DESIGNING, DEVELOPING AND ADAPTING EXPERIMENTS FOR A CHEMISTRY 2A LABORATORY MANUAL

Sabbatical Report
Submitted by
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Many people have provided invaluable support during my sabbatical leave from Mt. San Antonio College. I am grateful to the members of the Salary and Leaves Committee for granting me this opportunity to work on the project. I would like to give special thanks to Dr. Barbara Crane for generously sharing her time with me.

I have been fortunate to have the support of both the Natural Science Division and the Chemistry Department. I would especially like to thank Larry Redinger, Dean, Natural Sciences Division, and Deborah Boroch, Associate Dean, Natural Sciences Division not only for their support for this project but for their leadership over the years. My colleagues in the Chemistry Department have provided unwavering support. In particular I would like to express my sincere thanks and appreciation to Bob Roemmele for providing invaluable advice and guidance in developing the procedures and for testing three of the experiments with his students. Special thanks to Eileen DiMauro, Department Chair, for taking the time to discuss the philosophy and ideas reflected in the project and for allowing me a place on the agenda for Flex Day. Special thanks are also extended to Patricia Perez for editing parts of the lab manual. I would also like to acknowledge the support of Nicole Clark and thank her for allowing me use of the hood and bench space in the stockroom.

This project would not have been completed without the support of my family. Thank you for allowing me to take over the dining room table and occasionally turning the kitchen into a chemistry lab. I especially want to thank my husband, Keith Mason, for the illustrations and tables in the lab manual.

DESIGNING, DEVELOPING AND ADAPTING EXPERIMENTS FOR A CHEMISTRY 2A LABORATORY MANUAL

Sabbatical Proposal

Karelyn Hoover
Department of Chemistry

INTRODUCTION

The Chemistry Department has a long history of developing and implementing laboratory exercises and experiments utilizing cutting edge technology and the latest in chemical education in general and organic chemistry. The lone exception has been the introductory chemistry laboratory course, Chemistry 2A. Originally Chemistry 2A used a standard commercially published laboratory manual to accompany the text. These experiments were predictable, traditional and absent of any open ended experiments or the use of any technology. A few years ago, the Department came to the conclusion that it was time to update the laboratory portion of Chemistry 2A and a "custom published" lab manual was created, drawing on the resources of several lab manuals. While this has improved the quality of the lab experience for students, it has not been a sufficient solution. There are no open-ended experiments, the outcome of many experiments is predictable and there is a lack of consistency in the acquisition and analysis of data, as well as the quality of post laboratory exercises. The ultimate remedy is the creation of an "in house" lab manual, incorporating the latest in chemical education and taking full advantage of the resources available to students in the Mt. San Antonio College Chemistry Department.

PROPOSAL

As my sabbatical project, I propose to design and develop, or adapt from the literature, 10 experiments to be included in a Mt. San Antonio College Chemistry Department Laboratory Manual for Chemistry 2A students. My proposal includes 10 experiments as a workable number. With a few good experiments, carefully chosen from other sources, a complete laboratory manual can be created. The development process will include surveying all full time faculty of the Department as well as part-time faculty who have taught Chemistry 2A students within the past two semesters. The purpose of the survey is to determine conclusively the needs of the students, significance of the skills currently emphasized, and relevancy of the laboratory topics. This data will also be correlated with the Department's current exit skills for Chemistry 2A.

The next phase will include a search of current laboratory practices and experiments used in other facilities. The purpose of this is to learn the best practices, to become aware of the resources available and to learn about the innovations in teaching introductory chemistry. This will take place as a Web search and literature search (Journal of College Science Teaching, Journal of Chemical Education, Chemical Educator).

After collecting this data, I will begin developing the experiments. Among the 10 proposed labs five will concentrate on current needs: spectrophotometric methods of analysis (1), use of technology in the introductory laboratory (2), molecular model using molecular model kits (3), determination of the calorie content of food (4) and an open-ended exercise where the results are not simply verification of known facts or properties (5). The remaining five labs (6-10) will be determined based on the results of the survey and data gathered in the searches. The initial stages of development will begin with data collection. Once reproducible results are obtained, procedures will be written. All ten labs will have appropriate introductions on theory, procedure, data collection and analysis sheets and post lab exercises.

The complete set of ten labs will be ready for field testing in fall 2004. Two instructors will field test these labs while the remaining instructors will use the current lab manual. The first cohort of students will provide feedback in terms of readability, clarity of procedures and length of time to carry out the experiments and post laboratory exercises and appropriate level of difficulty. Although the sabbatical year will have passed, the process will continue. The labs will be revised as necessary with the end goal of a complete Department authored laboratory manual for Chemistry 2A.

The publication of a Department authored laboratory manual for Chemistry 2A is a strong need for our Department. The sole reason why this has not been accomplished previously is a lack of time. This project receives considerable support from the Chemistry Department. At our most recent Department meeting, I presented my proposal and it was met with enthusiasm by all full time faculty members in attendance.

PROJECT ACTIVITIES

1. Collect data by surveying department members
   Analyze data from surveys

2. Conduct literature search and search web sources

3. In conjunction with literature search, determine specific area needs for five additional labs (**experiments 6-10**)
   Cross match with exit skills

For each experiment:

4. Develop and test procedure

5. Prepare data collection and analysis sheets

6. Create post-laboratory exercises

7. Write introduction

TIMELINE

August/September 2003:	Activity 1 & 2
November 2003:	Activity 3 Activities 4-7 for **experiment 1**: spectrophotometric analysis
December 2003	Activities 4-7 for **experiment 2**: molecular models Activity 4 for **experiment 6**
January 2004:	Activities 4-7 for **experiment 3:** calorie content of food Activity 5 for **experiment 6** Activity 4 for **experiment 7**
February 2004:	Activities 4-7 for **experiment 4:** utilizing technology in the lab Activity 6 for **experiment 6** Activity 5 for **experiment 7** Activity 4 for **experiment 8**
March 2004:	Activities 4-7 for **experiment 5**: open-ended lab Activity 7 for **experiment 6** Activity 6 for **experiment 7** Activity 5 for **experiment 8** Activity 4 for **experiment 9**
April 2004:	Activity 7 for **experiment 7** Activity 6 & 7 for **experiment 8** Activity 5 for **experiment 9** Activity 4 for **experiment 10**
May 2004:	Activity 6 &7 for **experiment 9** Activity 5, 6 & 7 for **experiment 10**

June 2004: Submit copies to pilot instructors
(Pilot instructors will be Karelyn Hoover and Eileen DiMauro)
Write sabbatical report

BENEFITS TO THE COLLEGE

Mt. San Antonio College is a premier college. The Chemistry Department has certainly benefited from the College's support of innovation in the classroom and laboratory. We have instituted molecular modeling (a computer based activity) using SPARTAN software in General and Organic Chemistry, state-of-the-art calculator based laboratories, computer delivered instruction and the creation of an introductory chemistry course for nursing majors, to name a few. Many of these advances have been made possible by grants from the National Science Foundation and Department of Defense with Mt. San Antonio College support as well as by sabbatical leaves of other Department members. We currently use three Department authored laboratory manuals. Developing and publishing a Department authored laboratory manual for Chemistry 2A will allow our introductory course to keep pace with our other courses in terms of innovation and standards for excellence.

The Department's introductory course serves as a gateway to courses in other departments. By improving the quality of the laboratory portion of our course, we will better prepare students to continue their studies in other departments. Students who have experienced a more challenging laboratory curriculum can be expected to be better problem solvers – a benefit to all Chemistry 2A students as they pursue their educational goals here.

It has been the tradition of the Chemistry Department to donate royalties from Department authored laboratory manuals to our awards program. Each semester we honor the most accomplished students in each of our classes. Currently, our publisher provides us with $2.00 per lab manual sold. The ten experiments developed in this project are destined for publication and the royalties will be donated to the awards program. The current Chemistry 2A lab manual provides no royalties.

BENEFITS TO THE CHEMISTRY DEPARTMENT

A multitude of problems currently exists with the laboratory portion of the Chemistry 2A course. Many of the labs simply require students to verify existing known properties. In other words, the students know what the end result will be prior to performing the experiment. Many experiments can be difficult to complete in a three-hour time block leaving little time for post-laboratory discussions. Using a "custom published" lab manual has allowed us to introduce labs from a variety of sources but that creates its own set of problems. The labs vary in organization, data collection process and procedures for analysis. The current mix and match results in student confusion. They are also inconsistent in the quality and quantity of post laboratory questions and exercises.

By having a Department authored lab manual, the overall quality of the lab experiments will improve. The organization will be consistent which will reduce student confusion. The labs will be designed with our student population in mind and tailored for our equipment and facilities. Safety will be addressed by including safety, waste and disposal instructions specific to our facilities. Faculty will have an opportunity to verify the concepts and skills most important to our Department. The labs will provide experiences not available in our current lab manual. New pedagogies will be implemented based on current research. Communication skills will be improved by having students present their findings orally or with the use of PowerPoint. A Department authored lab manual will enhance flexibility to add or delete experiments in the future. Of course, our Department benefits by having well trained Chemistry 2A students, better prepared for the next level of chemistry.

BENEFITS TO ME PROFESSIONALLY

The obvious benefit is having a set of experiments that meets my needs as a teacher. I look forward to the time when data analysis instructions do not state "subtract line 3 from line 2" but challenge the student to actually analyze the data. I intend these labs to answer the often-asked question of "what is the point of this experiment?" Instead, students should be able to answer the question "what did you learn from this experiment?" I envision students engaged in the lab process.

In the past few years I have attended the Langford Quality Learning Seminar twice. I have been focused on how to bring quality learning into my classroom and now I look forward to creating the same kind of quality experience for my students in the lab. Within the last year I have been introduced to outcomes based teaching and learning. I have been working on implementing outcomes based learning in the classroom. By developing these experiments, I will be defining the outcomes for Chemistry 2A students in the laboratory. This is an opportunity for me to incorporate so many aspects of what I have learned about teaching and student learning and I am excited about the prospect.

Teaching, even if you are not a chemist, involves both art and science. I am passionate about teaching and this project is so interesting to me because it allows me to create something unique for my students. As a scientist, I enjoy the practice of science: literature searches, discovery, collecting and analyzing data, developing experimental procedures and drawing conclusions. I want my students to share with me this joy of practicing science.

ADDENDUM

Provide further explanation for the need of a year long leave

There are hundreds of lab manuals on the market today. One is not terribly different from the next. I am not interested in producing a copy of one of those manuals. I want to create something that is unique, that our students will find relevant and interesting. The process of designing and developing laboratory experiments is a complex one. Procedures must be tested and refined until they yield repeatable results that can be acquired by introductory chemistry students in one lab period. Accessibility and costs of equipment and chemicals must be considered. Health and safety concerns and the ultimate disposal of waste must be researched. Each laboratory experiment must have a meaningful introduction, procedures that are clear and easy to understand, as well as post laboratory exercises. All of this is a lengthy process.

For this project ten (10) student laboratory experiments will be designed, developed or adapted to complement the Introduction to Chemistry (Chem 2A) course. Five (1-5) topics have been chosen and five (6-10) additional topics will be determined following a literature search and faculty survey. I have selected the topics for the first five experiments based on my experience teaching Chemistry 2A for several years along with my professional interest in these topics. I have identified 16 additional topics that would be appropriate to include. The Chemistry Department faculty will be surveyed and ten total lab topics will be selected. The list of 16 topics should not be considered inclusive: faculty may propose experiment topics not listed and the literature search may reveal a novel laboratory that neither my colleagues nor I have considered. The final product will include ten laboratory experiments ready for field testing with students.

The first five (1-5) topics chosen, along with a brief explanation of why they were chosen, is given below:

Spectrophotometry

- Utilizes standard curves applicable in all areas of science as well as medicine
- Use of colored solutions results in high interest for students (most experiments use colorless solutions)

Molecular Models

- Needed for understanding bonding, shape, polarity
- Current experiments require little thinking or interpreting on the student's part (for example, molecules with all single bonds are grouped together, molecules with all double bonds are grouped together, and so on; students aren't required to distinguish between the groups)

Calorie Content of Food

- Creates high student interest in heat flow calculations

Utilizing Technology in the Laboratory

- Students are required to use various forms of technology in Chem 1A and 1B labs, but get no introduction to technology in Chem 2A

Open-ended Laboratory

- Provide an opportunity for students to discover a concept rather than follow a "recipe" to a known conclusion

How will the remaining experiments (6-10) be chosen?

The Chemistry Department faculty will be asked to rank a list of topics to be considered in terms of relevancy and skills. The list of topics includes:

Observations
Measurement
Density
Specific heat of a metal
Change of state
Properties of gases
Properties of water
Chemical reactions
Acids and bases
Neutralization titration
Cation/Anion analysis
Formula of a hydrate
Empirical formula
Ionic equations
Solutions
Oxidation/reduction

Experiment topics will not be limited to this list. Other topics as suggested by faculty may also be considered. I will conduct a literature search including the following sources: Journal of College Science Teaching, Journal of Chemical Education, Chemical Educator. The literature search may provide additional topics.

PROJECT ACTIVITIES

1. Select Experiment Topics

Survey Chemistry Department members to determine which topics are most relevant. The data from the surveys will be analyzed and the topics will be prioritized. The exit skills for Chemistry 2A laboratory will be reviewed. The topics will be correlated with the exit skills.

2. Conduct Literature/Web Search

Become knowledgeable about innovations in the Introductory Chemistry laboratory by searching available resources. Correlate findings with selected experiment topics. Adjust topic selection, if needed.

3. Design, Test and Refine Experimental Procedures

Develop experimental procedures for ten laboratory experiments. Test and refine procedures until data is reproducible.

4. Write Complete Laboratory Experiments

Each laboratory experiment will include:

- Introduction (purpose and objectives of experiment as well as introduction to theory of the concept explored)
- Procedure for Experiment (instructions on how to perform experiment, clean up, and waste disposal)
- Data and Analysis Sheets (provides format for students to record data as it is collected and instruction on analysis of data)
- Post-Laboratory Exercises (reflective exercises and/or questions that direct the students towards application and deeper understanding)

5. Laboratory Manual

- Compile a manual of ten laboratory experiments as described above.
- Submit copies to pilot instructors (Pilot instructors will be Karelyn Hoover and Eileen DiMauro)

6. Sabbatical Report

Write and submit a sabbatical report detailing the progress and results of the project.

TIMELINE

August/September 2003:
- Select Experiment Topics (Activity 1)
- Begin Literature Search (Activity 2)

November 2003 – February 2004:
- Continue Literature Search (Activity 2)
- Design, Test and Refine Experimental Procedures (Activity 3)

March – April 2004:
- Write Complete Laboratory Experiments (Activity 4)

May 2004:
- Compile Laboratory Manual (Activity 5)
- Submit copies to pilot instructors (Activity 5)
- Write Sabbatical Report (Activity 6)

Please further explain how profits from sale of the manual will be distributed

Three lab manuals currently in use in the Chemistry Department are "Department authored." No one in the Chemistry Department personally receives any royalties for the lab manuals. Instead, a check is given by the publisher to the Chemistry Department and deposited in an account with Auxillary Services. At the close of each semester, one student is selected from each Chemistry class and is honored as an Outstanding Chemistry Student. With the honor is a check from the Chemistry Department's account with Auxillary Services.

I referred to the laboratory manual I propose to create as "Department authored" to further indicate my intentions. The Chemistry 2A Laboratory Manual will be published by the same publisher and the royalties dispersed in the same manner. All royalties will be included in a check to the Chemistry Department, deposited in the Department's account with Auxillary Services and used to support the Outstanding Chemistry Student award. I have no interest in profiting personally from the laboratory manual. This is an "in house" publication sold only to Mt. San Antonio College students.

Correction to original proposal

Page 2 of the original proposal states "The complete set of ten labs will be ready for field testing in fall 2003." This is a typographical error and has been corrected to read "The complete set of ten labs will be ready for field testing in fall 2004."

STATEMENT OF PURPOSE

The purpose of this sabbatical project is to develop ten (10) laboratory experiments for the Introductory Chemistry course (Chemistry 2A). The ultimate purpose is to provide students with a laboratory experience that is meaningful and relevant to them now and useful in achieving their future goals.

Driving this project is a need for change in laboratory instruction. The need to revise Chemistry 2A has been a topic of discussion in Department meetings for several years. The laboratory curriculum in particular was identified as needing to be overhauled. The laboratories of both Chemistry 1A, 1B and 2B had been revised and updated but Chemistry 2A was left virtually untouched. The time had arrived to address the needs of Chemistry 2A.

A disconnect often exists between lecture instruction and laboratory instruction. It is possible for a student to correctly solve a mathematical problem involving the gas laws yet not connect that to the observable behavior of a gas in the laboratory. A disconnect also exists between laboratory experiments and everyday phenomena. A student might collect and process data in the laboratory on the behavior of a gas and have excellent results but not understand why a balloon popped on a very hot day or why soft drinks go "flat." In addition, the choice of materials often used in the procedures lacks a connection between the experiment and everyday substances with which students come in contact.

Students determining density in the lab do not find the density of a rubber stopper particularly relevant. Thus, a change is needed that will better connect laboratory with lecture and connect a formal chemistry course with everyday life.

Chemistry 2A laboratory periods are three hours in duration. During that time students are introduced to the experiment, collect data, analyze it and draw conclusions. Students hand in reports at the end of the lab session. There is not sufficient time to "debrief" many of the labs. Adequate time at the end of the procedures would allow students to discuss the lab results with each other and with their instructor. A change in the length and complexity of laboratory procedures is needed to allow time for the instructor to discuss the results and conclusions with the class.

Working with 28 novice chemists in a laboratory is quite challenging for instructors. Students learn laboratory techniques, data collection and analysis, how to draw appropriate conclusions, and safety and waste disposal procedures. Instructors actively engage students and monitor results. Changing experimental procedures and report sheets so that students must check their data and analysis with their instructor before proceeding to the next step provides another opportunity to engage students and re-teach skills as necessary. Students who find it difficult to ask questions will have support provided simply by following the class protocol and there will be no refuge for those students who prefer to "hide." A change in procedures, reporting and protocol will identify students who need help with particular skills and allow the instructor to more closely monitor the progress of the class.

Health, safety and waste disposal are major concerns in any instructional laboratory. Reducing the amount of hazardous waste produced and reducing the toxicity of the waste are ongoing priorities. Selecting laboratory reagents that are environmentally friendly and providing instructions on handling and disposal of waste help achieve these priorities.

This sabbatical project addresses all of these concerns in the developing of ten laboratory experiments. Students will conduct experiments that have been written to support their learning in lecture, that have connections to everyday phenomena and that have been designed to engage them. Wherever possible, materials and reagents have been chosen that are both familiar and environmentally friendly. Again, the ultimate purpose is to provide students with a laboratory experience that is meaningful and relevant to them now and helpful in meeting their future goals.

PROJECT ACTIVITIES

SELECT EXPERIMENT TOPICS

During August and September 2003 the process of selecting the topics that would eventually become laboratory experiments was initiated. Surveys were sent to all full-time faculty members and to part-time faculty who had recently taught Introductory Chemistry (Chemistry 2A). The survey included 17 potential topics and faculty were asked to rank them in order of importance (see survey in Appendix C). Some instructors had difficulty differentiating their priorities and ranked several labs equivalently. Three labs emerged as clear priorities: Observations, Measurements, and Chemical Reactions.

During the previous semester, the Chemistry Department made a decision to eliminate the use of mercury filled J-tubes due to the hazard posed by exposure to mercury. Two separate experiments (already in use) utilized the J-tubes in procedures for gas laws. Plastic syringes (filled with air instead of mercury) were purchased by the Department to replace the J-tubes but no experiments had been developed. This left the Department without any experimental procedures for gas laws and it was clear that topic should be added to the list.

During the fall 2003 semester various instructors approached me with requests for specific laboratory topics to be included in the project. Specific Heat, Paper Chromatography and revision of the Department authored laboratory, *Chemical Search*, were added to the list.

The list of topics was narrowed to Spectrophotometry, Molecular Models, Calorie Content of Food, Utilizing Technology in the Laboratory, Observations, Measurements, Chemical Reactions, Properties of Gases, Density, Cation/Anion Analysis, Specific Heat, Chemical Search, Acids and Bases, Ionic Equations/Reactions, Graphical Analysis, Solutions, and an experiment related to Coffee serving as an Open-Ended experiment. A literature search was initiated.

SEARCH THE LITERATURE

A literature search began and continued through February. As important as the topics themselves was the goal of developing laboratory experiments that:

- support lecture instruction
- students can connect with everyday occurrences
- students find relevant to their futures
- are engaging
- are environmentally friendly

One of the things I learned was that subscribers to the *Journal of Chemical Education* have access to a web search engine that is specific to instructional laboratories. A number of very promising laboratory experiments were considered. Most were rejected because they required expensive equipment, could not be completed in a typical three-hour lab period or could be completed in less than two hours, did not correlate well with lecture instruction, or used reagents that were not suitable for Chemistry 2A students. One disappointment was Coffee. Many students (and some faculty) practically live off coffee and it would have been interesting to have an experiment that determined the caffeine content of coffee or compared the caffeine in different brands or compared decaf to regular. The problem was that the experiments all used dichloromethane, a toxic chemical and suspected carcinogen. It did not seem appropriate to have novice chemistry students handling such a reagent. In the end, four laboratory experiments were selected to be adapted.

The first of the four experiments included a titration to determine vitamin C content in citrus fruits. Titration of vitamin C is not an uncommon laboratory exercise but this lab had a novel approach using actual fruit. Students were given a *Survivor*-type scenario in which they had to choose what type of citrus fruit to take with them on a sea voyage. The second exercise was based on a photographic essay depicting students engaged in unsafe laboratory practices. Students were to identify what was wrong in each photo. I thought this approach could be well adapted to revise the Chemical Search exercise currently used in Chemistry 2A laboratory. A third considered the color composition of M&M®s candies, which could be adapted for the Spectrophotometry lab. The last exercise was a qualitative analysis scheme of common household substances that could be used as the Open-Ended experiment.

Following the literature search, the lab topics considered for development were the original five identified in the Sabbatical Proposal (Spectrophotometry – using M&M®s candies, Molecular Models, Calorie Content of Food, Utilizing Technology in the Laboratory, Qualitative Analysis as an Open-Ended exercise), three identified from the faculty survey, (Observations, Measurements, Chemical Reactions), Properties of Gases because of a change in Department equipment, three specific requests from Department members (Specific Heat, revision of Chemical Search – adapting the photographic essay approach, Paper Chromatography), Titration of Vitamin C because of its novel approach and two additional experiments that interested me personally (Density, Graphical Analysis). Fifteen ideas for experiments were about five ideas too many. Realizing that some ideas look better on paper than in real life, I decided to begin the design and test phase and let some of the ideas eliminate themselves in the process of going from paper to actual lab procedure.

Conducting a literature search turned out to be very inspiring. Reading about innovations made me anxious to try out new ideas. I learned a number of things by reading so many articles: I felt most comfortable reading laboratory discussions and procedures that were written directly to the reader instead of using the typical scientific language that discourages the use of personal pronouns; simple language is more effective; less is not necessarily more when it comes to discussion. These ideas were incorporated in designing and writing the laboratory experiments.

