drop of 10\% \text{H}_2\text{O}_2 to well 7. Stir. Wait for dark blue color to fade to a light green or possibly violet color.

   The solution in well 7 is chromium(III), the same oxidation state as in well 1, but the solution may contain

   \[
   \begin{array}{c}
   \text{(H}_2\text{O)}_4\text{Cr} \quad \text{Cr(H}_2\text{O)}_4 \\
   \text{H} \quad \text{H} \\
   \text{O} \quad \text{O} \\
   \text{H} \quad \text{H}
   \end{array}
   \]

   either the violet \text{Cr(H}_2\text{O)}_6^{3+} ion or the green \text{Cr(H}_2\text{O)}_4\text{Cr(H}_2\text{O)}_4\text{H}_2\text{O}\text{O}\text{O}\text{H} dimer. The exact reaction conditions (e.g. temperature, rate of mixing ) determine which is produced.

\textbf{QUESTIONS:}

1. What is the purpose of the \text{H}_2\text{O}_2 in forming the contents of well 4?

2. What is the purpose of the \text{H}_2\text{O}_2 in forming the transient species in well 7? The final species?

3. Why is Cr(III) in \text{Cr(H}_2\text{O)}_6^{3+} a different color than Cr(III) in

   \[
   \begin{array}{c}
   \text{(H}_2\text{O)}_4\text{Cr} \quad \text{Cr(H}_2\text{O)}_4 \\
   \text{H} \quad \text{H} \\
   \text{O} \quad \text{O} \\
   \text{H} \quad \text{H}
   \end{array}
   \]

4. Why is it possible for chromium to exhibit so many colors in various compounds?
OBJECTIVE: To observe the colors of the precipitates and complexes of some metal ions.

CONCEPTS: Precipitation, Complex ions, Color, Electron configurations

INTRODUCTION: Solutions of eight different metal ions (each as the nitrate salt) and solutions of eight different anions (each as the sodium salt) are mixed in every possible combination to investigate the colors of precipitates and complexes of several metal ions. Color results from absorption of light from a portion of the visible spectrum. Light in the visible range is absorbed by the electrons in the compounds and ions in this experiment. The electrons may change their energy level within the atomic or molecular orbitals in an ion or molecule or may be transferred from one atom, ion, or molecule to another (charge transfer).

HAZARDS: Wear your goggles. Use caution when using any chemical. Rinse your clothing or skin with water if any chemicals are spilled on you.

PROCEDURE:

1. Place a 96 well plate on a sheet of white paper. Label the rows as A through H, and the columns as 1 through 9. (You will only be using 9 of the 12 columns.)

2. In rows A through H place three drops of the indicated solution in the first nine wells:
   - Row A, 0.1 M Cr(NO₃)₃
   - Row B, 0.1 M Fe(NO₃)₂
   - Row C, 0.1 M Fe(NO₃)₃
   - Row D, 0.1 M Co(NO₃)₂
   - Row E, 0.1 M Ni(NO₃)₂
   - Row F, 0.1 M Cu(NO₃)₂
   - Row G, 0.1 M Zn(NO₃)₂
   - Row H, 0.1 M Pb(NO₃)₂

   Record the color of each metal ion solution in the "control" row in the table below.

3. In columns 1 through 8 place three drops of the indicated solution in each of the 8 wells. Note that all of the anion solutions are colorless except Na₂CrO₄ which is yellow.
   - Column 1, 0.2 M NaOH
   - Column 2, 0.2 M NaCl
   - Column 3, 0.2 M NaI
   - Column 4, 0.1 M Na₂SO₄
   - Column 5, 0.1 M Na₃PO₄
   - Column 6, 0.1 M Na₂C₂O₄
   - Column 7, 0.1 M Na₂CrO₄
   - Column 8, 0.1 M Na₂S

4. Shake or agitate the plate gently to mix the solutions.

5. Note whether a precipitate is formed or color change occurs and record this information in the table. If you do not detect a change, record "NR".

6. Dispose of all chemicals in the waste container and clean out the well plate with water and cotton swabs.
QUESTIONS:
1. Summarize the solubility information in the table in a few phrases.

2. What are the two visual indications of a reaction?

3. Refer to a periodic table and note which metal ions have incompletely filled d orbitals (between 0 and 10 electrons). Do these ions tend to give colored or colorless compounds? Why is a solution of Ni(NO$_3$)$_2$ colored while a solution of Zn(NO$_3$)$_2$ is not?

4. Give one reason why the color of a transition metal ion complex usually changes when the ligands change, e.g. FeCl$_4^-$ is yellow while Fe(H$_2$O)$_6^{3+}$ is purple and Fe(OH)(H$_2$O)$_5^{2+}$ is yellow-brown.

5. Light absorption occurs by a charge transfer from S$^{2-}$ to M$^{2+}$ in PbS and ZnS, but PbS is colored while ZnS is white. Explain. (Hint: Which is the better oxidizing agent, Pb$^{2+}$ or Zn$^{2+}$?)

6. Did you observe any redox reactions? (Hint: I$^-$ and S$^{2-}$ are the best reducing agents in the table. Also, FeI$_3$, CuI$_2$, Fe$_2$S$_3$ and CuS do not exist.) If so, write the equations.

7. What did you like about this experiment and how would you improve it?
OBJECTIVE: To demonstrate the release or removal of ions from solution by chemical reactions.

CONCEPTS: Acid-base chemistry; Ionic reactions, Conductance by ions in solution, Dissociation of electrolytes

INTRODUCTION: In this experiment, the conductivity of solutions is determined with a buzzer connected to a battery through two copper electrodes. The electrodes are separated by a small distance. When immersed in a conducting solution the circuit is completed and the buzzer sounds. If the solution has a high conductance (high concentration of ions), the buzzer sounds loudly. If the solution has a low conductance (low concentration of ions), the buzzer sounds softly. It may barely vibrate. Ionic concentrations in solution can be increased by reaction of weak electrolytes to form strong electrolytes or can be decreased by combination of ions to form water (in neutralization reactions), weak electrolytes, or precipitates.

HAZARDS: If any of the chemicals spill on skin or clothes, wash with water. Wear goggles.

PROCEDURE:
1. Place 10 mL each of the four 0.1 M solutions in the table below in four 25-mL beakers. Test the solutions for conductivity and note in the table whether the solute is a strong, weak, or non-electrolyte. To prevent cross-contamination of solutions, be sure to rinse the electrodes with distilled water before inserting into a new solution. Mix the solutions together as indicated below and test the mixtures for conductivity.

<table>
<thead>
<tr>
<th>Solution</th>
<th>Strong, Weak, or Non</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.10 M HCl</td>
<td></td>
</tr>
<tr>
<td>0.10 M NaOH</td>
<td></td>
</tr>
<tr>
<td>mixture: HCl + NaOH</td>
<td></td>
</tr>
<tr>
<td>0.10 M CH₃COOH</td>
<td></td>
</tr>
<tr>
<td>0.10 M NH₃</td>
<td></td>
</tr>
<tr>
<td>mixture: CH₃COOH + NH₃</td>
<td></td>
</tr>
</tbody>
</table>

2. Place 10 mL each of 0.10 M H₂SO₄ and 0.10 M Ba(OH)₂ in two 25-mL beakers. Test each solution for conductivity. Add one drop of bromothymol blue indicator to the H₂SO₄ solution. While the electrodes are immersed in the H₂SO₄ solution, slowly add the Ba(OH)₂ solution with stirring until the yellow solution just changes to green or blue, indicating the equivalence point. Note what happens to the sound level of the buzzer during the titration.

3. Wash the electrodes and beakers thoroughly with distilled water. The wash water goes into a
waste container.

**QUESTIONS:**

1. Write total and net ionic equations for the reactions occurring in each of the two pairs of solutions mixed in part 1.

2. In the reaction of HCl and NaOH, Na\(^+\) and Cl\(^-\) are spectator ions. Why was there little or no measurable change in conductance when the HCl and NaOH solutions were mixed?

3. Why is there an easily measurable increase in conductivity on mixing solutions of ammonia and acetic acid?

4. Write the total and net ionic equations for the titration of sulfuric acid with barium hydroxide. Explain why the buzzer sound level changed as it did during the titration.

5. What did you like about this experiment and how would you improve it?
40. Copper Chameleon

Name ___________________________ Course Number ___________________________
Student ID _________________________ Section Number ___________________________
Date ______________________________ Instructor ________________________________

OBJECTIVE: To observe the color changes occurring when the water molecules on Cu(H_2O)_4^{2+} are replaced by Cl^-, OH^-, and NH_3.

CONCEPTS: Complex ions, Colors of transition metal ions

INTRODUCTION: Transition metal ions form coordination complexes with electron pair donors (Lewis bases). Coordination of ligands to the metal ion changes the energies of partially filled d-orbitals on the metal ion, changing the absorption spectrum and color of the ion. Most copper(II) ions are colored. The color depends upon the identity of the ligands. This experiment shows the effects of replacing the ligand H_2O with Cl^-, NH_3 or OH^- . Copper(II) has 9 electrons and 1 vacancy in the 3d orbitals. The d orbital energies are split by the electric field of the ligands. The absorption of light excites a d electron into the vacancy. The colors arise from this light absorption. Each ligand gives a characteristic splitting of d orbital energies and hence its own color. Color may also result from light absorption by charge transfer from ligand orbitals to metal orbitals.

HAZARDS: Wear goggles. Immediately wash any chemicals off skin or clothing.

PROCEDURE:

1. Place 10 drops of 0.1 M copper sulfate solution in each of 4 wells of a 24-well plate. Place the plate on a sheet of white paper to aid in seeing the colors. Well #1 will serve as a control.

2. To well #2, add 6 M HCl dropwise until a color change is noted. Note that a large excess of HCl is required to produce the color change.

3. To well #3, add 1 M NaOH solution dropwise with stirring until a precipitate appears. Two drops may be enough.

4. To well #4, add 1 M NH_3 solution until the reaction above is noted, then add ammonia solution until the precipitate dissolves.

5. Record the color observed in each well and write the equation for the reaction.

<table>
<thead>
<tr>
<th>Well #1</th>
<th>clear, light blue</th>
</tr>
</thead>
<tbody>
<tr>
<td>Well #2</td>
<td>Cu(H_2O)_4^{2+} + 4Cl^- =</td>
</tr>
<tr>
<td>Well #3</td>
<td>Cu(H_2O)_4^{2+} + 2OH^- =</td>
</tr>
<tr>
<td>Well #4</td>
<td>Cu(H_2O)_4^{2+} + 4NH_3 =</td>
</tr>
</tbody>
</table>
**QUESTIONS:**

1. In part 4 of the procedure, addition of ammonia first produces Cu(OH)$_2$(s), the same precipitate as produced in part 3 by addition of NaOH. Write an equation showing how ammonia acts as a Lowry-Brønsted base.

2. In the second reaction observed in part 4, does ammonia act as a Lowry-Brønsted or Lewis base? Write the equation for dissolution of the precipitate by ammonia.

3. Explain why copper(II) complexes are usually colored.

4. Anhydrous CuSO$_4$ is nearly colorless. Could it be used to test for H$_2$O(g) or NH$_3$(g)? Explain.

5. One way to predict the weather is to observe the color of a strip of paper impregnated with K$_3$CoCl$_4$. The paper is normally blue, but turns pink shortly before a rain. Explain with reactions. (Hint: The CoCl$_4^{2-}$ ion is blue, but Co(H$_2$O)$_6^{2+}$ is pink.)

5. What did you like about this experiment and how would you improve it?
41. Corn Syrup Cell

Name ____________________________ 
Identification Number ____________________ 
Date ____________________________ 
Course Number ____________________ 
Section Number ____________________ 
Instructor ____________________________

**OBJECTIVE:** To observe the effects of osmosis.

**CONCEPTS:** Osmosis, Osmotic pressure, Hypotonic, Hypertonic, Isotonic, Semipermeable membranes

**INTRODUCTION:** A semi-permeable membrane can separate solvent particles from a solution. In a process called osmosis, solvent molecules can flow through the membrane, while solute particles do not. Solution volume increases on the side of the membrane with the lowest solvent concentration. The driving force for osmosis (i.e. the osmotic pressure, \( \pi \)) is usually quantified by measuring the pressure across a membrane as shown in the figure.

The formula for osmotic pressure is \( \pi = iMRT \), where \( T \) is the absolute temperature, \( R \) is the gas constant, \( M \) is the solute molarity, and \( i \) is the number of particles produced in solution per formula unit of solute. Osmotic pressure is a colligative property depending only on the total concentration of solute particles and not on their nature. For example, \( i=1 \) for sugar, which does not dissociate, and \( i=2 \) for \( \text{NaCl} \), which is completely dissociated into \( \text{Na}^+ \) and \( \text{Cl}^- \) ions. The product \( iM \) is the osmolality, which expresses the total concentration of particles in a solution. In this experiment two different cells (not the biological kind) are made from dialysis tubing (a semipermeable membrane) and used to observe the effects of osmosis.

In living tissues, osmosis occurs through semi-permeable cell membranes. When the osmolarity of the solution surrounding a cell equals the osmolarity of the solution inside the cell, the surrounding solution is said to be isotonic. If the solution surrounding the cell has a lower osmolarity it is hypotonic. Cells in hypotonic solutions take up solvent and thus experience an increased internal pressure which may cause rupture or "lysis". Cells immersed in hypertonic solutions experience decreased internal pressure, which may cause shrinkage and crenation. Intravenous injections are carefully adjusted with salt solutions to ensure that they are isotonic with blood cells. Osmotic pressure is a substantial effect, as shown by the fact that osmosis accounts for the movement of water from the roots to the tops of trees.

**HAZARDS:** Do not taste the corn syrup. Always wear your goggles.

**PROCEDURE:** The TA’s will setup the first two dialysis tube samples for each day. So, you may not want to come to the lab in the first hour it is open each day.

**Step 1.** Locate the beaker of water with the dialysis tube full of water. This is labeled, “control.”

**Step 2.** Find a second beaker with 2–4 dialysis tubes (each contains corn syrup) which will be labeled by the time they were placed in water. New tubes should be prepared about every half hour. Write the approximate time each tube has been in the water in the table below. Draw a picture of the relative sizes in the boxes in the following table. You must record observations for at least two corn syrup filled tubes. If there are not any tubes in the water, see step 3 below. There may be fewer tubes in the beaker if you are one of the first students of the day. If there are fewer than four tubes in the beaker of water or if a tube is older than 2 hours, please follow the
instructions below (step 3 a–f as necessary).

**Time:**

<table>
<thead>
<tr>
<th>Control (water)</th>
<th>(in min.)</th>
<th>( in min.)</th>
<th>( in min.)</th>
<th>( in min.)</th>
</tr>
</thead>
</table>

Observation:

Step 3 – Keep the dialysis tubing in water as much as possible.

a. If replacing a tube, remove it from the water beaker, and open the clip which has the time label attached to it **over the sink**. Remove the time label and rinse the tube several times with water. Keep in mind that the content of the tube is syrup; with other experiments, solutions are never disposed of down the drain.

Follow the steps b–e to create/replace a new corn syrup cell.

b. If you are the first person of the day, locate the dialysis tubing in a beaker of water. A clamp the tubing about 4 cm from one end with the provided clamp.

c. Using a pipette or dispenser bottle provided, fill the tube about half full with corn syrup.

d. Squeeze out all the air and clamp the open end of the tubing, along with the provide time label.

e. Rinse/wipe off any corn syrup that may be on the outside of the tubing.

**Questions:**

1. What has happened to the tubes filled with corn syrup? What has happened to the water-filled tube? Explain.

2. Was the water outside the tubes (circle one) hypertonic, hypotonic, or isotonic to the corn syrup?

3. What would happen to blood cells if pure water was injected intravenously? If 100% corn syrup was injected?

4. If the corn syrup is 1.0 M, and the room temperature is 25°C, what is the theoretical osmotic pressure across the membrane? (R=0.08206 L atm mol⁻¹ K⁻¹.)

5. What concentration of sugar from photosynthesis is required to get water to the top of a 70 m (200 ft) tall tree at 25°C? (70 m of water = 6.8 atm of pressure.)
# 42. Daniell’s Battery

<table>
<thead>
<tr>
<th>Name</th>
<th>Course Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Student ID</td>
<td>Section Number</td>
</tr>
<tr>
<td>Date</td>
<td>Instructor</td>
</tr>
</tbody>
</table>

**OBJECTIVE:** To construct a voltaic pile capable of sounding an electric buzzer.

**CONCEPTS:** Electrochemistry, Galvanic cell, Half reactions, Battery

**INTRODUCTION:** The battery built in this experiment is similar to the ones first invented by Alessandro Volta in 1800 and further developed by J.F. Daniell in 1836. The standard potentials for the copper and zinc half-cell reactions used in this battery are:

\[
\begin{align*}
\text{Zn}^{2+}(aq) + 2e^- &\rightarrow \text{Zn}(s) \quad E^* = -0.76 \text{ V} \\
\text{Cu}^{2+}(aq) + 2e^- &\rightarrow \text{Cu}(s) \quad E^* = 0.34 \text{ V}
\end{align*}
\]

at 25°C when [Zn²⁺] = 1.0 M and [Cu²⁺] = 1.0 M. Zn is easily oxidized by Cu²⁺ ion. A copper-zinc cell has a 1.10 V potential when the Zn²⁺ and Cu²⁺ ions are the same concentration. In the battery built in this experiment, a filter paper soaked with 0.1 M Cu²⁺ and 0.1 M Zn²⁺ serves as the electrolyte reservoir in contact with the electrode surfaces. A stack of copper-zinc galvanic cells connected in series (a battery) will produce a multiple of the voltage of a single cell.

**HAZARDS:** Wash your hands after the experiment.

**PROCEDURE:**

1. Obtain one 2-cm x 1-cm and two 1-cm x 1-cm pieces of copper and three 1-cm x 1-cm pieces of zinc. Obtain three pieces of filter paper slightly larger than the squares of zinc. Clean the metal pieces with sandpaper.

2. Place the 2-cm x 1-cm piece of copper metal in a petri dish. Place a drop of 0.1 M CuSO₄ on the copper. Add a piece of filter paper. Place a drop of 0.1 M ZnSO₄ on the filter paper. The filter paper should now be thoroughly moist. Add a square of Zn. It is important that the copper and zinc be separated by the filter paper and not touch. Place a 1-cm x 1-cm square of copper on the zinc (these must touch), then add in succession 0.1 M CuSO₄, filter paper, 0.1 M ZnSO₄, and another square of zinc. Add a third layer of Cu, CuSO₄ solution, filter paper, ZnSO₄ solution, and Zn. You should have a voltaic pile (battery) as illustrated.
3. Touch the buzzer wires to the top zinc sheet and the exposed part of the larger piece of copper on the bottom. The buzzer should sound or vibrate.

4. Dispose of the filter paper in the waste container and clean the metal pieces with sandpaper.

**QUESTIONS:**

1. Write the reaction for the Daniell cell. Calculate $E$ for a single cell. Calculate the voltage of the battery constructed in this experiment.

2. Do the electrons enter the buzzer from the Zn or the Cu terminal of the battery? How do you know?

3. In which direction do the cations move within the pile when current is drawn from the pile? Upward or downward? Explain.

4. If the electrolyte-soaked filter paper were inserted between each zinc and copper piece instead of every other one, would you expect the buzzer to sound? Why or why not? What would be the voltage of the battery?

5. A black solid is deposited on both the Zn and Cu. What is the black solid and why does it deposit on both the Zn and Cu? Explain. (Note: ZnO is white.)

6. What did you like about this experiment and how would you improve it?
**43. How Acids Differ**

**OBJECTIVE:** To demonstrate how acid strength and molar mass are determined.

**CONCEPTS:** Acid-base chemistry, pH, Titration

**INTRODUCTION:** Acids are characterized by their strength and by their molar mass. Strong acids are completely ionized in aqueous solution. Weak acids are only partially ionized. Strong and weak acids can be distinguished by measuring the pH of their solutions. Solutions of strong acids have larger [H⁺] and therefore lower pH. The molar mass of an acid can be determined by titration of a known weight/volume solution of the acid with a strong base solution of known molarity (moles/liter). The purpose of the titration is to determine the number of moles of the acid in a given volume of the acid solution, i.e. the molarity of the acid solution. The ratio of grams per liter to the molarity (moles per liter) equals the molar mass (grams per mole). In this experiment, unknown acids are identified by determining the molar mass and whether they are strong or weak. The available unknown acids are (strong) HI, HBr, HCl, HNO₃, and (weak) B(OH)₃, H₃PO₄, formic acid (HCOOH), acetic acid (CH₃COOH), and benzoic acid (C₆H₅COOH). To identify your unknowns, compare the molar masses and strengths you determine to the molar masses and strengths of the acids on the list.

**HAZARDS:** Wear goggles. Avoid spills. Wash thoroughly and immediately if contact with eyes, skin, or clothes occurs.

**PROCEDURE:**

1. There are four 5% wt/vol (50 g/L) solutions of acids labeled 1–4. Two of the acids are strong and two are weak. Place 10 drops of each unknown acid solution in each of four wells in a 24-well plate.

2. Add 1 drop of thymol blue indicator solution to each acid solution. Record the color in the table.

3. Thymol blue changes from reddish-purple to yellow between pH 1 and 3 and from yellow to blue between pH 8 and 10.

4. Titrate each acid dropwise with 1.0 M NaOH until the blue color remains after 5 sec of stirring. Be sure to count the drops of NaOH required.

5. Record the number of drops of base needed to titrate each acid.

6. From your results, determine which acids are strong and which are weak.

7. Rinse the contents of the plate into the waste container and wipe up any spills on the bench.
Acid | Color at start | Drops of base | Molarity of acid* | Molar mass | Weak or strong | Formula
--- | --- | --- | --- | --- | --- | ---
1 | | | | | | 
2 | | | | | | 
3 | | | | | | 
4 | | | | | | *At the endpoint (molarity acid) x (drops acid) = (molarity base) x (drops base).

**Questions:**

1. Bromothymol blue indicator has only one change, from yellow to blue near pH 7. Why couldn’t this indicator be used to distinguish strong and weak acids?

2. Which has the higher molarity, a 5% solution of NaOH or a 5% solution of KCl?

3. What is the relationship between the %wt/vol and the molar mass?

4. Calculate the molar mass of each acid and record it in the table.

5. What did you like about this experiment and how would you improve it?
OBJECTIVE: To demonstrate the surface tension of water and the effects of surfactants.

CONCEPTS: Intermolecular forces, Dipole-dipole forces, Hydrogen bonding

INTRODUCTION: In the liquid phase, water molecules are bound together by hydrogen bonds, a form of dipole-dipole attraction. These intermolecular bonds hold the surface of water together. This surface bonding or tension can be demonstrated by floating objects more dense than water on the unbroken surface. Molecules that break the continuity of the surface bonding, i.e. surfactants or surface active agents, cause such objects to sink.

HAZARDS: Don’t get pepper or soap in your eyes.

PROCEDURE:

1. Half fill 4 clean petri dishes with distilled water. Observe and record what happens in each experiment.

2. Sprinkle some pepper on the water in one petri dish.

3. Sprinkle some talcum powder on the water in another petri dish.

4. Float a thin, very limp rubber band on the water in the third petri dish. Rub the rubber band along the side of your nose to oil it so it will float.

5. Float a small, fine-mesh, oiled wire screen or needle on the water in the fourth petri dish.

6. Add a drop of liquid detergent (a surfactant) solution to each petri dish. Place the drops inside the rubber band and on the screen.

7. Place a penny or other coin on a paper towel. Add several drops of water to the top of the coin. Note how the water piles up on the coin. Now add a drop of detergent solution.

8. Place the dishes and screen in a pan of water and clean the lab area.

QUESTIONS:

1. Why do the pepper and talcum powder behave differently? What can you say about their densities?
2. Explain what happened to the rubber band.

3. Why does the wire screen or needle float on clean water and sink on addition of detergent?

4. What holds the water on top of the coin?

5. Explain why water skippers are able to maneuver on top of water as they do.

6. Why are even small amounts of biodegradable detergents a serious water pollutant?

7. What did you like about this experiment and how would you improve it?
OBJECTIVE: To observe the solubilities of ionic solids in water.

CONCEPTS: Solubility rules, Solubility

INTRODUCTION: This experiment illustrates the solubility rules for some ionic salts. Eight different anions, all present as the sodium salts, are combined with 9 different cations, present as the chloride or nitrate salts. Each of the 72 combinations either does not react or produces an insoluble solid.

HAZARDS: Always wear your goggles. If you get any chemicals on you, wash immediately.

PROCEDURE:

1. Use rows A through H and columns 1 through 9 on a 96-well plate.

2. Put three drops of each of the following solutions into each of the first nine wells of the corresponding row. (See table on reverse.)
   - Row A: 0.2 M NaNO₃
   - Row B: 0.2 M NaF
   - Row C: 0.2 M NaCl
   - Row D: 0.2 M NaI
   - Row E: 0.2 M NaOH
   - Row F: 0.1 M Na₂SO₄
   - Row G: 0.1 M Na₂CO₃
   - Row H: 0.1 M Na₂S

3. Add three drops of each of the following solutions to each of the wells of the corresponding column:
   - Column 1: 0.2 M NH₄NO₃
   - Column 2: 0.1 M Mg(NO₃)₂
   - Column 3: 0.1 M Ca(NO₃)₂
   - Column 4: 0.1 M SrCl₂
   - Column 5: 0.1 M Mn(NO₃)₂
   - Column 6: 0.1 M Zn(NO₃)₂
   - Column 7: 0.07 M Al(NO₃)₃
   - Column 8: 0.2 M AgNO₃
   - Column 9: 0.1 M Pb(NO₃)₂

   Use one toothpick to stir all the wells.

4. Write formulas and colors (including white) of precipitates in the chart below. Write "soluble" where no precipitate forms.

5. Clean out the well plates by rinsing the chemicals into the waste containers and using a cotton swab to clean out all precipitates.
A | NH₄NO₃ | Mg(NO₃)₂ | Ca(NO₃)₂ | SrCl₂ | Mn(NO₃)₂ | Zn(NO₃)₂ | Al(NO₃)₃ | AgNO₃ | Pb(NO₃)₂
---|---------|----------|----------|-------|----------|---------|---------|-------|---------
B | NaNO₃  |          |          |       |          |         |         |       |         |
C | NaF     |          |          |       |          |         |         |       |         |
D | NaCl    |          |          |       |          |         |         |       |         |
E | NaI     |          |          |       |          |         |         |       |         |
F | NaOH    |          |          |       |          |         |         |       |         |
G | Na₂SO₄  |          |          |       |          |         |         |       |         |
H | Na₂CO₃  |          |          |       |          |         | *       |       |         |
I | Na₂S    |          |          |       |          |         | *       |       |         |

* Note that Al₂(CO₃)₃ does not exist and Al₂S₃ is unstable in water. Al(OH)₃ forms in basic aqueous solutions.

**QUESTIONS:**

1. Summarize your observations in a few short phrases.

2. Why were all sodium salts used for the test solutions of anions (rows A–H)?

3. Why were nitrate or chloride salts used for the test solutions of cations?

4. Write balanced, net-ionic equations for the precipitation reactions observed with iodide ion.

5. Write the equation for reaction of Al₂S₃ with H₂O.

6. What did you like about this experiment and how would you improve it?
46. Is It an Electrolyte, and If So, What Kind?

**Name** _____________________________________________  
**Student ID** ___________________________________________  
**Date** ____________________________  
**Course Number** ____________________________  
**Instructor** ____________________________  

**Section Number** ____________________________  

**OBJECTIVE:** To observe the different conductivities of strong and weak electrolytes in aqueous solutions.

**CONCEPTS:** Conductance by ions in solution, Strong and weak acids, Dissociation of electrolytes

**INTRODUCTION:** Some solutes dissociate into ions. Because there are ions present and the ions are mobile, a solution of an electrolyte conducts electricity. Strong electrolytes are completely dissociated into ions in solution and therefore these solutions conduct electricity well. Solutions of weak electrolytes contain mostly undisassociated molecules and only a few ions, and thus conduct electricity poorly. For example, ammonia reacts with water to only a slight extent: $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$. Non-electrolytes remain undisassociated in solution and therefore don’t conduct electricity. Pure water is so slightly ionized it is essentially a nonconductor.

In this experiment several different solutes are tested to determine whether they are strong, weak, or non-electrolytes. A buzzer connected to a 9 volt battery is used to test the solutions. When the circuit is complete the buzzer sounds. If the solution is a weak electrolyte, less current flows in the circuit and the buzzer sounds more softly than for a strong electrolyte. With very weak electrolytes the buzzer may only vibrate with no audible sound. Non-electrolytes give no response of the buzzer.

**HAZARDS:** Wear your goggles. Rinse off any chemicals spilled.

**PROCEDURE:**

1. Touch the two copper electrodes together to complete the circuit. The buzzer will sound loudly. This is your reference for maximum conductivity. Rinse the electrodes with distilled water from a wash bottle. Place 40 mL of distilled water in a clean 50 mL beaker and immerse the electrodes. Any sound or vibration of the buzzer indicates contamination, and the electrodes and beaker need to be washed again.

2. Place about 5 mL each of 1 M HCl, 1 M acetic acid (CH$_3$COOH), 1 M sugar (C$_{12}$H$_{22}$O$_{11}$), 1 M NaOH, 1 M NH$_3$, 1 M NaCl, 1 M urea (NH$_2$CONH$_2$), and tap water in each of eight wells of a 24-well plate. Keep a record of which solution is in which well.

3. Place the electrodes so they are barely touching the acetic acid solution. Do you hear anything? Touch the buzzer. Do you feel anything? Slowly move the electrodes further into the solution and note how the buzzer responds. Avoid cross contamination of the solutions by rinsing the electrodes with distilled H$_2$O each time before testing a new solution.

4. Repeat #3 with the NaOH solution and note the buzzer response.

5. Place the electrodes into each solution to determine whether the solute is a strong, weak, or non-electrolyte. Complete the table. Be sure to rinse the electrodes with distilled water before testing each solution.
<table>
<thead>
<tr>
<th>Solution</th>
<th>Strong, Weak, or Non-electrolyte?</th>
</tr>
</thead>
<tbody>
<tr>
<td>1 M CH₃COOH</td>
<td></td>
</tr>
<tr>
<td>1 M NaOH</td>
<td></td>
</tr>
<tr>
<td>1 M HCl</td>
<td></td>
</tr>
<tr>
<td>1 M NaCl</td>
<td></td>
</tr>
<tr>
<td>1 M C₁₂H₂₂O₁₁</td>
<td></td>
</tr>
<tr>
<td>1 M NH₃</td>
<td></td>
</tr>
<tr>
<td>1 M NH₃CONH₂</td>
<td></td>
</tr>
<tr>
<td>Tap water</td>
<td></td>
</tr>
</tbody>
</table>

**Questions:**

1. Why does the buzzer sound louder as the immersion depth of the electrodes increases in the acetic acid solution? Would the distance between the electrodes also affect the buzzer sound? Why?

2. Why does tap water conduct electricity better than distilled water?

3. Why does the buzzer sound louder with a strong electrolyte than with a weak electrolyte of the same concentration?

4. Why does the HCl solution conduct better than the acetic acid solution? Dissociation equations might be helpful.

5. List ammonia, urea, and NaOH in order of increasing basicity in aqueous solution.

6. What did you like about this experiment and how would you improve it?
47. Molecular Marathon

Name ___________________________ Course Number ___________________________
Student ID ___________________________ Section Number ___________________________
Date ___________________________ Instructor ___________________________

OBJECTIVE: To demonstrate Graham's law of diffusion of gases.

CONCEPTS: Kinetic molecular theory of gases

INTRODUCTION: Graham's law is an empirical description of the observation that the rate of diffusion or effusion of a gas is inversely proportional to the molar mass of the gas. Thus, at constant temperature

\[ \frac{R_A}{R_B} = \left(\frac{M_B}{M_A}\right)^{0.5} \]

where \( R \) is the transport rate of the gas and \( M \) is the molar mass. Graham's law is a basis for the tenet of the kinetic molecular theory that states that the average kinetic energy of the molecules of any gas is the same at the same temperature, i.e., \( KE = 1.5kT = 0.5m\bar{v}^2 \). This experiment demonstrates that lighter molecules travel faster than heavier molecules by determining the relative distances traveled in the same time by different gases in still air.

HAZARDS: Concentrated HCl and NH\(_3\) can damage lungs, skin and clothing. Keep containers closed. Immediately clean up any spills with large amounts of water.

PROCEDURE:

1. Place one drop each of concentrated HCl and NH\(_3\) about 10 cm apart on a plastic sheet.

2. Lightly blow between the drops to remove existing gases, and immediately cover the drops with a 15 cm petri dish lid.

3. Observe until a cloud of white NH\(_4\)Cl(s) particles forms in a line under the lid. Measure and record the distance from the white line to each drop.

   Distance of HCl drop from white line: ________________

   Distance of NH\(_3\) drop from white line: ________________

4. Wipe up both drops and the NH\(_4\)Cl(s) on the sheet and lid with a wet sponge.

QUESTIONS:

1. Which gas traveled faster and farther?
2. Which gas has the lower molecular mass?

3. Do your observations agree with kinetic molecular theory, i.e. does the ratio of molecular masses calculated from your data agree with the ratio of the known molecular masses of HCl (36.5) and NH$_3$ (17)?

4. Write the chemical equation for the reaction.

5. If the experiment was done with C$_2$H$_5$NH$_2$ instead of ammonia, how would the result differ, i.e. what would the relative distances be?

6. Explain how you could determine the molar mass of an unknown, volatile acid.

7. What did you like about this experiment and how would you improve it?
# 48. The Nose Knows

**Name**___________________________  **Course Number**__________________________

**Student ID**___________________________  **Section Number**__________________________

**Date**___________________________  **Instructor**__________________________

**OBJECTIVE:** To demonstrate that the olfactory sense can differentiate isomers.

**CONCEPTS:** Chirality, Isomers

**INTRODUCTION:** Physical properties of structural and geometric isomers differ; however, chiral isomers (enantiomers) have identical physical properties. Except for their interaction with polarized light, enantiomers respond differently only when they are in a chiral environment. Our olfactory sense can easily differentiate structural, geometric, and chiral isomers on the basis of the chemical interaction between the molecules of the olfactory receptors and the isomer.

**HAZARDS:** Never put laboratory substances in your mouth.

**PROCEDURE:** Do the following for each of 4 pairs of compounds. Fill in the chart.

<table>
<thead>
<tr>
<th>Name of isomer</th>
<th>Melting point, °C</th>
<th>Boiling point, °C</th>
<th>Do isomers have same or different polarity?</th>
<th>Structural, geometric, or chiral isomers?</th>
<th>Describe smell</th>
</tr>
</thead>
<tbody>
<tr>
<td>1st pair</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2nd pair</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3rd pair</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>4th pair</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
1. Record the names of the isomers.

2. Inspect the models and decide if each of the pairs are (a) structural isomers (same number and kind of atoms, but connected in different sequence), (b) geometric or cis/trans isomers (same sequence but different spatial arrangement, i.e. not mirror images) or (c) chiral isomers (nonsuperimposable mirror images).

3. Note the color and melting and boiling points in the chart.

4. Inspect the models of the isomers and estimate whether the isomers will have the same or different dipole moments.

5. Take off the caps of the vials and smell the contents. Describe the smells. Re-cap the vials!

**QUESTIONS:**

1. What do you conclude about the physical properties of isomers?

2. Do the chiral isomers have different smells? What does this tell you about olfactory receptors?

3. Do you think chiral isomers would taste different?

4. What did you like about this experiment and how would you improve it?
49. Properties of the Halogens

**OBJECTIVE:** To demonstrate the relative reduction potentials of and observe some physical properties of the halogens.

**CONCEPTS:** Oxidation-reduction reactions, Periodic properties

**INTRODUCTION:** The halogens, group 7A in the periodic table, are fluorine, chlorine, bromine, iodine, and astatine. At room temperature these elements each exist as a diatomic molecule; F₂ and Cl₂ as gases, Br₂ as a liquid, and I₂ and At₂ as volatile solids. Fluorine is very difficult to handle because of toxicity and extreme reactivity. Astatine is very rare and radioactive. The elemental forms of the other three members of this group are readily prepared by the following reactions.

\[
\text{ClO}^- + \text{Cl}^- + 2 \text{H}^+ \rightarrow \text{H}_2\text{O} + \text{Cl}_2 \quad \text{reaction i}
\]

\[
\text{Cl}_2 + 2 \text{Br}^- \rightarrow 2 \text{Cl}^- + \text{Br}_2 \quad \text{reaction ii}
\]

\[
\text{Cl}_2 + 2 \text{I}^- \rightarrow 2 \text{Cl}^- + \text{I}_2 \quad \text{reaction iii}
\]

**HAZARDS:** Hypochlorite bleach, Cl₂, Br₂, and I₂ can damage eyes, clothes and skin. Please avoid contact with all of the solutions involved in this experiment and be sure to wear goggles.

**PROCEDURE:**

1. Working over a sheet of white paper, place 5 drops each of solution A (a basic solution of ClO⁻ and Cl⁻), solution B, and solution C at the corners of a triangle in a small, plastic petri dish. The solutions should be separated about 2 cm.

2. Add 2 drops of 6M HCl to A, place a cover on the petri dish and wait a few minutes for color changes.

   Record the color of A _______________ B _______________ C _______________.

3. Add a drop of mineral oil to each of B and C and stir each with a toothpick.

   Record the color in the oil in B _______________ and in C _______________.

4. Rinse the contents of the petri dish into the appropriate waste container.

**QUESTIONS:**

1. Cl₂ is produced in A by reaction i. How does Cl₂ get from A to B and C?
2. Concerning solutions B and C, one of them is 0.5 M KBr and the other is 0.1 M KI. Most people have had sufficient experience with iodine to be able to use the results from Procedure 3 to determine which is the KBr solution and which is the KI solution. According to your observations, solution B is _________ and solution C is ____________. Explain how you decided.

2. From your observations, which is the better oxidizing agent, Cl₂ _____ or Br₂ _____? Cl₂ _____ or I₂ ____? Explain how you decided.

3. From the Δ values for the half-reaction X₂ + 2e⁻ → 2X⁻(aq) given in the box to the right, list the halogens X₂ in increasing order of oxidizing strengths (i.e. the one with the least oxidizing strength should be on the left).

<table>
<thead>
<tr>
<th>X</th>
<th>E°</th>
</tr>
</thead>
<tbody>
<tr>
<td>F</td>
<td>2.86 V</td>
</tr>
<tr>
<td>Cl</td>
<td>1.36 V</td>
</tr>
<tr>
<td>Br</td>
<td>1.06 V</td>
</tr>
<tr>
<td>I</td>
<td>0.54 V</td>
</tr>
</tbody>
</table>

Are the conclusions based on your observations and expressed in item 2 above consistent with your conclusions based on the E° values? Yes_____ No_____
50. Redox Wheel

Name ____________________________ Course Number ____________________________
Student ID _________________________ Section Number ____________________________
Date ______________________________ Instructor _______________________________

**OBJECTIVE:** To measure reduction potentials for some redox couples.

**CONCEPTS:** Electrochemistry, Reduction potentials

**INTRODUCTION:** The ability of a reducing agent to react with an oxidizing agent is measured by the standard reduction potential, $E^\circ$. These are defined by setting the standard reduction potential of the $\text{H}^+(\text{aq})/\text{H}_2(\text{g})$ couple equal to zero. Potentials of various half-reactions of some common metals are listed below.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>$E^\circ$</th>
<th>Reaction</th>
<th>$E^\circ$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{Ag}^+ + e^- \rightarrow \text{Ag}(s)$</td>
<td>0.80</td>
<td>$\text{Fe}^{2+} + 2e^- \rightarrow \text{Fe}(s)$</td>
<td>-0.41</td>
</tr>
<tr>
<td>$\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}(s)$</td>
<td>0.34</td>
<td>$\text{Zn}^{2+} + 2e^- \rightarrow \text{Zn}(s)$</td>
<td>-0.76</td>
</tr>
<tr>
<td>$\text{Pb}^{2+} + 2e^- \rightarrow \text{Pb}(s)$</td>
<td>-0.13</td>
<td>$\text{Mg}^{2+} + 2e^- \rightarrow \text{Mg}(s)$</td>
<td>-2.38</td>
</tr>
<tr>
<td>$\text{Sn}^{2+} + 2e^- \rightarrow \text{Sn}(s)$</td>
<td>-0.14</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Individual potentials cannot be measured directly. However, two oxidation-reduction couples can be paired as the anode and cathode of an electrochemical cell whose potential can be directly measured. The cell potential is related to the standard reduction potentials of the anode and cathode by the relation $E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}}$.

**HAZARDS:** Do not ingest any of the chemicals. Immediately clean up any spills with copious amounts of water. Wear goggles.

**PROCEDURE:**

1. Place an 11 cm circle of filter paper in the center of a clean, smooth, ceramic tile or piece of plastic wrap.


3. Obtain small pieces of 6 different metals. Your choices are Pb, Fe, Ag, Mg, Cu, Zn, and Sn. Cu should be one of your choices. Clean each metal with steel wool. The metal should look shiny. In the table below number the metals 1–6.

4. Place 4 drops of 0.5 M NaNO$_3$ in the center of the filter paper. The solution will spread towards the edge.

5. Place 2 drops of each of the metal ion solutions at the numbered places on the filter paper, matching the number assigned to the metal to the number on the filter paper. Refer to the table below. The different metal ion solutions must contact the NaNO$_3$ solution, but should not touch each other.
6. Place the corresponding metal strip in the middle of its metal ion solution.

7. Choose copper to be the reference metal. Place the black (–) electrode of a high impedance voltmeter on the copper. Place the red (+) electrode on one of the other metals and record the voltage in the table below. Do this for each of the metals, keeping the (–) on the Cu metal. Be careful to record the correct sign (+ or –). You now have a table of reduction potentials with the \( \text{Cu}^{2+}/\text{Cu} \) potential defined to be zero.

<table>
<thead>
<tr>
<th>Number</th>
<th>Metal</th>
<th>( E^\circ \text{ vs Cu}^{2+}/\text{Cu} )</th>
<th>Your ( E^\circ ) vs ( \text{H}^+\text{H}_2 )</th>
<th>( E^\circ ) from introduction</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>5</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>6</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

8. Given that \( E^\circ \) for \( \text{Cu}^{2+}/\text{Cu} \) vs. \( \text{H}^+\text{H}_2(g) \) is 0.34 V, calculate the potentials with \( \text{H}^+\text{H}_2(g) \) as the reference zero. Check your values against the values given in the introduction of this experiment.

QUESTIONS:
1. Is there a significant difference between your standard potentials and the accepted values given in the introduction? If so, why? Do your values increase in the same order as the accepted values?

2. How would your results differ if you placed the (+) electrode on the reference metal during your measurements?

3. Why were you advised to clean the metal with steel wool before taking any measurements?

4. What does a (–) sign on the electrode of a flashlight battery mean?

5. According to the sign convention used in the introduction, if a metal couple with a (+) \( E^\circ \) value reacts with a metal couple with a (–) \( E^\circ \) value, which metal ion will be reduced and which metal will be oxidized?

6. What did you like about this experiment and how would you improve it?
51. Relieving Stress

OBJECTIVE: To observe the shift in the position of a chemical equilibrium with changes in reactant concentrations.

CONCEPTS: Le Chatelier’s principle, Chemical equilibrium

INTRODUCTION: The intensity of a color in a solution is directly proportional to the concentration of the colored species. The blood-red FeSCN$^{2+}$ complex ion produced by the reaction

$$Fe^{3+} + SCN^- \rightleftharpoons FeSCN^{2+}$$

can be used to indicate the position of the equilibrium between ferric ion and thiocyanate ion in solution. If one of the reactants, e.g. Fe$^{3+}$, is added to an equilibrium mixture, the equilibrium position will shift to consume a portion of the added Fe$^{3+}$ and produce additional FeSCN$^{2+}$ and the solution will become darker colored. If Fe$^{3+}$ is removed from the mixture, the equilibrium position will shift to consume a portion of the FeSCN$^{2+}$ and the color will become lighter. In this experiment the concentrations of both Fe$^{3+}$ and SCN$^-$ are changed to demonstrate the resulting effect on the concentration of blood-red FeSCN$^{2+}$. The SCN$^-$ ion is colorless and Fe$^{3+}$ appears yellow because of the presence of a small amount of the hydrolysis product FeOH$^{2+}$.

HAZARDS: KSCN is a deadly poison; do not ingest. If contact is made, wash immediately. Carefully clean up spills and dispose of waste.

PROCEDURE:

1. Place 10 mL of water in a 50 mL beaker. Add 6 drops each of the 0.1 M Fe(NO$_3$)$_3$ and 0.1 M KSCN solutions. Mix. Place 15 drops of this solution in each of 5 wells in a 24-well plate. Place the plate over a sheet of white paper so the colors can easily be seen.

2. Reserve the solution in well #1 for a control. To well #2, add 5 drops of the 0.1 M Fe(NO$_3$)$_3$ solution. Mix and record your observations.

3. To well #3, add 5 drops of the 0.1 M KSCN solution. Mix and record your observations.

4. To well #4, add an equal amount of water. Mix and record your observations.

5. To well #5, add a small crystal of Na$_3$HPO$_4$. Mix and record your observations.
QUESTIONS:

1. Using Le Châtelier's principle, explain what happened in wells #2 and 3.


3. Using Le Châtelier's principle, explain what happened in well #5. Note that Fe(III) forms a colorless complex with hydrogen phosphate ion.
52. Ring Around the Bathtub

**OBJECTIVE:** To observe the differences between soaps and detergents and their actions in tap water and distilled water.

**CONCEPTS:** Precipitation, Solubility

**INTRODUCTION:** Much of the water we use comes from underground sources, where it is in contact with limestone and other rocks and minerals. "Hard water" contains traces of dissolved salts of calcium, magnesium, and iron. These cations form precipitates with soaps, which are sodium salts of fatty acids, wasting the soap and interfering with the cleaning process. Detergents are made from surfactants that do not precipitate with Ca\(^{2+}\), Mg\(^{2+}\), or Fe\(^{2+}\). The oxalate (C\(_2\)O\(_4\)\(^{2-}\)) and carbonate (CO\(_3\)\(^{2-}\)) salts of Ca\(^{2+}\) are insoluble and can be used to test for hard water.

**HAZARDS:** Don't get soap in your eyes. Wear goggles, clean spills, and do not ingest any chemicals including soap and detergent.

**PROCEDURE:**

1. **Test for Ca\(^{2+}\).** Add one drop of 1 M sodium oxalate (Na\(_2\)C\(_2\)O\(_4\)) to 5 mL distilled water, 5 mL tap water, and a solution of two drops of 0.1 M CaCl\(_2\) in 5 mL distilled water in separate 13x100 mm test tubes. (Five mL is half-filled.) Stopper and shake the tubes. Let stand several minutes and examine them closely. Record your observations in the table below.

2. Add five drops of soap solution to 5 mL samples of distilled water, tap water, and a solution of two drops of 0.1 M CaCl\(_2\) in 5 mL distilled water in clean test tubes. Stopper and shake the tubes. Record your observations in the table below.

3. Add five drops of detergent solution to 5 mL samples of distilled water, tap water, and a solution of two drops of 0.1 M CaCl\(_2\) in 5 mL distilled water. Stopper and shake the tubes. Record your observations in the table below.

**Results:**

<table>
<thead>
<tr>
<th>Addition</th>
<th>distilled water</th>
<th>tap water</th>
<th>5 mM CaCl(_2)</th>
</tr>
</thead>
<tbody>
<tr>
<td>oxalate ion</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>soap</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>detergent</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
QUESTIONS:

1. Can you distinguish between distilled and tap water merely by looking at it?

2. Write the chemical equation for the reaction observed with calcium and oxalate ions.

3. What obvious advantage do synthetic detergents have over soaps?

4. In pioneer days the housewife prepared soft water by adding washing soda (Na₂CO₃) to a barrel of washing water, stirring it, letting it stand, then using the water at the top of the barrel. Discuss the chemistry involved, with appropriate chemical equations.

5. What type of substance (soap or detergent) do you think is used in hair shampoo and dishwashing "soap"? Why?

6. People often conclude it is more difficult to rinse soap off in the shower with soft water than hard water because the soft water leaves the skin feeling slick while the hard water leaves the skin with a sticky feeling. Why is this conclusion wrong?

7. What did you like about this experiment and how would you improve it?
53. Signs of Chemical Change

OBJECTIVE: To observe changes that indicate a chemical reaction is occurring.

CONCEPTS: Chemical change, Physical properties

INTRODUCTION: Chemical reactions are always accompanied by changes in the physical properties as reactants are converted into products. Color change, gas evolution, precipitation, and heat production are all indications that a reaction is occurring.

HAZARDS: Immediately clean up any chemicals spilled or in contact with skin or clothing with generous amounts of water.