DESIGN, TEST AND REFINE EXPERIMENTAL PROCEDURES

With the literature search underway, it made sense to begin designing, testing and refining experimental procedures. Reproducible results were obtained for all experiments and the lab procedures written (and re-written) in language that was clear and user friendly. Procedures included specific instructions on safety, handling of materials and waste disposal.

During this phase it was expected that some experiments would be deemed more appropriate for the target audience than others and the process of selecting ten experiments from fifteen would occur naturally. As development proceeded I recognized that I was making a philosophical decision on behalf of the Department by the nature of the choices made to include some lab topics and exclude others. I was also making a pedagogical decision by the way in which procedures were being developed and eventually the context in which they would be written. This, of course, was the point of the sabbatical project to begin with and the project enjoyed full support of the chemistry faculty. Nonetheless, I did not want to overstep my bounds. I discussed these issues with the current Department Chair, Eileen DiMauro, on a continuing basis and presented my ideas to the Department at the Flex-Day meeting in January 2004.

At the Department meeting I was particularly interested in getting feedback on the following: choice of experimental topics to make lab more meaningful to students; limiting introductory lecture and procedures to 2.5 hours to allow 0.5 hours for debriefing; writing pre-laboratory questions for experiments correlated more directly with the introduction; instructor checkpoints to assist students as they progress during the lab; students working individually without the benefit of a partner. Faculty were very supportive of the choices for lab experiments and offered to pilot any labs that might be ready for spring 2004.

The idea of allowing adequate time at the end of the lab session to debrief was welcomed. I was convinced that by writing introductions that discussed each area of the experiment, instructors would require less time at the beginning of the lab session to set the stage and that would leave more time at the end of the session for meaningful discussions. Pre-laboratory questions would be directly related to the introductions. If lecture instructors had not yet discussed the material in the introductions, the students would still be prepared for the laboratory. If the lecture instructors had discussed the material, no lecture was needed and the students could skip to the pre-lab questions. If a student could answer all of the pre-lab questions, through classroom instruction and/or through reading the introduction, the student would be prepared for the lab experience. It would be necessary then for instructors to hold students accountable for the pre-laboratory assignments. This represented a change for some faculty. At the end of the departmental discussion the consensus was this was the right direction to take the experiments.

The discussion turned to checkpoints in the procedure. The idea that the progress of each student would be monitored was appealing. Checkpoints become more important if students work individually during lab. It had been the Department preference to have students work in pairs to increase their comfort zone and to experience being part of a team. The down side of this is that some students are less confident of their work because they are relying on their partner. Instructor checkpoints would alleviate the need to work in pairs and would still allow students to feel supported. Some concern was expressed that the process would create delays in the lab with students waiting in line for an instructor's signature. At the conclusion of the meeting the idea of including checkpoints in the procedures was met favorably.

The following is a list of experiments that were included in the final draft of the laboratory manual. A brief discussion of the development of each lab is included. See Appendix A for complete laboratory experiments.

Properties of Gases

This experiment was designed to study the relationship between pressure and volume of a gas (Boyle's law), to study the relationship between temperature and volume of a gas (Charles's law) and to quantitatively determine the value of the temperature absolute zero. In addition, a physiological application of Boyle's law was examined.

A plastic syringe filled with 25 cc of gas (air) was mounted in such a way that allowed textbooks to be placed on the syringe compressing the gas inside (1). As each textbook was placed on top of the syringe, the change in volume was measured. A graph of pressure (in "textbook units") vs. volume was plotted (the mass is proportional to the pressure). The graph showed an indirect relationship consistent with Boyle's law. The graph was a curve rather than a straight line. Air is not an ideal/perfect gas and a curve was predicted. It was possible this result was due to the varying mass of textbooks. Original instructions that accompanied the apparatus called for 1 kg masses to be used. Using 1 kg masses for even a single section of students would be cost prohibitive for the Department. The experiment was repeated using 1 kg masses to determine if better results could be obtained. The graph of pressure (1 kg masses) vs. volume also yielded a curve. It was concluded that the results were due to the non-ideal nature of air and were acceptable within experimental error.

A lung demonstrator model was used as a physiological application of Boyle's law (2). Balloons were used to simulate lungs. A change in pressure caused the "lungs" to change volume. This was incorporated into the laboratory to make the experience more relevant and meaningful to students. Many of the students in this course identify themselves as pre-pharmacy, pre-medicine and pre-physician assistant; therefore it is hoped a physiological application will peak student interest.

The second part of the experiment explores the relationship between the temperature of a gas and its volume. According to instructions included with the syringe (1), exposing the syringe to a heat source (150 watt bulb) will cause the air in the syringe to expand and the corresponding volume change can be measured directly from the syringe. This did not produce the expected results. The distance from the heat source to the syringe was varied. The syringe itself became very warm to the touch and the temperature of the surrounding air increased but the syringe did not move. A 250-watt IR bulb was substituted for the 150-watt bulb. It melted the apparatus.

It was concluded that the friction in the syringe prevented movement. A new apparatus was obtained and lubricated and the procedures repeated. No observable change in volume. The syringe was submerged in boiling water. The syringe moved 1 cc, barely observable. Theoretically, the volume of air should have expanded 6 cc, which would clearly be observable. The conclusion was too much friction within the syringe. To eliminate the possibility that the syringe used was somehow defective, a third syringe was obtained and the procedures repeated. The results were similar. This substantiates the conclusion.

An alternative method for observing the relationship between temperature and volume of a gas was chosen (3). This is a qualitative experiment rather than a quantitative one. Air was trapped in a capillary tube using a drop of dye or food coloring. The change in the volume of air was observed as the capillary tube was placed in ice water followed by warm water. The change in volume with changing temperature showed a direct relationship consistent with Charles's law. Results were easily reproducible.

The third part of this experiment is an instructor demonstration using an absolute zero demonstrator (4). The apparatus consists of a vessel that contains a gas (air) and a pressure gauge. The apparatus was alternately submerged in boiling water, ice water and slurry of dry ice and isopropyl alcohol. The apparatus measured the pressure and a plot was made of pressure vs. temperature. The line was extrapolated to reveal the temperature at which there was zero pressure. This is known as absolute zero. The graph produced acceptable results.

During development of this experiment it was noted that in addition to discussions regarding gas laws, the mathematical concepts of direct vs. indirect relationships would need to be discussed.

Measurements

The purposes of this experiment are to introduce students to the chemical laboratory, introduce them to the metric system, teach students how to make laboratory measurements and how uncertainty arises from measurements, teach students how to graph data and make predictions using their graph, and use their new found skills to solve a practical problem.

Typically, the first hands-on lab experiment was spent with students learning how to use simple laboratory equipment. Students measured temperature using a thermometer, measured the volume of liquids using a graduated cylinder or other volumetric device, determined the mass of an object with a balance and used rulers to measure length. The problem with this approach was that students collected data that was not used for any purpose other than to learn how to use the equipment. These introductory chemistry students were overwhelmed learning several new skills in one session; further, many of the skills were not used again for several lab sessions. The *Measurements* experiment developed differs in that only the equipment actually needed to collect data for the immediate experiment is introduced. The skills acquired were then used to solve a practical problem. Additional lab techniques will be acquired by students on a need-to-use basis in later experiments.

The key to this experiment was designing a practical problem for students to solve utilizing the skills learned. A scenario was created in which students are asked to determine if a wood sample used to make children's toys is tainted and if the toys should be ultimately recalled. Procedures were developed to measure the dimensions of wood blocks, calculate the volume of the blocks, measure the mass and plot the mass vs. volume. Data collection and graphical analysis were modeled after a series of exercises (5). The students use their skills learned in the procedures to determine if the unknown wood sample provided them is the same as the known wood blocks or if it is different and possibly tainted. Students decide if the wooden toys should be recalled. Results are reproducible. Graphs yield straight lines. Percentage differences are about 3%, which is acceptable within experimental error.

Utilizing a graph in this experiment allowed graphical analysis to be placed in context and there is no need for a separate experiment on graphical analysis. (One of the "extra" five experiments deleted)

During the development of this experiment it was decided to use the idea of case studies instead of examples or sample data. This is a small change but it is included to give the students a sense of the context of the experiment as opposed to collecting data for the sake of collecting data and to parallel the case study approach in the social sciences.

Density

This experiment utilized the graph of mass vs. volume from *Measurements* to introduce the concept of density as a ratio of mass to volume (5). The straight line of the graph illustrates that the blocks have the same density because they have the same ratio of mass to volume. This is a difficult concept for students to master. Students confuse size and density. Heavier samples of the same material are thought to be more dense. This lab is designed to help students conceptualize the nature of density. In addition, students learn two new skills for measurement: determining the volume of a liquid sample using a graduated cylinder and determining the volume of an irregular solid using the water displacement method. Both skills are used to solve practical problems.

As in the *Measurements* experiment, creating the scenario to put the skills learned into context was key. Two ideas were explored: Diet Coke™ and Coke™ might be distinguished from each another by virtue of their densities and gold was used as a coin metal because of its density.

Theoretical calculations indicated that Coke™ would be about 11% more dense than Diet Coke™ and therefore the two drinks could be distinguished from each other. Experimental results indicated that there is only about 1% difference in the densities. To explain the smaller percentage difference several factors were considered, including the sample size as compared to experimental error and the size and quantity of carbon dioxide bubbles within the samples. Changing the parameters of the experiment produced the same results – about 1% difference. It was concluded that one or more of the assumptions in the theoretical calculations was incorrect. A decision was made to keep the original procedures in the experiment. The scenario is used more as a question as to whether the two soft drinks can be distinguished on the basis of their densities and students are directed to conduct a taste test at the end of the experiment. The high profile of the soft drinks assures a connection between the laboratory and everyday life. The students are still required to utilize their laboratory skills to solve a practical problem.

Gold can be distinguished from other metals based on its density. The scenario was that the authenticity of a gold coin or gold nugget could be verified using the water displacement method to determine density. In order to preserve the departmental budget, the "gold" would be fake and the students would discover the fake. Coins painted gold and "fool's gold" were used to simulate gold coins and gold nuggets. Neither proved satisfactory. The samples were too small to produce observable volume changes. Increasing the sample size caused them to get stuck inside the graduated cylinder. A chain painted gold solved the problems and the scenario changed to determining the authenticity of the gold necklace. The results were reproducible.

Specific Heat of a Metal

The purpose of this experiment is to introduce students to the concept that heat transfer is dependent on the nature of the material being heated. Students are also introduced to the difference between heat and temperature.

The specific heats of four different metallic samples, copper, brass, iron and aluminum were measured. Data was collected using the metal samples and 150 mL of water. The results were within acceptable experimental error and reproducible.

The procedures require the students to determine the specific heat of a known sample of metal and check their results with their instructor. Once competency has been established, the students are issued an unknown and must determine its specific heat.

This experiment was completed during the spring 2004 semester in time to be given a trial run. The laboratory instructor provided feedback. The percentage error was unacceptable, averaging 20% - 30% (6). The procedures were repeated using 50 mL, 75 mL and 100 mL to determine the optimum volume of water. The greatest accuracy (lowest percentage error) was obtained using 75 mL of water. It was concluded that the larger volume of water produced the smallest temperature change and therefore the potential for the greatest error in reading the thermometer. It was also concluded that 75 mL of water generated a sufficient temperature change while allowing enough water to adequately cover the metal. The procedures were revised.

The laboratory instructor also provided feedback on checkpoints. He reported that they did not create delays in the lab, students were not left waiting for his signature and it was helpful in monitoring the progress of students.

For a complete description of the lab, please see Appendix A.

The Color of the Candy Coating (Spectrophotometry)

The goal of this exercise was to design a lab where students would investigate the chemistry of color. Materials that students see or use on a regular basis were sought. Kool-Aid® was considered but students in Chemistry 1A perform an experiment using Kool-Aid® and this would be too repetitive. I liked the idea of using M&M® candies and had discovered a laboratory exercise using M&M®'s during the literature search (7). The procedures have been adapted for *The Color of the Candy Coating.*

Spectrophotometers are generally used in the teaching laboratory to create standard curves and identify the concentration of unknown solutions. The maximum absorption of the solution is provided in the procedures. The chemistry of color is never mentioned or only touched on briefly. How maximum absorbance is determined is left unanswered. The purpose of this lab is to explore the properties of color: reflection, absorption, transmittance, and complementary colors. Maximum absorbance is determined by the student, not provided in the procedures.

The color was extracted from the candy coating of a brown M&M® candy using distilled water. The absorbance of the solution was measured from 340 nm to 720 nm (the entire visible region) in 20 nm increments using a Spectronic 20 spectrophotometer. The maximum absorbance was determined to be 480 nm. Some difficulty was encountered while collecting data. The procedures were repeated with a second Spectronic 20 and it was determined that the first machine was malfunctioning. The results were reproducible.

The development phase took a turn at this point. The procedures were repeated for solutions of other M&M® candies and were to be repeated using dyes as standards. The color composition of brown would be determined using the data collected. The problem was this was tedious and boring. Not the engaging lab I had planned. In addition, the correlation between the absorbance values of a solution made of red, yellow and blue dye with the absorbance values of the brown M&M® candy solution was not obvious. According to the literature (8), the correspondence between the peaks "is easily observed." I did not find this to be the case and it would not be observable for introductory chemistry students. Instead, I chose to use the absorbance values of the brown solution as an investigation into the nature of color.

Paper chromatography was added to the experiment. Solutions of brown, blue, green, orange, yellow, and red M&M® candy coatings were prepared. Solutions were concentrated by evaporation. Paper chromatography was performed on all solutions using distilled water as the developing solvent. Results of the paper chromatography were: yellow, red and blue showed no evidence of separation; orange showed only a slight separation of orange and yellow; green showed a clear separation of blue and yellow; brown separated into red, yellow and green. The separation of the green candy coating indicates that it is a mixture of blue and yellow dyes. The orange coating shows evidence of a mixture but the red dye was not clearly resolved. The separation of brown indicates that it is a mixture of red, yellow and blue dyes. The green color is unresolved blue and yellow dyes. The remaining coatings are composed of dyes that are pure substances as opposed to mixtures. Students are asked to correlate the results of the paper chromatography with the results of the spectrophotometry.

Incorporating spectrophotometers in this laboratory is the first time technology has been used for data collection in Chemistry 2A. As a result, a separate laboratory utilizing technology in the laboratory is not necessary.

Calorie Content of Food

Several years ago I came across a conceptual physics book. The text suggested that a "fun and impressive demonstration of the energy in food is to take a walnut meat, spear it onto a paper clip, and light it. It will quickly boil 50 mL of water" (8). I have performed this demonstration over the years for students in my chemistry classes and it has been a crowd pleaser. That demonstration forms the basis of this experiment.

The purpose of this experiment is to put heat flow calculations in a context that is interesting and relevant to students. Most students are interested in food and nutrition and are curious about the values in nutritional labels. Few students are aware that chemists also use the calorie unit and its relationship to the nutritional Calorie. *Calorie Content of Food* is designed to teach heat flow calculations, use of a calorimeter, the difference between the scientific calorie and the nutritional Calorie, the limitations of experimental procedures and the sources of error.

Several designs for the calorimeter were explored. The final design is featured in the full laboratory description. Several types of food items were burned, the resulting heat transferred to the water in the calorimeter and the Calorie content of the food was determined. Among the food items tested were candy bars, walnuts, peanuts and other types of nuts, goldfish crackers, and marshmallows. The candy items melted and did not burn completely. Goldfish crackers did not burn completely. This resulted in large percentage errors. Of the nuts tested, walnuts generated the most consistent heat. The best results using a marshmallow were obtained when a large marshmallow was cut in quarters. Presumably, this exposes more interior surface area to the flame and allows for more consistent burning. Percentage error was 20% - 30%, which is acceptable under the experimental conditions. Results were reproducible.

Students are asked to determine the Calorie content of a walnut and a marshmallow. These food items were chosen for two reasons: both produce consistent heat when burned and one is a natural food and one is human-made. This allows students to consider sources of error in working with a natural food. Quality control in human-made food means the samples will be consistent compared with natural food. This experiment has a number of sources of error that should be obvious to students. This is exploited to teach students about error in experimental procedures.

Chemical Reactions

The purpose of this experiment is twofold: first to teach students to recognize that a chemical reaction has occurred using four criteria (formation of an insoluble solid, release of bubbles, change in color of a solution, change in temperature) and second to predict whether a reaction will occur when two ionic solutions are mixed and confirm with experimental mixing of reactants. When students determine a chemical reaction has occurred, they predict the products of the reaction and write a balanced chemical equation.

The use of spot plates for multiple tests is common in clinical labs as well as teaching labs (9). Spot plates were chosen for this experiment to coordinate multiple testing. For the first part of the experiment two ionic solutions, Na_2CO_3 and $AgNO_3$; one acid, HCl; one base, NaOH; an indicator, phenolphthalein, and two solid metals, Mg and Zn, were chosen. The reagents were chosen because of their chemical properties. Three of the four criteria used to evaluate chemical reactions were observed from this set of reagents. Reproducible results were obtained.

For the second part of the experiment, four nitrate solutions, $Pb(NO_3)_2$, $Hg(NO_3)_2$, $Ca(NO_3)_2$, and $AgNO_3$, and three halide solutions, KBr, KI, and NaCl were used. Compounds containing potassium and sodium ions are generally soluble. Compounds containing nitrate ions are also generally soluble. Compounds containing halide ions tend to be soluble except those of silver, lead and mercury ions (10). These compounds were chosen to exploit these properties. The results allow students to examine the trends in the reactions and predict the products. Reproducible results were obtained.

Spot plates increased the efficiency of performing multiple tests. Their use was limited in determining a change of heat. Test tubes allow for observation of heat change by direct touch but this was not possible with spot plates. A decision was made to eliminate the use of temperature change as a criterion in favor of the more easily used spot plates.

This experiment requires extensive observation on the part of the student. It also requires that conclusions be drawn based on these observations. The experiment allows the placement of these concepts, observations vs. conclusions, into an experimental setting in lieu of having students perform a "theatrical" experiment.

Molecular Models

The goals of this laboratory exercise were to write electron dot (Lewis) structures from the chemical formulas, construct 3-dimensional models of molecules with single, double and triple bonds, make 3-D sketch of the molecule based on the model, and determine the molecular polarity from the structure and shape of the molecule.

This exercise was designed to be a departure from previous modeling exercises. Students must determine if the structure of the molecule contains single, double or triple bonds. Students are presented with two methods for determining the structure from the formula. Students are also presented with a way to determine polarity from the structure and introduced to the chemical consequences of polarity.

This experiment was completed during the spring 2004 semester in time to be given a trial run. The laboratory instructor provided feedback. The procedures had the desired effect of teaching students how to determine the number and type of bonds. However, there was insufficient practice. The procedures were revised.

What is a Mole?

The purpose of this experiment is to introduce students to the concept of a chemical mole and counting by weighing. Each semester I introduce Chemistry 2A students to these concepts by having them guess the number of jellybeans in a jar. That activity formed the basis of this experiment.

A jar full of jellybeans is presented to the class. Students are given access to the jar with jellybeans, an empty jar, a supply of jellybeans and a balance. They develop their own procedure for determining the number of jellybeans in the jar. Students report their findings to the class and the method of counting by weighing is established.

Students are then asked to weigh several containers with a dozen objects in each container. The containers contain common objects such as candies, paper clips, beads, etc. Students are then asked to predict what 100 objects would weigh. The results of this exercise are used to launch the concept of the mole.

This experiment was also prepared in time for a trial run in the spring 2004 semester. Student feedback was very favorable as reported by the laboratory instructor. It did accomplish the goal of introducing an abstract concept in a concrete method. Students reported to their lecture instructor that the laboratory experience was helpful in understanding the mole.

Qualitative Analysis

An open-ended experiment was needed to complete the set of ten experiments. I searched for a lab that would require the students to use what they had learned in previous experiments and challenge their critical thinking skills. The purpose of this experiment is to have the students analyze the properties of 14 common household substances and identify two unknown substances. The larger goals are to have students connect chemistry to everyday life, have students apply the skills they have acquired in previous procedures to solve problems and think through a logical scheme of separation. The reagents used are household substances. The laboratory is sufficiently challenging to meet the goals.

In a typical qualitative analysis scheme, ions are eliminated through a process of selective precipitation. The procedures for this experiment have been adapted from the literature (11) and differ from the typical in that the components are not eliminated through selective precipitation. Each of the 14 substances is tested individually for its chemical and physical properties. Observations are made and recorded. Students are required to identify two unknown substances by comparing their chemical and physical properties with the known substances. The first unknown has only one component; the second unknown has two components. Once students have identified their unknown, they report their findings.

In the original journal article (11), two unknowns were used: commercial antacid tablets and baking powder. Although the idea of using common household substances as unknowns is appealing, students can identify the unknowns by reading the labels of the products. The unknowns might be a mystery for the first couple of semesters, but soon after their identity would be obvious. To be able to use this lab for consecutive semesters, unknowns would have to be developed.

Problems arose during the creation of unknowns. Both water-soluble and insoluble carbonates were used in the separation scheme. Use of a carbonate as a component of a dual component unknown resulted in the incorrect identification of the second component. It was determined that carbonates could be used only as a single component. Sodium hydroxide was incompatible with other compounds and therefore could be used only as a single component unknown. Some tests interfered with the detection of the second unknown component. Four pairs of components were identified as workable for dual component unknowns. Any of the 14 common household substances analyzed were suitable for use as single component unknowns. Results were reproducible.

WRITE COMPLETE LABORATORY EXPERIMENTS

Following the design phase, complete laboratory experiments were written. During the writing process an emphasis was placed on user-friendly language directed at the student instead of the usual scientific language that avoids the use of personal pronouns. Introductions were written with rationales for studying the phenomenon in the experiment. Discussions were organized and written to allow the experiment to stand alone, independent of the text. Safety and waste disposal were included with instructions specific to waste collection in the Mt. San Antonio College Chemistry Department. Procedures were compiled in step-by-step format. Instructor checkpoints were included where appropriate and especially when a new technique was introduced. Report sheets were written to help students organize their data and analysis. Much attention was paid to post-laboratory questions developed to require students to apply and reflect on what they learned. Prelaboratory questions were coordinated with the discussion sections. Important discussion points had corresponding questions in the prelaboratory section. All experiments were developed and written with the Chemistry 2A student in mind.

Experiments were written with the following elements:

- Introduction to establish relevancy
- Discussion topics including background information necessary to perform all parts of the experiment independent of the text
- Safety and waste disposal
- Procedures
- Report sheets (data and analysis)
- Post-laboratory questions and problems
- Pre-laboratory questions and problems

For details, please see Appendix A.

LABORATORY MANUAL

A laboratory manual has been completed (see Appendix A). Copies have been submitted to the two pilot instructors, Jody Williams Tyler and Janet Truttman. The proposed pilot instructors, Karelyn Hoover and Eileen DiMauro, are unable to pilot the labs due to scheduling conflicts. Jody Williams Tyler and Janet Truttman volunteered to replace them.

CONCLUSIONS

I have achieved the goals proposed when I began this sabbatical project. Ten experiments were designed, developed or adapted. I am pleased with the quality of the experiments including the writing. I am satisfied that students will find the experiments both relevant and engaging.