PROCEDURE:
1. Place a small amount of CuCO$_3$(s) (about 1/4 the size of a garden pea) in a small test tube. Observe and note the physical properties in the table.
2. Add one drop of water. Note the absence of any sign of a chemical reaction.
3. Add 6 M HCl dropwise, no more than 5 or 6 drops. Note any changes in physical properties. Record the properties of the products of the reaction in the table.
   \[ \text{CuCO}_3(s) + 2\text{H}^+(aq) + 4\text{Cl}^-(aq) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(l) + \text{CuCl}_4^{2-}(aq) \]
4. Add 20 drops of water. Note any changes.
   \[ \text{CuCl}_4^{2-}(aq) + 4\text{H}_2\text{O}(l) \rightarrow \text{Cu(H}_2\text{O})_4^{2+}(aq) + 4\text{Cl}^-(aq) \]
5. Add about 6 drops of 6 M NH$_3$ solution dropwise, shaking the test tube carefully to mix the solution after each drop. Stop the addition when you observe a precipitate. Note any indication of heat being produced. Record the properties of the precipitate.
   \[ \text{NH}_3(aq) + \text{H}^+(aq) \rightarrow \text{NH}_4^+(aq) + \text{heat} \]
   \[ \text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightarrow \text{NH}_4^+(aq) + \text{OH}^- \]
   \[ \text{Cu(H}_2\text{O})_4^{2+}(aq) + 2\text{OH}^- \rightarrow \text{Cu(OH)}_2(s) + 2\text{H}_2\text{O}(l) \]
6. Continue addition of the ammonia solution dropwise, shaking the test tube after each drop, until the precipitate dissolves. (If the volume in the tube gets too large, transfer some of it to another test tube, and continue on with the first test tube.) Record the properties of the solution.
   \[ \text{Cu(OH)}_2(s) + 4\text{NH}_3(aq) \rightarrow \text{Cu(NH}_3)_4^{2+}(aq) + 2\text{OH}^-(aq) \]
7. Divide the contents of the test tube equally into two test tubes. Keep one as a control. Add granulated zinc metal, about 1/4 of a pea size, to the other. Stir slowly and watch for a change. What do you observe? (Hint: finely divided metals often appear black.)
   \[ \text{Cu(NH}_3)_4^{2+}(aq) + \text{Zn}(s) \rightarrow \text{Cu}(s) + \text{Zn(NH}_3)_4^{2+}(aq) \]
8. To the unused control from step #7, add 6 M HCl solution dropwise, to reverse steps 6 and 5 above. Note any changes.

9. Discard the test tubes, solutions, and solids in the proper containers and clean up any spills.

<table>
<thead>
<tr>
<th>Chemical species</th>
<th>Properties</th>
</tr>
</thead>
<tbody>
<tr>
<td>CuCO₃(s)</td>
<td></td>
</tr>
<tr>
<td>CO₂(g)</td>
<td></td>
</tr>
<tr>
<td>H₂O(l)</td>
<td></td>
</tr>
<tr>
<td>CuCl₂⁻(aq)</td>
<td></td>
</tr>
<tr>
<td>Cu(H₂O)₄²⁺(aq)</td>
<td></td>
</tr>
<tr>
<td>NH₃(aq)</td>
<td></td>
</tr>
<tr>
<td>NH₄⁺(aq)</td>
<td></td>
</tr>
<tr>
<td>Cu(OH)₂(s)</td>
<td></td>
</tr>
<tr>
<td>Cu(NH₃)₄²⁺(aq)</td>
<td></td>
</tr>
<tr>
<td>Zn(s)</td>
<td></td>
</tr>
<tr>
<td>Zn(NH₃)₄²⁺(aq)</td>
<td></td>
</tr>
<tr>
<td>Cu(s)</td>
<td></td>
</tr>
</tbody>
</table>

**QUESTIONS:**

1. Summarize the evidences of chemical reactions observed in this experiment.

2. Classify the reactions as precipitation, complexation, reduction-oxidation, etc.

3. What did you like about this experiment and how would you improve it?
OBJECTIVE: To observe and record the wavelengths for the colors of the visible spectrum.

CONCEPTS: Electromagnetic spectrum, Color, Color and energy

INTRODUCTION: Human eyes respond to light between the wavelengths of 400 and 700 nm, the visible part of the electromagnetic spectrum. Radiation from an incandescent source such as the sun or a tungsten light bulb appears to be white because it consists of all the colors of the visible spectrum. When a beam of electromagnetic radiation is passed through a prism or a diffraction grating, each wavelength is bent through a different angle and our eyes are able to detect a rainbow of colors. In a colorimeter, light from a lamp is passed through a prism or grating that can be rotated to shine different colors of light on the sample compartment. A Spectronic 20® or similar colorimeter with a test tube containing a small wad of crumpled aluminum foil in the sample compartment reflects the color selected by the wavelength knob up through the test tube for viewing.

HAZARDS: None that we know of, but be aware that other experiments may be in progress and all safety precautions are still in effect.

PROCEDURE:

1. Turn on the colorimeter. Set the wavelength at 600 nm. With your eye over the test tube, turn the test tube and move it up and down slightly to get the maximum light reflected. Rotate the wavelength knob and note the colors.

2. Start with a blue color and slowly adjust toward shorter wavelengths. The light intensity appears to decrease because your eyes become less responsive as the wavelength approaches the end of the visible region of the spectrum. Record the shortest wavelength that your eyes can detect.

3. Record the longest wavelength that your eyes can detect.

4. Pick out as many separate colors as you can and record the color and wavelength for each on the table.

5. Turn off the instrument.
<table>
<thead>
<tr>
<th>Color</th>
<th>Wavelength, nm</th>
</tr>
</thead>
<tbody>
<tr>
<td>Shortest detectable wavelength (near blue)</td>
<td></td>
</tr>
<tr>
<td>Longest detectable wavelength (near red)</td>
<td></td>
</tr>
</tbody>
</table>

Pertinent equations: \( E = h \nu \), \( c = \lambda \nu \), \( E = \frac{hc}{\lambda} \)

**QUESTIONS:**

1. Which color represents light of the greatest energy?

2. Which color represents light of the lowest energy?

3. Is there such a thing as white light of a given frequency?

4. If sunlight or light from an incandescent bulb actually consists of a number of colors, why does it appear white to us?

5. If you can, compare your data with that of other students. Why does it not always agree?

6. About 350 kJ is required to break a mole of covalent bonds. Does a photon of blue light at about 460 nm have enough energy to break such a covalent bond? \( h = 6.626 \times 10^{-34} \text{ J} \cdot \text{s} \), \( c = 3.00 \times 10^8 \text{ m} \cdot \text{s}^{-1} \), \( N_A = 6.02 \times 10^{23} \text{ molecules mol}^{-1} \).

7. What did you like about this experiment and how would you improve it?
55. Speed Limits: Determining a Rate Law

OBJECTIVE: To determine the order of a reactant in a rate law.

CONCEPTS: Reaction kinetics, Rate laws, Reaction order

INTRODUCTION: Thiosulfate ion is stable in neutral or basic solution, but decomposes in acidic solution to produce sulfurous acid and sulfur.

\[ 2H^+ + S_2O_3^{2-} \rightarrow H_2SO_3(aq) + S(s) \]

The dependence of the decomposition rate on the thiosulfate ion concentration is determined in this experiment. The rate law for the decomposition is

\[ \text{rate} = \frac{\Delta n_{S(s)}}{\Delta t} = k[H^+]^y[S_2O_3^{2-}]^x \]

where \( \Delta n \) is the moles of sulfur appearing as product in time interval \( \Delta t \); \( y \) and \( x \) are the orders with respect to \( H^+ \) and \( S_2O_3^{2-} \), respectively. If the hydrogen ion concentration, \( [H^+] \), is constant, the rate law becomes

\[ \text{rate} = \frac{\Delta n_{S(s)}}{\Delta t} = k'[S_2O_3^{2-}]^x. \]

To determine the rate of the reaction, the time required for enough solid sulfur to form to make the solution opaque is measured. A slower rate takes longer to yield the same amount of sulfur. In the procedure used in this experiment the amount of sulfur, \( \Delta n_{S(s)} \), is constant, thus the relative rate is given by \( 1/\Delta t \), i.e. the reciprocal of the time required for the solution to become opaque. The concentrations of \( H^+ \) and \( S_2O_3^{2-} \) change by a negligible amount during the reaction. The initial concentration of \( S_2O_3^{2-} \) differs in the six reaction mixtures. The variation of rate with \( [S_2O_3^{2-}] \) determines the reaction order, \( x \), with respect to thiosulfate.

HAZARDS: Immediately wash off any chemicals spilled on your skin or clothes. Wear goggles.

PROCEDURE:
1. Place two 24-well plates with their 6-well rows adjacent to each other as shown. Add reagents to the last row of plate A and the first row of plate B as shown.
2. Be ready to record the time of mixing to the nearest second! Place plate A up on the adjacent edge of plate B and tilt it over to allow each well's contents to run into the corresponding wells in Plate B. Set plate A aside. Swirl plate B briefly, but gently, to mix the contents of each well. Each well should have the same volume, that is, 27 drops.

3. Place the plate over a page with printed matter. Note that the wells start getting cloudy as sulfur forms. Record the time in seconds when a well gets so cloudy you can just barely make out the printing below it. Complete the following table:

<table>
<thead>
<tr>
<th>Well #</th>
<th>1</th>
<th>2</th>
<th>3</th>
<th>4</th>
<th>5</th>
<th>6</th>
</tr>
</thead>
<tbody>
<tr>
<td>drops S\textsubscript{2}O\textsubscript{3}\textsuperscript{2−}</td>
<td>21</td>
<td>18</td>
<td>15</td>
<td>12</td>
<td>9</td>
<td>6</td>
</tr>
<tr>
<td>time (s)</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>1/t (s\textsuperscript{−1})</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

After the last time is recorded, immediately clean the plates with a cotton swab because the sulfur becomes hard to remove after it settles.

4. Make a plot of 1/t (y axis) versus drops of S\textsubscript{2}O\textsubscript{3}\textsuperscript{2−} (x axis). Do the points fall on a straight line? Draw the best line. Relative rates that do not change with [S\textsubscript{2}O\textsubscript{3}\textsuperscript{2−}] indicates zero order with respect to S\textsubscript{2}O\textsubscript{3}\textsuperscript{2−}. A linear increase in rate with [S\textsubscript{2}O\textsubscript{3}\textsuperscript{2−}] indicates first order with respect to S\textsubscript{2}O\textsubscript{3}\textsuperscript{2−} or x=1. A quadratic curve (parabola) indicates x=2.

**QUESTIONS:**

1. Write the rate law you found.

2. Suppose you wanted to know the order with respect to hydrogen ion. Describe how you could determine this.

3. Suggest several sources of error in this experiment.

4. To obtain x, why is it not necessary to know the exact thiosulfate concentration, only ratios?

5. What did you like about this experiment and how would you improve it?
OBJECTIVE: To demonstrate the catalytic action of enzymes and the iodine test for starch.

CONCEPTS: Catalysis, Enzymes, Hydrolysis of starch to sugar

INTRODUCTION: Starch is a polymer of the sugar glucose. Enzymes are proteins present in living systems that catalyze chemical reactions. Hydrolysis of starch in food to sugar as catalyzed by the enzyme amylase in human saliva is demonstrated in this experiment. (Note that the names of enzymes end in -ase.) Glucose molecules in starch are joined head-to-tail through oxygen atoms. A thousand or more glucose molecules may be linked in this way. In aqueous solutions of amylase, the starch reacts with water to form glucose.

\[
\text{starch} \xrightarrow{+ nH_2O} n \text{glucose}
\]

Starch, but not sugar, forms an intensely blue-black colored complex with iodine. The infamous gangster Dillinger once escaped from prison by carving a potato into the shape of a gun and coloring it with tincture of iodine from a first aid kit. (While the story is probably apocryphal, it shows that chemistry has practical value.)

HAZARDS: Iodine can cause skin irritation and can stain clothing. Iodine is volatile, so vapors will be present over the solutions. Some people are allergic to iodine; contact may produce hives and other allergic reactions in these people. Use adequate ventilation.

PROCEDURE:

1. Test several food items for starch by placing a drop of 1% iodine solution (in alcohol) on each. (Menu suggestions: potato, apple, bread, crackers, rice, carrot, popcorn, candy, nuts, etc., whatever is in your lunchbox.)

<table>
<thead>
<tr>
<th>Food</th>
<th>Presence of starch (+/-)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Food</th>
<th>Presence of starch (+/-)</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

-112-
2. Place a small amount (about half the size of a garden pea) of corn starch in each of two small, disposable beakers. Add about 1 mL of water to one and twice as much to the other. Then add 1 drop of tincture of iodine to each, mix thoroughly, and note the color. Add about 1 mL of saliva to the beaker with less water. Use separate, new toothpicks to thoroughly mix the liquid with the corn starch. Observe and record the changes occurring over the next several minutes.

3. Because eating in the lab is not a good practice for safety reasons, this last part of the experiment must be done elsewhere. Chew a small piece of bread or cracker, but do not swallow it. Hold it in your mouth for several minutes. Note the change in flavor with time.

QUESTIONS:
1. What differences did you observe in the two beakers? Explain.

2. Why must separate toothpicks be used to mix the control and the saliva reaction mixtures?

3. How does the flavor of a starchy food change with time when mixed with saliva and held in the mouth? Explain.

4. What is the significance of amylase to living systems?

5. There is an enzyme in mammalian blood (and in many other organisms) called carbonic anhydrase that catalyzes the reaction

\[ \text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}). \]

Why is it important for organisms to have such an enzyme?

6. Can you describe a simple experiment to show the presence or absence of carbonic anhydrase?

7. What did you like about this experiment and how would you improve it?
OBJECTIVE: To generate and observe vanadium ions in various oxidation states.

CONCEPTS: Oxidation-reduction reactions, Colors of transition metal ions

INTRODUCTION: Vanadium ions of various oxidation states are easily produced in aqueous solution. Vanadium in +2, +3, +4, and +5 oxidation states is produced in this experiment. Each of these ions has a characteristic color in water. The VO$_2^+$ and VO$_3^+$ ions contain covalent metal-oxygen bonds. Because V(V) polymerizes in aqueous solution, and some of the polymerization reactions are extremely slow, the actual form of the V(V) ion (represented here as VO$_2^+$) depends on the history of the solution. The form most often encountered is the intensely yellow polymer, H$_4$V$_{10}$O$_{28}$$^{2-}$ (decavanadate). Zinc metal is used to reduce VO$_2^+$ in a stepwise manner and produce the different colored ions.

(+5) VO$_2^+$ (the experiment begins with a polymeric form of the dioxovanadium(V) ion)
(+4) 2VO$_2^+$ + Zn(s) + 4H$^+$ = 2VO$^{2+}$ + Zn$^{2+}$ + 2H$_2$O VO$^{2+}$ (vanadyl ion, sky blue)
(+3) 2VO$^{3+}$ + Zn(s) + 4H$^+$ = 2V$^{2+}$ + Zn$^{2+}$ + 2H$_2$O V$^{2+}$ (vanadium(III) ion, green)
(+2) 2V$^{3+}$ + Zn(s) = 2V$^{2+}$ + Zn$^{2+}$ V$^{2+}$ (vanadium(II) ion, violet)

HAZARDS: The solutions contain strong acid and will irritate skin and damage clothing. If a spill occurs, immediately rinse with copious amounts of water. To protect others, carefully clean up spills.

PROCEDURE:

1. Obtain a 24-well plate and place a piece of mossy zinc metal in one well. The piece of Zn should cover about one-half the bottom of the well.

2. Cover the Zn with the 0.2 M VO$_2^+$ and 2M H$_2$SO$_4$ solution until the well is nearly full. Fill a second well about one-third full with the same VO$_2^+$ solution. The VO$_2^+$ in the second well serves as a reference. White paper under the plate helps visualize the colors.

3. Stir the solution with the zinc for about 5 min. The yellow color first changes to green, a mixture of yellow and blue. Wait until the color is a clear sky-blue.

4. With a pipet, draw off about one-third of the blue solution and place in a third well.

5. Continue to stir the solution with the zinc metal until it turns a green color. When the solution is green, remove another third and place in a fourth well.

6. Continue to stir the solution over the zinc metal. Stir another 5–10 min. The solution will turn violet or purple.

7. Remove the zinc with tweezers and place in a waste-zinc beaker.
8. Clean the plate by rinsing into a waste container. Rinse the pipets and discard the toothpick.

**Questions:**

1. What is the gas produced during the reactions?

2. Why do transition metal ions have different colors in different oxidation states?

3. What do you think would happen if you mixed the solutions of vanadium(IV) and vanadium(II)? Write an equation for the reaction.

4. Note the electron configuration of vanadium(V). Is vanadium(VI) an expected oxidation state? Why or why not?

5. What did you like about this experiment and how would you improve it?
58. Warm It Up If You Are in a Hurry

**OBJECTIVE:** To demonstrate the effect of temperature on the rate of a chemical reaction.

**CONCEPTS:** Activation energy, Rate laws, Absolute temperature

**INTRODUCTION:** As temperature increases, molecular kinetic energy increases, the fraction of molecules with sufficient kinetic energy to react increases (about doubles for a 10 K temperature rise), the rate of collisions increases (about 7% for a 10 K temperature rise), and the reaction goes faster. The minimum kinetic energy a collision must have if a reaction is to occur is called the activation energy. Because of the relation between molecular kinetic energy and absolute temperature, the activation energy can be determined by measuring the rate of a reaction as a function of temperature. The rate of a reaction is described by a rate law specific for every reaction, i.e. rate = k times a function of the reactant concentrations. Svante Arrhenius showed that k is related to the absolute temperature by the equation $k = Ae^{-\frac{E_a}{RT}}$ where A is a constant, $E_a$ is the activation energy, R is the gas constant (8.315 J mol$^{-1}$ K$^{-1}$), and T is the Kelvin temperature. Thus k increases exponentially with T. As a rule of thumb, biological reaction rates double for every 10 K increase of temperature. This rule is exactly true for any reaction having $E_a = 53$ kJ mol$^{-1}$ at 25°C.

In this experiment, the time required for an antacid tablet to dissolve in vinegar (4% acetic acid in water) at different temperatures will be measured. The rate law for this reaction is

$$\text{average rate} = \frac{\Delta n}{\Delta t} = k \text{[acetic acid]}$$

where $\Delta n$ is the number of moles of antacid in the tablet, $\Delta t$ is the time required for the tablet to dissolve, k is the rate constant and [acetic acid] is the molarity of the acid. Because $\Delta n$, the number of moles in one tablet, and acetic acid concentration are kept constant, the average rate is proportional to $1/\Delta t$.

$$\text{average rate } \propto \frac{1}{\Delta t} = k \text{[acetic acid]}/\Delta n$$

**HAZARDS:** Avoid burns from the hot plate. Wear safety goggles as others around you may be using hazardous chemicals.

**PROCEDURE:**

1. To save time, three runs will be made simultaneously, with one antacid tablet used in each. Prepare beakers of 100 mL of vinegar. Cool the first to 15°C, warm the second to 30°C and warm the third to about 50°C. When all three are ready, drop one antacid tablet in each beaker. Use a watch or clock with a second hand to determine the time required for complete dissolution, and record time for each beaker. Measure the temperature, but do not stir the solution during the reaction.
Run | T, °C | T, K | Δt, s | rate = 1/Δt
--- | --- | --- | --- | ---
1  
2  
3

**QUESTIONS:**

1. Suggest two reasons why the rates of chemical reactions increase with increasing temperature.  
   1.  
   2.  

2. Which of these two should predominate? Why?

3. By what factor did you observe the rate of dissolution of the tablets to increase per 10 K increase in temperature?  
   a. From 15 to 30°C?  
   b. From 30 to 50°C?  
   c. From 15 to 50°C?

4. List several reactions that occur more rapidly at elevated temperatures. (Hint: consider cooking, gluing, painting, plants growing, lizards running, hair drying, etc.)

5. What did you like about this experiment and how would you improve it?
OBJECTIVE: To observe fluorescence.

CONCEPTS: Fluorescence, Optical brighteners

INTRODUCTION: Molecules are excited to a higher electronic energy state by visible and ultraviolet light. Most molecules convert the electronic excitation energy into vibrational energy, i.e. heat. However, some molecules, typically flat, rigid molecules with several \( \pi \) bonds, relax by light emission, i.e. fluorescence. Because some of the absorbed energy is converted into heat, the emitted light is at a lower energy and longer wavelength.

The laundry detergent industry uses fluorescence to their advantage by use of agents (optical brighteners) that absorb invisible ultraviolet and emit visible blue light. These optical brighteners cover the yellowish color in old whites by completing the spectrum. You see white light, i.e., white clothes.

(Note that flour, not fluor, is used to make bread, and that many flowers fluoresce in ultraviolet light.)

HAZARDS: Always wear your goggles while in the lab. Ultraviolet radiation is dangerous to your eyes! Do not look directly into the light. If any of the samples get on your hands, wash them thoroughly.

PROCEDURE:

1. Turn the UV lamp on with the switch on the cord. Push the button on the top of the light marked Longwave UV.

2. Choose ten of the samples, remove the lids, and place each sample under the light. If the sample fluoresces, record the color and relative intensity of the fluorescence in the chart. If it does not fluoresce, write "no" in the "color of fluorescence" column.

3. Many things we contact fluoresce. Put your hands under the light and see if there is anything on your hands that fluoresces.

4. Wash your hands with liquid laundry soap. Place your hands under the light. How well did the soap wash off?

5. Place the piece of paper marked with sunscreen under the light and look for fluorescence.

6. Turn off the light using the switch on the cord.
Questions:

1. Did you observe anything on your hands? If so, what? What happened after you washed your hands?

2. What did you observe with the sunscreen? How do you think sunscreen works?

3. What is the advantage of having an optical brightener in laundry detergent?

4. What other products do you think might have optical brighteners?

5. Can a molecule absorb infrared light and re-emit it in the visible spectrum? Why or why not?

6. Why do you think bees have better vision in the ultraviolet than visible region of the spectrum?

7. What did you like about this experiment and how would you improve it?
60. The Zinc Iodine Cycle

OBJECTIVE: To observe a redox chemical reaction.

CONCEPTS: Chemical change, Electrochemistry, Oxidation-reduction reactions, Physical properties

INTRODUCTION: Zinc and iodine react in a spontaneous oxidation-reduction reaction. Zinc is oxidized and iodine is reduced. The reverse reaction in which zinc ions are reduced and iodide ions are oxidized is not spontaneous and requires an external energy source to drive the reaction. Both the forward and reverse reactions are demonstrated in this experiment.

HAZARDS: The reaction that takes place in procedure 4 is somewhat violent and produces a lot of heat. Because of the violence of the reaction some reactants may be ejected from the test tube during the reaction. Wear your goggles! If you come in contact with any of the reactants, thoroughly rinse the affected area immediately with water. The test tube should be placed in the test tube rack before carrying out procedure 4.

PROCEDURE:

1. Add about 0.3 g zinc dust to a test tube. (See display.) Describe the physical properties of zinc.

2. Add about 0.2 g iodine crystals to a separate test tube. (See display.) Describe the physical properties of iodine.

3. Pour the iodine into the test tube containing the zinc. Describe any reaction and the physical properties of the mixture of solids.

4. Place the test tube with the solid mixture in a test tube rack, and carefully add a few drops of water. Describe the reaction and the changes you observed. As part of your description mention whether the reaction is endothermic or exothermic.

5. Use a 10 mL graduated cylinder to add 3-4 mL water to the test tube. Mix vigorously to clean off the sides of the test tube. Place the test tube in the test tube rack and let it stand until the solid has settled and the solution above the solid is clear. What are the physical properties of the solid at the bottom of the test tube? Which is the excess reactant, Zn or I₂? Explain how you decided which reactant was in excess.
6. Decant as much liquid as possible into a petri dish without allowing any solid into the petri dish. Place clean copper electrodes connected to a battery into the solution in the petri dish.
   a. Describe what is happening at the positive electrode.

   b. Describe what is happening at the negative electrode.

7. Without stirring, add one crystal of iodine to a colorless portion of the solution in the petri dish. Describe what happens when you add the crystal of iodine.

**Questions:**

1. Write a balanced equation for the reaction of solid zinc with solid iodine.

2. Write a balanced equation for the reaction that occurs when the product of the balanced equation in Question 1 is mixed with water.

3. When I\(_2\)(s) dissolves in a solution containing I\(^{-}\), the triiodide ion, I\(_3\)\(^{-}\), forms. What color is I\(_3\)\(^{-}\) in aqueous solution? Explain how you know.

4. Write a balanced half-reaction for what happened at each of the electrodes in Procedure 6.

   Half-reaction at the positive electrode:

   Half-reaction at the negative electrode:

5. The positive electrode is the (check one) anode , cathode . Explain how you decided.
OBJECTIVE: To demonstrate the acid-base properties of salts.

CONCEPTS: Acids and bases, Hydrolysis, pH

INTRODUCTION: Salts can be considered to be formed by the reaction of an acid with a base. Salts are classified as neutral, basic, or acidic according to the effect they have on pH when dissolved in water. A salt formed from a strong acid and a strong base will be neutral, i.e. will give a neutral solution when dissolved in water. When a weak acid reacts with a strong base, the salt produced is basic. When a weak base reacts with a strong acid, the salt produced is acidic.

To summarize:

Strong acid + strong base → neutral solution

Weak acid + strong base → basic solution

Strong acid + weak base → acidic solution

To explain: When a weak acid (HA) is neutralized, the conjugate base (A⁻) is produced. The conjugate base reacts with water to produce OH⁻.

\[ A^- + H_2O \rightleftharpoons HA + OH^- \]

Similarly, when a weak base (B) is neutralized, the conjugate acid (HB⁺) is produced and reacts

\[ HB^+ + H_2O \rightleftharpoons B + H_3O^+ \]

Acidic solutions are also produced by salts of small, highly charged cations such as Al³⁺, Fe³⁺, and Cr³⁺ with strong acid anions. In aqueous solution, these metal ions are surrounded by six water molecules. The high charge density on the cation attracts electrons away from the O-H bonds in the water molecules, allowing ionization of H⁺.

\[ Al(H_2O)_6^{3+} + H_2O \rightarrow Al(H_2O)_5OH^{2+} + H_3O^+ \]

HAZARDS: Avoid ingestion or contact with the salts, several are toxic.

PROCEDURE:
1. Put a small amount, about the size of a small pea, of each salt listed in the table in the wells of a 24-well plate.
2. Add ten drops of water to each salt and to one empty well as a control. Gently swirl the plate to dissolve the salts.
3. Add one drop of universal indicator to each well. For each solution, record the color of the indicator in the table.
4. Empty the solutions into the waste container and rinse the plate with distilled water when you are finished.
Universal indicator is red in acid, blue or purple in base and goes through a rainbow of colors between, with greenish yellow being neutral. Record your estimate of the pH in the table.

<table>
<thead>
<tr>
<th>color</th>
<th>red-orange</th>
<th>orange</th>
<th>yellow</th>
<th>green</th>
<th>purple</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH</td>
<td>2</td>
<td>4</td>
<td>6</td>
<td>8</td>
<td>10</td>
</tr>
</tbody>
</table>

The remaining columns of the table will help you understand your observations. Complete the table:

<table>
<thead>
<tr>
<th>Salt</th>
<th>Universal Indicator color</th>
<th>pH</th>
<th>Formula of Cation</th>
<th>Formula of products of reaction, if any, of cation with water</th>
<th>Formula of Anion</th>
<th>Formula of products of reaction, if any, of anion with water</th>
<th>Salt: Acidic or Basic?</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaCl</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaC₂H₃O₂</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>AlCl₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NH₄Cl</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ca(NO₃)₂</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na₂CO₃</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaF</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>NaI</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na₃PO₄</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Questions:**
1. Is ammonium nitrate an acidic or basic salt? Why?
2. Is potassium acetate an acidic or basic salt? Why?
3. Is FeCl₃ an acidic salt or a basic salt? Why?
4. Al₂(SO₄)₃ (alum) is used as a sizing agent in inexpensive paper such as newsprint, but not in “acid-free” papers. Explain why not.
5. CaCO₃ is the main ingredient in most antacid medicines. Write an equation showing how this salt neutralizes stomach acid, HCl.
6. What did you learn from this experiment? How would you improve the experiment?
62. Avogadro’s Hypothesis

**OBJECTIVE:** To demonstrate that equal volumes of gases at the same temperature and pressure contain the same number of molecules.

**CONCEPTS:** Gas properties, Mole, Molecular mass

**AVOGADRO’S HYPOTHESIS INTRODUCTION:** Amedeo Avogadro observed that the ratios of volumes of reactant gases involved in chemical reactions are ratios of small whole numbers. He hypothesized that equal volumes of gases under the same conditions of pressure and temperature contain equal numbers of molecules; but due to opposition, his hypothesis was not accepted until 50 years after his death.

**INTRODUCTION:** This hypothesis is not an obvious one because molecules differ in size. We find that in one mole of liquid water there are about 18 mL, roughly 1 tablespoon, while a mole of liquid pentane occupies about 115 mL, or 7.7 tablespoons. In the liquid phase the molecules are in close contact with one another. But, because of the large volume of space between molecules in the gas phase, the volume of the molecules becomes insignificant so that in the gas phase one mole of both water and pentane occupy about 22 L (at 1 atm and 25°C). One mole of each takes up the same amount of volume in the gas phase, but in the liquid phase the volumes are quite different.

In this experiment the same volume of several different gases under the same P and T conditions are weighed; and the number of molecules of each gas in that same volume is calculated from the molecular weight. Remember that a mole of gas is $6.02 \times 10^{23}$ molecules (or atoms if the gas is monoatomic).

**HAZARDS:** If misused, gas cylinders can be extremely dangerous. The main valves on the gas cylinders will be adjusted by a lab assistant. You only need to turn the small gas plastic delivery valve on and off. Pentane and hydrogen are flammable.

**PROCEDURE:**

1. Locate a 50 mL syringe with a small hole drilled in the plunger handle. The syringe is designed to be filled to a certain volume. The syringe should have a small bolt with it, and a hole in the syringe body to fit the bolt. Be sure to use the same syringe and bolt for all measurements.

2. Make a control measurement by pushing the plunger all the way in and sealing the syringe with a syringe cap. Next, pull the plunger out (it should be difficult), and insert the bolt through both the hole in the plunger and the syringe body (this is the volume to which you will fill your syringe with each of the gases). You now have a vacuum (nothing) in the syringe. Weigh the empty syringe and write the result here: __________
   All digits the analytical balance gives you are significant. You should have 6 significant figures.

3. Next, you will weigh gases that are available such as $\text{H}_2$, $\text{O}_2$, $\text{N}_2$, $\text{He}$, $\text{Ar}$ $\text{CH}_4$, and $\text{CO}_2$, which will be delivered from either gas cylinders, lab spouts, or can be drawn off of dry ice. Air is also available in plenty in the lab, and is a mixture of mostly nitrogen (78%), oxygen (21%), and argon (1%) with an average molar mass of 29.0 g mol$^{-1}$. Note: Always push the plunger in against the bolt so you have
the same volume in the syringe.
• Remove the bolt and turn the cap and then completely depress the plunger into the syringe body.
• Open the valve to the desired gas.
• Fill the syringe by connecting the syringe tube to the desired gas line (or CO₂ cooler spout). Slowly pull out the plunger to draw in the desired gas.
• Then, when just over 50 mL of gas are in the syringe, seal the syringe cap, disconnect the gas line and insert the bolt through the hole. If the bolt is not resting against the syringe body, push the plunger in all the way and quickly twist the cap open then closed. This process ensures a reproducible volume of 50 mL.
• Weigh the filled syringe and record the mass in the table below.
• Repeat and fill in the table below.

<table>
<thead>
<tr>
<th>Gas Identity</th>
<th>Weight of syringe and gas (g)</th>
<th>Mass of 50.0 mL of gas (g)</th>
<th>Molar mass of gas (g mol⁻¹)</th>
<th>Molecules in 50.0 mL of gas</th>
</tr>
</thead>
<tbody>
<tr>
<td>Air</td>
<td></td>
<td></td>
<td>29.0</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**QUESTIONS:**
1. Show your calculations for the number of gas molecules in 50.0 mL of air (include units).

2. Do your results support Avogadro’s hypothesis within experimental error? Yes___ No___ Explain how you decided.

3. A. Which has more mass, 50 mL of air or 50 mL of krypton (circle one)?
   B. Which has more molecules, 50 mL of air or 50 mL of krypton (circle one)?
63. Boiling Water with Ice

Name ____________________________ Course Number _______________________
Student ID ________________________ Section Number _______________________
Date _____________________________ Instructor ____________________________

**OBJECTIVE:** To demonstrate that boiling point depends on the external pressure.

**CONCEPTS:** Boiling point, Vapor pressure, Temperature

**INTRODUCTION:** In general, the boiling point is the temperature at which the vapor pressure of a liquid equals the external pressure whatever it may be. The *normal* boiling point is the temperature at which the vapor pressure of a liquid equals one atmosphere.

**HAZARDS:** Use a towel or tongs to handle hot glassware. Be careful not to dislodge the stopper from the side arm of the flask or to allow hot water to leak onto your hands or body. Do not use a thin wall flask; it might implode during the experiment.

**PROCEDURE:**
1. Place 60 mL of water in a 125 mL, heavy wall, Erlenmeyer filter flask with a tightly stoppered side arm.

2. Bring the water to a rolling boil on a hot plate, and measure the temperature of the water. The temperature is ____________.

3. Using a towel, stopper the flask while removing it from the hot plate.

4. Invert the flask, stand it on the large stopper, and carefully place a handful of crushed ice on the bottom of the flask.

5. Observe the liquid in the flask. When the liquid is boiling steadily, turn the flask upright, remove the stopper, and again measure the temperature. The temperature is ____________.

6. Clean up your work area.

**QUESTIONS:**
1. The normal boiling point of water is 100°C. Does the temperature you recorded in Procedure 2 agree with this? Yes___ No___ If the temperature doesn’t agree, explain why.
2. How is it you were able to get the water to boil at a temperature below the temperature you recorded in Procedure 2? What role did the ice play?

3. With ice on the flask, the water will boil for a time and then stop. Why does it stop?
64. Breathalyzer Reaction

**OBJECTIVE:** To observe the reaction used as a test for alcohol in the Breathalyzer.

**CONCEPTS:** Oxidation-reduction reactions

**INTRODUCTION:** The reduction of dichromate ion is used to test for alcohols.

\[
\text{Cr}_2\text{O}_7^{2-} + 3 \text{C}_2\text{H}_5\text{OH} + 8 \text{H}^+ \rightarrow 2 \text{Cr}^{3+} + 3 \text{C}_2\text{H}_4\text{O} + 7 \text{H}_2\text{O}
\]

The dichromate ion is orange. Addition of alcohol produces a green Cr\(^{3+}\) complex. The color change is the basis of the test. Most alcohols will give a positive test. Ethanol is the alcohol used in beverages. (The ethanol used in chemistry labs is denatured by addition of a toxic material, do not drink the alcohol!) Because many other compounds, such as the aldehydes in a diabetic’s breath, also give a positive test, a positive test must be verified by a more specific analysis.

The Breathalyzer uses concentrated sulfuric acid, and the addition of water vapor from the breath generates enough heat to warm the reactants so the reaction proceeds rapidly. Because of the low acid concentration used in this experiment, the mixture must be heated so the reaction will go in a reasonable length of time.

**HAZARDS:** The chromium solution is toxic and corrosive. Avoid ingestion or contact with skin or clothes. The ethanol solution is toxic.

**PROCEDURE:**
1. Put 5 ml of 0.1 M K\(_2\)Cr\(_2\)O\(_7\) in 1 M H\(_2\)SO\(_4\) solution in each of two 25 ml beakers.
2. Add 2 ml of 50% ethanol (C\(_2\)H\(_5\)OH) to the solution in one beaker and 2 mL of water to the other as a control.
3. Put the beakers on a hot plate set to medium heat.
4. Observe the color changes in the solutions. When the solutions come to a boil, remove the beakers from the heat. CAUTION: Don’t burn yourself or spill the solutions.
5. Pour the solutions into the chemical waste container and rinse out the beakers with distilled water.
6. Neldon the nerd was driving home from a late chemistry study session when he was pulled over by a police officer. Neldon knew he hadn’t been driving well because his mind was on his chemistry and not the road. To assure the officer that he hadn’t been drinking, he quickly swished some mouthwash to make his breath nice and fresh for the breathalyzer. To see if this was a good idea, repeat the above procedure, substituting mouthwash for ethanol.
7. Put used solutions in waste containers and rinse out the beaker. Make sure the hot plate is turned off.

**Questions:**

1. What is the oxidation number of each of the underlined atoms in the equation:

\[ \text{Cr}_2\text{O}_7^{2-} + 3 \text{C}_2\text{H}_5\text{OH} + 8 \text{H}^+ \rightarrow 2 \text{Cr}^{3+} + 3 \text{C}_2\text{H}_4\text{O} + 7 \text{H}_2\text{O} \]

2. Which compound is the reducing agent? (Which element is oxidized?)

3. Which compound is the oxidizing agent? (Which element is reduced?)

4. The acetaldehyde, \( \text{C}_2\text{H}_4\text{O} \), produced in the above reaction also reacts with dichromate ion in acid to produce acetic acid, \( \text{C}_2\text{H}_4\text{O}_2 \). Write the net ionic equation for this reaction.

5. Did Neldon pass the test? What other common household materials have the potential of getting Neldon in trouble?

6. What did you learn from this experiment? How would you improve the experiment?
65. Chemical Periodicity

OBJECTIVE: To demonstrate the periodic law.

CONCEPTS: Periodic law, Periodic properties, Group properties

INTRODUCTION: In the 1860's about 60 elements were known. From their physical properties, the elements were classified as metals, nonmetals, or metalloids. Metals have low electrical resistance (high conductivity), have a metallic luster, and are malleable. Non-metals have high resistance, low luster, and are brittle. Metalloids have properties between those of non-metals and metals. Chemists of the time further classified the elements into groups. The elements in a particular group have similar chemical and/or physical properties. This classification led to the discovery of the periodic table and the periodic law. In the periodic table elements are assigned to a period and a group, groups are the columns and the periods are the rows of the table. The periodic law states that the properties of the chemical elements vary with the atomic number in a systematic way. We now understand that the properties depend on the electronic structures of the atoms.

In this exploratory you will observe some physical properties of one group of elements, and determine if the physical properties vary systematically down the group. The group starts with carbon and ends with lead. This group contains one nonmetal, two metalloids and two metals. The higher oxides of this group have the same chemical formula, that is CO₂, SiO₂, GeO₂, SnO₂, and PbO₂.

HAZARDS: Use care when handling chemicals.

PROCEDURE:
1. Samples of these five elements are in bottles and in petri dishes. Shake the bottles and examine the elements to determine their relative hardness and other properties. List similarities and differences.

2. Examine the five vials with the oxides of each element. List the differences and similarities.

3. With the meter provided, measure the resistance between two probes spaced a centimeter apart on the surface of each element. Use the ohm, Ω, setting on the meter. Record the readings in the table below. Because of oxides on the surface, it may be difficult to get consistent readings; take several measurements. Record the smallest value you observe.
4. Write the electronic configurations for carbon, silicon, germanium, tin and lead in the table.

<table>
<thead>
<tr>
<th>Element</th>
<th>Z</th>
<th>Atomic Mass</th>
<th>Electronic Configuration</th>
<th>Density g cm$^{-3}$</th>
<th>Melting Point,°C</th>
<th>Boiling Point,°C</th>
<th>Resistance, ohms</th>
</tr>
</thead>
<tbody>
<tr>
<td>Carbon *</td>
<td>6</td>
<td>12.011</td>
<td>diamond 3.51 graphite 2.25</td>
<td>2.34</td>
<td>&gt;3550</td>
<td>N/A</td>
<td></td>
</tr>
<tr>
<td>Silicon</td>
<td>14</td>
<td>28.086</td>
<td></td>
<td>1410</td>
<td>2355</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Germanium</td>
<td>32</td>
<td>72.61</td>
<td></td>
<td>937</td>
<td>2830</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Tin</td>
<td>50</td>
<td>118.71</td>
<td></td>
<td>232</td>
<td>2260</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Lead</td>
<td>82</td>
<td>207.2</td>
<td></td>
<td>328</td>
<td>1740</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

*Diamond is the form of carbon that has the same structure as silicon and germanium. A diamond may not be available for this experiment. Diamond has a very large electrical resistance whereas the allotrope graphite has a molecular structure that gives low resistance.

**QUESTIONS:**

1. Refer to a periodic table and predict which has the higher density, Te or Se?

2. Refer to a periodic table and predict which has the lower melting point, As or Sb?

3. Aluminum forms an oxide with formula Al$_2$O$_3$. Predict the formula for indium oxide.

4. Why is carbon classified as a non-metal and tin as a metal?

5. Why are germanium and silicon classified as metalloids? What properties do they have which distinguish them as metalloids?

6. For more information about the elements and the periodic table, go to www.webelements.com.
66. Chloride Ion in Gatorade®

<table>
<thead>
<tr>
<th>Name</th>
<th>Course Number</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Student ID</th>
<th>Section Number</th>
<th>Date</th>
<th>Instructor</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**OBJECTIVE:** To demonstrate a precipitation titration.

**CONCEPTS:** Titration, Precipitation

**INTRODUCTION:** Gatorade® is a drink designed to restore electrolytes and other chemicals that are lost from the body by sweating. In this experiment you will perform a simple analysis to determine the concentration of chloride ion in Gatorade®. Silver ion reacts with Cl\(^-\) to form insoluble AgCl. The concentration of Cl\(^-\) in a solution can be determined by measuring the quantity of Ag\(^+\) required to react with the Cl\(^-\) in a known volume of the solution. The end point of the titration is indicated by the first permanent appearance of Ag\(_2\)CrO\(_4\), a red solid. Ag\(_2\)CrO\(_4\) will not precipitate until essentially all the AgCl has precipitated because AgCl is less soluble than Ag\(_2\)CrO\(_4\).

**HAZARDS:** Chromate ion is toxic. If you come in contact with a solution containing chromate ion, wash the affected area thoroughly with soap and water. If you come in contact with a solution containing Ag\(^+\), the affected area will turn black after exposure to sunlight. This black stain is not harmful and will probably disappear after a few days.

**PROCEDURE:**

1. Count about 20 drops of Gatorade® into a test tube. Record the number of drops in the table below.
2. Add an amount of powdered NaHCO\(_3\) about the size of a grain of rice to make the solution slightly alkaline and then add 1 drop of 1.0 M K\(_2\)CrO\(_4\) indicator solution.
3. Add AgNO\(_3\) solution dropwise, stirring between drops, until the first permanent appearance of a red precipitate of Ag\(_2\)CrO\(_4\). Record the number of drops added in the table below.
4. Repeat steps 1-3 until you are confident your results are reproducible.
5. Dispose of used solution in the waste container. Rinse the test tubes with distilled water.
6. Calculate the molar concentration of chloride ion in Gatorade®. [Hint: a drop can be considered a unit of volume just as L (liter) is a unit of volume.] Record your results in the table below.

<table>
<thead>
<tr>
<th>Titration</th>
<th>Vol. Gatorade® (drops)</th>
<th>Vol. AgNO(_3) Solution (drops)</th>
<th>[Cl(^-)] in Gatorade®</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>4</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>5</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
7. Using data from one of your titrations, show how you determined $[\text{Cl}^-]$ in Gatorade®.

8. Assume you are a chemist working for Pepsico (the owner of Gatorade®). You are asked to report $[\text{Cl}^-]$ in Gatorade®. Write the value you would report here: ______________

**QUESTIONS:**

1. Write a balanced equation that shows why adding $\text{NaHCO}_3(s)$ to a solution makes the solution slightly basic.

2. Chromate ion reacts with hydrogen ion to produce dichromate ion according to the equation $2\text{CrO}_4^{2-} + 2\text{H}^+ = \text{Cr}_2\text{O}_7^{2-} + \text{H}_2\text{O}$. Silver dichromate is soluble in water. Would the titration work properly if the addition of $\text{NaHCO}_3$ in Step 2 of the procedure above was omitted? Yes___ No___ Explain why your answer is correct.

3. You must complete Question 2 above before proceeding further. Once you have answered Question 2, devise an experiment to test whether your answer to Question 2 is correct. Report your procedure and results in the space below.

4. Do the results of your experiment support your answer to Question 2? Yes___ No___ If the results of your experiment do support your answer, congratulations; if the results do not support your answer, suggest reasons why.
67. Chromatography of Ions

OBJECTIVE: To demonstrate separation of ions by chromatography

CONCEPTS: Separations, Chromatography, Transition metal ions

INTRODUCTION: Chromatography is a common means for separation and identification of molecules and ions. The sample to be analyzed is placed on a spot at one end of a sheet or column of an immobile phase (paper in this example), then a fluid (the mobile phase) is passed over or through the immobile phase. Because different ions equilibrate differently between the mobile and immobile phases, the ions move at different rates and become separated. Tentative identification of an ion can be made by measuring the rate of movement of the ion.

HAZARDS: Some of the solutions are toxic and/or corrosive. Avoid ingestion or skin contact and wash your hands before leaving the lab.

PROCEDURE:

1. With a capillary tube, place small amounts of 1M solutions of NiSO₄, CuSO₄, CoSO₄, FeSO₄, Fe₂(SO₄)₃ and a mixture of all five salts, evenly spaced in a line about 1 cm from the edge of a 10 x 10 cm rectangle of filter paper. Allow the spots to dry and add a second volume to each spot. Label the spots with a pencil, not a pen.

2. Fold the filter paper in thirds as shown.

3. Place the paper so it stands upright in a beaker in about 0.5 cm depth of HCl/H₂O/acetone solution.

4. Seal the beaker with Al foil.

5. Allow the paper to stand in the solution until the solution has moved two-thirds of the way up the paper.

6. Remove the paper from the beaker and allow it to dry. Replace the Al foil on the beaker.

7. Locate the spots of metal ions and circle them with a pencil.
8. Place the paper in a beaker with an opened vial of concentrated ammonia. Cover the beaker with a watch glass or Al foil to keep the NH\textsubscript{3}(g) contained. The ammonia will complex with some ions and will precipitate hydroxides of others, thus making the spots more visible. Observe and note what happens to the metal ion spots after a few moments. Wait for the color to fully develop, 3-5 minutes.

9. Place the filter paper in the waste container provided. Stopper the vial of concentrated NH\textsubscript{3}, and clean up your work area.

QUESTIONS:

1. Were the cations in the mixture separated?

2. Arrange the ions in the order of the rate of movement, fastest first.

3. Based on the rates of movement, which ion is most strongly absorbed to the paper?

4. What causes the colors of the spots to change when exposed to ammonia vapors?

5. Why can’t a pen be used to mark the paper?

6. What did you learn from this experiment? How would you improve the experiment?
68. Cobalt(II) Complexes

Objectives: To observe complexation of a transition metal ion and Le Chatelier's principle in an equilibrium system.

Concepts: Equilibrium, Le Chatelier's principle, Temperature effects on equilibrium, Complex ions

Introduction: The salts of cobalt have been used in pigments since the middle ages. Not until the late 18th century was the individual element discovered. Cobalt is an important component of strong magnets. Radioactive isotopes of cobalt are used in the sterilization of foods. Cobalt salts are still used in pigments as well as in the coloring of glass, porcelain, pottery, and tiles. This element is a transition metal, which means that it can readily form coordination complexes in which the ion is surrounded by several atoms or molecules (the molecules bonded to the metal ion are generally referred to as ligands). Most complex ions have four or six ligands. The ligands can be neutral molecules such as H₂O or anions such as Cl⁻. Cobalt(II) ion forms complexes with H₂O and Cl⁻ as shown in the equilibrium reaction

\[ \text{CoCl}_4^{2-} + 6\text{H}_2\text{O} \rightleftharpoons \text{Co(H}_2\text{O)}_6^{2+} + 4\text{Cl}^- \]

The origin of the complex ions color is due to the interaction of the charged ligands and the metal ion d-orbital electrons. The color results from the fact that d-orbitals electrons can absorb photons and be promoted to higher energy levels. For example, CoCl₄²⁻ is a deep blue color in solution, because reddish colors were absorbed by the compound.

Equilibrium is also a factor in the color. Addition of Cl⁻ ions shifts the reaction to the left to produce the blue CoCl₄²⁻ complex. Adding water dilutes the Cl⁻ ion and shifts the reaction to the right producing the pink Co(H₂O)₆²⁺ complex. Temperature also affects the position of equilibrium, as heat can cause a shift in equilibrium. In this lab you will observe these effects.

Hazards: Avoid contact with the solutions. Rinse with water if a spill occurs.

Procedure:
1. Add 0.1 M CoCl₂ to a 10x75 mm test tube until it is one-fourth full. Note the color.

2. Add saturated LiCl solution drop-wise to the test tube until the solution in the test tube just turns blue. Agitate the solution between drop-wise additions to make sure the whole solution is of uniform color.

3. Divide the solution into two equal parts, using a second test tube. Place the second solution aside for comparison in steps 4 and 5.
4. Test the solution in both the ice/water mixture and hot water. Dip the test tube into each of the water baths, and observe for a minute or two. Does temperature cause a change in the color of the solution?

5. Next, add water to the solution, mixing frequently. Is there any change in the color of this solution as water is added?

6. Dispose of the solutions in the waste container provided and rinse the test tubes with distilled water when you are finished.

QUESTIONS:

1. Using your observations from above, decide whether the equilibrium we studied here is: (check one) exothermic___or endothermic___. Rewrite the reaction below and include heat as one of the species in the reaction.