I have had input from all department members both formally, in department meetings, and informally, in discussions with individual faculty members. The Department was very supportive throughout the entire process. Three of the experiments were completed in time for trials in the spring 2004 semester. The feedback was very positive and the suggestions for improvements were incorporated in the final drafts. The experiments are on target for field-testing with four lab sections during fall 2004. Revisions will be made and tested again in spring 2005. The Department will have the opportunity to have input before the final adoption in fall 2005.

Personal Benefits

This was one of the most rewarding experiences of my career. I am very proud of the finished product. I believe it hits just the right mark with format, style and content. I appreciate the opportunity to work on a project of this scope. I consider it a privilege to have been given this time. I am pleased to be able to make this contribution to the Chemistry Department's instructional program.

I return to the Chemistry Department refreshed and invigorated, looking forward to working with students and colleagues. I am even looking forward to Department meetings.

Benefits to the Students

The most obvious benefit is the creation of ten laboratory experiments designed with their needs and interests in mind. Students will perform procedures designed to help them solve practical problems. A bridge has been built to span the concepts of chemistry and everyday occurrences. I envision students asking, "What in the world ***isn't*** chemistry?"

During every phase of the project, I had the opportunity to reflect on how the final product would affect students. Checkpoints have been added so that students can verify their results before proceeding. Students who are reluctant to seek help when needed will find the checkpoints more in their comfort zone. Everyone's progress will be monitored and no one need feel different. Complete discussion sections have been included so that students have all the necessary background material available to them to be successful in the laboratory. Prelaboratory exercises were written so that students can check their understanding. Postlaboratory exercises challenge students to apply what they have learned. Case studies have been used to place examples into context.

For the first time Chemistry 2A students will use technology in the laboratory. Spectrophotometers are used to teach students about the chemical properties of color in *The Color of the Candy Coating*. This is a logical introduction to the use of spectrophotometers in Chemistry 1A where students explore the properties of solutions by exploiting their color. For the first time Chemistry 2A students will have an open-ended exercise. In *Qualitative Analysis* students will determine which tests to perform and develop their own scheme to identify an unknown substance. Waste and disposal instructions have been included in the laboratories for the first time. Materials for all experiments have been chosen to reflect the trend of using more environmentally friendly reagents. The use of mercury in J-tubes open to the surroundings has been eliminated.

Benefits to the Instructional Program

The need to revise Chemistry 2A has been an ongoing discussion in the Chemistry Department for more than a decade. The fall 2003-spring 2004 academic year brought those discussions to the action phase. The completion of this project dovetails with the Chemistry Department's plans for revision. The experiments reflect the change in pedagogical philosophy and department members had input from the selection of the experiments to the organization of the written labs.

Direct benefits to the instructional program include lab sessions that will now have sufficient time to debrief students; making the lab experience more meaningful; experiments that have been specifically tailored for the Mt. SAC Chemistry Department in terms of needs, equipment and facilities; access to ongoing revisions not possible with a commercially prepared lab manual; elimination of the use of mercury; reduction in the use of toxic materials; instructions to the students on safety and waste disposal. These are all things specifically requested or endorsed by the Department.

Appendix A

Complete Laboratory Manual

MEASUREMENTS

INTRODUCTION

Why would you want to know about measurements? Would you buy clothes without knowing your size? Engines can be measured in cubic inches or liters. Does the volume (size) of the engine make a difference to you? If you are getting your hair cut and your hairdresser suggests a two-inch trim, would you like to know how long your hair will be after its cut? Or you may be getting your hair buzzed in which case you might want to know the difference between the one-inch comb, the half-inch comb and the quarter-inch comb. If you replace the cabinets in your kitchen or the carpet in your living room, you will need to measure.

If you were born in the United States, you are probably familiar with the English system of measurement. Inches, feet and yards are used for most measurements of length. The yard is the oldest unit of length and is based on the body dimensions of a 12th century English king. The yard was the distance from the tip of his nose to the tip of his outstretched arm. What if smaller measurements needed to be taken? Another 12th century king approved a smaller unit called the inch, which is the length of three grains of barley laid end to end. The foot came along later, which was to be shorter than the yard but longer than the inch. Since there are twelve inches in one foot and three feet in one yard, measuring in these units can involve mutliplying or dividing by three or twelve. If the measurements involve part of an inch, you will have to use fractions, such as a16th or 32nd of an inch. These units are not very convenient. In addition, few countries of the world outside the United States currently use the dead king's English system. This makes it difficult for the U.S. to communicate with other countries.

The metric system or the International System of Units (SI) was developed in France and has been adopted by most countries (including England) for everyday use. Many companies in the United States use both the dead king's English system and the metric system for the purpose of trade. Scientists exclusively use the metric system. Your prescriptions are measured in milligrams or cubic centimeters (cc). Some of your everyday products are packaged using the metric system. Wine typically comes in 750 milliliter bottles and soda in 2 liter bottles. One of the advantages (aside from being able to communicate easily with other countries) of the metric system is that it is a decimal system. Units are related by multiples of ten. It is much simpler to multiply and divide by ten than by the fractions used in the dead king's English system.

In this experiment, you will measure the dimensions of several wood blocks and calculate the volume of each from its dimensions. You will also learn how to determine the uncertainty in a measurement and apply uncertainty to significant figures. You will measure the mass of the wood blocks and construct a graph of mass vs. volume and use the graph to determine the mass of an unknown. You will also use the graph to solve a practical problem.

METRIC PREFIXES

The unit of length in the metric or SI system is the **meter** (a little more than an English yard). It can be abbreviated by **m**. A meter could be used to measure your height or the length and width of the room. It may be too long for some measurements such as determining your waist size (some clothing is sold according to waist size). For this you would use the **centimeter**, which is one hundredth of a meter. What if you needed to measure something even smaller, such as sizing your finger for a ring? You could use the **millimeter**, which is one thousandth of a meter. Note that the units are related to one another by factors of ten. If you wanted to measure the distance from your home to Mt. SAC, a much longer quantity, you would use the **kilometer**, which is one thousand meters.

The table below shows the relationships among some of the metric prefixes and the basic quantity.

Table of Metric Prefixes

Prefix	Symbol	Multiple	Example
Mega-	M	1,000,000 or 10^6	1,000,000 bucks = 1 Megabuck
kilo-	k	1,000 or 10^3	1,000 mockingbirds = 1 kilomockingbird
deci-	d	1/10 or 0.1 or 10^{-1}	0.1 mate = 1 decimate
centi-	c	1/100 or 0.01 or 10^{-2}	0.01 pede = 1 centipede
milli-	m	1/1000 or 0.001 or 10^{-3}	0.001 vanilli = 1 millivanilli

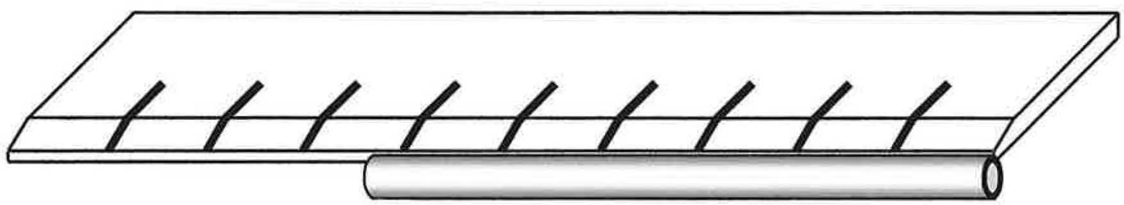
There are other prefixes in addition to the ones in the table above. These entries are the most common prefixes and the ones you will most likely use in the laboratory.

UNIT OF LENGTH & USING A RULER

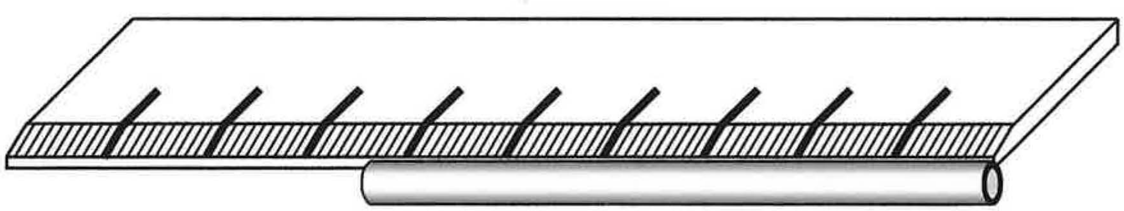
In this experiment, you will use a meter stick, ruler or tape measure to collect data on the length of various objects. The fundamental SI unit of length is the **meter** and you will use the metric prefixes to modify this unit.

Case Study

A student needs to measure glass tubing. The drawing below shows a section of glass tubing next to a metric ruler. Each division of the ruler represents one centimeter. There are no markings between the centimeters on this ruler so when measuring you will have to make a guess (estimate) between the divisions. To do that you imagine (estimate) ten intermediate markings between the centimeter markings. The glass tubing is between six and seven centimeters long. The first marking represents 6.0 cm and the imaginary markings represent 6.1, 6.2, 6.3, 6.4, 6.5, 6.6, 6.7, 6.8, and 6.9 cm. The end of the tubing lies about three-tenths past six, so you record the length as 6.3 cm. This measurement has been made to the nearest 0.1 cm.



Case Study
The same student may use a different ruler to measure glass tubing. The drawing below shows the same section of glass tubing next to a different metric ruler. This ruler is divided into centimeters and each centimeter has ten intermediate markings. Again, the glass tubing is between six and seven centimeters long. The end of the tubing lies between the third and fourth intermediate markings. Again, you estimate between divisions, this time between the intermediate markings. In the example above, you imagined ten markings between divisions. There is not enough space between the intermediate markings to imagine ten more divisions. You can imagine one additional marking. If the glass tubing is on the 6.3 cm mark, record the length as 6.30 cm. If it is on the 6.4 cm mark, record the length as 6.40 cm. If the length is between the two marks, record it as 6.35 cm. Since the glass tubing is between the two marks, record the length as 6.35 cm. This quantity has been measured to the nearest 0.05 cm.



UNCERTAINTY IN MEASUREMENTS

In every measurement there is some **uncertainty** (guesstimation). In the first example above, you were certain about the 6 cm but you made a guess (estimate) at the 0.3 cm. This is the part of the measurement that is uncertain. When you record your data, you may record all of the numbers about which you are certain and **one guess**. Again, for the glass tubing, you were certain about the 6 cm and made a guess of the 0.3 cm. You recorded two digits, called **significant figures** (or **sig figs** for short).

In the second example, you were certain of the 6 cm and the 0.3 cm. The guess was the 0.05 cm. Again, this is the part of the measurement that is **uncertain**. How many sig figs are recorded with this data? There are three sig figs: the two digits about which you are certain and one that is a guess.

When you record your data for any measurement, you must always include all of the numbers about which you are certain and one guess.

Summary

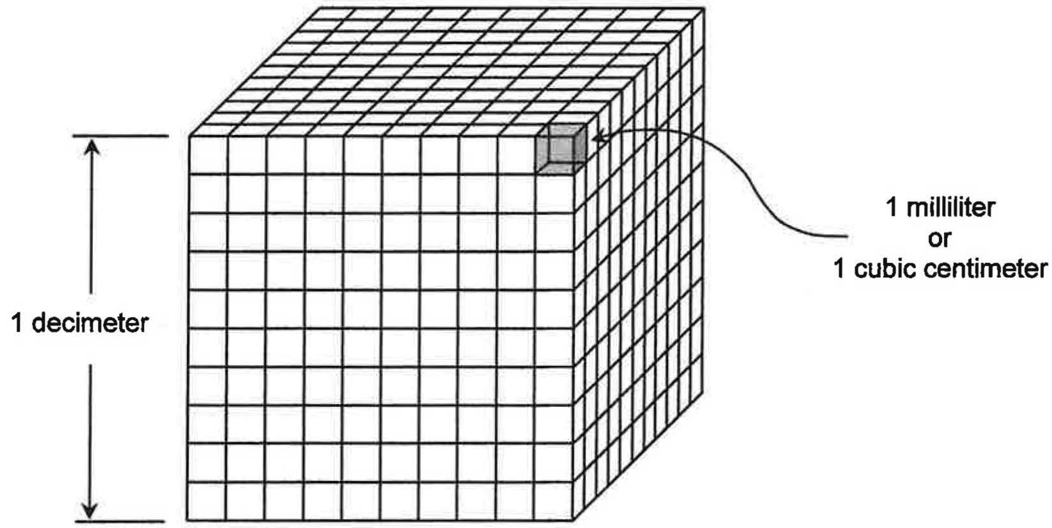
1. There is uncertainty in all measurements
2. You should estimate between the markings, even if they are closely spaced
3. If possible, imagine 10 divisions between markings
4. If the markings are too close together, imagine one division half-way between the markings
5. Record all measurements using all of the certain numbers and one guess
6. Use units when recording measurements

USING A BALANCE & MEASURING MASS

Your instructor will demonstrate how to use the balance and which balances are appropriate for this experiment. When recording the data for mass, use all of the digits available. Do not round off before calculating.

VOLUME OF A REGULAR SOLID

Volume is different from length and mass because it is a **derived measurement**, meaning it is based on another measurement. One **liter** (**L**), a bit less than an English quart, is defined as a cubic decimeter. If you imagine a cube, such as the one in the drawing below, and measure it on all sides, the length, width and height would all measure one decimeter. That is a liter. One **milliliter** (**mL**) is one-thousandth (1/1000) of a liter. It is also the volume of a cube that measures 1 centimeter on each side. That is why a milliliter is also called a **cubic centimeter or cc**.



Case Study
To measure the volume of a regular solid, you would measure the length, width and height of the solid and calculate its volume. Consider the following: a gold metal bar is 10.1 cm in length, 5.3 cm in width and 2.4 cm in height. What is the volume of the gold bar?

Begin with the equation:

volume = length x width x height

Substitute in the data:

volume = (10.1 cm) x (5.3 cm) x (2.4 cm)
= **128.472 cm^3**
= **130 cm^3** (two sig figs)

The calculated volume is **128.472 cm^3**. There is something wrong with this mathematical answer. From the data you can see that the uncertainty in the measurement is in the tenths place (0.1 cm). The answer implies that the uncertainty is in the thousandths place (0.001 cm). Review the rules in your text for using **significant figures** in calculations. For this problem, there are three sig figs in the length measurement, two sig figs in both the width and height measurements. Your answer must be expressed with only two sig figs. The volume is **130 cm^3**, or 1.3 X 10^2 cm^3, expressed with two sig figs. (The zero in **not** significant – it is a trailing zero).

GRAPHICAL ANALYSIS

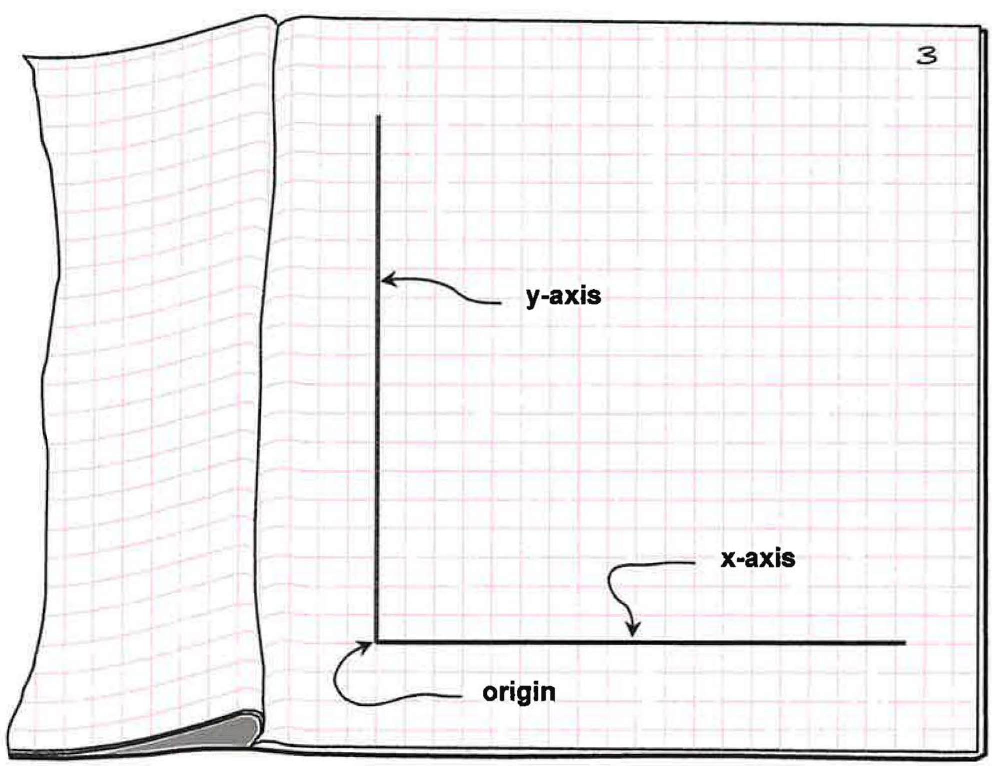
Numbers on a data sheet can be difficult to interpret. To get a picture (an overall visual) of the data, you can construct a graph. The graph will help smooth out experimental errors.

Case Study

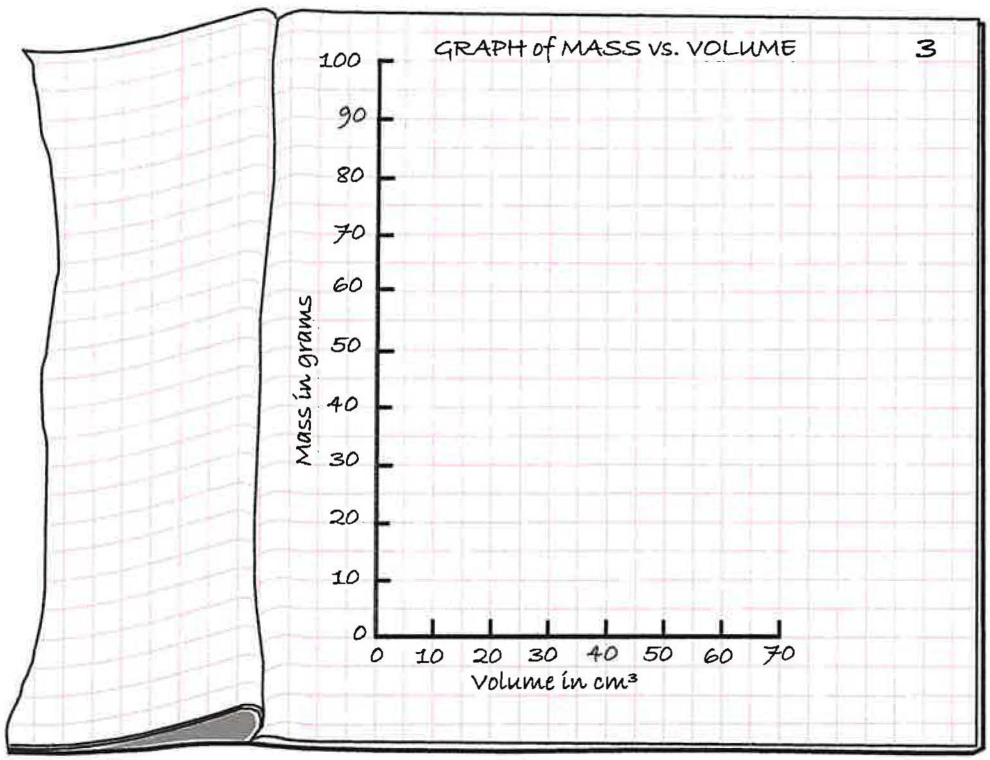
You have collected the following data and now must construct a graph of the mass and volume values of wood blocks. The graph that you will construct can be used to predict the mass of a wood block if you know its volume or the volume of a wood block if its mass is known. It can also be used to determine if an unknown wood block is made from the same wood as the known wood blocks..

Block Color	***Volume (cm^3)***	***Mass (g)***
Purple	61	94
Green	45	68
Pink	33	45
Brown	19	27

Begin with a piece of graph paper and draw a vertical (up and down) line near the top of the paper to within two or three centimeters of the bottom of the paper. The line should be drawn two or three centimeters in from the left-hand margin. This line is called the **y-axis**. Draw a horizontal (left and right) line that begins at the bottom of the vertical line and extends to within two or three centimeters of the right-hand margin. The line should be drawn two or three centimeters up from the bottom margin. This line is called the **x-axis**. The point where the two lines meet is called the **origin**.



You need to set up a logical, numerical scale along the y-axis that will include your experimental mass values. Your largest mass value is 94 g so your scale should go from 0 to 100 g. On this graph, there are ten lines between each division on the scale. Label the y-axis "mass in g". Likewise, set up a scale for the x-axis. The largest volume value is 61 cm^3 so the scale should go from 0 to 70 cm^3. Again, there are ten lines between each division on the scale. Label the x-axis "volume in cm^3". Give your graph a title, "Graph of Mass vs. Volume."

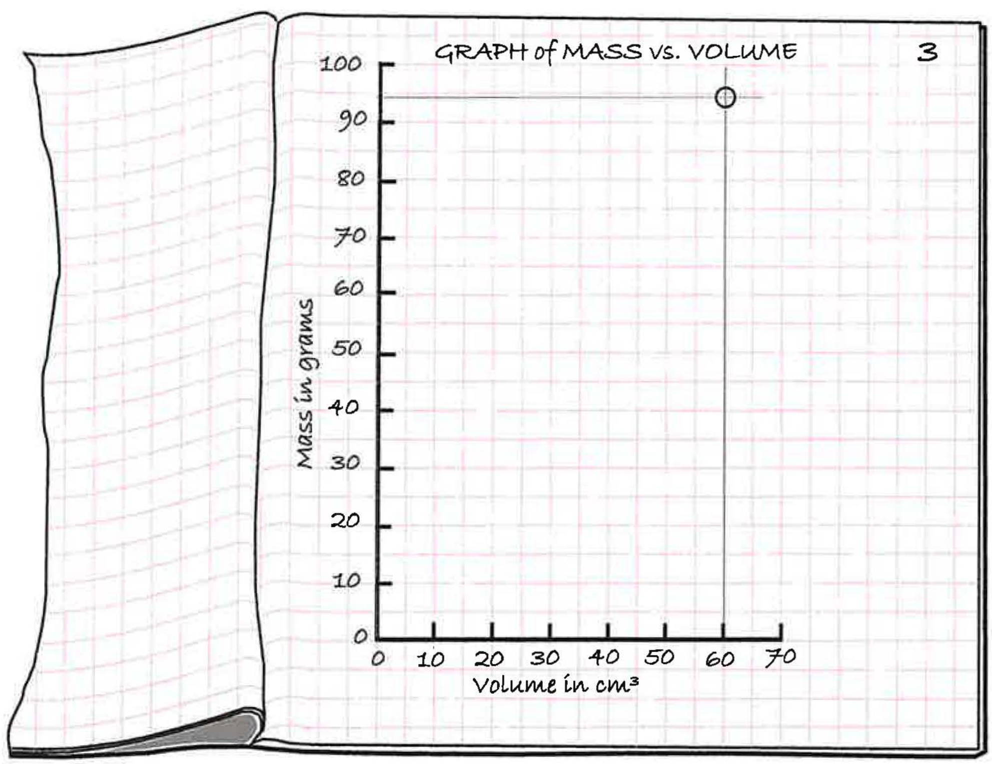


The purple block has a volume of 61 cm^3 and a mass of 94 g. This set of data is plotted as a point on a graph. Imagine a line drawn upward from the 61 cm^3 mark on the x-axis crossing a line drawn from the 94 g mark on the y-axis. This is the data point. You don't actually draw the lines but you draw a point where the two imaginary lines intersect (cross). The point should be the smallest point that you can make with your (very sharp) pencil. Draw a circle around the point so that you don't lose track of the small pencil mark. Your graph should look like the one below.