2. Water is the solvent, but the addition of the water causes a shift in the equilibrium. How can the addition of water, cause a shift in equilibrium? [Hint: Write out the equilibrium constant for the reaction on the front page (you may to use your text as a resource). How will the addition of water affect variables in the equilibrium constant expression?]
OBJECTIVE: To observe the color components of white light and the effect of color filters on reflected and transmitted light.

CONCEPTS: Visible spectrum, Color

INTRODUCTION: White light contains the colors of the rainbow. These colors are seen when light is passed through a prism. When white light is passed through a colored filter, part of the visible spectrum is absorbed by the filter and the eye perceives the non-absorbed parts of the spectrum as the complementary color. Color is one of the most easily observed and most distinguishing characteristics of substances.

HAZARDS: Do not look directly into the light source, the bright light could damage your eyes.

PROCEDURE:
1. Align a slide projector (the light source) with a glass prism so the light beam is separated into a visible spectrum (rainbow) and projected onto a white screen. In the projector are five slides masked with tape to produce a thin vertical slit. Slides 2 through 5 also have a colored filter covering half the slit. Record your observations below and in the table as you go through the slides.

2. Place slide 1 in the light path and write the colors that you can distinguish in the order that they appear.
   Which color is bent the most in passing through the prism?
   Which color is bent the least?
   Is the angle of “bending” related to the wavelength?

3. Place slide 2 (containing a blue filter) in the light path. What colors are transmitted through the blue filter?

4. Place slide 3 (containing a red filter) in the light path. What colors are transmitted through the red filter?

5. Place slide 4 (containing a yellow filter) in the light path. What colors are transmitted through the yellow filter?

6. Place slide 5 (containing both blue and yellow filters) in the light path. What colors are transmitted through both filters?

7. Place the bottle of potassium permanganate (KMnO₄) solution provided in the light path. What colors are transmitted through the KMnO₄ solution?
<table>
<thead>
<tr>
<th>Conditions</th>
<th>Colors Transmitted</th>
<th>Colors Absorbed</th>
</tr>
</thead>
<tbody>
<tr>
<td>White light</td>
<td></td>
<td></td>
</tr>
<tr>
<td>White light with blue filter</td>
<td></td>
<td></td>
</tr>
<tr>
<td>White light with red filter</td>
<td></td>
<td></td>
</tr>
<tr>
<td>White light with yellow filter</td>
<td></td>
<td></td>
</tr>
<tr>
<td>White light with blue and yellow filters</td>
<td></td>
<td></td>
</tr>
<tr>
<td>White light with permanganate solution</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**QUESTIONS:**

1. What color would a red sweater appear under blue light?

2. What color would a blue sweater appear under the same blue light?

3. Explain in your own words why you saw what you saw in step #6 above.

4. How is color related to wavelength?

5. Sketch the absorption spectrum for potassium permanganate.

6. What did you learn from this experiment? How would you improve the experiment?
70. Common Ion Effect

**Objective:** To observe the common ion effect on solubility equilibria.

**Concepts:** Le Chatelier’s principle, Common ion effect, Equilibrium, Solubility, Acids and bases

**Introduction:** The common ion effect is an example of le Chatelier’s principle which states that a system at equilibrium, subjected to a disturbance or stress, adjusts to reduce the effect of the disturbance. The common ion effect describes what happens when an ion common to both the solution and solid phase is added to a system at equilibrium. For example, Mg(OH)₂(s) in equilibrium with Mg²⁺ and OH⁻ is described by the balanced equation Mg(OH)₂(s) = Mg²⁺ + 2OH⁻. If Mg²⁺ is added to this equilibrium system, reaction will proceed to the left to decrease [Mg²⁺]. As can be seen from the balanced equation, this also causes a decrease in [OH⁻]. This decrease in [OH⁻] disturbs the equilibrium H₂O = H⁺ + OH⁻, causing reaction to the right to increase [OH⁻] and also [H⁺]. Thus addition of Mg²⁺ ends up causing an increase in [H⁺]. An increase in [H⁺] results in a decrease in pH, so the consequence of adding Mg²⁺ can be detected by monitoring the pH of the system. The pH indicator thymol blue will be used to detect changes in pH. Thymol blue is yellow at pH 8.0 and below, blue at pH 9.6 and above, and a mixture of yellow and blue in the range 8.0 to 9.6. In this experiment you will deal with the system described above and also with a system consisting of Li₂CO₃(s) in equilibrium with a solution of its constituent ions, Li⁺ and CO₃²⁻.

**Hazards:** Avoid contact with and ingestion of the solutions. In case of contact, wash with water.

**Procedure:**

1. Half fill two wells of a 24-well plate with a saturated solution of Mg(OH)₂.
2. Add 2-3 drops thymol blue indicator to each of the wells. Stir the solutions and record the color of the solutions in the table below.
3. Add distilled water to one of the solutions until the well is full. Stir. Record the color.
4. Add 5 M MgCl₂ to the other solution until the well is full. Stir. Record the color.
5. Half fill two other wells with a saturated solution of Li₂CO₃.
6. Add 2-3 drops thymol blue indicator to each of the wells. Stir the solutions and record the color.
7. Add distilled water to one of the two Li₂CO₃ solutions until the well is full. Stir. Record any changes.
8. Add 10 M LiCl to the second solution until the well is full. Stir. Record any changes.

<table>
<thead>
<tr>
<th>Solution</th>
<th>saturated Mg(OH)₂</th>
<th>Mg(OH)₂ and water</th>
<th>Mg(OH)₂ and 5 M MgCl₂</th>
<th>saturated Li₂CO₃</th>
<th>Li₂CO₃ and water</th>
<th>Li₂CO₃ and 10 M LiCl</th>
</tr>
</thead>
<tbody>
<tr>
<td>Observations or color</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

8. Dispose of your solutions in the waste container. Rinse the 24-well plate with distilled water.
Questions:

1. What is the color of thymol blue in an acidic solution ________, a neutral solution ________, a basic solution with pH > 9.6 ________?

2. Since Mg(OH)₂ contains hydroxide ion it seems obvious that a saturated solution of Mg(OH)₂ will be basic. Does the color of the thymol blue indicator in the saturated Mg(OH)₂ solution confirm this conclusion? Yes___ No___

3. According to statements in the introduction, the pH of a saturated Mg(OH)₂ solution should decrease when 5 M MgCl₂ is added. Did you observe anything that confirms this claim when you added 5 M MgCl₂ to the saturated Mg(OH)₂ solution in the 24-well plate? Yes___ No___ If you answered “Yes”, tell what you observed and how it confirms the claim made in the introduction. If you answered “No”, suggest at least one reason why the claim was not confirmed.

4. According to the color of the thymol blue indicator that you observed in the saturated Li₂CO₃ solution, a saturated Li₂CO₃ solution is acidic (or neutral)___ basic___. If you answered “acidic (or neutral),” explain whether the solution is more likely acidic or more likely neutral. If you answered “basic,” explain, using a relevant balanced chemical equation, why a saturated Li₂CO₃ solution is basic.

5. According to your observations recorded in the table, was there a shift in equilibrium when water was added to the saturated Li₂CO₃ solution? Yes___ No___ If you answered “Yes,” explain how your observations justify a “Yes” answer and explain the shift in terms of le Chatelier’s principle.

6. According to your observations recorded in the table, was there a shift in equilibrium when 10 M LiCl was added to the saturated Li₂CO₃ solution? Yes___ No___ If you answered “Yes,” explain how your observations justify a “Yes” answer and explain the shift in terms of le Chatelier’s principle.
OBJECTIVE: To observe the effects of complexation and acid-base reactions on solubility of ionic compounds.

CONCEPTS: Acids and bases, Complex ions, Equilibrium, Solubility

INTRODUCTION: Ionic compounds dissolve in water to give the hydrated ions. For example:

\[
\text{AgCl(s)} \rightleftharpoons \text{Ag}^+(aq) + \text{Cl}^-(aq)
\]

At equilibrium, the solubility is described by the solubility product equation:

\[K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.6 \times 10^{-10} \text{ at 25°C}\]

where the brackets indicate molar concentrations. Thus, if another reaction occurs that decreases [Ag⁺], more AgCl(s) will dissolve to increase [Cl⁻]. In this experiment [Ag⁺] is decreased by reaction with NH₃ to form a complex ion, Ag(NH₃)₂⁺.

If the anion in the solid is the anion of a weak acid, solubility can be increased by acidifying the solution. For example:

\[
\text{CaCO}_3(s) \rightleftharpoons \text{Ca}^{2+}(aq) + \text{CO}_3^{2-}(aq) \quad K_{sp} = [\text{Ca}^{2+}(aq)][\text{CO}_3^{2-}(aq)].
\]

Addition of acid decreases the [CO₃²⁻(aq)] by formation of HCO₃⁻, i.e.

\[
\text{CO}_3^{2-}(aq) + \text{H}^+(aq) \rightleftharpoons \text{HCO}_3^-(aq).
\]

HAZARDS: Avoid contact with the reagents. In case of a spill, wash with water. Contact of AgNO₃ solution with skin will cause a temporary black stain, but no permanent harm.

PROCEDURE: After each addition of a new reagent, record your observation and complete or balance the chemical equation for the reaction that you observed.

1. Add 10 drops of 0.01 M NaCl to a test tube.

2. Add an equal volume of 0.01 M AgNO₃ to the tube.
   Complete and balance the precipitation equation: \( \text{Ag}^+ + \text{Cl}^- \rightarrow \)

3. Add 1 M NH₃ dropwise, gently shaking the test tube between drops to mix the solution, until the AgCl just dissolves. Balance the complexation reaction:

\[
\text{AgCl(s)} + \text{NH}_3(aq) \rightarrow \text{Ag(NH}_3)_2^+(aq) + \text{Cl}^-(aq)
\]

4. Add 1 M HNO₃ to the tube with shaking until the AgCl precipitate reappears.
   Complete and balance the precipitation equation: \( \text{Ag(NH}_3)_2^+ + 2\text{H}^+ + \text{Cl}^- \rightarrow \)

5. Add 1 M NaOH to the tube with shaking until the precipitate redissolves.
Complete the dissolution equation: \( \text{AgCl}(s) + 2 \text{NH}_4^+ + 2 \text{OH}^- \rightarrow \)

6. Add 10 drops of 0.01 M KBr to the tube. 
Complete the precipitation equation: \( \text{Ag}(\text{NH}_3)_2^+ + \text{Br}^- \rightarrow \)

7. Add 1 M Na$_2$S$_2$O$_3$ dropwise until the precipitate just dissolves. 
Complete the dissolution equation: \( \text{AgBr}(s) + 2 \text{S}_2\text{O}_3^{2-} \rightarrow \)

8. Add 10 drops of 0.01 M KI to the tube. 
Complete the precipitation equation: \( \text{Ag}(\text{S}_2\text{O}_3)_2^{3+} + \text{I}^- \rightarrow \)

9. Place a small chip of marble, CaCO$_3$ in an empty test tube.

10. Add 20 drops of water. Do you observe any reaction?

11. Add 20 drops of 2 M HCl. What sign of a reaction do you observe? 
Complete the dissolution equation: \( \text{CaCO}_3(s) + 2 \text{H}^+ \rightarrow \)
And the gas-forming equation: \( \text{H}_2\text{CO}_3 \rightarrow \)


**QUESTIONS:**

1. Which is more soluble, AgCl or AgBr? How do you know?

2. Which is more soluble, AgBr or AgI? How do you know?

3. Which complexes Ag$^+$ more strongly, NH$_3$ or S$_2$O$_3^{2-}$? How do you know?

4. Explain why the marble sculptures and buildings in industrialized parts of the world are rapidly (compared to pre-twentieth century rates) disintegrating. Where does the acid come from and what acids are the culprits?

5. What did you learn from this experiment? How would you improve the experiment?
OBJECTIVE: To observe phase changes.

CONCEPTS: Pressure - temperature - volume behavior of materials, Phase diagrams, Critical temperature

INTRODUCTION: The phase diagram of CO₂ is given below.

CO₂ sublimes at -78°C and 1 atm. CO₂ thus vaporizes from the solid phase directly to the gas phase without passing through a liquid phase. But if the pressure of the CO₂ is above 5 atm, CO₂ melts from the solid phase to the liquid phase. The critical temperature for CO₂ is 31°C, above this temperature there is no distinction between the liquid and gas phase, there is only a supercritical fluid. In the supercritical state, CO₂ cannot be liquefied or solidified by applying pressure. If solid CO₂ is placed in a sealed container and allowed to warm, conditions inside the container are described by the solid-gas (at T < -57°C) and liquid-gas (at T > -57°C) phase boundaries.

HAZARDS: Handle dry ice with tongs. Dry ice is extremely cold and can cause frostbite. Hold the pipet bulb under water at all times to protect yourself and others in case it ruptures.

PROCEDURE:
1. Wrap a small piece of dry ice about the size of a large pea in a sheet of paper and crush the solid into a fine powder with a hammer.

2. Obtain a Beral pipet with the tip cut off so that 1 to 2 cm of stem remain. Place the crushed dry ice into the pipet bulb. The dry ice can be forced into the pipet by pressing the open end into the finely crushed solid.

3. Hold the open end of the pipet with the pliers provided. Clamp the open end tight enough to seal the end, but not hard enough to cut off the end of the pipet. Submerge the bulb into a large
beaker of cool water.

4. Watch the solid in the pipet. After a minute or two the solid will melt and then start to boil. At this point, release the gas in the pipet by releasing the pressure on the pliers. The liquid will solidify.

**QUESTIONS:**

1. Was the pressure inside the bulb greater than 5 atm when you saw the liquid?

2. Write reactions describing the phase changes in step 4 with the water below 30°C.

   $$\text{CO}_2(s) \rightarrow$$

3. Write reactions describing the phase changes that would occur in step 4 if the water was above 40°C.

   $$\text{CO}_2 (s) \rightarrow$$

4. Assuming you had 44 milligrams of CO₂ at 25°C and that the pipet bulb is 2 mL, estimate the pressure in the bulb with the ideal gas law. According to the phase diagram, what phases should be present under these conditions?

5. Why did the CO₂ solidify when you released the pressure?

6. If you shake a CO₂ fire extinguisher at 25°C, you hear or feel liquid sloshing in the container. Would you feel any liquid sloshing if the temperature was 35°C? Why or why not?

7. What did you learn from this experiment? How would you improve the experiment?
OBJECTIVE: To demonstrate a model for the dynamic equilibria that occur in chemical reactions.

CONCEPTS: Chemical equilibrium

INTRODUCTION: Many chemical reactions do not go to completion. Instead they reach an equilibrium state in which reaction appears to stop. In this state, however, reactions in both directions are still occurring at the molecular level, but the rates of the forward and the reverse reactions are equal. Equilibrium can be attained from either direction, that is, from only reactants or from only products.

As an example, there are many natural organic acids that are weak, that is, they dissociate only partly in water. A common example is acetic acid, found in vinegar. The reaction of acetic acid with water

\[ \text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{C}_2\text{H}_3\text{O}_2^- \]

produces an equilibrium mixture of \( \text{HC}_2\text{H}_3\text{O}_2, \text{H}_2\text{O}, \text{H}_3\text{O}^+, \text{C}_2\text{H}_3\text{O}_2^- \) when the forward reaction rate to produce \( \text{H}_3\text{O}^+ \) and \( \text{C}_2\text{H}_3\text{O}_2^- \) equals the reverse reaction rate to produce \( \text{HC}_2\text{H}_3\text{O}_2 \) and \( \text{H}_2\text{O} \). This equilibrium is dynamic because on the molecular level the reactions continue even though the concentrations of all species remain constant.

In the following demonstration, the amounts of water in the two vials represent the concentrations of reactants and products of a chemical reaction. As water is transferred cycle by cycle from each vial to the other, the reaction approaches equilibrium and the level of the water in each vial remains constant. At this point, the amounts of water transferred in each direction per cycle have become equal.

HAZARDS: Don’t put the straws in your mouth.

PROCEDURE:

1. Begin with two equal-size, flat-bottom vials, one nearly filled with water, the other empty.

2. Select two straws of unequal diameter and place one straw in each vial.

3. Take a straw in each hand, hold the straws against the bottoms of the vials so they fill to the depths in the vials, and seal the straws with fingers over the ends to retain the water inside each straw.

4. Transfer the water in the straws to the opposite vial by lifting the straws and crossing your wrists.

5. Release the liquid and return the empty straws to their original vials.

6. Continue to transfer water until the water levels in the vials stop changing. Measure the height of water in each vial.

   Height of water in left vial _______________

   Height of water in right vial _______________
7. Switch the straws in your hands and repeat the transfer process using the vials just as they were left at the end of step 6 until a new equilibrium position is reached.

Height of water in left vial ________________

Height of water in right vial ________________

8. Leaving the straws and water as they are, add a drop of blue food coloring to one tube and a drop of yellow food coloring to the other. Continue the transfers and note what happens to the colors of the tubes.

Initial color in left vial ____________, right vial ____________.

Final color in left vial ____________, right vial ____________.

Did the height of the water in the vials change? Yes____ No____

9. Clean up the area before you leave.

QUESTIONS:
1. The rate of transfer obviously depends on the diameter of the straw. What else does the rate of transfer depend on?

2. What can you say about the forward and reverse rates when the water levels stop changing?

3. At equilibrium the height of the water is always greater in the vial from which water is being removed with the smaller diameter straw. Explain why.

4. What does the dye experiment prove.
74. How Strong Is Your Bleach?

OBJECTIVE: To demonstrate that the amount of heat produced is proportional to the amount of reaction.

CONCEPTS: Heat of reaction, Heat capacity, Concentration

INTRODUCTION: Heat evolution or absorption accompanies chemical reactions. The amount of heat produced is proportional to the amount of reaction. The heat produced per mole of reaction at constant pressure is called the enthalpy change, $\Delta H$. Measurement of the heat of reaction can be used as a general method for determination of the amount of reaction. In this experiment the concentration of hypochlorite ion, $\text{ClO}^\text{-1}$, in a commercial chlorine bleach is determined by measuring the amount of heat produced when an excess of acetone, $\text{C}_3\text{H}_6\text{O}$, is added to the bleach solution.

$$3\text{ClO}^\text{-1} + \text{C}_3\text{H}_6\text{O}(\ell) = \text{CHCl}_3(\ell) + \text{CH}_3\text{COO}^\text{-1} + 2\text{OH}^\text{-1} \quad \Delta H = -519 \text{ kJ mol}^{-1}$$

HAZARDS: Acetone is volatile and flammable, keep the container closed and away from flames. Avoid eye or skin contact with the bleach solution.

PROCEDURE:
1. The calorimeter body consists of a 6-dram vial in a foam insulator.
2. Measure 10.0 mL of bleach solution into the 6-dram vial.
3. Use the stir bar and magnetic stir plate to begin stirring the bleach solution.
   Measure and record the temperature of the bleach solution. _____________
4. Add 1 mL acetone and observe the temperature as it increases.
   Record the maximum temperature reached. _____________
5. Show your work in the space provided for each of the calculations that follow. Circle your answers.
6. For the solution and calorimeter combined, $C = 57 \text{ J °C}^{-1}$. Calculate $q$, the amount of heat produced.
7. From the thermochemical equation in the introduction, note that 3 moles $\text{ClO}^\text{-1}$ produce 519 kJ. Use this information and the value for $q$ above to calculate the amount of $\text{ClO}^\text{-1}$ that reacted in your experiment.
8. Calculate the concentration of ClO$^–$ and the weight percent NaClO (74.4 g mol$^{-1}$) in the bleach. The density of the bleach solution is 1.030 g mL$^{-1}$.

9. Dispose of the solution in the waste container provided. Be careful that you don’t drop the stir bar into the waste container. Rinse the glassware, thermometer, and stir bar with distilled water.

**QUESTIONS:**

1. The correct wt% NaClO is given on the bleach container. Calculate the percent error in your determination of wt% NaClO in bleach.

2. In your opinion, what is the largest potential source of error in the experiment?

3. The heat of reaction results from some bonds being broken and new bonds being formed. Calculate $\Delta H$ for the reaction in the introduction from the following bond energy data: C-C, 331 kJ mol$^{-1}$; C-H, 414 kJ mol$^{-1}$; C-Cl, 325 kJ mol$^{-1}$; C-O, 326 kJ mol$^{-1}$; Cl-O, 205 kJ mol$^{-1}$ and O-H, 464 kJ mol$^{-1}$.

4. Calculate the percent error in your result from item 3 above.

5. What is the main source of error in your result from item 3 above? [Hint: What are the conditions for which bond-energy data are valid?]
75. Hydrogen and Oxygen Chemistry

OBJECTIVE: To observe several compounds of hydrogen and oxygen and some reactions for producing hydrogen gas and oxygen gas.

CONCEPTS: Hydrogen, Oxygen

INTRODUCTION: Hydrogen and oxygen are two of the most common elements on earth. They both occur in water, the most abundant compound on earth. Oxygen also makes up about 21% of the air and occurs in oxides of most other elements and in common oxyanions such as carbonate ion and silicate ion. Hydrogen occurs mostly in compounds. Most organic compounds contain hydrogen.

HAZARDS: Some of the reactions are very exothermic so exercise caution not to get burned. Some of the reagents are corrosive, so don’t let them contact your skin or clothes. Wash with water if contact occurs. Hydrogen peroxide will cause a temporary bleaching of the skin and a permanent bleaching of clothing.

PROCEDURES:
1. Pour 5 mL of 6 M NaOH into the first well of a six well plate. Add 2 or 3 small pieces of aluminum foil to the well. Stir. What evidence of a chemical reaction do you observe?

2. Pour 5 mL of 6 M HCl into the second well. Add one piece of mossy zinc to the well. Gently swirl the plate. What evidence of a chemical reaction do you observe?

3. Pour 5 mL of water into the third well. With tweezers or a spatula, add a piece of CaH₂ about the size of half a grain of rice. What evidence of a chemical reaction do you observe?

4. Pour 5 mL of 10% H₂O₂ into the fourth well. Add an amount of MnO₂ about the size of one-fourth of a grain of rice. What evidence of a chemical reaction do you observe?

5. Pour 5 mL of 1 M Na₂SO₄ into the fifth well. Add two drops of universal acid-base indicator. Insert electrodes (pencils connected to a 9 volt battery). Observe the production of H₂ and O₂. Observe color changes in the vicinity of the electrodes.

Color and pH relationships for universal indicator are given in the following chart.

<table>
<thead>
<tr>
<th>color</th>
<th>red-orange</th>
<th>orange</th>
<th>yellow</th>
<th>green</th>
<th>purple</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH</td>
<td>2</td>
<td>4</td>
<td>6</td>
<td>8</td>
<td>10</td>
</tr>
</tbody>
</table>

6. Clean up: Put the zinc in the zinc waste container, everything else in the main waste container.
**Questions:**

1. Write a balanced equation for the reaction that occurred in Procedure 1.

2. Write a balanced equation for the reaction that occurred in Procedure 2.

3. What is the oxidation state of hydrogen in CaH₂? _____
   Write a balanced equation for the reaction that occurred in Procedure 3.

4. H₂O₂ reduces MnO₂ to MnO. Write a balanced equation for this reaction.

5. H₂O₂ oxidizes MnO to MnO₂. Write a balanced equation for this reaction.

6. Write a balanced equation for the overall reaction that occurred in Procedure 4.

7. In Procedure 4 MnO₂ serves as a _____________.

8. The following questions apply to Procedure 5.
   Write a balanced equation for the half-reaction that occurs at the anode.

   Write a balanced equation for the half-reaction that occurs at the cathode.

   Write a balanced equation for the overall reaction.

Which terminal of the battery is the anode? positive terminal_____ negative terminal_____ Explain how the colors produced around the electrodes reveal which is the anode and which is the cathode.
76. It’s A Colorful World

OBJECTIVE: To demonstrate that many naturally occurring colored substances in plants are weak acids that change color on ionization.

CONCEPTS: Acids and bases, Indicators, pH

INTRODUCTION: Many pigments in plants are weak acids. These colored compounds are natural pH indicators, changing color when they are converted to the conjugate base. The juice or extract from colored grapes, blackberries, cherries, beets, red cabbage and many other colored fruits and vegetables all contain such pigments. Some of these natural indicators exhibit several color changes because they contain several different acids, or because they contain polyprotic acids with several pK values.

HAZARDS: Avoid contact with solutions. Rinse with water if a spill occurs.

PROCEDURES:
1. Pick a substance from those provided. Place several drops in each of 5 wells in a 24-well plate. Dilute each sample with 5 drops of water. You should be able to see through the solutions when they are placed over a white paper.
2. To well #1 add 2 drops of 6 M HCl solution.
   To well #2 add 2 drops of 1 M acetic acid.
   To well #4 add 2 drops of 1 M NH₃ solution.
   To well #5 add 2 drops of 6 M NaOH solution.
   Record the colors in the table below.
3. Repeat and record results for two other substances.
4. Dispose of solution in the waste container when you are finished. Clean up your work area.