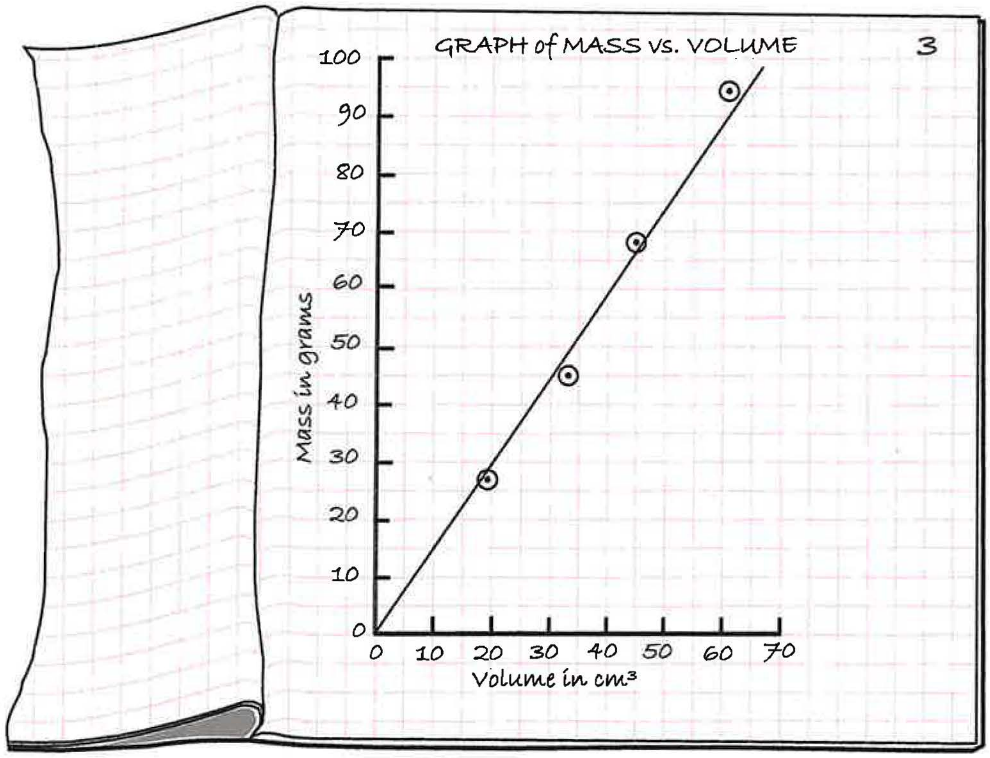
Plot the other three points in the same way. Now, using a ruler and a pencil, draw the best straight line through the four data points. The line will not go through all of the points – some points will lie slightly above the line, some slightly below. The best line will have the same number of points below the line as those above the line.

A couple of questions about the completed graph:
Why aren't all the data points on the line? Your experimental values are not quite accurate. This is one of the reasons scientists graph data – to smooth out the errors.

Should the line pass through the origin?
Sometimes it should and sometimes it should not. In this case, a block with zero mass will have zero volume so the line should pass through the origin.



Your completed graph should now look like the one below.



You have now found the mass/volume ratio of four wood blocks. Since the blocks are made from the same wood, the graph of mass vs. volume yielded a straight line. If the blocks had been made from different kinds of wood, the graph would look more like a "connect the dots" picture than a straight line.

Summary

1. Draw and label the axes.
2. Choose appropriate scales.
3. Plot each point carefully and draw a small circle around the point.
4. Draw the best straight line through the points.
5. Not all the points are necessarily on the line due to experimental error.
6. The line should pass through the origin if the data suggests that is reasonable (in this case a wood block with zero mass will have zero volume so it is reasonable).
7. A graph of mass vs. volume will yield a straight line if the objects measured are made out of the same material.
8. A graph of mass vs. volume can be used to determine either the mass or volume of an unknown object, provided the unknown is made of the same material. For example, if the mass of an unknown block is determined by weighing, the volume of the unknown block can be determined from the graph. Likewise, if the volume is measured, the mass can be determined by the graph.

PERCENTAGE DIFFERENCE

You will be asked to make a comparison between two experimental values, for example, the mass of a wood block as measured by a balance and the mass of a wood block determined from a graph of mass vs. volume. How can you compare the two values? One way is to calculate the percentage difference. To do this you will use the relationship:

$$\text{percentage difference} = \frac{\text{the difference between the values}}{\text{the average of the two values}} \times 100\%$$

Case Study

You weigh a wood block and find that the mass is 53.7 g; the mass obtained from the graph is 55.1 g. Compare the two values.

You begin with the relationship

$$\text{percentage difference} = \frac{\text{the difference between the values}}{\text{the average of the two values}} \times 100\%$$

Substitute in the data

$$\text{percentage difference} = \frac{55.1\text{g} - 53.7\text{g}}{1/2(55.1 + 53.7)} \times 100\%$$

$$= \frac{1.4\text{g}}{54.4\text{g}} \times 100\%$$

$$= 2.5735\% \text{ or } 2.57\% \text{ (three sig figs)}$$

(NOTE: Do not use the % key on your calculator. Multiplying the value by 100 converts the product to a percentage.)

SAFETY AND WASTE DISPOSAL

The materials used in this experiment are not toxic. The experiment is conducted in a chemical laboratory so be sure to wash hands with soap and water before leaving and especially before eating or drinking.

No special disposal is required. Return all materials to the hood.

PROCEDURE

Using Metric Prefixes

1. Obtain a tape measure. Find the markings that represent millimeters. Use the tape measure to locate a feature on your body, such as a freckle, that is 1 mm in length. Use this feature as a reference. As an alternative, you could measure the width of a pencil mark.

2. Find the markings on the tape measure that represent centimeters. Use the tape measure to locate a feature on your body, such as the width of one of your fingers or your fingernail, that is 1 cm. Use this feature as a reference.

3. Find the markings on the tape measure that represents decimeters. Use the tape measure to locate a feature on your body, such as the width of your hand, that is 1 dm in length. Use this feature as a reference.

4. Measure your height in meters. Record your height on the data sheet.

5. Measure (do not calculate) your height in centimeters. Record.

6. Measure your height in decimeters. Record.

Using a Ruler

1. Obtain a ruler. Locate the centimeter divisions on the ruler. Count the number of marks between the centimeter divisions. Record this on the data sheet.

2. If there is sufficient room, imagine 10 divisions between the markings. If there is not enough room for 10 divisions, imagine one division. Record the number of divisions ***imagined*** between markings.

3. From the data collected above, determine if your ruler will measure to the nearest 0.1 cm or the nearest 0.05 centimeter. Record your result.

4. Bring your ruler and data sheet to your instructor for approval before going on to the next step.

Volume of Wood Blocks

1. Obtain a blue block from the hood. Measure the length, width and height of a blue block. Record. Keep the blue block for the next part of the experiment.

2. Repeat step 1, using a red and then an orange block. Record. Keep the red and orange blocks for the next part of the experiment.

3. Calculate the volume of each of the wood blocks. Neatly show all work on the report sheet. Record.

Mass of Wood Blocks

1. Weigh the blue block (from above) using the balance indicated by your instructor. Record your data using all of the digits on the balance.

2. Repeat using red and orange blocks (from above). Use the same balance for each weighing. Record data.

3. Show your data sheet to your instructor for approval before going on to the next step.

4. Return wood blocks.

Graphical Analysis of Data

1. Using the graph paper provided, construct a graph of mass vs. volume of wood blocks.

2. Obtain an unknown wood block from your instructor. Record the number of the unknown on your data sheet.

3. Measure the length, width and height of the wood block. Record data.

4. Calculate the volume of the wood block. Show all work on the report sheet.

5. Using the graph from step 1, determine the mass of the unknown wood block. To do this, find the calculated volume on the x-axis. Trace an imaginary line from the x-axis to the graphed line. This is the data point for the mass and volume of your unknown wood block. Trace a second imaginary line from the graphed line over to the y-axis. This is the mass of the unknown wood block. Record this value.

6. Weigh the unknown wood block on the same balance used earlier. Record your data.

7. Compare the two values using percentage difference. Record on report sheet.

WHAT KIND OF WOOD IS IT?

A shipment of children's wood blocks may have been made with wood that is toxic if ingested. Children do chew on and sometimes eat their toys, so this presents a health hazard. If the blocks were made of the same wood as the blocks you used, the children's blocks are safe. If they are made from a different wood, they will need further testing. How could you determine if the children's wood blocks are made of the same wood as the ones used in this lab? Think about it for a minute. If the block is made of the same wood, its mass and volume will represent a point on the line of your graph. If the wood is different, it will not be on the line.

1. Obtain one of the "suspicious" children's wood blocks from your instructor. Record the number of the unknown on your data sheet.

2. Measure the length, width and height of the wood block. Record data.

3. Calculate the volume of the wood block. Record. Show all work on the report sheet.

4. Weigh the unknown wood block on the balance used earlier. Record your data.

5. Using the mass and volume as a data point, plot the point on your graph.

6. Decide whether the block is safe or requires further testing.

Name________________

MEASUREMENTS DATA SHEET

Using Metric Prefixes

1. Height in meters __________ m
2. Height in centimeters __________ cm
3. Height in decimeters __________ dm

Using a Ruler

1. Number of divisions between centimeters __________
2. Number of markings imagined __________
3. Measurements will be made to the nearest __________
4. Instructor signature __________

Volume of Wood Blocks

Block Color	***Length (cm)***	***Width (cm)***	***Height (cm)***
Blue			
Red			
Orange			

Mass of Wood Blocks

Block Color	***Mass (g)***
Blue	
Red	
Orange	

Instructor signature __________

Graphical Analysis of Data
Determining Mass of Unknown

Unknown Number	***Length (cm)***	***Width (cm)***	***Height (cm)***

Mass of unknown from balance __________ g

WHAT KIND OF WOOD IS IT?

Suspicious Block	***Length (cm)***	***Width (cm)***	***Height (cm)***

Mass of block from balance __________ g

Name________________

MEASUREMENTS
REPORT SHEET

Volume of Wood Blocks

Calculate the volume of each wood block. Show only one sample calculation and record the volume for all blocks in the table below. Begin with the equation for volume.

*Answer*__________

Volume of Wood Blocks

Block Color	***Volume (cm^3)***
Blue	
Red	
Orange	

Graphical Analysis of Data

Calculate the volume of the unknown wood block. Begin with the equation for volume.

*Answer*__________

Mass of unknown from graph __________ g

Mass of unknown from balance __________ g

Compare the two values using percentage difference. Begin with the equation.

*Answer*__________

WHAT KIND OF WOOD IS IT?

Calculate volume of "suspicious" sample. Begin with equation.

*Answer*__________

Name______________

MEASUREMENTS POST-LABORATORY REPORT

1. What is the relationship between your height in meters, centimeters and decimeters?

2. Are all of the data points in your graph of mass vs. volume on the line? What does this indicate about your data points?

3. Describe the line in your graph of mass vs. volume (is it a straight line, a curve, etc.). What does this indicate about the wood blocks?

4. What is the percentage difference between the mass of the unknown as measured by the balance and the mass as determined by graphical analysis?

   Is this difference significant?

5. Does the data point of mass and volume of the "suspicious" wood block fit on the line of your graph? What does this indicate about the nature of the wood? Should this wood be tested further?

6. The graph you made of mass vs. volume is called a standard line (curve). Can you think of any other uses for standard lines? Your instructor will give you hints for this question.

Name________________

MEASUREMENTS
PRE-LABORATORY QUESTIONS AND PROBLEMS

1. Give two reasons why the metric system of measurement is more convenient than the English system.

2. Express $1,000 using one of the metric prefixes.

3. If you need to measure out 9.85 cm of glass tubing, which ruler would you use?
   (a) ruler I
   (b) ruler II
   (c) it doesn't matter if you use ruler I or ruler II

4. If you used ruler I to measure out exactly ten cm of glass tubing, how would you record your data?
   (a) 10 cm
   (b) 10.0 cm
   (c) 10.00 cm

5. Which of the following units is a derived unit?
   (a) length
   (b) mass
   (c) volume

6. The volume of a gold bar with length 19.7 cm, width 10.3 cm and height 5.6 cm is (expressed with the correct number of sig figs)
   (a) 1136.296 cm^3
   (b) 1136 cm^3
   (c) 1140 cm^3
   (d) 1100 cm^3

7. Construct a graph of the following data.

Block Color	***Volume (cm^3)***	***Mass (g)***
Purple	24.2	19.3
Green	44.9	36.6
Pink	61.7	52.8

8. Compare 43.9 mL with 45.6 mL using the percentage difference relationship. Show all work. Begin with the written equation.

9. State all safety concerns in this laboratory.

BONUS

Bring in two labels or items that give both the metric and English measurements (e.g. candy wrappers, coffee cans, etc.).

DENSITY

INTRODUCTION

In the Measurements experiment, you prepared a graph of mass vs. volume using four wood blocks as samples. The graph yielded a straight line because the blocks were made from the same wood. The blocks have the same ***density***. (More information on density later).

In this experiment, you will measure the volume and mass of an irregular solid (an irregular solid does not have a uniform shape so you cannot measure its length, width and height). You will learn how to use a graduated cylinder to measure volume of a liquid and volume by displacement. Finally, you will use the volume and mass to determine density.

You will also solve two practical problems using density. You will attempt to identify Diet Coke™ and Coke™ by their densities and then check your results with a taste test. In the last procedure, you will use density to determine if a gold chain is the real deal or a fake.

USING A GRADUATED CYLINDER & MEASURING VOLUME

There is more than one way to measure volume depending on the state of the object you are measuring. Are you measuring a liquid or are you measuring a solid? Is the shape of the solid regular, like a gold bar, or irregular, like a gold nugget?

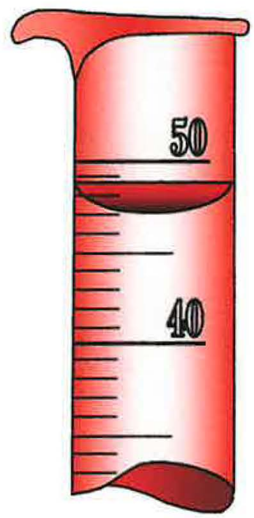
Volume of a Liquid

If the sample is a liquid, you will most often use a **graduated cylinder** to measure its volume. The cylinder is not graduated in the sense that it has a college diploma but it refers to the markings. It is graduated in the same way a ruler is graduated.

The drawing below shows a liquid in a 50 mL graduated cylinder. What volume does each numbered marking represent? Each of the markings represents 10 mL. How many divisions are between the markings? There are ten divisions in between so each marking represents 1 mL. There is not enough room between the divisions to imagine ten intermediate markings, but you can imagine one intermediate marking half-way between the divisions. Using this graduated cylinder, you can measure to the nearest 0.5 mL. All measurements with this graduated cylinder will include one decimal place.

Notice that the top of the liquid has a bowl shape to it caused by the pull of gravity on the liquid and its attraction to the glass. (Why don't you see this in your water glass at home? It has to do with the shape of the graduated cylinder). This bowl shape is called a **meniscus** and the graduated cylinder must be read at the bottom center of the meniscus. To do this you must keep your eye level with the bottom of the meniscus.

To measure the volume of the liquid in the cylinder, first note that it is between 40 and 50 mL. The bottom of the meniscus lies between the seventh and eighth intermediate markings so the volume is between 47 and 48 mL. You estimate between the intermediate markings. If the bottom of the meniscus is on the 47 mL mark, record the volume as 47.0 mL. If it is on the 48 mL mark, record the volume as 48.0 mL. If the meniscus is between the two marks, record the volume as 47.5 mL. Since the liquid's meniscus is at the 47 mL mark, record the volume as 47.0 mL. This volume has been measured to the nearest 0.5 mL.



Volume of an Irregular Solid

Say you wanted to measure the volume of a gold nugget instead of a gold bar. If you have ever sat in a bathtub full of water, you know the water rises when you submerse an object in it. This is called **displacement**. You can actually measure the volume of an object by water displacement. To measure the volume of an irregular solid, such as a gold nugget, you place the nugget in a graduated cylinder that has water in it. The gold nugget will displace a volume of water equal to the volume of the gold nugget. The key to determining the volume of the gold nugget (or any irregular solid) is to carefully observe the volume of water in the graduated cylinder before submerging the gold nugget, to take care that the nugget is completely submerged (no part of the nugget should stick out of the water) and to carefully read the new volume. The difference in the volume before and after submersion is the amount of water displaced, which is equal to the volume of the gold nugget.

Summary

1. Volume is a derived unit based on length.
2. The volume of a regular solid object can be determined by measuring the length, width and height of the object and performing the appropriate mathematical calculation.
3. Graduated cylinders are used to measure the volume of liquids.
4. Graduated cylinders are read from the bottom of the meniscus at eye level.
5. Water displacement can be used to determine the volume of an irregular solid.

THE RATIO OF MASS TO VOLUME: DENSITY

Which has a greater density: a solid pure gold ring or a heavy, solid pure gold bar?

If you answered the gold bar, you are incorrect. If you answered the gold ring, you are also incorrect. Because both the ring and the bar are pure gold, they have the same density.

Density is the ratio of mass to volume. The gold bar has a greater mass than the gold ring, but it also has a greater volume. The gold ring has a much smaller mass, but the volume is also smaller. The ratio will be the same for both the ring and the gold bar. The equation below shows the relationship between mass and volume:

$$\text{DENSITY (D)} = \frac{\text{MASS (M)}}{\text{VOLUME (V)}}$$

For the ring, the mass is small and so is the volume:

$$D = \frac{\scriptstyle M}{\scriptstyle V}$$

For the gold bar, the mass is many times larger as is the volume:

$$D = \frac{\Huge M}{\Huge V}$$

In both cases, the density remains the same because it is a **ratio** of mass to volume. This is why density is considered a characteristic property of a substance and density can be used to identify a substance. In fact, gold is one of the densest elements on earth (19.3 g/mL) and was used as currency because it could be verified by its density. If you tried to buy goods with gold, the merchant could weigh the gold, measure its volume using a graduated cylinder and calculate the density. If the density of your coin matched the density of gold, the merchant would accept the coin as payment.

PERCENTAGE ERROR

In the measurement experiment, you compared values by using the percentage difference between them. In this experiment you will be asked to compare an experimental value, the density of the gold chain, with an accepted value, the density of gold as reported in your text or *Handbook of Chemistry and Physics*. When a comparison is made between experimental and accepted values, ***percentage error*** is used. To do this you will use the relationship:

$$\text{percentage error} = \left|\frac{\text{accepted value} - \text{experimental value}}{\text{accepted value}}\right| \text{ X } 100\%$$

The symbols | | mean "absolute value", which is the difference without +/- sign.

Case Study

You determine the density of a sample of aluminum to be 2.89 g/mL. The accepted value is 2.70 g/mL. Compare the two values.

You begin with the relationship

$$\text{percentage error} = \left|\frac{\text{accepted value} - \text{experimental value}}{\text{accepted value}}\right| \text{ X } 100\%$$

Substitute in the data

$$\text{percentage error} = \left|\frac{2.70 \text{ g/mL} - 2.89 \text{ g/mL}}{2.70 \text{ g/mL}}\right| \text{ X } 100\%$$

$$= \frac{0.19 \text{ g/mL}}{2.70 \text{ g/mL}} \text{ X } 100\%$$

$$= 7.04\%$$

(NOTE: The absolute value of –0.19 is used.)

A percentage error between 5% and 10% is acceptable. This error can usually be attributed to student measuring lab techniques and will improve as students practice techniques. One way to reduce your error is to take care particular care with all measurements.

IS IT COKE™ OR DIET COKE™?

Coke™ and Diet Coke™ have very different tastes. Why? Coke™ contains sugar and Diet Coke™ contains an artificial sweetener (note that sugar and artificial sweeteners are not the same). In fact, a 240 mL serving of Coke™ contains 27 g of sugar and the same serving of Diet Coke™ contains 0 g of sugar. The volume of the two servings is the same but Coke™ has sugar dissolved in it. Will this affect the density of Coke™? In this experiment, you will measure the density of Coke™ and of Diet Coke™ to determine if the two liquids can be distinguished by their densities. **DO NOT TASTE THE DRINKS.** At the end of the procedure, you will make a taste test to compare with your laboratory results.

IS IT GOLD OR IS IT FAKE?

Is the "gold" chain the real deal? You can determine the volume of the gold chain by water displacement and you can weigh the chain to determine its mass. After calculating the density of the gold chain, you can compare its experimental value to the accepted density of gold and determine if the chain is real or fake. In the second part of this experiment, you will measure the volume and mass of a gold chain, calculate its density and compare it with the accepted density and determine if the gold is real.

SAFETY AND WASTE DISPOSAL

The materials used in this experiment are not toxic. The experiment is conducted in a chemical laboratory so be sure to wash hands with soap and water before leaving and especially before eating or drinking.

No special disposal is required. Return all materials to the hood.

PROCEDURE

You must begin this experiment with clean glassware. To clean your glassware:

1. Fill your largest (400 mL) beaker about two-thirds full of warm water. Add a small amount of liquid soap. Put your test tube brush in the beaker and stir.
2. Spread out a few paper towels on your bench top so that your clean glassware can drain.
3. Scrub the needed glassware with the warm soapy water and the test tube brush. You will need a 150 mL beaker, a 250 mL beaker and a 50 mL graduated cylinder.
4. Rinse your glassware with tap water.
5. Rinse your glassware a second time with distilled water from your wash bottle. There are two distilled water taps in each lab. Fill your wash bottle with distilled water, take it back to your lab bench and rinse your glassware in the sink nearest your lab bench.
   [CAUTION: The distilled water is dispensed with a great deal of pressure – enough to knock the wash bottle out of your hand. Take precautions by firmly holding your wash bottle.]
6. Place the glassware rinsed with distilled water on the clean paper towels on your lab bench so that it can drain.
7. Dry the glassware with clean paper towels.

Density of Coke™ and Diet Coke™

1. Weigh a clean, **dry** 50 mL graduated cylinder on the top loading balance. (Your instructor will indicate which balance to use). Record the mass on your data sheet.

2. In the hood, pour approximately 50 mL of soft drink **A** into your clean, dry 50 mL beaker. Take the beaker to your lab bench to perform the rest of the procedure. This will allow more room for other students to work as needed in the hood.

3. Pour between 40 mL and 50 mL of soft drink **A** into the preweighed 50 mL graduated cylinder. Read the volume from the bottom of the meniscus and record it.

4. Weigh the cylinder with soft drink **A** on the same balance that you used in step 1. Record the mass.

5. Calculate the density of soft drink **A**.

6. Repeat steps 1-4 with soft drink **B**.

7. Soft drinks **A** and **B** can be poured down the sink followed by a little tap water.

8. Once the entire class has completed the procedure, your instructor will give you samples for a taste test.

Is it Gold or is it Fake?

1. Obtain a length of "gold" chain from the hood. Weigh the chain on the top loading balance. (Your instructor will indicate which balance to use). Record the mass on your data sheet.

2. Fill a clean 50 mL graduated cylinder approximately half full of tap water. Measure the volume from the bottom of the meniscus and record.

3. Carefully place the chain in the graduated cylinder. ***The chain must be completely submerged. If it is not, remove the chain, add more water and begin again with step 2.*** Measure the volume from the bottom of the meniscus and record.

4. Remove the chain from the graduated cylinder and dry it. Place the chain back in the hood so that it may be used by another student.

5. Calculate the density of the chain.

Name________________

DENSITY DATA SHEET

Density of Coke™ and Diet Coke™

Soft Drink **A**

1. Mass of empty graduated cylinder __________ g
2. Volume of soft drink **A** __________ mL
3. Mass of graduated cylinder and soft drink **A** __________ g
4. Mass of soft drink **A** __________ g

Soft Drink **B**

1. Mass of empty graduated cylinder __________ g
2. Volume of soft drink **B** __________ mL
3. Mass of graduated cylinder and soft drink **B** __________ g
4. Mass of soft drink **B** __________ g

Is it Gold?

1. Mass of gold chain __________ g
2. Volume of water in graduated cylinder __________ mL
3. Volume of water and submerged chain __________ mL
4. Volume of chain __________ mL

Name________________

DENSITY REPORT SHEET

Density of Coke™ and Diet Coke™

1. Calculate density of soft drink **A** and **B**. Show only one sample calculation and record the density for both soft drinks in the table below. Begin with the formula.

Density of Soft Drinks

Soft Drink	***Density (g/mL)***
Soft Drink A	
Soft Drink B	

2. Which soft drink has the greater density?

3. Which soft drink do you predict is Coke™?

4. Which soft drink do you predict is Diet Coke™?

Instructor signature______________________________

Results of taste test:
Soft drink **A** is ___________

Soft drink **B** is ___________

Is it Gold?

Calculate density of the gold chain. Begin with the formula.

Density of gold from text __________ g/mL

Compare the density of the gold chain with real gold. Use the percentage error calculation.

*Answer*__________

Name_______________

DENSITY POST-LABORATORY QUESTIONS

1. Were you able to identify Coke™ and Diet Coke™ from their densities? Briefly explain why. If not, do you think it is because of the densities of the two drinks, or is it because of the data?

2. Is the gold chain the real deal or is it fake? Explain your answer.

3. Which has a greater density: an aluminum soda can or a solid aluminum bicycle frame? Explain your answer.

4. Osmium is one of the densest elements on earth. Which has a greater density: 1 kilogram of osmium or the whole earth? Explain your answer.

5. The density of water is 1.00 g/mL. The density of gasoline is 0.70 g/mL. Will gasoline float on top of water or will water float on top of gasoline? Explain your answer.

6. In the movie, ***Raiders of the Lost Ark***, the character Indiana Jones tries to steal a gold statue. Indiana Jones replaces the gold statue with a bag of sand that is supposed to be the same volume (size) as the statue. This does not work and the theft of the gold statue is detected. Why didn't the same volume of sand work?

Name_______________

DENSITY PRE-LABORATORY QUESTIONS

1. When using a graduated cylinder
   (a) measure the liquid carefully
   (b) read the liquid level from the bottom of the meniscus
   (c) congratulate the cylinder on its graduation
   (d) read the liquid level from the top of the meniscus

2. To determine the volume of a gold nugget
   (a) use the water displacement method
   (b) measure the length, width and height of the nugget and calculate its volume
   (c) place the nugget in an empty graduated cylinder

3. If you measured exactly 39 mL of water in a 50-mL graduated cylinder, you would record your data as
   (a) 39 mL
   (b) 39.0 mL
   (c) 39.00 mL

4. A platinum necklace is found to have a mass of 59.6 g and a volume of 2.61 mL. What is the experimental density of the platinum necklace? Neatly show all work.

*Answer*__________

5. If the accepted density of platinum is 21.4 g/mL, is the necklace in problem 4 pure platinum? Why?

*Answer*__________

Determine the percentage error. Is the error acceptable?

*Answer*__________

CALORIE CONTENT OF FOOD

INTRODUCTION

Have you ever read a nutritional label? How is the number of Calories per serving determined? Why is the word "calorie" spelled with a capitol letter?

When you are exercising, you might say that you have "burned" calories. This is actually the way calorie content is determined. Food is burned and the energy released is used to warm water. The energy provided to warm the water is equal to the energy released from the food.

In this experiment, you will burn a walnut. The heat produced will be used to warm a known amount of water. By knowing the mass, temperature change and specific heat of the water you can calculate the energy released from the walnut. You will repeat the experiment with a marshmallow.

CALORIMETER

In a previous experiment, a styrofoam cup was used as a calorimeter. The styrofoam provided insulation to keep the heat from escaping. In this experiment, a soft drink can will be used. The aluminum can will actually conduct heat and not act as an insulator. Why use a conductor instead of an insulator? Heat from the burning walnut or marshmallow must be transferred to the water to cause its temperature to change. The most effective way to transfer heat is to use a heat conducting material such as aluminum. Although this will allow the heat to be transferred it will not insulate the warm water. Some of the heat from the warm water will "leak" out (escape). This will introduce some error but the results will still be acceptable.

HOW TO CALCULATE THE CALORIE CONTENT

Burning a walnut or marshmallow will produce heat. The heat energy flowing out of the food will be used to warm up water in the calorimeter.

heat released from food = heat gained by water

If you know the mass of the water in the calorimeter, the beginning and ending temperature of the water and the specific heat of the water, you can calculate the energy used to warm the water. This is equal to the energy released from the food:

heat gained = mass x Δt x specific heat of water

The change in temperature can be symbolized by Δt, where Δ is read "the change in".

Experimental values for mass and temperature change are used; the specific heat of water is already known (you can look it up in the appropriate table). The heat gained by the water is equal to the heat released from the food. This is the calorie content of the food burned. The walnut (or marshmallow) will not completely burn. You will weigh the walnut before burning and the residue after burning. By subtracting the two values, you will know the mass of walnut that actually burned. Dividing the heat released by the mass burned gives the calories per gram which can be compared to the values on the label.

Case Study:
A student burned a marshmallow weighing 0.916 grams. The burning marshmallow raised the temperature of 100.0 mL of water in a calorimeter from 24.0 °C to 35.5 °C. After burning, the residue of the marshmallow weighed 0.596 grams. What is the Calorie content of the marshmallow?

Solution:
Begin with the relationship:

$$\begin{aligned}
\text{heat gained} &= (\text{mass}) \times (\Delta t) \times (\text{specific heat of water}) \\
&= (100.0\ \text{g})(11.5\ ^\circ\text{C})(1.00\ \text{cal/g}^\circ\text{C}) \\
&= 1150\ \text{cal} \\
&= 1150\ \text{cal} \times \frac{1\ \text{Calorie}}{1000\ \text{cal}} \\
&= 1.15\ \text{Calories}
\end{aligned}$$

NOTE: The density of water is 1.00 g/mL. In other words, 100.0 g of water occupies 100.0 mL.

The nutritional or food calorie, the Calorie, is equal to 1,000 chemical calories or 1 kcal. When you read the nutritional labels, you are looking at the number of kilocalories in the food.

Not all marshmallows are created equal. Some marshmallows will be larger than others. In addition, not all of the marshmallow burned. Therefore, you must compare the number of Calories in each gram rather than in each marshmallow.

To find the number of Calories in each gram of marshmallow burned, first find the number of grams burned:

mass burned = mass of original marshmallow – mass of residue after burning
= 0.916 g – 0.596 g
= 0.320 g

Divide the number of Calories produced by the number of grams burned:

$$\text{Calories/grams} = \frac{\text{Calories produced}}{\text{grams burned}}$$

$$= \frac{1.15 \text{ Calories}}{0.320 \text{ grams}}$$

$$= \frac{3.59 \text{ Calories}}{\text{gram}}$$

COMPARING EXPERIMENTAL VALUES WITH NUTRITIONAL LABELS

How does the value above, 3.59 Calories/gram, compare with the value on the nutritional label? The package will give the number of Calories in each serving. The serving size is given in grams, usually in parenthesis following the mass in the English system. Dividing the number of Calories by the mass in grams will give the Calories/gram. This is the accepted value and can be compared to the experimental value.

For this particular brand of marshmallows, there are 110 Calories in a 34 gram serving.

$$\text{Calories/grams} = \frac{110 \text{ Calories}}{34 \text{ grams}}$$

$$= \frac{3.2 \text{ Calories}}{\text{gram}}$$

In a previous experiment, you learned that when a comparison is made between experimental and accepted values, ***percentage error*** is used. To do this you use the relationship:

$$\text{percentage error} = \frac{|\text{accepted value} - \text{experimental value}|}{\text{accepted value}} \times 100\%$$

The symbols | | mean "absolute value" (You can check an algebra text if you do not remember what this means)

For this Case Study:

$$\text{percentage error} = \frac{|\text{accepted value} - \text{experimental value}|}{\text{accepted value}} \times 100\%$$

$$= \frac{|3.2\ \text{Cal/g} - 3.59\ \text{Cal/g}|}{3.2\ \text{Cal/g}} \times 100\%$$

$$= \frac{0.35\ \text{Cal/g}}{3.2\ \text{Cal/g}} \times 100\%$$

$$= 11\ \%$$

Is this percentage error acceptable? Usually 5% - 10% error is acceptable. For this experiment the error will probably be larger than 10%.

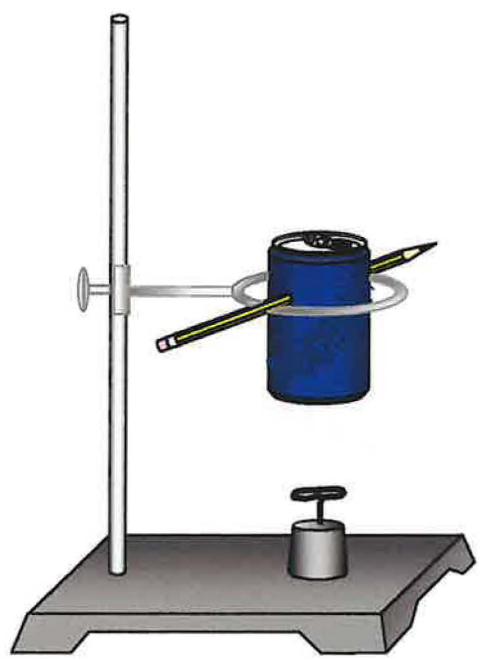
CALORIE CONTENT OF FOOD
PROCEDURE

Calorie Content of a Walnut

1. Obtain a soft drink can and a 100 mL graduated cylinder from the hood.

2. Carefully measure out 100.0 mL of tap water and pour it into the soft drink can. (Incorrect measurement of the volume of water can result in error). Record volume of water on data sheet. When you have finished with the graduated cylinder, return it to the hood to be used by other students.

3. Set up the calorimeter apparatus according to the figure below. Use a ring stand and iron ring to suspend/support the soft drink can. Place a pencil or stirring rod through the can (the can will already have two holes in the sides). Insert the can into the iron ring, allowing the pencil ends to overlap the iron ring and suspend/support the can.



4. Your instructor will direct you to the type of balance you should use. Note the number of the balance you weigh the walnut on and use the same balance when you weigh the burned residue. (Using a different balance will introduce error into your data). Weigh a walnut half. Record mass on data sheet.

5. Set up the apparatus to hold the walnut according to the figure below. Open up one end of the paper clip and push the end into the cork. Bend the other end of the paper clip to act as a secure platform to hold the walnut.

6. Place the walnut apparatus underneath the soft drink can. You will need to adjust the height of the soft drink can so that it is just above the walnut.

7. Measure the temperature of the water. Record this temperature as ***initial temperature*** on data sheet.

8. Ignite the walnut with a match. Once the walnut begins to burn, extinguish the match. The flame from the walnut should touch the bottom of the soft drink can. Adjust the height of the can if necessary. **You must wear goggles while you are burning the walnut.** (You must always wear goggles when you are working or visiting the lab).

9. Gently stir the water in the can while the walnut continues to burn. Take care not to break the thermometer.

10. When the walnut is no longer burning (it will not burn completely), stir the water in the can one more time. Record the highest temperature as the ***final temperature*** of the water on data sheet.

11. Weigh the residue of the walnut. Be sure to use the same balance used in step 4. Record mass on data sheet.

12. Empty the soft drink can. There will be soot on the bottom of the can. Clean this off with a paper towel. Keep the can and the apparatus set up for the next experiment.

13. Record the Calories per serving and the grams per serving from the package label.

14. Calculate the calorie content of the walnut.

15. Compare the experimental value and the actual value using percentage error.

16. Obtain your instructor's signature before going on to the next step.

Calorie Content of a Marshmallow

1. Use the soft drink can from above.

2. Obtain a 100 mL graduated cylinder from the hood and carefully measure out 100.0 mL of tap water and pour it into the soft drink can. Record volume of water on data sheet. When you have finished with the graduated cylinder, return it to the hood to be used by other students.

3. Use the calorimeter apparatus set up from above and insert the soft drink can.

4. Cut a marshmallow in half (this will help the marshmallow burn). Open up one end of the paper clip and push it into the marshmallow. Weigh the marshmallow and the paper clip together (once the marshmallow has burned, it will be difficult to separate the paper clip from it). Note which balance you weighed the marshmallow on and use the same balance when you weigh the burned residue. Record mass on data sheet.

5. Assemble the marshmallow apparatus by inserting the end opposite the marshmallow into the cork.

6. Place the marshmallow apparatus underneath the soft drink can. You will need to adjust the height of the soft drink can so that it is just above the marshmallow.

7. Measure the temperature of the water. Record temperature as ***initial temperature*** on data sheet.

8. Ignite the marshmallow with a match. Once it begins to burn, extinguish the match. The flame from the marshmallow should touch the bottom of the soft drink can. Adjust the height of the can if necessary. **You must wear goggles while you are burning the marshmallow.**

9. Gently stir the water in the can while the marshmallow continues to burn. Take care not to break the thermometer.

10. When the marshmallow is no longer burning (it will not burn completely), stir the water in the can one more time. Record the highest temperature as the ***final temperature*** of the water on data sheet.

11. Remove the paper clip and marshmallow from the cork. Weigh the residue of the marshmallow with the paper clip. Be sure to use the same balance used in step 4. Record mass on data sheet.

12. Empty the soft drink can. There will be soot on the bottom of the can. Clean this off with a paper towel. Return the soft drink can to the hood.

13. Record the Calories per serving and the grams per serving from the package label.

14. Calculate the calorie content of the walnut.

15. Compare the experimental value and the actual value using percentage error.

Name________________

CALORIE CONTENT OF FOOD
REPORT SHEET

CALORIE CONTENT OF A WALNUT

DATA AND ANALYSIS

1. Volume of water in soft drink can __________ mL
2. Initial temperature of water __________ °C
3. Final temperature of water __________ °C
4. Temperature change __________ °C
5. Mass of walnut before burning __________ g
6. Mass of residue __________ g
7. Mass of walnut burned __________ g

From the label

8. Calories per serving __________ Cal
9. Grams per serving __________ g

CALCULATIONS

Calculate the experimental calorie content of a walnut in Calories/g. Begin with the equation.

*Answer*__________

Calculate the actual calorie content of a walnut based on information in the label.

*Answer*__________

Compare the two values using percentage error.

*Answer*__________

Instructor's Signature ____

CALORIE CONTENT OF A MARSHMALLOW
DATA AND ANALYSIS

1. Volume of water in soft drink can __________ mL
2. Initial temperature of water __________ °C
3. Final temperature of water __________ °C
4. Temperature change __________ °C
5. Mass of walnut before burning __________ g
6. Mass of residue __________ g
7. Mass of walnut burned __________ g

From the label

8. Calories per serving __________ Cal
9. Grams per serving __________ g

CALCULATIONS

Calculate the experimental calorie content of a marshmallow in Calories/g. Begin with the equation.

*Answer*__________

Calculate the actual calorie content of a marshmallow based on information in the label.

*Answer*__________

Compare the two values using percentage error.

*Answer*__________

Name________________

CALORIE CONTENT OF FOOD
POSTLABORATORY REPORT SHEET

1. What was your percentage error in determining the Calorie content of a walnut? Is the error within acceptable experimental range? Explain.

2. What was your percentage error in determining the Calorie content of a marshmallow? Is the error within acceptable experimental range? Explain.

3. Which gave the lower experimental error: the walnut or the marshmallow?

4. Marshmallows are human-made according to a recipe. Walnuts, on the other hand, are grown naturally. Do you expect a difference in their percentage errors? Explain.

5. Some potential sources of error were mentioned in the procedure. List at least five (5) sources of experimental error.

Name________________

CALORIE CONTENT OF FOOD
PRELABORATORY QUESTIONS AND PROBLEMS

1. What is the difference between a calorie and a Calorie?

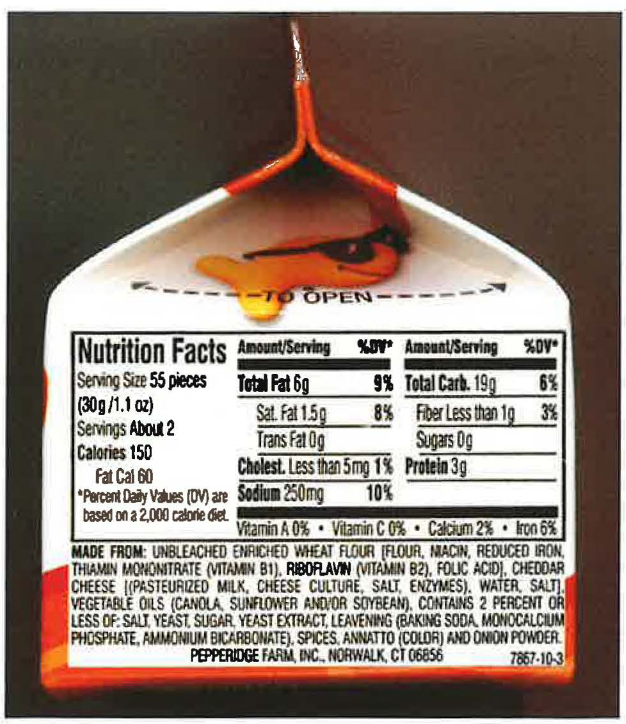
2. What are the advantages of using an aluminum can as a calorimeter? What are the disadvantages?

3. Use the label below from Goldfish™ crackers to answer the following questions:

   How many Calories per serving? __________

   How many grams per serving? __________

   Calculate the Calories/gram __________



4. A student performed this experiment using a Goldfish™ cracker and obtained the following data:

Volume of water	100.0 mL
Initial temperature of water	20.5°C
Mass of Goldfish™ cracker	0.59 g
Final temperature of water	23.0°C
Mass of residue	0.43 g

Calculate the energy released from burning the Goldfish™ cracker and express your results in Calories/g.

*Answer*__________

5. Compare the student's results in question 4 with the results from the calculations in question 3. Use percentage error in your comparison. Are the results acceptable? Explain.

*Answer*__________

6. List all safety precautions to be taken in this experiment.

PROPERTIES OF GASES

INTRODUCTION

The properties of gases are not just important to chemists but are exploited everyday by people in all walks of life. You live at the bottom of a sea of gases (the atmosphere); you breathe in a gas (oxygen) and exhale a different gas (carbon dioxide). Plants also breathe in and out gases: plants breathe in carbon dioxide and exhale oxygen. You use gases to fill your tires and operate a piston engine. You are warned not to put air in the tires when they are warm and not to toss cans into a fire. Open a can of soda and you will hear a hiss. Leave the can of soda out and it will "go flat". Exploring the properties of gases will help you understand these and other everyday occurrences.

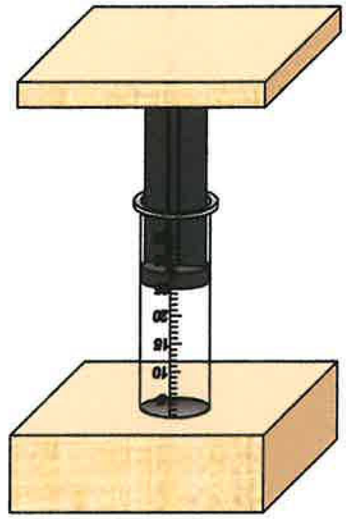
THE GAS LAWS

Gases have an indefinite shape (takes the shape of the container). Consider a balloon: the gas (air) in your lungs takes the shape of the balloon, not the shape of your lungs. Different sizes and shapes of balloons can be filled with the gas from your lungs. Gases also have indefinite volume (compressible and expandable so that it fills any size container). You may detect perfume from someone across the room. The gas (perfume) has expanded to the volume of the room. Gases share other properties described by the gas laws. Gases obey Boyle's Law, Charles' Law and Henry's Law (provided the gases are at normal to high temperature and normal to low pressure).

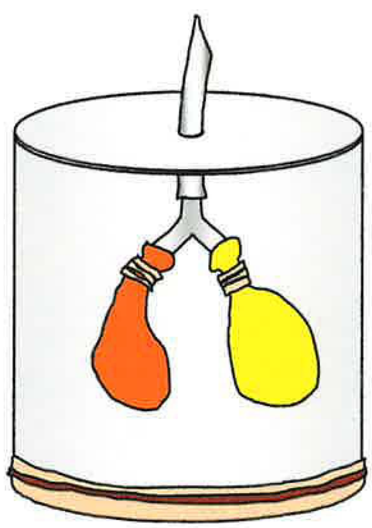
BOYLE'S LAW: THE RELATIONSHIP BETWEEN PRESSURE AND VOLUME

Boyle discovered that if you increase the pressure on a gas, the volume decreases and if you decrease the pressure on a gas, the volume increases. This works if the temperature is kept constant. Think about it: if the volume decreases, there is less space for the gas molecules to move around in so they collide with the container wall more often and increase pressure; if the volume increases, there is more space to move around in, fewer collisions and less pressure. This relationship will be graphically illustrated in the first part of this experiment.

A syringe filled with air is mounted on a block of wood. A second block of wood is placed on top of the syringe. When books are placed on the second block of wood, you will be able to observe the effects of changing pressure on the volume of air in the syringe. Pressure is the force exerted by on object divided by the area of the object. The force exerted by the book is its weight. You don't need to weigh the books or measure their area. You can simply use a "book unit" of pressure. The data will be graphed and you will be able to determine if the relationship is inverse or direct.



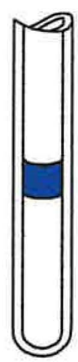
In the second part of this experiment, an application of Boyle's Law is illustrated with a lung demonstration model. The model consists of two balloons (simulating two expandable lungs) connected to a Y-tube contained in a cylinder (chest cavity). The Y-tube is open to the air so that air can enter and exit the balloons in the same way air can enter and exit your lungs. A rubber sheet stretched across the bottom of the cylinder seals it so that no air can enter or leave it. The rubber sheet can be pressed in or pulled away to simulate the action of the diaphragm. In this way, the air that is contained within the cylinder can change in volume and pressure.