<table>
<thead>
<tr>
<th>Substance</th>
<th>6 M HCl</th>
<th>1 M Acetic Acid</th>
<th>Natural Color</th>
<th>1 M NH₃</th>
<th>6 M NaOH</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Name ___________________________ Course Number ___________________________
Student ID ______________________ Section Number ___________________________
Date ___________________________ Instructor _________________________________
**Questions:**

1. Are most fruits and vegetables acidic or basic? How do you know from what was done in this experiment?

2. What would “red” cabbage be called if it was basic?

   What would “purple” grapes be called if the fruit were basic?

3. Why should florists be concerned about the pH of the soil in which certain flowers are grown? (Hint: Some species of hydrangea can have either pink or blue blossoms.)

4. Why is the pH of the solution equal to the $pK_a$ of the indicator when the indicator is half in the acid form and half in the basic form? (Hint: write the equation for the ionization constant of a weak acid.)

5. What did you learn from this experiment? How would you improve the experiment?
**77. Metal Alloy Analysis**

<table>
<thead>
<tr>
<th>Metal</th>
<th>Electrolyte</th>
<th>Reagent</th>
<th>Positive Test</th>
</tr>
</thead>
<tbody>
<tr>
<td>Aluminum</td>
<td>1M HCl</td>
<td>10g/l Alizarin Yellow</td>
<td>Yellow/Orange ppt.*</td>
</tr>
<tr>
<td>Chromium</td>
<td>1M H₂SO₄</td>
<td>0.02M Dithizone in isopropyl alcohol</td>
<td>Yellow complex</td>
</tr>
<tr>
<td>Copper</td>
<td>0.2M K₄Fe(CN)₆</td>
<td>0.2M K₄Fe(CN)₆ Yellow/Green complex</td>
<td></td>
</tr>
<tr>
<td>Gold</td>
<td>1M H₂SO₄</td>
<td>0.005M Rhodamine B</td>
<td>Violet ppt.</td>
</tr>
<tr>
<td>Iron</td>
<td>1M HCl</td>
<td>0.5M NH₄SCN</td>
<td>Red complex</td>
</tr>
<tr>
<td>Nickel</td>
<td>1M NH₃</td>
<td>0.2M Dimethylglyoxime</td>
<td>Pink ppt.</td>
</tr>
<tr>
<td>Silver</td>
<td>0.2M K₂CrO₄</td>
<td>0.2M K₂CrO₄</td>
<td>Red ppt.</td>
</tr>
</tbody>
</table>

*ppt. = precipitate*

**PROCEDURE:** Place a sheet of aluminum (3 cm x 3 cm) in a plastic dish. Place a piece of filter paper on the aluminum. Place one drop of electrolyte solution and one drop of reagent solution or one drop of a combined electrolyte-reagent solution on the filter paper. Place the metal to be tested over the reagent spot on the paper. Do not allow the test metal to touch the aluminum sheet. Attach the positive lead from a 9V battery to the test metal piece. Attach the other lead to the Al sheet. Press the assembly together for 30 seconds to a minute and then observe the color of the spot on the filter paper. Repeat with the next electrolyte and reagent. Test several known metals first to become familiar with positive and negative results, then test two or three coins.
RESULTS:

<table>
<thead>
<tr>
<th>U.S. coins</th>
<th>composition</th>
</tr>
</thead>
<tbody>
<tr>
<td>1¢ post-1982</td>
<td></td>
</tr>
<tr>
<td>1¢ pre-1982</td>
<td></td>
</tr>
<tr>
<td>5¢</td>
<td></td>
</tr>
<tr>
<td>10¢</td>
<td></td>
</tr>
<tr>
<td>25¢</td>
<td></td>
</tr>
<tr>
<td>50¢</td>
<td></td>
</tr>
<tr>
<td>$1</td>
<td></td>
</tr>
<tr>
<td>other</td>
<td></td>
</tr>
</tbody>
</table>

QUESTIONS:

1. In each electrolysis, hydrogen is produced at the aluminum cathode. Complete the equation.

   \[ 2H^+ + 2e^- \rightarrow \text{(in acid)} \]

   or \[ 2H_2O + 2e^- \rightarrow \text{(neutral or base)} \]

2. Write the equation for the reaction at the anode during electrolysis of Cu.

   \[ \text{Cu} \rightarrow \]

3. What are the characteristics of a useful reaction for a spot test?

4. What did you learn from this experiment? How would you improve the experiment?
78. Molar Volume of Carbon Dioxide Gas

OBJECTIVE: To demonstrate that one mole of a gas occupies about 24 liters at room temperature and pressure.

CONCEPTS: Molar volume, Molecular mass, Gas properties

INTRODUCTION: All gases occupy very close to 22.4 L per mole of gas at 0°C and 1 atm of pressure. The volume occupied by a mole of gas is almost independent of the molecular mass of the gas, but depends strongly on the temperature and pressure. In comparison, the volume occupied by one mole of a solid or liquid is only about 0.1% as large, depends mostly on the size of the molecules, and is only slightly dependent on temperature and pressure. In this experiment, you will have the opportunity to observe the change in volume as a solid converts to a gas.

HAZARDS: Dry ice can cause severe frostbite (similar to tissue damage from a burn). Avoid direct contact with your skin.

PROCEDURE:
1. Weigh a small dry Ziploc® bag. Add a piece of dry ice (solid CO₂) about the size of a pea and weigh again. Quickly flatten the bag to remove as much air as possible, and seal the bag. Record the weight of the dry ice. (Do not use a larger piece of dry ice. The volume of the bag may be exceeded. You need approximately 0.5 g of dry ice.)
   ______________________ grams, empty bag
   ______________________ grams, bag and dry ice
   ______________________ grams of dry ice

2. While the dry ice sublimes, half-fill a 1 L graduated beaker with water. Record the volume of water in the beaker by reading the graduated marks.

3. After the dry ice has completely sublimed, submerge the inflated bag in the water using a potato masher, a screen cut to fit into the beaker, or similar tool and again record the volume of the water in the beaker.
   ______________________ volume of water displaced by the inflated bag

4. Calculate the number of moles of carbon dioxide.

   ( ____________ g of dry ice)/(44 g mol⁻¹) = ________________ mole CO₂
5. Calculate the molar volume of CO₂(g).

\[
\frac{\text{L of CO}_2}{\text{mole of CO}_2} = \text{__________ L mol}^{-1}
\]

6. Clean up your work area.

**QUESTIONS:**

1. Is your answer close to 24 L per mole? If not, why not?

2. In question 1, why was 24 L used instead of the usual textbook number of 22.4 L?

3. Why is the molar volume for all gases approximately equal?

4. Why are the molar volumes of liquid water (H₂O(l), 18 mL mol⁻¹) and gasoline (mostly C₈H₁₈(l), 163 mL mol⁻¹) so different?

5. Estimate the molar volumes of H₂O(g) and C₈H₁₈(g). Why are these the same?

6. What did you learn from this experiment? How would you improve it?
79. Organic Functional Groups

OBJECTIVE: To observe the structures of organic functional groups.

CONCEPTS: Functional groups

INTRODUCTION: The chemistry of organic compounds can be organized according to the functional groups present. Alkanes and benzene generally react only with difficulty or at high temperatures. Functional groups are introduced when hydrogen is removed from an alkane to form double or triple bonds, or when another element is substituted for one or more hydrogens. The reactivity greatly increases and is characteristic of the functional group. Functionalized hydrocarbons also have properties that make them useful plastics, cleaners, solvents, textiles, paints, etc. In this exercise, models of compounds with several functional groups will be constructed.

HAZARDS: Don’t step on the models.

PROCEDURE: Construct or locate models of the suggested molecules.

1. As an example of an alkane, construct a model of propane, CH₃CH₂CH₃.

2. Alkenes and alkynes are unsaturated, they contain less hydrogen than the alkane with the same number of carbons. Construct models of propene, CH₃C=CH, and propyne, H₃CC≡CH.

3. Construct models of an alcohol and a carboxylic acid: ethanol, CH₃CH₂OH, and ethanoic acid, CH₃COH.

4. Construct models of an aldehyde and a ketone: propanal, CH₃CH₂CH, and 2-propanone, CH₃CCH₃.

5. Construct models of ammonia, NH₃, ammonium ion, NH₄⁺, and methylamine, H₃CNH₂.

6. Construct models of the cyclic compounds benzene, C₆H₆, and cyclohexane, C₆H₁₂.

7. Leave the assembled models on the bench top.
**Questions:**
Use the models you have made to help you answer the following questions.

1. Alkanes react very slowly with Br₂, but alkenes and alkynes react very rapidly. Write equations for the reactions of Br₂ with propene, propyne, and cyclohexene.

2. Explain why carboxylic acids RCOH are more acidic than alcohols.

3. Aldehydes (such as propanal) are easily oxidized to carboxylic acids but ketones (such as 2-propanone) are not. Explain. (Hint: carbon-carbon bonds are quite strong.)

4. A lone pair of electrons on an N atom makes the molecule act as a base. An H⁺ attached to an N lone pair makes a molecule acidic. Which of the N-containing molecules you made are basic? Which are acidic? Which are neither?

5. What did you like about this experiment and how would you improve it?
OBJECTIVE: To observe the relation between the presence of unpaired electrons in a compound and the attraction to a magnet.

CONCEPTS: Electron configurations, Magnetism

INTRODUCTION: Diamagnetic materials are very weakly repelled from a magnetic field. Because this repulsion is very weak, very strong magnetic fields and sensitive equipment are required to observe diamagnetism. The molecules, ions, or atoms in diamagnetic materials contain only spin paired electrons. Paramagnetic materials are weakly attracted to a magnetic field because the molecules, ions, or atoms in paramagnetic materials have at least one electron per particle that is not spin paired. The force of the attraction to a magnetic field is proportional to the number of unpaired electrons. Ferromagnetic materials are strongly attracted to a magnetic field because they contain unpaired electrons whose spins are aligned over many atoms.

HAZARDS: Avoid contact with chemicals. Rinse immediately with water if any spills occur.

PROCEDURE:
Stoppered test tubes filled with samples are suspended from a horizontal rod by threads so the samples can swing freely just above the bench top. Estimate the relative strength (on a scale of 1 to 5, 5 being the strongest) of the attraction of each sample for a magnet by sliding the magnet slowly toward the sample.

<table>
<thead>
<tr>
<th>Material</th>
<th>Number of unpaired electrons</th>
<th>Relative attraction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fe filings</td>
<td>4</td>
<td></td>
</tr>
<tr>
<td>NiSO₄</td>
<td>2</td>
<td></td>
</tr>
<tr>
<td>Cr₂(SO₄)₃</td>
<td>3</td>
<td></td>
</tr>
<tr>
<td>FeSO₄</td>
<td>4</td>
<td></td>
</tr>
<tr>
<td>K₄[Fe(CN)₆]</td>
<td>4 or 0</td>
<td></td>
</tr>
<tr>
<td>Fe₂(SO₄)₃</td>
<td>5</td>
<td></td>
</tr>
<tr>
<td>K₃[Fe(CN)₆]</td>
<td>5 or 1</td>
<td></td>
</tr>
<tr>
<td>MnCl₂</td>
<td>5 or 1</td>
<td></td>
</tr>
</tbody>
</table>
**Questions:**

1. Which sample is ferromagnetic? Which of the samples are paramagnetic? Which are diamagnetic?

2. How would you go about showing that O₂ molecules contain unpaired electrons.

3. If you have studied the ligand field theory of bonding in transition metal compounds, you should be able to answer this question: Write the electronic configuration of Fe(II). How could FeSO₄ have a different number of unpaired electrons than K₄[Fe(CN)₆] even though they both contain Fe(II)? Hint: the Fe(II) is in an octahedral ligand field.

4. What did you learn from this experiment? How would you improve it?
OBJECTIVE: To observe the effect of a reaction on the pH of the solution.

CONCEPTS: pH, Oxidation-reduction reactions, Indicators

INTRODUCTION: Many redox reactions change the pH of the solution because H⁺ or OH⁻ is a reactant or product in the reaction. For example, hydrogen peroxide reacts with iodide ion (I⁻) to produce OH⁻.

\[ H_2O_2 + 2I^- \rightarrow 2OH^- + I_2 \]

When thiosulfate ion is also present, it reacts with the iodine (I₂) to produce tetrathionate and regenerate I⁻.

\[ I_2 + 2S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-} \]

Because it is unchanged in the overall process, I⁻ acts as a catalyst in this process.

The pH of the solution can be followed by observing the color of universal pH indicator. Universal indicator is red in acid, blue or purple in base and goes through a rainbow between, with greenish yellow being neutral.

<table>
<thead>
<tr>
<th>color</th>
<th>red-orange</th>
<th>orange</th>
<th>yellow</th>
<th>green</th>
<th>purple</th>
</tr>
</thead>
<tbody>
<tr>
<td>pH</td>
<td>2</td>
<td>4</td>
<td>6</td>
<td>8</td>
<td>10</td>
</tr>
</tbody>
</table>

HAZARDS: Avoid contact with solutions. Rinse with water if a spill occurs.

PROCEDURE:

1. In a 25-mL beaker mix 5 mL of Solution A and 5 mL of 7.5 mM Na₂S₂O₃. (Solution A contains 0.25 M KI, 0.8 mM acetic acid, and universal indicator.)

2. Add 2 drops of 10% hydrogen peroxide.

3. Swirl the beaker or stir to mix the contents and observe the color changes. The changes will be more apparent if the beaker is placed on a sheet of white paper.

4. Dispose of the solution in the waste container provided and rinse the beaker with distilled water when you are finished.
Questions:
1. What is the overall reaction of hydrogen peroxide with iodide and thiosulfate ions?

2. Define “catalyst.”

3. What color changes did you observe in the solution? What was the pH at the start of the reaction? What was the pH at the end of the reaction?

4. In your own words, explain why the pH of the solution increased during the reaction.

5. If 30% peroxide was used instead of 10%, what would you predict would happen to the rate of reaction?

6. If a reaction started with purple and ended with red colored indicator, what could be concluded about the reaction?

7. What did you learn from this experiment? How would you improve the experiment?
82. Photoelectric Effect

OBJECTIVE: To demonstrate the photoelectric effect.

CONCEPTS: Ionization energy, Photoelectric effects, Quantization

INTRODUCTION: The nature of the photoelectric effect proves that light energy is quantized. The ability of light to excite electrons enough for the electron to leave a metal depends on the wavelength of the light but is independent of the intensity of the light. The kinetic energy of the electron ejected from the metal by a photon of light is \( E = h\nu - w \) where \( h\nu \) is the energy of the photon and \( w \) is the work required to just remove an electron from the surface of the metal. The work function, \( w \), is related to the ionization energy for the metal which is the energy required to just remove an electron from an isolated atom. The value of \( w \) depends on the metal, e.g. Zn and Cu are different. Albert Einstein received a Nobel prize for showing that this phenomenon could be explained by assuming light exists as quantized particles called photons.

An electroscope is a simple instrument that can detect the presence of only a few thousand excess electrons.

HAZARDS: Do not look directly at the UV lamp. Ultraviolet light can rapidly damage the eyes.

PROCEDURE:
1. Add electrons to the surface of a rubber rod by rubbing it with dry animal fur.

2. Transfer the electrons from the rod to the electroscope by passing the rod along the wire on top of the electroscope. The aluminum foil strips inside the electroscope should be moved apart by electrostatic repulsion of the extra electrons now on the strips.

3. Discharge the electroscope by touching the wire with your finger.

4. Hang a small zinc strip on the wire on top of the electroscope. The zinc strip should first be sanded with fine sandpaper to remove the oxide coating that forms when zinc is exposed to air.

5. Charge the electroscope with the rubber rod.

6. Shine the light from a reading lamp on the zinc strip, and note what happens. Recharge the electroscope if necessary.

7. Shine the long wavelength UV light on the zinc strip and note what happens. Recharge the electroscope if necessary.
8. Shine the short wavelength UV light on the zinc strip and note what happens. (Note: If the photons are more energetic than $w_{Zn}$, electrons will be ejected from the zinc into the air, and the electroscope will then discharge electrons into the zinc.)

9. Replace the zinc strip with a copper strip and repeat the experiment.

**QUESTIONS:**
1. Arrange the light sources, i.e. visible, long wavelength UV, short wavelength UV, in order of increasing energy of the photons.

2. According to your observations, which has the larger work function ($w$), Zn or Cu? Explain.

3. Of the metals Cs and Au, which would you expect to have the larger work function? (Hint: Consider the standard reduction potentials.)

4. How would the brightness (i.e. photons per second) of the light affect the results?

5. What commercial devices can you name that make use of the photoelectric effect?

6. What did you learn from this experiment? How would you improve the experiment?
83. Polarity of Molecules

**OBJECTIVE:** To demonstrate that molecules of some materials are dipolar.

**CONCEPTS:** Bond polarity, Molecular shapes, Electrostatic interactions

**INTRODUCTION:** The molecules in many materials behave like electric dipoles, i.e. the charges in the molecule are unevenly distributed. To have a dipole moment, a molecule must have polar bonds and must have the bonds arranged so the bond dipoles do not all cancel when added as vectors. Because the dipoles in the liquid can be aligned by an electric field, a liquid composed of polar molecules is attracted to a charged object, whether the object is positively or negatively charged.

**HAZARDS:** Isopropyl alcohol is toxic if ingested and can harm the eyes.

**PROCEDURE:**

1. Hang a small ball of aluminum foil or a pith ball from a ringstand with a cotton sewing thread.

2. Place a static charge on a glass rod by rubbing it with a sheet of polyethylene plastic. What happens when you bring the glass rod near the ball of aluminum foil? What happens when you touch the ball?

3. Suspend a polyethylene rod from a thread. Use a wire cradle to hold the rod in a horizontal position so you can observe rotation of the rod. Rub the polyethylene rod with a piece of rabbit or cat fur just before placing it in the cradle. Rub the glass rod with the polyethylene sheet and bring the rod near one end of the polyethylene rod. What happens?

4. Open the stopcock on a water-filled buret so that a steady stream (not drops) runs out. Place the end of a charged (use the fur) polyethylene rod near the stream. What happens? Is the stream attracted to or repelled by the rod?

5. Place the end of a charged (use the plastic sheet) glass rod near the water stream. What happens? Is the stream attracted or repelled?

6. Repeat 4 and 5 except with a stream of isopropyl alcohol. (Collect the alcohol in a beaker for reuse.) What happens? Is the stream attracted or repelled as strongly as water?

7. Repeat 4 and 5 except with a stream of baby oil. (Collect the oil in a beaker for reuse. Also, the oil is more viscous than water, so use a buret with a large hole in the tip or remove the tip entirely to get a steady stream.) What happens in this case?
8. A small bottle was partially filled with water, several disks of paper coated on one side with graphite from a pencil were floated on the water, and oil was added to the bottle. Shake the bottle. Is there anything notable about the orientation of the paper punches?

9. Return the materials to their place. Refill the burets. Use care to not mix liquids.

**Questions:**

1. Explain your observations of the motion of the aluminum foil or pith ball. Hint: electrons are transferred from the glass rod to the polyethylene sheet.

2. From your observations, what can you deduce about the polarity of the bonds in the water molecule? What can you deduce about the shape of the water molecule?

3. The structural formula of isopropyl alcohol is CH₃CH(OH)CH₃. What is the geometry around each carbon atom and the oxygen atom? Where are the most polar bonds in the molecule? Are isopropyl alcohol molecules polar?

4. The formula of oil is approximately C₁₆H₃₄. Why are oil molecules essentially nonpolar?

5. Paper is mostly cellulose which contains long chains of C atoms with an -H and an -OH bonded to most carbons. Graphite is a form of elemental carbon with the atoms arranged in sheets. Why do the paper disks align themselves with the graphite toward the oil phase in the one bottle? Graphite is apparently wetted by oil, but not by water. Why not? Is graphite polar or nonpolar?

6. What did you learn from this experiment? How would you improve the experiment?
84. Popcorn Hypothesis

OBJECTIVE: To illustrate how problems are solved by the scientific method.

CONCEPTS: Observation, Hypothesis, Critical thinking, Scientific method

INTRODUCTION: The scientific method is based on discovering how things work by observing what happens in carefully controlled experiments, forming new hypothesis from these observations, and doing further experiments to test the validity of the hypotheses.

An authoritative book on popcorn states, "The two factors controlling the explosion ... are moisture in the kernel and an intact pericarp. Upon heating, the moisture inside the kernel expands to a point where enough pressure is built up to burst the pericarp." [1] In this experiment you will test these two hypotheses by trying to pop corn that has been dried and corn with a damaged pericarp hull.

HAZARDS: Be careful not to burn yourself.

PROCEDURE:

1. (Control observation) Test ten kernels from the package by popping them in a beaker on a hot plate. Do this without oil. Use tongs or a hot pad to gently shake the beaker while the corn is heating. Record the percentage popped.

2. Test the hypothesis that the intact pericarp is necessary for popping by making the following observations.
   a. Use ten kernels from the package. Rub each kernel on a piece of sandpaper to sand off a portion of the pericarp. Subject the kernels to the same popping procedure as before and determine the percentage that pop.
   b. Take another ten kernels and sand or cut off the pointed ends. Determine the percentage that pop.
   c. Take another ten kernels and using a hammer and a towel on the floor shatter the kernels into approximately one-fourth kernel pieces. Determine the percentage that pop.

What can you conclude (hypothesize) from your observations?

3. Test the hypothesis that moisture in the kernel is necessary for popping by making the following observation.
   a. Take ten kernels that have been previously dried (either in a warm oven or in a desiccator). What percentage popped? What can you conclude from this observation?

4. (Observation) Estimate the ratio of the volume of a popped kernel to the volume of an unpopped kernel, i.e. by what factor does the volume increase when a kernel pops?

5. Place the used (popped and unpopped) kernels in the trash can.

QUESTIONS:

1. Which hypotheses do your observations support? Which do they not support?

2. If your observations do not support a hypothesis, can you propose a better hypothesis that does explain your observations?

3. The literature [1] states that 14% moisture is optimum for popping corn. Assuming expansion of liquid water to water vapor is the process that makes the kernels explode, and that the volume of water vapor is 1000 times larger than the volume of the same amount of liquid water, estimate the theoretical (maximum) ratio of the volume of popped corn to the volume of unpopped corn.

4. Is your estimate of the volumes consistent with your observation of the increase in volume when corn pops?

5. Polymeric foams such as Styrofoam® and foam rubber are made by a process very similar to popping corn, except that blowing agents other than water are usually used. Butane and nitrogen are common blowing agents for synthetic foams. Why is nitrogen rather than water used as the blowing agent for plastic and rubber?

6. What did you learn from this experiment? How would you improve the experiment?
85. Precipitation Pictures

Name ___________________________ Course Number ___________________________
Student ID ___________________________ Section Number ___________________________
Date ___________________________ Instructor ___________________________

OBJECTIVE: To demonstrate silver-based photography.

CONCEPTS: Precipitation, Photochemistry

INTRODUCTION: The silver halides, water insoluble compounds, are the basis of many photographic processes. Most silver compounds, including silver chloride, are light sensitive. White AgCl turns dark from formation of metallic Ag when exposed to light. In this experiment the AgCl not exposed remains white, creating a negative image. A single photon may catalyze decomposition of an entire AgCl micro crystal.

HAZARDS: Getting AgNO₃ or AgCl on your hands or clothes will stain them. You (and everybody else) will be able to tell how good or bad your lab technique is by how many stains you get doing this experiment. Stains will wear off your skin after a few days.

PROCEDURE:

1. Pour a small amount of 0.5 M NaCl solution into a petri dish (just barely enough to cover the bottom).
2. With a pair of tweezers, dip a 5 cm circle of filter paper into the NaCl solution. Wet the paper evenly, but not to excess. Allow any excess liquid to drain from the paper.
3. Place the filter paper in a clean, dry petri dish.
4. Evenly space eight drops of 0.1 M AgNO₃ solution on the filter paper. The drops should run together. The objective is to cover the filter paper with a thin, even layer of AgCl.
5. Place 2 or 3 small plastic objects on the filter paper (the photographic film), being careful not to touch the paper with your fingers. Do not use metal objects.
6. Place the filter paper in the petri dish under a bright light or in direct sunlight.
7. Answer the questions while you wait for the “film” to be exposed. After 5 to 10 minutes, turn off the light and remove the plastic objects with tweezers.
8. When you are finished, dispose of the filter paper and NaCl solution in the waste container provided.
Questions:

1. Write the net ionic equation for the reaction of NaCl(aq) and AgNO₃(aq).

2. Write the chemical equation for the reaction that AgCl(s) undergoes when exposed to light.

\[ \text{AgCl}(s) \xrightarrow{hv} \]

3. Why can black and white film be developed in a room lit by a dim, red light, but not in a room with white light? (Hint: What colors of light have photons of the greatest energy?)