CHARLES'S LAW: THE RELATIONSHIP BETWEEN VOLUME AND TEMPERATURE

Charles determined that when the temperature of a gas increases, the volume increased and when the temperature of a gas decreased (cooled), the volume also decreased – this works if the temperature is measured in Kelvin and the pressure is held constant. Think about it: by increasing the temperature, the gas particles moves faster and collide more with the container; if the pressure is to be kept constant, the molecules must have more room so they don't collide as often.

A drop of green dye has been placed in a capillary tube that is sealed at one end. The drop of dye traps a small amount of air in the tube. Since the drop of green dye is free to move in the tube, it will act like a movable piston. The volume in the capillary tube can change but the pressure will remain constant. By warming up and cooling down the air inside the capillary tube, you can observe the effects of temperature on volume.



TEMPERATURE MEASUREMENT IN GAS LAWS: CELCIUS VS. KELVIN

When determining the effect of temperature on a gas, the Kelvin temperature scale must be used. Why is Kelvin used instead of Celsius? The zero points in both the Celsius and Fahrenheit scales were chosen. The zero point in the Kelvin scale was not chosen – the Kelvin scale is "nature's scale" and begins with the coldest possible value for its zero point. This is why 0K (zero Kelvin) is known as *absolute zero*. [NOTE: the degree symbol (°) is not used with the Kelvin scale to reinforce that its zero point was not chosen]. Consider this: suppose you order hot apple pie and the waitress brings you pie straight from the refrigerator at 0°C. You tell her you want your pie hotter, that you'd like your pie twice as hot. What temperature should the pie be? If you use the Celsius scale, you might state 0°C but you would realize the pie would not be hotter. The Celsius scale is not useful in this case.

Temperature is a measure of the average kinetic energy of the molecules. The more the molecules move the higher the temperature. What temperature should a gas be when the molecules are no longer moving? Is it possible to have a temperature below this point? If there is no movement, then there will be no pressure. Think about it: if the gas molecules are not moving, there are no collisions with the container and therefore, no pressure.

In this part of the experiment, your instructor will demonstrate the relationship between pressure and temperature and you will use the data to determine absolute zero. This would be the temperature at which there is no movement and no pressure.

The absolute zero demonstrator consists of a pressure gauge attached to a hollow steel ball (see figure 5). The hollow ball is filled with air. The ball will be immersed in liquids at three different temperatures and the pressure of the air inside the ball will be measured by the pressure gauge. The data points will be plotted and the best-fitting straight line will be drawn. The line will be extrapolated to the point where the pressure is zero. This point will be the absolute zero.

DIRECT VS. INDIRECT RELATIONSHIPS
In each of these experiments two quantities (variables) are compared. One of the variables is changed and the effect on the other variable is observed. In part I, the pressure is changed and the effect on the volume is observed. The relationship between the two variables can be described as direct or indirect. If the first variable is increased and the second variable also increases, the relationship is a direct relationship. For example, if you earn $10 per hour, the more hours you work, the more money you earn. Likewise, the fewer hours you work, the less money you earn. This is a direct relationship. In indirect relationships, when one variable is increased, the other variable is decreased. The more hours you work, the less free time you have available. Likewise, the fewer hours you work, the more free time you have available. You will characterize the relationships studied in this experiment as either direct or indirect.

SAFETY AND WASTE DISPOSAL
Avoid contact with isopropyl alcohol and dry ice. These reagents will be used by your instructor during a demonstration. Wear goggles while observing the demonstration. The reagents you will use in this experiment are not toxic.

Take care when using a flame for the water bath. Do not place more than six textbooks on the syringe and take care not to let them fall off.

Wash hands with soap and water before leaving and especially before eating or drinking.

No special disposal is required. Water may be disposed of down the sink. Return all materials to the hood.

PROCEDURE

I. BOYLE'S LAW: THE RELATIONSHIP BETWEEN PRESSURE AND VOLUME

A. Volume-Pressure Relationship

1. Set up the apparatus as illustrated in figure 1. Remove the rubber tip from the syringe and pull the plunger out to about the 30-cc mark. Insert the top of the plunger into the hole in the top block. Push the plunger in to the 25-cc mark and replace the rubber tip. Insert the bottom of the syringe assembly into the base block.

2. Gently tap the block until a consistent volume reading is obtained. Record this initial volume reading as zero pressure ("book units") in the data table.

3. Carefully place one book on top of the top block. You will need to steady the book. Gently tap the book until a consistent volume reading is obtained. Record this volume reading in the data table.

4. Repeat step 6 until six books have been placed on the block. Take care to steady the books each time. **DO NOT PLACE MORE THAN 6 BOOKS ON THE BLOCK.**

5. Repeat the process of recording volume readings as the books are removed one by one. Remember to tap the top book until a consistent volume reading is obtained. Steady the remaining books.

6. Using the graph paper provided, label the vertical axis (y-axis) pressure ("book units") and the horizontal axis (x-axis) volume. Volume is measured in cubic centimeters (cc).

7. Number the major divisions along each axis. Choose a scale so that the graph will cover at least one-half page.

8. Plot the data. Draw a smooth curve between the data points.

B. Physiological Application of Boyle's Law

1. Obtain the lung demonstration model from the hood.

2. Pull out the rubber sheet as illustrated in figure 2. Observe the behavior of the balloons. Take care not to puncture the rubber sheet.

3. Release the rubber sheet and allow it to return to its original position. Observe the behavior of the balloons.

II. CHARLES'S LAW: THE RELATIONSHIP BETWEEN VOLUME AND TEMPERATURE

1. Prepare a warm water bath in a 150-mL beaker. The water bath should be heated to 60°C.

2. While the water bath is warming up, obtain the apparatus for demonstrating Charles' Law from the hood (see figure 3).

3. Using a wax pencil, mark the capillary tube at the lower edge of the green dye.

4. Immerse the capillary tube, open end up, in the warm water bath. Allow the gas to warm up to the temperature of the warm water bath. Remove the tube and observe any change in the volume of air trapped below the dye.

5. Use your thumb to mark the new position of dye and remove the tube from the water bath.

6. Dry the tube and use the wax pencil to mark your thumb position on the tube. Record your observations on the data sheet.

7. Repeat steps 4, 5, and 6 with an ice water bath.

8. Wipe the tube with a paper towel to remove the wax pencil lines. Return the capillary tube with the dye inside to the hood.

III. ABSOLUTE ZERO

1. Your instructor will immerse the hollow bulb of the absolute zero demonstrator in ice water (0°C), boiling water (100°C) and a mixture of dry ice and isopropyl alcohol (-78°C). Record the temperature and pressure for each of the immersions.

2. Using the graph paper provided, label the vertical axis (y-axis) pressure (lbs/in^2) and the horizontal axis (x-axis) temperature (°C)

3. Number the major divisions along each axis. Choose a scale so that the graph will cover at least one-half page.

4. Draw a best-fitting straight line (a straight line which comes closest to passing through all three data points).

5. Extend the straight line until it crosses the x-axis. Determine the temperature at which the pressure is zero – this is the absolute zero.

Name________________

REPORT FOR PROPERTIES OF GASES

I. BOYLE'S LAW: THE RELATIONSHIP BETWEEN PRESSURE AND VOLUME

A. Volume-Pressure Relationship

1. Data

Pressure adding books ("book units")	Volume of air in syringe (cc)	Pressure removing books ("book units")	Volume of air in syringe (cc)
________	________	________	________
________	________	________	________
________	________	________	________
________	________	________	________
________	________	________	________
________	________	________	________
________	________	________	________

2. What does your graph indicate about the relationship between pressure and volume? Is it direct or indirect?

3. What happens to volume if pressure is increased? What happens to volume if pressure is decreased?

4. If the pressure is increased, does the number of gas molecules increase, decrease or remain the same?

5. Using your graph, predict the volume if the pressure is 3.5 “book units”.

B. *Physiological Application of Boyle’s Law*

1. What happened when you pulled out the rubber diaphragm on the lung demonstrator?

2. Explain your observations.

3. What happened when you released the rubber diaphragm on the lung demonstrator?

4. Explain your observations.

II. CHARLES’S LAW: THE RELATIONSHIP BETWEEN VOLUME AND TEMPERATURE

1. What happened to the volume of air trapped in the capillary tube when the temperature was increased?

2. What happened to the volume of air trapped when the temperature was decreased?

3. What is the relationship between temperature and volume? Is it direct or indirect?

4. If the temperature is increased, does the number of gas molecules increase, decrease or remain the same?

III. ABSOLUTE ZERO

1. Data

Temperature (°C)	Pressure (lbs/in^2)
________	________
________	________
________	________

2. According to your graph, what is the temperature of absolute zero?

3. How does this value compare with the accepted value of absolute zero?

4. Why is this temperature considered the lowest temperature possible?

Name________________

PROPERTIES OF GASES POSTLABORATORY REPORT

1. Based on your observations from this experiment, will it take more gas to fill an empty propane tank in the summer or in the winter (assume the propane tank has the same volume in each case)? Explain your answer.

2. A popular movie depicted an explosion on an airplane that left a gaping hole in the passenger cabin. According to the movie, everyone and everything that wasn't anchored down was swept out of the opening. Use the gas laws to explain how this could happen.

3. Most aerosol cans (spray cans) provide a warning not to incinerate (burn). Use the gas laws to explain why you would not want to put an aerosol can in a fire.

4. A piece of metal has a temperature of 0°C. A second piece of identical metal is twice as hot. What is the temperature of the second piece of metal?

5. In your own words, why is it necessary to use absolute temperatures, or the Kelvin scale, when referring to gases?

Name________________

PROPERTIES OF GASES
PRELABORATORY QUESTIONS

1. In your own words, describe Boyle's law.

2. In Part I. A., how will you measure a change in volume?

3. In Part I. B., what parts of the human body are represented by the pieces of the lung demonstration model below?

   The two rubber balloons

   The Y-tube

   The open end of the Y-tube

   Plastic cylinder

   Rubber sheeting

4. In your own words, describe Charles's law.

5. In Part II., how will you measure a change in volume?

6. List all safety precautions to be taken in this lab.

SPECIFIC HEAT OF A METAL

INTRODUCTION

Have you ever stirred a cup of hot cocoa and found that the spoon is hot but the cocoa is barely warm? Both have had the same heat applied to them – why such a difference? Have you ever walked on the beach and found the sand too hot to walk on but the water too cool to swim in? The sun shines down on the sand and the water – why is the sand much warmer than the water? Suppose you have a kettle of boiling water and a cup of boiling water. Do they have the same temperature? Do they contain the same quantity of heat? This experiment will help you answer these questions.

HEAT VS. TEMPERATURE

Temperature is a measure of the average kinetic energy of the particles of matter in the sample. In other words, the temperature of a substance is related to the speed of its atoms or molecules; the faster they move the higher the temperature. How can the speed of the molecules be measured? It cannot be measured directly so an indirect method must be devised. Most substances expand when their temperature is raised. The amount of expansion can be measured and this is how mercury and alcohol thermometers work. Mercury and alcohol expand when heated and contract when cooled. (Unlike silvery mercury, alcohol is actually colorless so a red or blue dye is usually added to aid in reading).

Heat is a form of energy. How are heat and temperature related? Heat flows from an area (or object) of high temperature to an area (or object) of low temperature. In other words, heat flows from a hot object to a cold one. If you set a cold drink on your desk, heat flows from the surroundings to the drink and warms it up (the surroundings cool down but this is not detectable because of the huge mass of the surroundings). If you set a cup of hot coffee on your desk, heat flows from the coffee to the surroundings and the coffee cools down (the surroundings warm up but again this is not detectable because of the mass of the surroundings).

UNIT OF HEAT

The basic unit of mass is a kilogram and the basic unit of length is a meter. These are things we can touch. The unit of heat is not something we can touch but it can be defined. The basic unit of heat is the calorie and it is defined as the amount of heat required to raise the temperature of one gram of water one degree Celsius (1°C). Note that the definition of the calorie involves mass, temperature change and a particular substance, water.

The calorie (abbreviated cal) is actually a very small unit and is not useful in measuring the heat released in chemical reactions. For these measurements, the kilocalorie (kcal) is used. The calorie is also too small to use for the energy contained in food and the Calorie (Cal) is used, which is equal to the kilocalorie. Note that the lower case is used for the chemical calorie and the upper case is used to express the food calorie, which is a kilocalorie.

The unit of heat energy most often used in chemistry is the SI unit called the joule. The calorie can be defined in terms of the joule:

$$1 \text{ cal} = 4.184 \text{ joule (J)}$$

SPECIFIC HEAT (HEAT CAPACITY)

If you mixed equal amounts of hot (100°C) and cold (0°C) water and measured the temperature after stirring the mixture for a couple of minutes, would you predict the final temperature of the mixture to be closer to 100°C, closer to 0°C, or about 50°C? The mixture would be about 50°C. The heat flowing out of the hot water would flow into the cold water and warm it up and the resulting mixture would be about 50°C. What if you doubled the amount of hot water? Would you predict the final temperature of the mixture to be above 50°C, below 50°C, or about 50°C? The mixture would be above 50°C. The heat flowing out of the hot water would warm up the cold water as in the first example. The difference is the mass. This time there is twice the mass of hot water and therefore twice as much heat energy. That heat will flow to the cold water and the temperature of the resulting mixture would be above 50°C.

By now you realize that the starting and ending temperatures and the mass of the substance are two factors determining the amount of heat transferred. Is there a third factor? What if you substituted alcohol for the cold water? What would be the final temperature of the mixture? It would be difficult to predict without knowing something about alcohol itself. The third factor needed to determine the amount of heat transferred is the sample's **specific heat** (also called heat capacity). Recall that one calorie is the amount of heat required to raise the temperature of one gram of water one degree Celsius. Less than one calorie is needed to raise the temperature of one gram of alcohol one degree Celsius because alcohol has a lower specific heat than water. The specific heat of a substance is defined as the number of calories required to raise the temperature of one gram of that substance one degree Celsius. The specific heat of alcohol is 0.55 cal/g°C, nearly half that of water. The amount of heat flowing into or out of a substance depends on the mass of the substance, the change in temperature and the specific heat of the substance.

Using the SI units, the specific heat of water is 4.184 J/g°C and the specific heat of alcohol is 2.30 J/g°C.

DETERMINING THE SPECIFIC HEAT OF A METAL

Boiling point, melting point and density of a substance are characteristic properties of that substance. Likewise, the specific heat of a substance is a characteristic property. The specific heat of a known quantity of a metal can be determined by warming up the metal and placing it in cold water. The metal will cool down and the water will warm up. In this heat transfer process the heat flowing into the water is equal to the heat flowing out of the metal and the final temperature of the components of the mixture (metal and water) are the same. By calculating the heat flowing into the water, you can determine the specific heat of the metal.

Think about it: if the water warms up, heat must be flowing into it; likewise, if the metal cools down heat must be flowing out of it. The heat flow out of the metal must equal the heat flow into the water:

heat flow out of metal (heat lost) = heat flow into water (heat gained)

The heat flow into the water can be calculated using an algebraic equation relating the three factors determining the amount of heat transferred: temperature change, mass of substance and specific heat of substance.

$$\text{heat gained} = (\text{mass}) \times (\Delta T) \times (\text{specific heat of water})$$

The change in temperature can be symbolized by ΔT, where Δ is read "the change in".

The heat flow out of the metal can be calculated by using the same equation, substituting the mass of the metal, the change in temperature of the metal, and specific heat of the metal:

$$\text{heat lost} = (\text{mass}) \times (\Delta T) \times (\text{specific heat of metal})$$

Experimental values for mass and temperature change and the given value of specific heat of water are used. In the transfer of heat, the heat gained by the water is equal to the heat lost by the metal. Now you have sufficient data to calculate the specific heat of the metal.

Case Study:

A 50.0-g sample of a metal at 100.0°C is added to 100.0 g of water at 25.0°C. The final temperature of the mixture (water and metal) is 28.3°C. What is the specific heat of the metal?

Solution:
The heat flow from the metal is equal to the heat flow into the water. First calculate the heat flow into the water.

$$\begin{aligned}\text{heat gained} &= (\text{mass}) \times (\Delta T) \times (\text{specific heat of water}) \\ &= (100.0\text{ g}) \times (28.3°C - 25.0°C) \times (4.184\text{ J/g °C}) \\ &= 1380.72\text{ J} \\ &= 1380\text{ J (three sig figs)}\end{aligned}$$

The heat lost is equal to the heat gained.

$$\text{heat lost} = (\text{mass}) \times (\Delta T) \times (\text{specific heat of metal})$$

Isolate the unknown quantity, specific heat of metal

$$\text{Specific heat of metal} = \frac{\text{heat lost}}{(\text{mass})(\Delta T)}$$

Substitute the mass of the metal and the temperature change from the data.

$$\begin{aligned}\text{specific heat of metal} &= \frac{1380\text{ J}}{(50.0\text{ g})(100.0°C - 28.3°C)} \\ &= \frac{1380\text{ J}}{(50.0\text{ g})(71.7°C)} \\ &= \frac{1380\text{ J}}{(3590\text{g°C})}\end{aligned}$$

[NOTE: 50.0 x 71.7 = 3585.0 on the calculator; expressing the result to three sig figs it becomes 3590]

$$\text{specific heat of metal} = 0.384\text{ J/g°C}$$

[NOTE: 1380/3590 = 0.3844011 mathematically; expressing the correct scientific result to three sig figs, it becomes 0.384

Notice that the units for specific heat are J/g°C.

CALORIMETER

When the hot metal and the cold water are mixed together in a container, it is assumed that the heat lost by the metal equals the heat gained by the water. Is this assumption entirely correct? Does some of the heat escape up to the air or even through the walls of the container? Some heat does escape and that heat cannot be transferred to the cold water.

To minimize this heat loss, a well-insulated container is used. This type of container is called a calorimeter. Styrofoam is a good heat insulating material and will be used in this experiment instead of glass beakers.

SAFETY AND WASTE DISPOSAL

The reagents used in this experiment, solid metal samples and water, are not toxic. After heating, the metal samples will be about 100°C. Take care when handling the hot samples. Wear goggles at all times. Wash hands with soap and water before leaving and especially before eating or drinking.

No special disposal is required. Water may be disposed of down the sink. Return metal samples to the instructor.

PROCEDURE

1. Fill a 400 mL beaker approximately 3/4 full of tap water. Place the beaker on a wire gauze supported by a ring attached to a ring stand and heat the water to boiling.

2. While the water is heating, obtain a sample of aluminum from your instructor and weigh the sample to the nearest 0.01 g. Record the mass of the sample on your data sheet.

3. Place a styrofoam cup inside a 250 mL beaker (to keep it from tipping over) and add 75.0 mL of distilled water. Record the volume of water on the data sheet.

4. Immerse the aluminum in the boiling water. Allow the aluminum to remain in the water bath for about 5 minutes. Measure the temperature of the boiling water to the nearest 0.1°C and record the temperature on the data sheet. (You may not assume that the water is 100.0°C because your thermometer may not be accurate). Assume the aluminum sample is the same temperature as the hot water and record the temperature of the aluminum sample to the nearest 0.1°C.

5. Measure and record the temperature of the cold water, to the nearest 0.1°C, just before adding the hot aluminum sample.

   Using crucible tongs, remove the aluminum sample from the boiling water and place it in the styrofoam cup containing the cold water.

6. Carefully stir the water with your thermometer, taking care not to break it. Note the highest temperature reached and record this temperature to the nearest 0.1°C.

7. Dry off the known aluminum sample and return it to your instructor. Perform the specific heat calculations on the report sheet. Perform the percentage error calculation on the report sheet. Show your results to your instructor and obtain your instructor's signature on the report sheet. If your results are within experimental error, your instructor will give you an unknown sample. If your experimental error is too large, you will repeat the procedure with the aluminum sample before going forward with the unknown.

8. Once you have obtained the unknown sample from your instructor, repeat steps 1-7 with the unknown.

10. Dry off the metal sample and return it to your instructor. Return the styrofoam cup to the hood (do not throw it away).

Name________________

REPORT FOR SPECIFIC HEAT OF A METAL

SPECIFIC HEAT OF ALUMINUM

	Trial 1	Trial 2 (if needed)
1. Mass of aluminum sample	__________g	__________g
2. Volume of water in styrofoam cup	__________mL	__________mL
3. Temperature of boiling water	__________°C	__________°C
4. Temperature of **cold** water **before** adding aluminum sample	__________°C	__________°C
5. Temperature of **warm** water **after** adding aluminum sample	__________°C	__________°C

CALCULATIONS

Mass of water in Styrofoam cup

*Answer*__________

Specific heat of aluminum

*Answer*__________

Percentage error

(See your textbook for the theoretical value of the specific heat of aluminum)

$$\text{percentage error} = \left| \frac{\text{theoretical value} - \text{experimental value}}{\text{theoretical value}} \right| \times 100\%$$

*Answer*__________

Instructor's Signature ____________________________________

Name________________

REPORT FOR SPECIFIC HEAT OF A METAL

SPECIFIC HEAT OF UNKNOWN METAL

	Trial 1	Trial 2 (if needed)
1. Unknown number	__________	__________
2. Mass of unknown sample	__________g	__________g
3. Volume of water in styrofoam cup	__________mL	__________mL
4. Temperature of boiling water	__________°C	__________°C
5. Temperature of **cold** water **before** adding unknown sample	__________°C	__________°C
6. Temperature of **warm** water **after** adding unknown sample	__________°C	__________°C

CALCULATIONS

Mass of water in Styrofoam cup

*Answer*__________

Specific heat of unknown

*Answer*__________

Name_______________

SPECIFIC HEAT OF A METAL POST LABORATORY REPORT

1. If you have ever eaten hot pizza, you know that the sauce can burn your mouth but the crust, at the same temperature, does not. Explain.

2. If 100 g of water at 10°C are poured into 100 g of pancake batter at 50°C (the batter contains flour, sugar, eggs, oil) what would be the temperature of the resulting mixture? Circle your answer. [NOTE: You don't know the specific heat of the batter. This question is a thought question, not a calculation question]

   (a) about 50°C
   (b) about 30°C
   (c) above 30°C
   (d) below 30°C

   Explain your choice.

3. If 100.0 g of water at 70.0°C is mixed with 50.0 g of water at 20.0°C and the mixture comes to a constant temperature, what is the temperature of the mixture?

   *Answer*__________

4. The specific heat of alcohol is 0.55 cal/g°C and the specific heat of mercury is 0.033 cal/g°C. Based on this information, which would respond to a change in temperature more quickly, an alcohol thermometer or a mercury thermometer?

5. Will 100 mL of hot chocolate at 70°C cool more quickly in a cup made of copper (specific heat = 0.093 cal/g°C) or a cup made of silver (specific heat = 0.056 cal/g°C)? Explain your reasoning.

6. In an effort to lose weight, a student submerged his body in a bathtub containing 50 L of water and in an hour managed to raise the temperature of the water from 31.5°C to 35.0°C. How many kilocalories did the student give off?

*Answer*__________

Name_______________

SPECIFIC HEAT OF A METAL
PRELABORATORY QUESTIONS AND PROBLEMS

1. The word *thermometer* comes from two Greek words: *therm*, which means heat, and *meter*, which means measure. Does the thermometer measure heat? Explain your answer.

2. A Calorie is equal to (circle all of the correct answers)
   (a) 4.184 J
   (b) 1 kilocalorie
   (c) a food calorie
   (d) 1000 calories

3. A full kettle of water is heated to 100°C. A cup of water is poured from the kettle of water at 100°C and set alongside the kettle. Which of the following are correct statements?
   (a) The kettle of water and the cup of water cool at the same rate.
   (b) The kettle of water cools faster than the cup of water.
   (c) The cup of water cools faster than the kettle of water.

4. If you mix 100 mL of water at 100°C and 100 mL of water at 0°C, what will be the final temperature of the mixture?
   (a) above 50°C
   (b) below 50°C
   (c) about 50°C

5. If you mix 100 mL of water at 100°C and 50 mL of water at 0°C, what will be the final temperature of the mixture?
   (a) above 50°C
   (b) below 50°C
   (c) about 50°C

6. Why is a styrofoam cup used as the calorimeter?

7. What are the three factors affecting the transfer of heat?

8. If 50.0 mL of water at 40.0°C are added to 200.0 g of crushed rock at 20.0°C, the temperature of the resulting mixture is 32.5°C. What is the specific heat of the rock?

*Answer*__________

9. List all safety precautions to be taken in this lab.

MOLECULAR MODELS

INTRODUCTION

Models can be used to describe everything from weather patterns to economic trends. Scientists utilize mathematical and conceptual models to interpret data. Chemists use physical models, the type of models you will build in the lab, to illustrate the number and kind of bonds formed between atoms in molecules. Three-dimensional models were critical to Watson and Crick's discovery of the structure of DNA. Constructing models can also help clarify the nature of the molecule: is it polar or nonpolar? If the bonds in a molecule are polar, does that necessarily mean that the molecule itself is polar? The 3-D structure of a molecule is difficult to visualize from its formula. Creating molecular models will allow you to see the shape of a molecule. Why would you be interested in the shape of a molecule? Shape has a great deal of biological, as well as chemical, importance.

Did you know that there is a "handedness" to the shape of some biological structures as well as molecules? Helical seashells are right handed, all but one of the 20 amino acids that make up naturally occurring proteins are left handed and for evolutionary reasons (far from understood), most people are right handed. Most of the natural sugars are right handed and DNA turns to the right. The left hand and right hand versions of a molecule called limonene are responsible for the odor of oranges and the odor of lemons.

Prior to 1963 the drug thalidomide was used to alleviate the symptoms of morning sickness in pregnant women. In 1963 it was discovered that thalidomide was the cause of horrible birth defects in many children born to mothers who had used the drug. The right handed version of thalidomide had the intended effect of curing morning sickness but the left handed version (also present in the drug in equal amounts) caused the birth defects.

On a less dramatic scale, the left-hand version of ibuprofen exerts its effect in 12 minutes. A mixture of left and right handed ibuprofen takes 30 minutes.

In this experiment, you will write out the Lewis structure of various compounds, construct models of the compounds, determine the number of bonds (single, double, triple) between any two atoms, draw the 3-D structure of the compound, and identify the compound as polar or nonpolar.

POLAR VS. NONPOLAR

What is *polar*? The word pole can mean either end of an axis, such as the two poles of the earth or two opposing forces or parts, such as the poles of a magnet or the terminals of a battery. Molecules can also have poles: positive and negative regions.

The positive and negative regions of a molecule arise from a difference in electronegativity (the attraction an atom has for the electrons shared in a chemical bond). If the two atoms bonded together have different electronegativities, the electrons will be more strongly attracted to the atom with the higher electronegativity. The result is that one region of the molecule is negative and the other region is positive. Think about it: if electrons are negative and are pulled in closer to one atom, that region of the molecule will have a partial negative charge. Why not a full negative charge? Remember that in ionic bonds, electrons are transferred from one atom to another. The complete transfer of an electron gives a full negative charge. In covalent bonds the electrons are shared, not transferred, so the unequal sharing results in a partial charge. Likewise, the unequal sharing results in a partial positive charge for the other atom in the molecule. Think about it: if the negative electrons are pulled away from an atom, it will be left with a partial positive charge. The two opposing charges are the poles and the molecule is said to be polar.

Nonpolar bonds occur when there is little or no electronegativity difference. Oxygen, O_2, is completely nonpolar. The electronegativity difference is zero. Methane, CH_4, is nonpolar because the electronegativity difference between carbon and hydrogen is small.

Why would you want to know about polar and nonpolar molecules? Most of what is inside your body's cells is polar and therefore water soluble (dissolveable in water). When you took a shower this morning your skin did not dissolve. That's because your cells are wrapped in a nonpolar coating (membrane). The nonpolar membrane keeps the water out and the cellular material in. You also take advantage of polarity when you do your laundry. Oil stains are nonpolar and water is polar. Have you heard of the saying "oil and water don't mix"? You cannot dissolve the grease stain in water. Instead, you use a detergent that has a nonpolar portion to wrap around the grease and a polar portion to dissolve in water. It helps to know about polarity when using solvents around your house. Fingernail polish won't come off with water – you'll need a nonpolar solvent to remove it. Chemists use polarity in everything from selecting a solvent to predicting the products of a reaction (we also use it to remove our nail polish and clean our clothes).

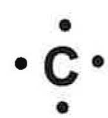
ELECTRON DOT (OR LEWIS) STRUCTURES

Before you can construct a model of a molecule, you must use the chemical formula to draw the electron dot structure. These structures are used to show the *valence* electrons of an atom (the outermost electron in the highest numbered energy level that are involved in chemical bonding). A single dot is used to represent a single electron. The dots are placed on each side of the element's symbol and are first placed singly, then paired for any element with more than four valence electrons. The number of valence electrons for the representative elements (groups IA – VIIIA) corresponds to the group number on the periodic table. For example, sodium is in group 1A and has one valence electron; likewise, iodine is in group VIIA and has seven valence electrons.

Covalent bonds are formed when two nonmetals share their valence electrons by orbital overlap. In electron dot structures, covalent bonds are represented by a pair of dots placed between the two bonding atoms. A double bond would be illustrated with two pairs of electrons and a triple bond with three pairs.

How to draw the electron dot structure of carbon

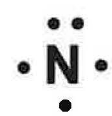
1. Locate it on the periodic table. Carbon is in group IVA. How many valence electrons does it have? (4)
2. Draw the symbol for carbon. Draw one dot for each electron. Place the electrons on each "side" of the symbol.



3. Notice that carbon has four unpaired electrons. Carbon can make four bonds by sharing its four electrons with one or more atoms. When carbon has made four bonds, it will have an octet of electrons and all of its valence orbitals will be filled.

How to draw the electron dot structure of nitrogen

1. Locate it on the periodic table. Nitrogen is in group VA. How many valence electrons does it have? (5)
2. Draw the symbol for nitrogen. Draw one dot for each electron. Place the electrons on each "side" of the element's symbol. Nitrogen has more than four electrons so two of the electrons must be paired. These are called a "lone pair."



3. Notice that nitrogen has three unpaired electrons and one lone pair. How many bonds can nitrogen make? (3) What happens to the lone pair? (It contributes to the overall shape of a molecule containing a nitrogen atom).

How to draw the electron dot structure of chlorine, Cl_2

1. Find chlorine on the periodic table. It is in group VIIA. How many valence electrons are in chlorine? (7)
2. Draw the symbol for chlorine. Draw one dot for each electron. Place the electrons on each "side" of the symbol. Chlorine has more than four electrons so some of the electrons must be paired. There are three lone pairs in each chlorine atom. How many bonds can chlorine make? (1)

3. Draw two chlorine atoms side by side. Draw in three lone pairs for each chlorine atom and draw a pair of electrons shared between the two atoms. The shared pair of electrons represents a single bond. Each chlorine atom now has an octet. Is chlorine polar or nonpolar? (Nonpolar because the electronegativity difference is zero).

An* alternate *method to draw the structure chlorine

1. Draw the skeleton structure of chlorine. (Write the symbols of the atoms that will be bonded to each other – in this case, simply write them next to each other).

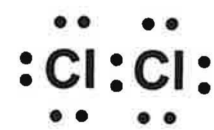
Cl Cl

2. Calculate the total number of valence electrons. (7 for each chlorine atom for a total of 14).

3. Draw a pair of electrons (2) between the two chlorine atoms to represent a bond.

Cl:Cl

4. Subtract the two electrons from the total calculated (14-2 = 12). Distribute the remaining electrons among the atoms so that each has no more than eight electrons.

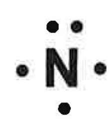


5. Check to see that all atoms in the molecule have an octet (a duet for hydrogen).

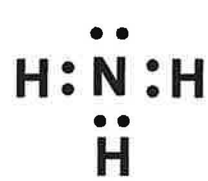
How to draw the electron dot structure of NH_3 (ammonia)

1. You need to determine which atom is the central atom, the atom to which other atoms bond. You know that nitrogen can make three bonds. How many bonds can hydrogen make? (1) It seems most likely that nitrogen is the central atom and not hydrogen. Think about it: if hydrogen can form only one bond, it can bond once to another hydrogen to form H_2 (does exist) or once to nitrogen to form NH (does not exist) but it cannot be the central atom and bond three times to form NH_3. In general, the first atom in the formula is the central atom and hydrogen is never a central atom.

2. Draw the symbol for nitrogen with three unpaired electrons and one lone pair.

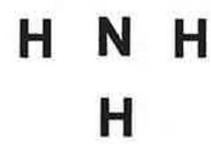


3. Draw three hydrogen atoms on the sides with single electrons. Each hydrogen atom has a single electron. Draw in hydrogen's electrons as dots and show a pair of electrons shared between the hydrogen and nitrogen atoms. The shared pair of electrons represents a single bond. The nitrogen atom now has an octet and each hydrogen atom has a duet. Is NH_3 polar or nonpolar? (Polar because the electronegativity difference between nitrogen and hydrogen and the shape of NH_3.)



Alternate method for NH_3

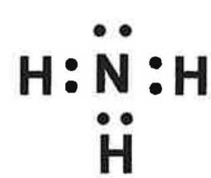
1. Draw the skeleton structure of NH_3.



2. Calculate the total number of valence electrons. (5 for the nitrogen atom, one for each hydrogen, for a total of 8).

3. Draw a pair of electrons (2) between the atoms that have covalent bonds.

4. Subtract the bonding electrons from the total calculated (18-6 = 2). Distribute the remaining electrons among the atoms so that each has no more than eight electrons (no more than two for hydrogen).



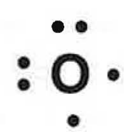
5. Check to see that all atoms in the molecule have an octet (a duet for hydrogen).

What happens if there are double or triple bonds between atoms?

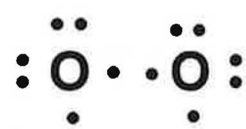
How to draw the electron dot structure of O_2

1. Determine the number of valence electrons in oxygen. (6)

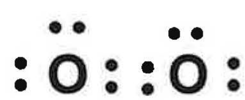
2. Draw the symbol for oxygen. Draw one dot for each electron. Place the electrons on each "side" of the symbol. Oxygen has more than four electrons so some of the electrons must be paired. There are two lone pairs in each oxygen atom. How many unpaired electrons does oxygen have? (2) How many bonds can oxygen make? (2)



3. Draw a second oxygen atom next to the first. Draw in the lone pairs and the single electrons for the second oxygen atom.



How can the two oxygens bond together? Each oxygen can form two bonds. The two oxygen atoms form a double bond. Is oxygen polar or nonpolar? (Nonpolar because the electronegativity difference is zero).



Alternate method for O_2

1. Draw the skeleton structure of O_2.

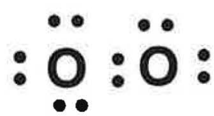
O O

2. Calculate the total number of valence electrons. (6 for each oxygen atom, for a total of 12).

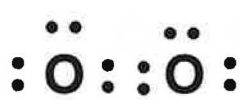
3. Draw a pair of electrons (2) between the atoms that have covalent bonds.

O : O

4. Subtract the bonding electrons from the total calculated (12-2 = 10). Distribute the remaining electrons among the atoms so that each has no more than eight electrons.



5. Check to see that all atoms in the molecule have an octet. In this case one oxygen atom has only six electrons surrounding it. You cannot simply add more electrons because the total number of valence electrons is only 12. If an atom has fewer than eight electrons, you can move a lone pair to form a double bond.



6. Now each oxygen atom has an octet.

What happens if you don't know there are double or triple bonds present?

In the previous example, you knew there would be a double bond. There is a method to determine if there are multiple bonds. Consider the molecule C_2H_4.

1. Calculate the total number of valence electrons. (4 for each carbon atom, 1 for each hydrogen, 12 electrons total).

2. If the number of valence electrons is equal to 6N + 2, where N = the number of atoms other than hydrogen in the formula, the molecule has only single bonds. If the number of valence electrons is two less than 6N + 2, there is one double bond in the molecule (it could also mean there is a ring structure). If the number of valence electrons is four less than 6N + 2, there is one triple bond or two double bonds in the molecule. The next step then, is to apply the "6N + 2 rule" to C_2H_4. In this case, 6N + 2 = 14, where N = 2. The total number of valence electrons is two less than this (14-12 = 2) so the molecule contains a double bond.

3. Draw the skeleton structure. Start the same way as NH_3 by determining the central atom(s). Hydrogen cannot be the central atom so try placing the two carbon atoms as the central atoms. Where do you place the hydrogens? Many molecules are symmetrical so try drawing each carbon with two hydrogens bonded to it instead of three hydrogens bonded to one carbon and one hydrogen bonded to the second. (Note also that if three hydrogens are bonded to one carbon, the second carbon does not obey the octet rule).

H C C H
H H

4. Draw two pairs of electrons between the carbon atoms to represent the double bond (hydrogen atoms can only form single bonds so the double bond must be between the two carbon atoms). Draw in a pair of electrons between each carbon and hydrogen to signify single bonds.

5. Check to see that all atoms in the molecule have an octet (a duet for hydrogen).

6. Is C_2H_4 polar or nonpolar? (Nonpolar because the electronegativity difference between carbon and hydrogen is very small).

How to draw the electron dot structure of C_2H_2
(NOTE: this formula is different from the one above; look at the subscript for hydrogen)

1. Calculate the total number of valence electrons. (4 for each carbon atom, 1 for each hydrogen, 10 electrons total).

2. Apply the "6N + 2 rule". In this case, 6N + 2 = 14, where N = 2. The total number of valence electrons is four less than this (14-10 = 4) so the molecule contains one triple bond or two double bonds.

3. Draw the skeleton structure.

H C C H

4. Is there a triple bond or are there two double bonds? Again, hydrogen cannot form multiple bonds, only the two carbon atoms. There must be a triple bond between the two carbon atoms. Draw three pairs of electrons between the carbon atoms to represent the triple bond. Draw in a pair of electrons between each carbon and hydrogen to signify single bonds.

H:C:::C:H

5. Check to see that all atoms in the molecule have an octet (a duet for hydrogen).

6. Is C_2H_2 polar or nonpolar? (Nonpolar because the electronegativity difference is very small).

How to draw the electron dot structure of N_2

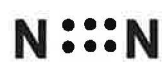
1. Calculate the total number of valence electrons. (5 for each nitrogen atom, 10 electrons total).

2. Apply the "6N + 2 rule". In this case, 6N + 2 = 14, where N = 2. The total number of valence electrons is four less than this (14-10 = 4) so the molecule contains one triple bond or two double bonds.

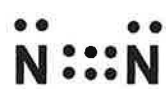
3. Draw the skeleton structure.

N N

4. Is there a triple bond or are there two double bonds? There are only two atoms present - there must be a triple bond between the two nitrogen atoms. Draw three pairs of electrons between the nitrogen atoms to represent the triple bond.



5. Check to see that all atoms in the molecule have an octet (a duet for hydrogen). Neither nitrogen atom has an octet. There are 10 electrons available, only six have been used for bonding. Distribute the remaining electrons so that each atom has an octet.



6. Is N_2 polar or nonpolar? (Nonpolar because the electronegativity difference is zero).

How do you draw the electron dot structure of an ion?

Ions are drawn in much the same way as molecules. The difference is in calculating the number of valence electrons. The charge must be taken into consideration when tabulating the number of valence electrons. If the ion has a negative charge, electrons are added to the total number of valence electrons. If the ion has a positive charge, electrons are deducted.

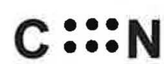
Consider CN^-:

C N

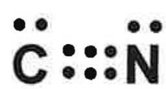
1. Calculate the total number of valence electrons. (4 for carbon, 5 for nitrogen, one additional electron due to the negative one charge, 10 electrons total).

2. Apply the "6N + 2 rule". In this case, 6N + 2 = 14, where N = 2. The total number of valence electrons is four less than this (14-10 = 4) so the molecule contains one triple bond or two double bonds.

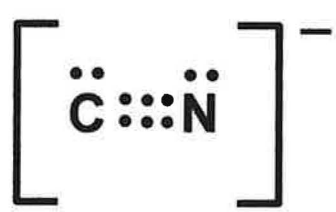
3. Draw the skeleton structure.
4. Is there a triple bond or are there two double bonds? There are only two atoms present - there must be a triple bond between the carbon and nitrogen atoms. Draw three pairs of electrons between the nitrogen atoms to represent the triple bond.



5. Check to see that all atoms in the molecule have an octet (a duet for hydrogen). Neither atom has an octet. There are 10 electrons available, only six have been used for bonding. Distribute the remaining electrons so that each atom has an octet.



6. Because this is an ion, the structure is surrounding by brackets and the charge is placed outside the brackets.

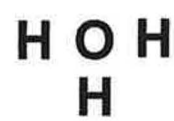


What if the ion is positively charged? Consider H_3O^+

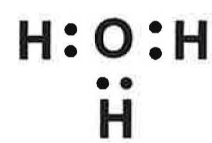
1. Calculate the total number of valence electrons. (6 for oxygen, 1 for each hydrogen, one electron deducted for the positive one charge, 8 electrons total).

2. Apply the "6N + 2 rule". In this case, 6N + 2 = 8, where N = 1. The total number of valence electrons is equal to this so the molecule contains only single bonds.

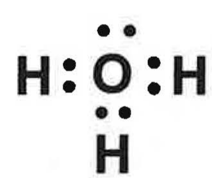
3. Draw the skeleton structure.



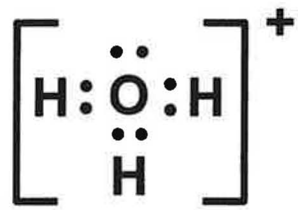
4. Draw in a pair of electrons between the oxygen atom and each hydrogen.



5. Check to see that all atoms in the molecule have an octet (a duet for hydrogen). Oxygen has only six electrons. There are eight electrons available. Distribute the remaining electrons so that has an octet.



6. Enclose the structure with brackets and place the charge outside.



Using Molecular Models

The model kit provided by your instructor contains different colored balls that represent different atoms. The balls have holes in them and each hole represents a single electron available for bonding (lone pairs can also be represented). Links are used to connect two balls. The medium gray rigid links are used to connect two atoms together simulating a single covalent bond. The long gray flexible links are used for double and triple bonds. Two (or three) links connecting two balls signifies a double (or triple) bond. The short white links are used to represent lone pairs. If a model is constructed properly, all the holes will have a link.

The different colored balls are used to represent different atoms according to the following matrix:

White ball – hydrogen, one hole
Black ball – carbon, four holes
Red ball – oxygen, four holes
Blue ball – nitrogen, four holes
Green ball – halogen, four holes

PROCEDURE

1. The data tables provide the molecular formulas for the molecules to be constructed.

2. Draw the electron dot structure from the molecular formula following the examples provided. Record in the data table.

3. Construct a model of the molecule. Choose the correct color and number of holes for your atoms. Use white links to represent lone pairs. Use medium gray rigid links to connect two atoms (balls) together in a single bond. Long gray flexible links can be used to connect two atoms together in double (two links) or triple (three links) bonds.

4. Sketch the 3-D molecular model you have constructed. Record in the data table.

5. Draw the structural formula from your model. When drawing the flat 2-D structural formula, use the symbol of the atom and lines connecting atoms to indicate bonds. Use dots to represent lone pairs. Record in data table. For example:

NH_3

```
H-N-H
  |
  H
```

6. Determine the molecule's polarity. Record either "polar" or "nonpolar" in the data table.

Name________________

REPORT FOR MOLECULAR MODELS

I. CONSTRUCTION OF MOLECULAR MODELS
DATA TABLE

MOLECULAR FORMULA	ELECTRON DOT STRUCTURE	3-D SKETCH OF MODEL	STRUCTURAL FORMULA	POLARITY
HBr				
CH_2Cl_2				
NH_2OH				
C_2H_3Cl				

Name_______________

MOLECULAR FORMULA	ELECTRON DOT STRUCTURE	3-D SKETCH OF MODEL	STRUCTURAL FORMULA	POLARITY
H_2CO				
HCN				
C_2HCl				

Name________________

II. UNKNOWN MOLECULAR MODELS
DATA TABLE

MOLECULAR FORMULA	ELECTRON DOT STRUCTURE	3-D SKETCH OF MODEL	STRUCTURAL FORMULA	POLARITY
#1				
#2				
#3				
#4				
#5				

Name________________

MOLECULAR MODELS
POST-LABORATORY REPORT

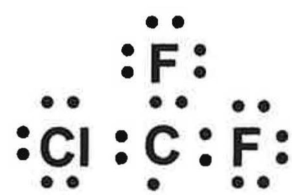
1. If a molecule contains a polar bond, is the entire molecule polar

2. $CHCl_3$ is polar while CCl_4 is nonpolar. Explain.

3. Draw the electron dot structure for the following ions:

ION	ELECTRON DOT FORMULA
NH_4^+	
OH^-	
CO_3^{-2}	

4. Chlorofluorocarbons (CFCs) called *freons* have been used as refrigerants, solvents and propellants in aerosol cans since the 1930s. In 1995 P.J. Crutzen, M.J. Molina and F.S. Rowland were awarded the Nobel Prize in Chemistry for their discovery showing that freons can disturb the natural ozone balance in the stratosphere. In this process, freons first form compounds called *radicals*. The structure below shows dichlorodifluoromethane, called *Freon-12* , as a radical. What makes this structure unusual?



Name________________

MOLECULAR MODELS
PRE-LABORATORY QUESTIONS

1. What is the relation between valence electrons and the number of covalent bonds an element is able to form?

2. Why is a nitrogen atom able to share three pairs of electrons in a nitrogen molecule?

3. Draw the electron dot structure for HCl.

4. Is HCl polar of nonpolar? How can you predict the polarity of HCl?

Name________________

WHAT IS A MOLE?

PART I: HOW MANY JELLY BEANS ARE IN THE JAR?

How many jelly beans are in the jar? You may not count them. How could you figure this out without counting them?

Working in a group of four students total, design an experiment to determine the number of jelly beans in the jar. You will have access to a balance, the jar with jelly beans, an empty jar and 10 jelly beans. Once you have designed your experiment, write out your procedure (in pencil) and show it to your instructor for approval. After you have obtained your instructor's approval, carry out your experiment. Keep your results confidential until your instructor asks for them. Record your procedure and data on the data sheet.