4. Why were plastic objects used for the photographic images instead of metal ones? (Hint: Ag⁺ is a fairly good oxidizing agent.)

5. What did you learn from this experiment? How would you improve the experiment?
INTRODUCTION:
You may remember being fooled as a child by the riddle, “which weighs more, a pound of feathers, or a pound of bricks?” Hopefully, you know by now that the answer lies in examining the differences in mass/volume ratios, or density. This lab is designed to provide other situations where the results may be unexpected. Hands-on experience will help you build your understanding of the principle of density.

CONCEPTS: Solutions, Density

BACKGROUND:
The density of a material depends partly on how closely the molecules of that material are packed. As molecules are more tightly arranged, a material becomes more dense. For objects within a fluid, substances that are more dense will sink to the bottom, while things that are less dense float to the top. For example, a rock sinks in water, but a piece of wood will float.

PROCEDURE:
1. Find the station with the balloons and a bucket of water. Test each of the balloons in the bucket and record the results or each item in a bucket of room temperature water.
   - Balloon of hot water. _______ sink, or _______ swim
   - Balloon of cold water. _______ sink, or _______ swim
   - Balloon of salt water. _______ sink, or _______ swim

2. Now drop the soda cans into the room temperature water bucket. Record observations in the table (part A). Then calculate the densities using parts B-E as a guide.

<table>
<thead>
<tr>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>Non-diet bottle</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Diet bottle</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Non-diet can</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Diet can</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

3. Now, go to the station with a beaker of oil and ice. Record in the space below what happens when an ice cube is placed in the beaker of oil and also in a beaker of water.

4. Around 300 B.C., the Italian King Heiro gave a goldsmith a certain weight of pure gold to use in making him a crown. When the crown was finished, Heiro was suspicious of the goldsmith, thinking that the goldsmith had withheld some of the gold for himself in the forging of the crown. The Greek scientist Archimedes was given the task of determining the integrity of the goldsmith without destroying the crown.

Archimedes was inspired with a solution to the problem when he stepped into a full tub of water in the public baths. As he stepped in, he noticed that water began to flow over the side. His
excitement was so great he forgot about his immodest state, and ran through the streets, shouting, “Eureka, eureka, I found it!” Archimedes went on to show that the crown displaced more water than an equal mass of pure gold. The crown had a lower density than that of gold and the goldsmith reluctantly returned the gold that should have been in the crown.

5. Two crowns made of different metals are provided, and you get to play the part of Archimedes. First, you must find the mass of the crowns by using a balance. Next, you will need to find the volume. This is done by water displacement in a two customized plastic beakers with spouts drilled into the sides; one for each crown. First, fill up the plastic beaker with water past the spout allowing the water to drain out until the water level is even with the spout. Next, prepare to catch displaced water with a graduated cylinder underneath the spout and then place the crown in the beaker while collecting the water (50 ml graduated cylinder for the small crown, at least two 100 ml graduated cylinders are needed for the large crown). Lastly, measure the volume on the side of the graduated cylinder(s), add volume up if necessary.

Using the table determine the identity of each crown.

<table>
<thead>
<tr>
<th></th>
<th>Mass</th>
<th>Volume</th>
<th>Density</th>
<th>Identity</th>
</tr>
</thead>
<tbody>
<tr>
<td>Large Crown</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Small Crown</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Density (g/cm³)</th>
<th>Metal</th>
<th>Density (g/cm³)</th>
<th>Metal</th>
<th>Density (g/cm³)</th>
<th>Metal</th>
<th>Density (g/cm³)</th>
<th>Metal</th>
</tr>
</thead>
<tbody>
<tr>
<td>2.70</td>
<td>Al</td>
<td>7.14</td>
<td>Zn</td>
<td>8.92</td>
<td>Cu</td>
<td>11.34</td>
<td>Pb</td>
</tr>
<tr>
<td>4.51</td>
<td>Ti</td>
<td>7.47</td>
<td>Mn</td>
<td>10.22</td>
<td>Mo</td>
<td>19.30</td>
<td>Au</td>
</tr>
<tr>
<td>6.11</td>
<td>V</td>
<td>7.87</td>
<td>Fe</td>
<td>10.49</td>
<td>Ag</td>
<td>21.46</td>
<td>Pt</td>
</tr>
</tbody>
</table>

QUESTIONS:
1. Water molecules are farther apart in (check one) cold water______, ice______. Explain how you decided.

2. Given your experiments, how does the density of ice compare to the density of oil? Is this what you would have expected?

3. In your calculations for the crown, how close were your calculations to the values listed? What are possible sources of error, if any, in determining density?
OBJECTIVE: To demonstrate Graham’s law of effusion of gases.

CONCEPTS: Kinetic molecular theory of gases, Molecular mass

INTRODUCTION: Equal volumes of different gases at the same temperature and pressure contain the same number of molecules. Because the average kinetic energy of gas molecules, \( KE = \frac{1}{2}mv^2 \), depends only on the temperature, the average velocity is proportional to the square root of the molecular mass. Graham’s law of effusion describing the rate at which a gas flows into a vacuum through a pinhole, is thus

\[
\text{Rate} = \frac{k}{\sqrt{\text{molecular mass}}}
\]

where \( k \) depends on the gas temperature, pressure, and the size and shape of the orifice; all variables which are held constant in this experiment. The greater the molecular mass of the gas, the slower the rate of escape through the orifice.

HAZARDS: Gas cylinders can be extremely dangerous if mishandled. Only the teaching assistants should adjust the handles of the gas regulators. You only need to turn the small delivery knob on and off.

PROCEDURE:
1. Assemble the apparatus as shown in the figure.
2. Connect the vacuum tubing to the vacuum. Turn the pump or lab vacuum valve on and leave it on throughout the experiment.
3. Fill the syringe with a gas. Begin with air. If cylinder gases are used, attach a short rubber tube to the delivery line from the gas regulator, open the small delivery knob to produce a slow flow of gas, place the end of the emptied syringe in the rubber tube, and withdraw the plunger slowly to fill the syringe. Close the delivery knob. The syringe may be filled with methane, \( \text{CH}_4 \), using the natural gas cock instead of a gas cylinder.
4. Connect the syringe to the top connector on the valve of the flask.
5. Turn the small, 3-way valve to connect the syringe to the flask. With a stopwatch, measure the time required for the syringe to empty.
6. Record the time in the table below.

7. Repeat with three other gases.

8. When you finish, make sure gas and vacuum valves are off.

<table>
<thead>
<tr>
<th>Gas</th>
<th>Time(s)</th>
<th>Molecular Mass (amu)</th>
<th>(\frac{1}{\sqrt{\text{mol. mass}}})</th>
<th>Rate of Effusion = (\frac{1}{\text{Time} (s^{-1})})</th>
</tr>
</thead>
<tbody>
<tr>
<td>Air</td>
<td></td>
<td>29.0</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

9. Plot the rate of effusion versus the inverse square roots of the molecular masses of the gases.

**QUESTIONS:**

1. Do your results agree with Graham’s Law? Explain.

2. Describe how you would determine the molecular mass of an unknown gas by this method?

3. What did you learn from this experiment? How would you improve the experiment?
88. Something Smells

OBJECTIVE: To observe the characteristic odors of organic compounds.
CONCEPTS: Functional groups, Odors

INTRODUCTION: The odor sensors in your nose discriminate among organic functional groups. For example, most amines smell “bad” and are produced when protein decays (i.e. is metabolized by bacteria). Esters smell “good” and are the odorants in ripe fruit. A compound must be volatile to be smelled, although amounts as small as a few molecules of some substances are detectible, even by humans. Ionic compounds are generally not volatile enough to be detected. For example, we use the acids in lemon juice to convert volatile amines, RH₂, to nonvolatile ammonium salts, RHN⁺X⁻, and thereby reduce the “fishy” smell of cooked fish. Many chemists can identify organic compounds merely by smelling them.

HAZARDS: To safely smell an organic compound, first smell the lid without opening the bottle (there is often enough vapor leakage to smell). If no or only a slight odor is detected, remove the lid or stopper and cautiously smell the inside of the lid. Only if the odor is still weak should you smell the open bottle. Do this by waving your hand over the bottle to carry some of the vapors from the bottle to your nose. Never directly smell an open bottle. Clean up any spills with copious amounts of water.

PROCEDURE: Smell two or three of each of the classes of compounds and describe their odors as best you can. R represents any hydrocarbon group or hydrogen.

1. Amines, RH₃N

2. Mercaptans, RSH

3. Aldehydes and ketones, RH₂C O

4. Alcohols, ROH

\[ \text{RCOH} \]

5. Carboxylic acids, O

\[ \text{RCOR} \]

6. Esters, O

---

-176-
Questions:
1. Does each functional group have a characteristic odor?

2. To what extent does the odor depend on the remainder of the molecule containing the functional group?

3. What can you conclude about odor receptors from this experiment?

4. Just as people vary in their perception of color (color blindness), people also inherit variations in their perception of odor (odor blindness). For example, most people find the odor of butyric acid very objectionable, but most people of Scandinavian descent find the odor somewhat pleasant. Given that butyric acid occurs in rancid butter and many cheeses (e.g. Havarti) and that prehistoric Scandinavians depended on dairy products, particularly during winter and times of famine, as a food source, can you think of an explanation for the existence of this particular odor blindness?

5. What did you learn from this experiment? How would you improve the experiment?
OBJECTIVE: To demonstrate the properties of hydrophylic polymers.

CONCEPTS: Polymerization, Hydrogen bonding, Dehydration

INTRODUCTION: Polymers are large molecules made by attaching many small molecules together. Polyvinyl alcohol (PVA) is a typical synthetic polymer used in water-based glues and paints.

Polyvinyl alcohol (PVA)

H2C\(\text{C}\)\(\text{C}\)\(\text{C}\)\(\text{C}\)\(\cdot\)\(\text{OH}\)

n

\[\text{H}_2\text{C}\cdot\text{C}\cdot\text{OH}\]

\[\longrightarrow\]

\[\text{H}_\text{n}\quad\text{H}_\text{C}\cdot\text{C}\cdot\text{OH}\]

\[\text{H}_2\text{C}\cdot\text{C}\cdot\text{OH}\]

\[\text{H}_\text{n}\quad\text{H}_\text{C}\cdot\text{C}\cdot\text{OH}\]

polyvinyl alcohol (PVA)

Vinyl alcohol

Polyacrylates, polymers of polyacrylic acid (PAA) and its derivatives, are widely used in glues, paints, textiles, and disposable diapers. Both PVA and PAA have many -OH groups in a polymer molecule. Because these groups are polar and can hydrogen bond, these polymers have a great affinity for water.

Polyacrylic acid

HAZARDS: Don’t get acetone or the PVA solution in your eyes. First aid: wash thoroughly with water. Clean up any spills with water. Acetone is flammable.

PROCEDURE:
1. One-fourth fill a 20 mL beaker with 10% PVA solution in water.

2. Carefully, so as to avoid mixing, add enough acetone to the beaker to completely cover the PVA solution. Note the layer of dehydrated PVA film that forms at the interface. The acetone dissolves water, but not PVA, thus leaving a layer of relatively dry PVA at the interface between the two liquids.

3. With tweezers, slowly pull a strand of PVA vertically from the center of the interface.
4. Place the strand on a piece of Saran® wrap and allow it to dry.

5. Observe the properties of the strand as it dries.

6. If the strand becomes stiff and hard, dip it in water and again observe the properties.

7. Place equal amounts about the size of a grain of rice of Hydrogel® (or stuffing from a disposable diaper) in three test tubes.

8. Add 5 mL of water to the first tube, 5 mL of 0.1 M NaCl to the second tube, and 5 mL of 1 M NaCl to the third tube.

9. After a few minutes, decant (pour off) any excess liquid and compare the volumes of the swelled polymer in the tubes.

10. When you are finished, dispose of all solutions in the waste container. Rinse the beaker and test tubes with distilled water.

**QUESTIONS:**

1. Acetone and water are completely miscible (i.e. they mix in all proportions). What is the purpose of the acetone in this experiment?

2. Plasticizers are often added to polymeric materials to make them softer and more flexible.

   \[
   \begin{array}{c}
   \text{CH}_3
   \\
   \text{CH}_3
   \end{array}
   \]

   Why do you think the water molecule is a good plasticizer for PVA and acetone / \text{is not} /?

3. Why do you think PAA is swelled by water?

4. How did the volume of water absorbed by PAA vary with increasing NaCl concentration in parts 8 and 9 of the procedure? Why do you think NaCl diminishes the swelling of PAA?

5. What did you learn from this experiment? How would you improve the experiment?
OBJECTIVE: To observe the effects of viscosity on the motion of liquids and relate viscosity to intermolecular attractive forces.

CONCEPTS: Viscosity, Hydrogen bonding, Intermolecular forces, Polarity

INTRODUCTION: Viscosity describes how easily a liquid can be made to change its shape. Very viscous liquids like honey resist changes in shape and change very slowly. Liquids with lower viscosity, like water, can change their shape very easily. Viscosity depends on how readily the molecules can change partners within the liquid. Viscous liquids have strong intermolecular forces. In this experiment, viscosity is used as an indication of the strength and extent of intermolecular attractions.

Intermolecular forces can be classified as London forces, dipole-dipole forces and hydrogen bonding. For small molecules, hydrogen bonding generally dominates.

HAZARDS: Do not open the flasks. Work in a well-ventilated area. In case of a spill avoid contact and ingestion of the liquids and breathing of the vapors. Ether and toluene are also flammable.

PROCEDURE: Swirl each flask. Note how easy it is to get the liquid to rotate inside the flask and form a vortex. Note how quickly the liquid motion ceases after you stop swirling the flask. Use a scale of 1 (easiest) to 6 to indicate the relative viscosities of the liquids.

<table>
<thead>
<tr>
<th>Liquid</th>
<th>Molecular structure</th>
<th>Relative viscosity</th>
</tr>
</thead>
<tbody>
<tr>
<td>toluene</td>
<td><img src="" alt="Toluene structure" /></td>
<td></td>
</tr>
<tr>
<td>diethylether</td>
<td><img src="" alt="Diethylether structure" /></td>
<td></td>
</tr>
<tr>
<td>water</td>
<td><img src="" alt="Water structure" /></td>
<td></td>
</tr>
<tr>
<td>ethylene glycol</td>
<td><img src="" alt="Ethylene glycol structure" /></td>
<td></td>
</tr>
</tbody>
</table>
glycerine

corn syrup (a very concentrated solution of glucose in water)

QUESTIONS:

1. Which liquids have intermolecular hydrogen bonding? Which are polar?

2. Which is more viscous, diethylether or water? Why do you think they differ in viscosity?

3. Why do you think glycerine is more viscous than ethylene glycol?

4. Why do you think corn syrup is so viscous?

5. What did you learn from this experiment? How would you improve it?
91. Classification Tests for Organic Functional Groups

OBJECTIVE: To demonstrate tests used to identify organic functional groups.

CONCEPTS: Functional groups, Oxidation-reduction reactions

INTRODUCTION: The reactions of functional groups can be used to identify the groups present in an unknown organic compound. The reactions of alcohols, aldehydes, ketones, and carboxylic acids with dichromate ion [Cr$_2$O$_7^{2-}$] in acidic solution and with diamminesilver(I) ion [Ag(NH$_3$)$_2^{+}$] in basic solution are demonstrated in this experiment. Alcohols and aldehydes both rapidly reduce the orange dichromate ion to green Cr(III) ion. Aldehydes, but not alcohols, will reduce the diaminesilver(I) ion to a metallic silver mirror. Carboxylic acids and ketones will not reduce either reagent.

HAZARDS: The chromium solution in sulfuric acid is toxic and corrosive. Avoid ingestion and skin contact. Some of the organic compounds are also toxic. The silver solution will stain skin and clothing. If contact with any of the reagents occurs, wash thoroughly with water.

PROCEDURE:

Four known compounds (an alcohol, an aldehyde, a ketone, and a carboxylic acid) are each tested with both reagents to observe the reaction (or lack of reaction). Then three unknowns (labeled A, B, and C) are tested.

Place 14 test tubes numbered 1-14 in a test tube rack. Add 20 drops (= 1mL) of a solution that is 0.1 M K$_2$Cr$_2$O$_7$ and 1 M H$_2$SO$_4$ to tubes 1-4 and 9-11. Add 20 drops of Ag(NH$_3$)$_2^{+}$ solution (Tollen’s reagent) to tubes 5-8 and 12-14. Add 5 drops of 10% ethanol to tubes 1 and 5, of 10% acetic acid to tubes 2 and 6, of 10% dextrose to tubes 3 and 7, and of 10% acetone to tubes 4 and 8. Let stand for 3 minutes. Record your observations.

Observations on reactions of known compounds
Tube 1: Cr$_2$O$_7^{2-}$ + C$_2$H$_5$OH
Tube 2: Cr$_2$O$_7^{2-}$ + CH$_3$COOH
Tube 3: Cr$_2$O$_7^{2-}$ + C$_6$H$_{12}$O$_6$
Tube 4: Cr$_2$O$_7^{2-}$ + CH$_3$COCH$_3$
Tube 5: Ag(NH$_3$)$_2^{+}$ + C$_2$H$_5$OH
Tube 6: Ag(NH$_3$)$_2^{+}$ + CH$_3$COOH
Tube 7: Ag(NH$_3$)$_2^{+}$ + C$_6$H$_{12}$O$_6$
Tube 8: Ag(NH$_3$)$_2^{+}$ + CH$_3$COCH$_3$

Add 5 drops of unknown A to tubes 9 and 12, of B to 10 and 13, and of C to 11 and 14. Wait 3 minutes and record your observations.
Observations on reactions of unknowns.
Tube 9: $\text{Cr}_2\text{O}_7^{2-} + \text{A}$
Tube 10: $\text{Cr}_2\text{O}_7^{2-} + \text{B}$
Tube 11: $\text{Cr}_2\text{O}_7^{2-} + \text{C}$
Tube 12: $\text{Ag(NH}_3\text{)}_2^+ + \text{A}$
Tube 13: $\text{Ag(NH}_3\text{)}_2^+ + \text{B}$
Tube 14: $\text{Ag(NH}_3\text{)}_2^+ + \text{C}$

Empty the test tubes into the appropriate, labeled waste containers provided. Place the test tubes into a waste container. If you wish, you may keep the tubes with silver mirrors after you have thoroughly rinsed them with water. Clean up the bench area of any spills.

QUESTIONS:

1. What class of compounds is oxidized by Tollén’s reagent?

2. What classes of compounds are not oxidized by either dichromate ion or Tollén’s reagent?

3. If a compound reduces dichromate ion, but not Tollén’s reagent, what kind of functional group must it have?

4. Match the unknowns A, B, and C to their chemical class.

   ______ alcohol

   ______ aldehyde

   ______ acid or ketone

5. What did you learn from this experiment? How would you improve it?
OBJECTIVE: To demonstrate the effect of temperature on the rate of a chemical reaction.

CONCEPTS: Activation energy, Rate laws, Absolute temperature

INTRODUCTION: As temperature increases, molecular kinetic energy increases, the fraction of molecules with sufficient kinetic energy to react increases (for many reactions, this roughly doubles every 10 K increase), the rate of collisions increases (about 7% for a 10 K temperature rise), and the reaction goes faster. The minimum kinetic energy a collision must have if a reaction is to occur is called the activation energy. Because of the relationship between molecular kinetic energy and absolute temperature, the activation energy can be determined by measuring the rate of a reaction as a function of temperature. The rate of a reaction is described by a rate law specific for every reaction, i.e. rate = k times a function of the reactant concentrations. Svante Arrhenius showed that k is related to the absolute temperature by the equation $k = A e^{-E_a/RT}$ where A is a constant, $E_a$ is the activation energy, R is the gas constant (8.315 J mol$^{-1}$ K$^{-1}$), and T is the Kelvin temperature. Thus k increases exponentially with T. As a rule of thumb, biological reaction rates double for every 10 K increase of temperature. This rule is exactly true at 25°C for any reaction having $E_a = 53$ kJ mol$^{-1}$.

In this experiment, the time required for thiosulfate ion to decompose in an acidic solution at different temperatures is measured. The rate law for this reaction is

$$\text{rate} = \frac{\Delta n}{\Delta t} = k[S_2O_3^{2-}]^y[H^+]^x$$

where $\Delta n$ is the number of moles of $S_8(s)$ produced during time $\Delta t$, k is the rate constant, and $[S_2O_3^{2-}]$ and $[H^+]$ are the concentrations of thiosulfate and hydrogen ions. If $\Delta n$, $[S_2O_3^{2-}]$, and $[H^+]$ are held constant, the rate constant is proportional to $1/\Delta t$.

$$\text{rate} \quad 1/\Delta t = k \{ [S_2O_3^{2-}]^y[H^+]^x/\Delta n \} = k \cdot \text{constant}$$

The basis of the procedure is that the $S_2O_3^{2-}$ ion is stable in neutral to basic solution, but decomposes rapidly by the reaction

$$16H^+ + 8S_2O_3^{2-} \rightarrow 8H_2SO_3(aq) + S_8(s)$$

in acidic solutions.

HAZARDS: Wear safety goggles as others around you may be using hazardous chemicals.

PROCEDURE:
To save time, make three simultaneous determinations, one at about 10°C, one at about 20°C and one at about 30°C. Prepare 3 beakers with about 100 mL of H$_2$O at each temperature by using ice and hot and
cold tap water. When all three temperature baths are ready, quickly mix the solutions, place the test tubes in the beakers, and measure the times for the solutions to become opaque. Placing a piece of printed material behind the beakers and test tubes will help. Take the time when the print becomes unreadable through the tubes as the endpoint.

Solution: Mix 1 mL of 0.1 M HCl with 1 mL of 0.1 M Na₂S₂O₃ in each test tube

<table>
<thead>
<tr>
<th>Run</th>
<th>T, °C</th>
<th>T, K</th>
<th>Δt, s</th>
<th>rate = 1/Δt</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Questions:

1. Suggest two reasons why the rates of chemical reactions increase with increasing temperature.
   A.
   B.

2. Which of these two predominates? Why?

3. By what factor did the rate increase per 10 K increase in temperature?
   a. From 10 to 20°C?
   b. From 20 to 30°C?
   c. From 10 to 30°C?

4. List several reactions that occur more rapidly at elevated temperatures. (Hint: consider cooking, gluing, painting, plants growing, lizards running, hair drying, etc.)

5. What did you like about this experiment and how would you improve it?
OBJECTIVE: To demonstrate the effect of temperature on the rate of a chemical reaction.

CONCEPTS: Activation energy, Rate laws, Absolute temperature

INTRODUCTION: As temperature increases, molecular kinetic energy increases, the fraction of molecules with sufficient kinetic energy to react increases (for many reactions, this roughly doubles every 10 K increase), the rate of collisions increases (about 7% for a 10 K temperature rise), and the reaction goes faster. The minimum kinetic energy a collision must have if a reaction is to occur is called the activation energy. Because of the relation between molecular kinetic energy and absolute temperature, the activation energy can be determined by measuring the rate of a reaction as a function of temperature. The rate of a reaction is described by a rate law specific for every reaction, i.e. rate = k times a function of the reactant concentrations. Svante Arrhenius showed that k is related to the absolute temperature by the equation 
\[ k = A e^{-\frac{E_A}{RT}} \]
where A is a constant, \( E_A \) is the activation energy, R is the gas constant (8.315 J mol\(^{-1}\) K\(^{-1}\)), and T is the Kelvin temperature. Thus k increases exponentially with T. As a rule of thumb, biological reaction rates double for every 10 K increase of temperature. This rule is exactly true at 25°C for any reaction having \( E_A = 53 \text{ kJ mol}^{-1} \).

In this experiment, the time required for hydrolysis of an alkyl halide is measured. The reaction is

\[ (\text{CH}_3)_3\text{CCl} + \text{OH}^- \rightarrow (\text{CH}_3)_3\text{COH} + \text{Cl}^- \]

The rate law is

\[ \text{rate} = \frac{\Delta n}{\Delta t} = k[\text{(CH}_3)_3\text{CCl}][\text{OH}^-] \]

where \( \Delta n \) is the number of moles of reaction occurring during time \( \Delta t \), k is the rate constant, and \( [\text{(CH}_3)_3\text{CCl}] \) and \( [\text{OH}^-] \) are the concentrations of tertiary butyl chloride and sodium hydroxide. If \( \Delta n \) and the concentrations of both reactants are held constant, the rate constant is proportional to \( 1/\Delta t \).

\[ \text{rate} \cdot \frac{1}{\Delta t} = k \frac{[\text{(CH}_3)_3\text{CCl}][\text{OH}^-]}{\Delta n} = k \cdot \text{constant} \]

HAZARDS: Wear safety goggles as others around you may be using hazardous chemicals.

PROCEDURE:
To save time, make three simultaneous determinations, one at about 10°C, one at about 20°C and one at about 30°C. Prepare 3 beakers with about 100 mL of H\(_2\)O at each temperature by using ice and hot and cold tap water. When all three temperature baths are ready, quickly mix the
solutions, place the test tubes in the beakers, and measure the times for the solutions to change color.

Solution: Mix 1 mL of 2% by volume tertiary-butyl chloride in isopropyl alcohol with 1 mL of 0.01 M NaOH/thymol blue in each test tube.

<table>
<thead>
<tr>
<th>Run</th>
<th>T, °C</th>
<th>T, K</th>
<th>Δt, s</th>
<th>rate = 1/Δt</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Questions:

1. Suggest two reasons why the rates of chemical reactions increase with increasing temperature.  
   A.  
   B.  

2. Which of these two predominates? Why?

3. By what factor did the rate increase per 10 K increase in temperature?  
   a. From 10 to 20°C?  
   b. From 20 to 30°C?  
   c. From 10 to 30°C?  

4. List several reactions that occur more rapidly at elevated temperatures. (Hint: consider cooking, gluing, painting, plants growing, lizards running, hair drying, etc.)

5. What did you like about this experiment and how would you improve it?
94. Measurement of Radioactivity

Name: ________________________________ Course Number: ____________________
Student ID: __________________________ Section Number: ____________________
Date: ___________________________ Instructor: ______________________________

**OBJECTIVE:** Learn about radiation properties, shielding, and half-life.

**CONCEPTS:** Radioactivity, Decay rate, Shielding of radiation

**HAZARDS:** Exposure to radiation is measured in units of “rem.” Having an X-ray image taken of your body will expose you to a range of 1 to 80 millirem (mrem) depending on the part of your body imaged. On average in Provo, from both manmade and natural sources of radiation, a person is exposed to about 1 mrem of radiation per day. By following the procedures in the lab, you will be exposed to less than 0.5 mrem to your hands and forearms. The maximum you could be exposed to if you were completely irresponsible and held your hand in front of the aperture for 30 minutes is about 2 mrem, but don’t do it.

Treat the hole of the radioactive sources as you would a high powered laser—keep the hole pointed down the bench (not towards you) and never look in the hole. As with all other labs, follow standard practices of cleanliness and wash your hands when you leave.

**INTRODUCTION:** Our society has utilized radioactive sources for over one hundred years in applications such as smoke detectors (americinium-241), glass lantern mantles (thorium-232), cancer treatments (iodine-131), medical imaging (technecium-99), and nuclear power plants (uranium or plutonium). Unstable atomic nuclei spontaneously decompose to emit α-particles (He\(^{2+}\)), β-particles (e\(^-\)), γ-rays, or some combination of this set.

Radioactivity can be measured with a Geiger counter. The sensor in a Geiger counter is a gas-filled metal tube with a thin mica window. When α and β particles and γ rays pass into the tube at various locations, gas in the tube ionizes. Every time the gas particles are ionized, electricity flows through a circuit which signals the presence of radioactivity. The circuit in the Geiger counter counts the events, and displays as a radiation level on the readout. Keep in mind that radiation is emitted from the radioactive source in all directions. Unfortunately, Geiger counters can only measure a fraction of the radiation because the sensor is small. If the radiation from a source could be measured in all directions, the activity of the sample could be measured, and a half-life calculated. To account for this difference, you will need to apply a correction factor by multiplying the measured radiation level by 216 (when the detector is touching the hole of the source).\(^1\) This correction factor will only be needed in question 5.

A common concern with radiation is protecting humans from exposure. For high radiation levels federal regulations allow workers to only be exposed for short amounts of time. One practice in the storage of radioactive waste is to store it over 2000 feet underground in remote areas to reduce exposure by increasing distance. Another method to minimize radiation exposure is shielding, which you might remember as the heavy shield placed on your body before an x-ray at the dentist office.

---

\(^1\) For those interested in how this conversion factor was determined, see the Appendix.
**PROCEDURE:**

**OPERATION OF GEIGER COUNTER**

Depending on the position of the selector switch (illustrated in the figure at right), the Gamma-Scout® Geiger counter detects different combinations of α, β, and γ radiation. For example, in the far right position, all radiation types are allowed to pass to the detector. Do NOT touch the detector or shielding, they can be easily damaged!

1. Locate a yellow Geiger counter (without black leather cover). On the Geiger counter input pad, locate either the “cps” or the “Bq” button. “Bq” stands for Becquerels, or the number of decay events in one second. “Cps” stands for counts per second, or the number of events detected (Bq and cps are equivalent for our purposes here). Press “cps”/“Bq” to change the display to read an average over time of what is being measured, in counts/sec. Whenever you want to start a new running average press “cps”/“Bq,” which will reset the counter and a new measurement will begin as # flashes on the display. For your measurements today, allow the counter to read for 30–35 seconds and then read the result.

2. What does the counter read when it is away from the radioactive sources? That is, what is the background measurement in the lab? Take a “control” measurement by moving the Geiger counter to the other side of the laboratory, away from the radioactive samples. Reset the counter and take a 30–35 second measurement of the average radiation level. __________ counts/sec.

**SHIELDING**

3. Locate the depleted uranium radioactive source with its snap-on shield. Carefully remove the shield; it will take some effort. Point the top of the Geiger counter at the radiation source aperture and place the entrance to the Geiger counter sensor about 5 cm away from the aperture of the source. Measure the α+β+γ radiation rate, and record your result on the table at right. Next, place one sheet of cardboard between the sample and Geiger counter. The thickness of the cardboard is 0.78 mm. Measure the rate, and record it on the table. Repeat for 2, 3, and 4 sheets of cardboard. Next, repeat this process using the laminated 0.1mm lead sheets.

<table>
<thead>
<tr>
<th>Cardboard</th>
<th>rate</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sheets</td>
<td>(counts/sec)</td>
</tr>
<tr>
<td>0</td>
<td></td>
</tr>
<tr>
<td>1</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
</tr>
<tr>
<td>Lead</td>
<td>rate</td>
</tr>
<tr>
<td>Sheets</td>
<td>(counts/sec)</td>
</tr>
<tr>
<td>1</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
</tr>
<tr>
<td>4</td>
<td></td>
</tr>
</tbody>
</table>

**TYPES OF RADIATION**

4. Next, leaving the equipment behind, move to a station with a black Geiger Counter (please do not move stations) and encased natural uranium source. Place the detector and source aperture directly in contact. Measure the counts/sec for all three positions of the selector switch: Write in your results here:

α, β, and γ __________  β, and γ __________  γ __________
**Radiation Levels and Distance**

5. Set the Geiger counter to measure $\beta+\gamma$ (leftmost position). Place the detector and sample at various distances from each other (up to 20 cm). Remember to add 3.35 cm to your distance measurements because the sample sits back inside the casing (this means you will not have a point at 0 cm). Plot at least 6 points. Add the numbers to the y axis as appropriate.

*On the metal box I am using, it shows that I am using source (circle one) #1 #2*

**Questions**

1. When you took the Geiger counter away from the sample, was the radiation level zero? ______ Explain.

2. Based on your results, determine the percent of each type of radiation from the natural uranium source.

   $\alpha \quad \% \quad \beta \quad \% \quad \gamma \quad \%$

3. Approximately what thickness of cardboard had or would have the same level of shielding as 0.1mm lead?

4. Explain your observations from the section “Radiation Levels and Distance.” In this experiment, the sample was the same, and source was located in the same place. Why is the measured rate changing if the sample conditions (that is location, size, and shielding) remain the same?

5. Use the conversion factor given in the introduction to determine the total gamma radiation rate from the source in counts/sec. This is called the activity of the sample.
6. Determine the number of uranium atoms in your natural source. The mass of uranium in sample #1 is 0.0917g, and the mass of uranium in sample #2 is 0.0810g (these masses represent about 2.5% of the total mass of each rock). There are $2.53 \times 10^{21}$ uranium atoms/gram.

7. Total radiation levels are described as decay rates with $\lambda = \frac{\text{rate}}{N}$, where $\lambda$ is the rate constant of the isotope and $N$ is the number of radioactive particles (nucleotides) in the sample. The rate here is the answer you found for question #5. Often the rate is expressed as the time needed for half of the original particles in the sample to decay, or half-life which is given as $t_{1/2} = \frac{0.693}{\lambda}$. Solve for the half-life of the sample in years using the equations given.

**APPENDIX (OPTIONAL INFO)**

Because radiation is emitted in all directions we need to take into account that only a part of the total rate is measured and the rest is shielded. The fraction of radiation total to the radiation measured is equivalent to the area of radiation total compared to that measured or $\frac{\text{Rate}_{\text{total}}}{\text{Rate}_{\text{measured}}} = \frac{\text{Area}_{\text{tot}}}{\text{Area}_{\text{meas}}}$. We say radiation is emitted as a sphere with a surface area of $4\pi r_1^2$, where $r_1$ is the distance from the center of the sample to the detector, and the detector is a circular shape with an area of $\pi r_2^2$ and is smaller than the source hole (see figure). Thus $\frac{\text{Rate}_{\text{total}}}{\text{Rate}_{\text{measured}}} \approx \frac{4\pi r_1^2}{\pi r_2^2}$. For these sources $r_1 = 3.35\text{cm} = 3.35 \times 10^{-2} \text{ m}$, and for these detectors $r_2 = 4.55\text{mm} = 4.55 \times 10^{-3} \text{ m}$. We assume all radiation from the hole is detected. With a little algebraic rearrangement and substitution we get $\text{Rate}_{\text{total}} = 216 \times \text{Rate}_{\text{measured}}$.
**95. Molecular Models**

**OBJECTIVE:** To make and observe models of molecular shapes.

**CONCEPTS:** Molecular shapes, Electron configurations

**INTRODUCTION:**
How can diet soda still taste sweet without sugar? Your tongue is covered in receptors that recognize sweetness. Some molecules have geometries that better fit sweet flavor receptors than sucrose (table sugar). Because of a better fit, 200–600 times less sweetener can be used for a drastic reduction in calories. Molecular shapes and geometries are used in pharmaceuticals, in thousands of chemical reactions in our bodies, and in hemoglobin to reduce the chance of carbon monoxide poisoning while still allowing O\(_2\) to bind. Though we may not always recognize it, molecular geometry is useful in many aspects of our lives. Today you will use molecular models to build and name basic molecular shapes.

Molecular shapes can be used to predict both physical properties such as polarity, boiling point and viscosity, as well as reactivity and reaction products. In biochemical reactions, enzymes will only interact with molecules of specific shapes. Basic shapes are named based on the positions of outer nuclei around a central atom. Nearly all polyatomic molecule and ion shapes are related to a linear, trigonal planar, tetrahedral, trigonal bipyramidal or octahedral arrangement of bonding and lone pairs of electrons around the central atom.

Models of CO\(_2\) (linear), BF\(_3\) (triangular planar), CH\(_4\) (tetrahedral), PF\(_5\) (trigonal bipyramidal), and SF\(_6\) (octahedral) will be displayed to show you how they are made. Model sets will be available for you to work with. Different colored balls represent different atomic nuclei. Connecting tubes represent pairs of electrons in bonds, and leftover stubs on the central ball represent lone pairs of electrons.

**PROCEDURE:**

I. Make models of and sketch (using element symbols) the following molecules.

<table>
<thead>
<tr>
<th>Linear</th>
<th>Trigonal planar</th>
<th>Tetrahedral</th>
<th>Trigonal bi-pyramidal</th>
<th>Octahedral</th>
</tr>
</thead>
<tbody>
<tr>
<td>(BeCl(_2))</td>
<td>(BCl(_3))</td>
<td>(SiCl(_4))</td>
<td>(PCl(_3))</td>
<td>(SF(_6))</td>
</tr>
</tbody>
</table>

II. Using the models you built in part I, remove one ball and its corresponding tube from each model (except linear). Note there are two ways to remove a ball from the trigonal bi-pyramidal, from one of the three equatorial or one of the two axial positions. Only molecules with vacant equatorial positions are ever formed. Sketch these models, but do not draw lone pairs because they are not considered in naming molecular shapes.

<table>
<thead>
<tr>
<th>Bent (SO(_2))</th>
<th>Trigonal pyramidal (NH(_3))</th>
<th>See-saw (SF(_4))</th>
<th>Square pyramidal (BrF(_5))</th>
</tr>
</thead>
</table>
III. Using the models you built in part II, remove one ball and its corresponding tube from each model (except bent). Note there are now two ways to remove a ball from square pyramidal one of four basal positions or the axial position. Remove an atom that will maximize the distance between lone pairs. Sketch these models. Note the angle of the bent molecule is 109.5° instead of 120°.

<table>
<thead>
<tr>
<th>Model</th>
<th>Shape</th>
<th># Total bonding pairs</th>
<th># Lone pairs around central atom</th>
</tr>
</thead>
<tbody>
<tr>
<td>Bent (H₂O)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>T-shaped (ClF₃)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Square planar (XeF₄)</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**QUESTIONS:**

1. Predict the shapes and number of bonding and lone pairs of electrons in each molecule or ion. Note that when fluorides or chlorides are the surrounding atoms, the central atom can exceed the octet rule to accommodate 10 or 12 electrons.

2. a) Draw (make models if necessary) of the two molecules that commonly bind to hemoglobin, taking note of the lone electron pairs. These molecules bind to the hemoglobin by their lone electron pairs. b) If hemoglobin has more space to bind at an angle rather than straight up, explain which molecule is favored to bind and why.

3. Optional: What did you like about this experiment and how could it be improved?
96. Molecular Shapes and Types of Bonds

Name ___________________________________________ Course Number ____________________________
Student ID _______________________________________________ Section Number __________________________
Date ___________________________ Instructor __________________________

OBJECTIVE: To model molecular shapes and observe the relation of the molecular shape to the bonds in molecules with atoms that obey the octet rule.

CONCEPT: Molecular shapes, Bond types

INTRODUCTION: The hybridization of the central atom in a molecule to its neighboring atoms determines the shape of the molecule. If the Lewis structure for the central atom does not contain any double bonds, then the atomic s and p orbitals hybridize to give four sp³ orbitals directed toward the corners of a tetrahedron. These orbitals form s-bonds or are occupied by an unshared pair of electrons. If the Lewis structure for the central atom contains one double bond, then the atomic s and p orbitals give sp² hybridization to give three sp² orbitals for s-bond(s) or lone pairs and one atomic p orbital for the p-bond. The sp² orbitals are directed at the corners of a planar, equilateral triangle. If the Lewis structure for the central atom contains two double bonds or a triple bond, then the s and p orbitals give sp hybridization to give two sp hybrids for s-bond(s) or lone pairs and two atomic p orbitals for p-bonds. The sp hybrid orbitals are directed along a line passing through the central atom. The atoms adjacent to the central atom can be analyzed in similar manner to determine the type of orbitals used in bonding (except for hydrogen which uses only the 1 s orbital).

Shapes of molecules determine the way in which three-dimensional addition of bond dipole moments gives rise to presence or absence of a molecular dipole moment. The molecular dipole moment correlates with physical properties such as melting point, boiling point, vapor pressure, solubility in polar or non-polar solvents, etc. The shapes of molecules are described in terms of the geometric object defined by the positions of the atoms around a central atom. The central atom is at the center of the object. Most molecular models use different colored balls to represent different atomic nuclei and a direct connecting rod to indicate a s-bond and two rods, one above and one below the direct connecting rod to indicate p-bonds. In this experiment, model sets with tubes representing s-bonds, p-bonds, and lone pairs are used to make models of linear, bent, trigonal planar, trigonal pyramidal, and tetrahedral molecules.

PROCEDURE:

For the molecules in the table:
1. Determine the number of bonds in the molecule.
2. Write a Lewis structure for the molecule representing each bond between the atoms with a line. Indicate unshared pairs of electrons with two dots.
3. Determine the type of hybrid orbitals on each atom and list in the table.
4. Make a model of the molecule showing both the s-bonds and p-bonds.
5. Determine the number of unshared pairs of electrons on each atom and list these in the table.
6. Describe the molecular shape in one or two words.
<table>
<thead>
<tr>
<th>Molecule</th>
<th>Draw Lewis Structure – Label bonds with $\sigma$ or $\pi$.</th>
<th>Hybridization of each atom</th>
<th>Number of unshared pairs of electrons on each atom</th>
<th>Molecular shape</th>
</tr>
</thead>
<tbody>
<tr>
<td>$N_2$</td>
<td>N</td>
<td>N</td>
<td>N</td>
<td></td>
</tr>
<tr>
<td>$CO_2$</td>
<td>C</td>
<td>C</td>
<td>O</td>
<td></td>
</tr>
<tr>
<td></td>
<td>O</td>
<td>O</td>
<td>O</td>
<td></td>
</tr>
<tr>
<td>$H_2CO$</td>
<td>C</td>
<td>H</td>
<td>H</td>
<td></td>
</tr>
<tr>
<td></td>
<td>O</td>
<td>C</td>
<td>O</td>
<td></td>
</tr>
<tr>
<td>$NH_3$</td>
<td>N</td>
<td>N</td>
<td>H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>H</td>
<td>H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>H</td>
<td>H</td>
<td></td>
</tr>
<tr>
<td>$CH_4$</td>
<td>C</td>
<td>C</td>
<td>H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>H</td>
<td>H</td>
<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>H</td>
<td>H</td>
<td></td>
</tr>
<tr>
<td>$H_2O$</td>
<td>O</td>
<td>H</td>
<td>H</td>
<td></td>
</tr>
</tbody>
</table>

What did you like about this experiment and how could it be improved?
97. Thermodynamics of a Galvanic Cell

OBJECTIVE: To determine the thermodynamic parameters for a reaction from electrical measurements.

CONCEPTS: Gibbs Free Energy, Enthalpy change, Entropy change, Oxidation-reduction reactions, Galvanic cell

INTRODUCTION: The relationship between the change in the Gibbs free energy of a thermodynamic system and the changes in entropy and enthalpy of that system is

\[ \Delta G = \Delta H - T\Delta S \]

where \( G \) = Gibbs free energy (the maximum useful work obtainable from the system), \( H \) = enthalpy, \( S \) = entropy, and \( T \) = absolute temperature of the system.

In a system involving a redox reaction the useful obtainable work is measured as the potential difference across the system, denoted as \( E \) = electromotive force. The relationship between Gibbs free energy, \( G \), and electromotive force, \( E \), is given as

\[ \Delta G = -nFE \]

with \( n \) = the moles of electrons involved in the reaction and \( F \) = coulombs of charge/mole of electrons.

Combining the two expressions we see that

\[ E = \frac{\Delta S}{nF}(T) - \frac{\Delta H^\circ}{nF} \]

Thus a plot of \( E \) for a cell against \( T \) should yield a straight line with a slope containing the entropy change and an intercept containing the enthalpy change for the reaction.

You will measure \( E \) at three temperatures for the redox reaction

\[ \text{Zn} + 2 \text{Fe(CN)}_6^{3-}(aq) \rightleftharpoons \text{Zn}^{2+}(aq) + 2 \text{Fe(CN)}_6^{4-}(aq) \]

The data will be plotted and used to estimate values at 25°C for \( \Delta H \) and \( \Delta S \), which will then be combined to give a value of \( \Delta G \) at the same temperature.
**PROCEDURE:**
1. Three cells should already be set up. Check that the cold bath has both ice and water in the beaker, and that the warm bath has sufficient water to surround the test tubes.
2. Once the baths are set and thermally equilibrated record the temperature of each in the table.
3. Use the voltmeter provided to measure the potential of each cell in millivolts. Connect the black lead on the meter to the zinc electrode, the red lead to the carbon electrode.
4. Record the voltage of each cell.
5. Convert the temperatures to Kelvin.
6. Plot the voltages against the temperatures.
7. Use a spreadsheet, your calculator or the graph to determine the slope and the intercept for the line of best fit on the plot.
8. Work out the value of $n$ for the reaction and look up the value of $F$ in your text book.
9. Calculate the values for $\Delta H$ and $\Delta S$ and combine them to estimate $\Delta G$.

<table>
<thead>
<tr>
<th>Variable</th>
<th>Cold</th>
<th>Room</th>
<th>Warm</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^\circ$C</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>K</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>E</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**QUESTIONS:**
1. Calculate $\Delta G$ from $\Delta G = -nFE$.
   How does this value compare with the one you obtained from $\Delta H - T\Delta S$?
2. How would you explain the difference between the two values (if a difference exists)?
3. What did you learn from this experiment? How would you improve the experiment?
OBJECTIVE: To learn how to determine the concentration of a weak acid by titration with a strong base.

CONCEPTS: Acids and bases, Concentration, Titration

INTRODUCTION: Titration is a technique used to determine the concentration of a solution. From titration data (i.e. the volume of an acid solution required to just neutralize a certain volume of a base solution, or vice versa) and the known concentration of one of the solutions (either the acid or the base), the concentration of the other solution can be determined. In this experiment you will determine the concentration of an acetic acid (CH₃COOH) solution by titration with a sodium hydroxide (NaOH) solution of known concentration.

\[
\text{CH}_3\text{COOH(aq)} + \text{OH}^- \rightarrow \text{CH}_3\text{COO}^- + \text{H}_2\text{O}
\]

You will use 2 different indicators to determine the end points of your titrations, methyl red (red to yellow over the pH range 4.8-6.0) and thymol blue (red to yellow over the pH range 1.2-2.8 and yellow to blue over the pH range 8.0-9.6). The end point of the titration occurs when the color of the indicator changes from the color it is at the lower end of the pH scale to the color it is at the higher end of the pH scale.

HAZARDS: Acids and bases are strong skin irritants. If acid or base solutions come in contact with your skin, thoroughly rinse the affected area immediately with water. As always, wear your safety goggles.

PROCEDURE:
1. Obtain a 24-well plate and rinse it thoroughly with distilled water. To each of 4 wells transfer about 20 drops of CH₃COOH solution. Record the number of drops added to each well in the table below.
2. Add one drop of methyl red indicator solution to each of wells 1 and 2 and 1 drop of thymol blue indicator solution to each of wells 3 and 4.
3. Titrate Well 1 by adding 0.20 M NaOH solution dropwise with stirring after each drop until you reach the end point. Record the number of drops added in the table below.
4. Titrate each of the other wells with 0.20 M NaOH. From the first titration you know about how many drops are required to reach the end point. So these titrations can be done rapidly by adding NaOH solution dropwise without stirring up to 2 or 3 drops from the expected end point. Then stir the solution in the well and finish the titration carefully by stirring after each additional drop as in item 3 above. Record in the table below the number of drops required to reach the end point for each well. Duplicate titrations with a given indicator should agree within 1 drop.
5. Use the data from each of your titrations to determine the concentration of the CH₃COOH solution and record the results for each well in the table below. Do additional titrations if you like with whichever indicator you choose. Dispose of waste in the proper waste container.
Well Number | 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8
---|---|---|---|---|---|---|---|---
drops CH₃COOH | | | | | | | | |
indicator | methyl red | methyl red | thymol blue | thymol blue |
drops NaOH | | | | | | | | |
[CH₃COOH] | | | | | | | | |

**Questions:**
1. To obtain accurate results, must the size of the drop delivered from the acid container be the same as the size of the drop delivered from the base container? Yes___ No___ Explain your answer.

2. Using data from your titration of Well 1, show your calculations for determining the concentration of the CH₃COOH solution.

3. Must the wells used in this experiment be dry before the CH₃COOH solution is added? Yes___ No___ Explain your answer.

4. The **end point** of the titration occurred when sufficient NaOH solution was added to cause a permanent color change in the indicator. The **equivalence point** of the titration occurred when sufficient NaOH solution was added to exactly neutralize the CH₃COOH. The accuracy of the titration depends on how close the end point is to the equivalence point.
   a. For the titration done in this experiment, determine the pH of the solution at the equivalence point.

   b. If you did the experiment correctly, the number of drops required to reach the end point using methyl red differs from the number of drops required to reach the end point using thymol blue. Which result yields the more accurate [CH₃COOH], methyl red___ or thymol blue___? Explain why your answer is correct.

   c. Record what you believe to be the correct [CH₃COOH] here: __________________________
Future updates

8- Collapsing Can. Changes to procedure and questions to discuss gas laws more than the strength of the can. DONE

17 – Hungry Pyromaniac. Including photograph of experiment setup. DONE

21 (Not included) – M&Ms – Moles and Molecules. Changes to introduction to make more comprehensible. Changes to data recording to help students think more about measurements.

24 – pH indicators. Changes to procedure to help students be more successful. Changes to a question to make less confusing.

62 – Avogadro’s Hypothesis. Minor changes on gases available.

68 – Cobalt Complexes. Minor change to question to make less confusing. Minor change to procedure to help students be more successful.

72 (Not included) – Dry Ice Under Duress – Adding figures to procedure to help students be more successful. Adding new phase diagram. Fixing questions to make less confusing so students reproducibly get the “right answer”.