PROCEDURE

Instructor's Approval____

DATA (include units)

CALCULATIONS (show all work)

Final Answer ______________

Name________________

PART II: CHEAPER BY THE DOZEN

You "counted" the number of jelly beans in the jar by weighing them. Likewise, chemists also count by weighing. We count atoms this way because they are very small and there are too many of them to count.

There are four containers, each with one dozen objects. Weigh an empty container and then weigh each of the four containers. Determine and record the weight of each set of objects. You may assume that all containers weigh the same. ***Do not open the containers.***

DATA:

Empty container ___________g

Container #1 ___________g Weight of objects ___________g

Container #2 ___________g Weight of objects ___________g

Container #3 ___________g Weight of objects ___________g

Container #4 ___________g Weight of objects ___________g

QUESTIONS

1. What do all of the containers have in common?

2. Do the four sets of objects weigh the same?

3. Does the type of object weighed make a difference in the weight data?

Name________________

Counting by Weighing

A dozen beads weigh 28.94 g. A "gross" is a dozen dozen, or 144. What would 144 beads weigh? Show your work for calculations *or* describe your thinking below.

*Answer*__________

To solve this problem, you probably multiplied the weight of a dozen (12) beads by 12. You can set this problem up by using the weight of a dozen beads as a conversion factor:

$$\frac{28.94 \text{ g beads}}{1 \text{ dozen beads}}$$

Multiply 12 dozen by the conversion factor:

$$12 \text{ dozen} \times \frac{28.94 \text{ g beads}}{1 \text{ dozen beads}}$$

The calculator answer is 347.28 g; the correct scientific answer is347.3 g, using four sig figs. (The number 12 is an "exact number" and not a measured one, so it is not considered with sig figs).

If you wanted 1,000 beads for a project, would you have to count out 1,000 beads or could you "weigh out" 1,000 beads? Which method would you choose? Explain how you would carry out your choice.

Name________________

III. COUNTING BY WEIGHING AND MOLES

1. Predict the weight of 100 objects from container #4 using a conversion factor. Show how you arrived at your answer.

*Answer*__________

2. Count out 100 of the objects from step 1. (Do not open the container you weighed from Part II. Locate the large container of objects in the hood. It will be marked for you.)

3. Weigh the 100 objects from step 2 on a balance. What is the mass?

   Mass of 100 objects__________ g

4. Does this weight, from the balance, agree with your prediction from step 1?

5. Does counting by weighing work? Explain.

What does a dozen have to do with a mole?
You use the quantity one dozen to buy everything from eggs to donuts. Likewise, chemists use a quantity for counting but our quantity is huge because atoms are so small. Our quantity for counting is the ***mole***. There are 6.02×10^{23} atoms in one mole.

How much does a mole weigh? Does the type of particle you are weighing make a difference? In Part I, one dozen pencils did not weigh the same as one dozen candies. Likewise, one mole of silver atoms will not weigh the same as one mole of helium atoms. One mole of atoms of any monatomic element is exactly equal to the atomic weight in grams of that element.

For example, one calcium atom has an atomic weight of 40.1 amu (atomic mass units) so one mole of calcium weighs 40.1 grams and contains 6.02×10^{23} atoms of calcium. Round off atomic weights to nearest 0.1 gram (unless told otherwise by your instructor).

Previously in this experiment, you used the relationship between one dozen beads and their mass as a conversion factor to determine what 144 beads would weigh. Likewise, the relationship between one mole of a substance and its mass can be used as a conversion factor. The relationship between one mole and the number of atoms in a mole, 6.02×10^{23}, can also be used as a conversion factor.

Case Study:
An experiment you are conducting requires 3.70 moles of magnesium. How many grams of magnesium should you weigh out? How many atoms is this?

To solve this problem, you need to know the atomic weight of magnesium. Magnesium is number 12 on the periodic table and its symbol is Mg. Underneath the symbol is the atomic mass, 24.3050. This number expressed in grams is the mass of one mole of magnesium.

Now you can use this relationship as a conversion factor:

$$1 \text{ mole Mg} = 24.3 \text{ grams}$$

$$\frac{1 \text{ mole Mg}}{24.3 \text{ g Mg}} \quad \text{or} \quad \frac{24.3 \text{ g Mg}}{1 \text{ mole Mg}}$$

To find the mass of 3.70 moles:

$$3.70 \text{ moles Mg} \times \frac{24.3 \text{ g Mg}}{1 \text{ mole Mg}} = 89.9 \text{ g Mg}$$

To find the number of atoms in 3.70 moles:

$$3.70 \text{ moles Mg} \times \frac{6.02 \times 10^{23} \text{ atoms}}{1 \text{ mole Mg}} = 2.23 \times 10^{24} \text{ atoms Mg}$$

NOTE: If you are using a scientific calculator, you have a function that will help you with exponents. Look for the **exp** key or **EE** key on your calculator. Enter the first part of the expression, for example 6.02, into your calculator. Press the **exp** key or **EE** key and your calculator will display 00 at the far right. This is for the exponent. The "X 10" part of the expression is already entered. Now enter the exponent, in this case 23. The display at the far right will change from 00 to 23. Remember that this is read "6.02 X 10^{23}. Following the example above, multiply by 3.70 and press "=". The display should now read "2.2274" with "24" at the far right. This is read "2.2272 X 10^{24}". Express your answer to the correct number of sig figs.

Case Study

A student has 27.8 grams of helium and requires 6.89 moles for an experiment. Does the student have enough helium for the experiment?

To solve this problem, find the number of moles in 27.8 grams and compare it to the 6.89 moles available to determine if the student has sufficient helium for the reaction.

$$27.8 \text{ grams He} \times \frac{1 \text{ mole He}}{4.00 \text{ g He}} = 6.95 \text{ moles He}$$

Comparing the available 6.95 moles with the needed 6.89 moles, you can conclude that the student has enough He for the reaction, with 0.06 moles left over.

Note that the conversion factor between moles and grams can be expressed two ways:

$$\frac{1 \text{ mole}}{\text{atomic mass in grams}} \quad \text{or} \quad \frac{\text{atomic mass in grams}}{1 \text{ mole}}$$

How do you know which conversion factor to use? Consider which unit you need and how units cancel in a setup. If you need to know the number of moles, use the conversion factor with moles in the numerator (on top). If you need to know the number of grams, use the conversion factor with grams in the numerator.

Name________________

QUESTIONS

1. In your own words, what is a mole?

2. Explain how you can count by weighing.

3. One secret admirer leaves you 4.56 moles of silver while another secret admirer leaves you 3.69 moles of gold. Which one left you the larger gift in grams?

*Answer*__________

4. If 9.36 moles of sodium are required for a reaction, how many grams should be weighed out?

*Answer*__________

5. Small amounts of platinum are used in automotive spark plugs. What is the mass of 3.21×10^{24} atoms of platinum?

*Answer*__________

QUALITATIVE ANALYSIS OF HOUSEHOLD CHEMICALS

INTRODUCTION

What is in your kitchen cabinet? Do you have cornstarch (a thickener), glucose (a sweetener), sucrose (table sugar), potassium bitartrate (cream of tartar), sodium bicarbonate (baking soda), monosodium glutamate or MSG (flavor enhancer), acetic acid (vinegar) or sodium chloride (table salt)?

How about your medicine cabinet? Do you use iodine (tincture of iodine), isopropyl alcohol (rubbing alcohol) or magnesium sulfate (Epsom salt)?

If you looked in your garage would you find drain cleaner (sodium hydroxide) or Plaster of Paris (calcium sulfate) or root killer (copper(II) sulfate)? Do you have a pool or spa? Do you use muriatic acid (32% hydrochloric acid) to lower the pH? Do you use chalk (calcium carbonate) to mark off measurements?

Do you have washing soda (sodium carbonate) or laundry booster (sodium borate also called borax) in your laundry room?

Is there any red cabbage (anthocyanin) in your refrigerator?

All of these everyday household substances will be used in this experiment. You will identify the substances by their physical and chemical properties and use your observations and skills as a chemist to identify two unknowns.

QUALITATIVE ANALYSIS

There are two types of chemical analysis: quantitative and qualitative analysis. In **quantitative analysis** the chemist determines how much of a substance is present in the sample. In **qualitative analysis** the chemists isn't concerned with how much substance is present but rather what that substance is. Qualitative analysis is a process of testing an unknown sample and selectively eliminating possibilities until the substance is identified. A flow chart like the one on the following page is used to describe the steps in the qualitative analysis scheme.

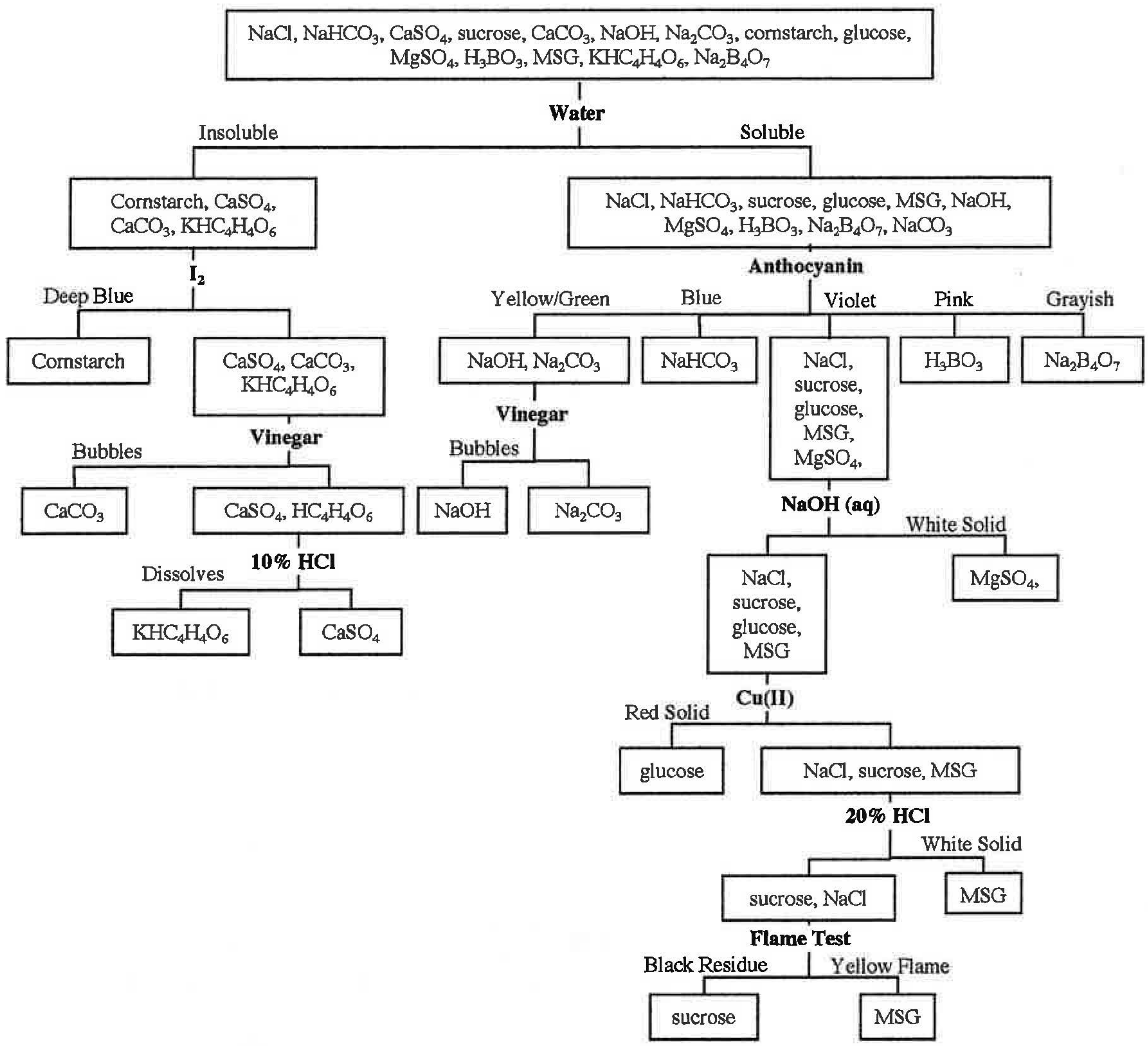


Table 1 on the following page lists the 14 white solid substances used in this scheme. All 14 substances are common household compounds and their uses and sources are also listed. Table 2 on the following page lists the reagents and their sources. Reagents are shown in bold print in the flow chart. Anthocyanin can be extracted from red cabbage using isopropyl alcohol. This will be prepared for you together with the hydrochloric acid solutions.

Unknown	Molecular Formula	Use /Source
Boric acid	H_3BO_3	Antifungal agent
Calcium carbonate	$CaCO_3$	Chalk or Tablets
Calcium sulfate	$CaSO_4$	Plaster of Paris
Cornstarch	Glucose polymer	Thickener
Fructose	$C_6H_{12}O_6$	Sweetener
Glucose (dextrose)	$C_6H_{12}O_6$	Sweetener
Magnesium sulfate	$MgS_4\_7H_2O$	Epson salt
Monosodium glutamate	$C_5H_8NNaO_4\_H_2O$	Flavor enhancer
Potassium bitartrate	$KHC_4H_4O_6$	Cream of tartar
Sodium bicarbonate	$NaHCO_3$	Baking Soda
Sodium borate (borax)	$Na_2B_4O_7\_10H_2O$	Laundry booster
Sodium carbonate	Na_2CO_3	Washing soda
Sodium chloride	NaCl	Table salt (noniodized)
Sodium hydroxide	NaOH	Lye / drain cleaner
Sucrose	$C_{12}H_{22}O_{11}$	Table sugar

Reagent	Source
Hydrochloride Acid	20% or 30% muriatic acid
Sodium hydroxide	Lye / drain cleaner
Iodine	Tincture of iodine
Anthocyanin	Red cabbage extract
5% acetic acid	Vinegar
Cu^{2+} ($CuSO_4$)	Root killer
Isopropyl alcohol	Rubbing alcohol

In this experiment you will identify the components of two unknowns. The first unknown will be one of 14 household compounds listed in Table 1 and the second will be a mixture of two of the 14 substances.

How do you identify an unknown compound?

Begin with the 14 known compounds and perform physical and chemical tests on each of them. Carefully record all observations in detail. When you have completed your tests on the known compounds, you will be given an unknown compound. Perform the same tests on the unknown and again carefully make and record all observations. Compare the physical and chemical properties of the unknown to the known substances to identify your first unknown.

Testing Known Substances

The qualitative analysis scheme (flow chart) shows the 14 known substances divided into two groups: soluble and insoluble in water. Begin by placing a small amount of the substance in a test tube and adding distilled water. If the substance is insoluble (does not dissolve in water), perform the iodine test. If the substance is soluble (does dissolve in water), perform the anthocyanin test. Continue testing according to the scheme.

Case Study – Potassium Bitartrate

Consider potassium bitartrate, or cream of tartar, $KHC_4H_4O_6$. According to the scheme, $KHC_4H_4O_6$ is insoluble in water. To check this, place a small amount (the size of a pea) of $KHC_4H_4O_6$ in a test tube. Add 5 mL of distilled water, stopper and shake. You should observe $KHC_4H_4O_6$ at the bottom of the test tube. It is insoluble. Using the same test tube (with the distilled water), add two drops of I_2 (tincture of iodine). $KHC_4H_4O_6$ does not react with I_2. Instead of a deep blue colored solution (indicating the presence of starch), the mixture will appear brown or discolored by the I_2. This is a negative result (expected for $KHC_4H_4O_6$). The next step in the scheme is the vinegar test for carbonates. Begin with a clean test tube and a fresh sample of $KHC_4H_4O_6$. The same test tube from the I_2 test cannot be used; the sample has been discolored and it will be difficult to make further observations. Use the same quantity of sample as in the solubility test, add 5 mL of distilled water, stopper and shake. Add 1 mL of vinegar. The $KHC_4H_4O_6$ will not produce bubbles, a negative result. For the final test, a fresh sample of $KHC_4H_4O_6$ is used. Add 5 mL of 10% HCl. $KHC_4H_4O_6$ does not dissolve in water but it will dissolve in 10% HCl. Some solid residue may remain but you can easily distinguish between the results of the water solubility test and the solubility in 10% HCl.

Do you need to try every test on each known?
No. You know from the schematic which known substances will react with which reagents. You must perform all of the tests that give a positive result for a particular known.

Do you need to start with a fresh sample for each test?
Yes. Many of the tests rely on color changes that will be masked if the same sample in the same test tube is used. The *exception* to this is the water solubility test. After you have tested a compound for water solubility, the sample may be used in (only) the next test.

Test Tube Calibration

Sample sizes and volumes of reagents do not have to be precise. You will calibrate your test tube so that you do not have to weigh out samples or measure the volume of reagents. This will save time on the analysis. To do this you will mark a control test tube at 1 mL, 5 mL and 10 mL volumes. Then you can measure out samples and reagents by comparing the level of sample or reagent in a working test tube with the marks on your calibrated test tube.

SAFETY

Take care when handling **all** reagents. HCl (muriatic acid) is a strong acid and NaOH (lye/drain cleaner) is a strong base; take care when handling these two reagents. The copper(II) reduction powder also contains NaOH. Wear safety goggles at all times. Wash hands with soap and water before leaving and especially before eating or drinking.

It may be tempting to taste the chemicals to attempt to identify your unknown. Do not taste any of the chemicals even if you believe they are nontoxic.

Dispose of chemicals according to instructions.

QUALITATIVE ANALYSIS

PROCEDURE
Test Tube Calibration

1. Measure 1.0 mL of distilled water in a small, 10 mL graduated cylinder. Pour the water into a test tube (one of your 12 in the rack). Use a wax pencil (or other marker that is not permanent) to mark the test tube at the 1.0 mL level. Pour out the water.

2. Measure 5.0 mL of distilled water in a small, 10 mL graduated cylinder. Pour the water into a test tube. Use a wax pencil to mark the test tube at the 5.0 mL level. Pour out the water.

3. Repeat the process for 10.0 mL.

Part I: Testing Knowns

Water Solubility

1. Place a small amount (the size of a pea) of the sample to be tested in a test tube.

2. Place the test tube with the sample in a test tube rack. Place the calibrated test tube next to it. Add 5 mL of distilled water to the solid sample: Add enough water to the test tube with the sample so that the water level is the same as the 5 mL mark on the calibrated test tube. You will have added about 5 mL of water to the sample. (Use this process later in the procedure to add 1 mL, 5 mL or 10 mL of reagent to the sample).

3. Stopper with a clean stopper, then shake the test tube. If a solution forms (the compound dissolves), the substance is soluble. If a mixture forms (the compound does not dissolve), the substance is insoluble. Record observations on data sheet.

4. Continue testing using the mixture or solution from the solubility test.

NOTE: Some residue may remain for some of the soluble compounds (e.g., borax). You should be able to distinguish which are soluble and which are not. If you use a very small sample of the substance, it may appear to dissolve when it is only slightly soluble (e.g., potassium bitartrate). You may want to repeat some of the solubility tests if there is any confusion.

Iodine Test
Iodine reacts with the amylose in starch to form a deep blue product. The deep blue color is a positive result and indicates the presence of starch. Iodine itself may appear as a range of colors in solution from yellow to purple. This is ***not*** a positive test. Test cornstarch for a known positive result and test one other insoluble compound for a known negative result.

$$\text{Starch} + I_2 \rightarrow \text{Starch-}I_2 \text{ (deep blue)}$$

1. Place a small amount (the size of a pea) of the sample to be tested in a test tube. Add 5 mL of distilled water, stopper and shake.

2. Add 2 drops of tincture of iodine. Use your wash bottle (filled with distilled water) to wash down the drops and shake. Record observations on the data sheet.

10% HCl Test
This test is only used to distinguish potassium bitartrate from calcium sulfate. Both are insoluble in water but potassium bitartrate dissolves in 10% HCl and calcium sulfate does not. Test both compounds to observe positive and negative test results.

1. Place a small amount (the size of a pea) of the sample to be tested in a test tube. Add 5 mL of distilled water, stopper and shake.

2. Add 1mL of 10% HCl. Take care to use the 10% HCl solution and ***not the 20% HCl** solution*. Record observations on the data sheet.

pH Test – Anthocyanin
Anthocyanin is a universal indicator derived from red cabbage. It changes color with changing pH:

pH Range	Color
1-3	Red
4	Pink
5-6	Violet-pink
7	Violet
9	Blue-green
10-12	Green-blue
>12	Yellow-green

Water soluble compounds are tested with anthocyanin and identified by their color. Sodium borate, which is basic, would be expected to turn yellow-green but it forms a borate ester with anthocyanin and turns grayish pink. The violet and pink colors can be distinguished from each other through careful observation. Anthocyanin itself is purple in color and will appear violet when diluted with water. It is helpful to prepare a blank sample with water and anthocyanin for comparison. A positive test will appear a different color from the blank.

1. Prepare a blank sample for comparison. Pour 5 mL of distilled water into a test tube. Add 3-5 drops of anthocyanin to the test tube. Mix by gently tapping the bottom side of the test tube.

2. Place a small amount (the size of a pea) of the sample to be tested in a test tube. Add 5 mL of distilled water, stopper and shake.

3. Add 3-5 drops of anthocyanin to the test tube. Gently mix. Compare the test tube with the sample to the blank test tube. (Place a sheet of white paper behind the test tubes to aid in distinguishing the colors). Record your observations on the data sheet.

Vinegar Test

Carbonates (CO_3^{-2}) and bicarbonates (HCO_3^{-2}) react with acids to produce carbon dioxide gas (CO_2). The presence of these ions can be detected by reacting them with vinegar, a weak acid. Bubbles of carbon dioxide gas indicate a positive test. $CaSO_4$, the main ingredient of Plaster of Paris, may give a positive test for carbonate due to a small amount of carbonate impurity. The bubbles produced from the impurity can be distinguished from the vigorous bubbling of a carbonate compound.

$$CO_3^{-2}{}_{(aq)} + 2H^+{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$$

Or

$$HCO_3^{-2}{}_{(aq)} + H^+{}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$$

1. Place a small amount (the size of a pea) of the sample to be tested in a test tube. Add 5 mL of distilled water, stopper and shake.

2. Add 1mL (about 20 drops) of vinegar. Record observations on the data sheet.

$NaOH_{(aq)}$ Test

Magnesium sulfate reacts with aqueous NaOH to produce magnesium hydroxide. Formation of a white precipitate confirms the presence of magnesium hydroxide and is a positive test for magnesium sulfate.

$$MgSO_{4(aq)} + 2NaOH_{(aq)} \rightarrow Mg(OH)_{2(s)} + Na_2SO_{4(aq)}$$

1. Place a small amount (the size of a pea) of the sample to be tested in a test tube. Add 5 mL of distilled water, stopper and shake.

2. Add 1 mL (about 20 drops) of aqueous NaOH. Record observations on the data sheet.

Cu(II) Reduction Test

Copper(II) ion is an oxidizing agent. Glucose is a reducing agent. When the two react, Cu^{+2} is reduced to Cu^{+1} and changes color from blue to brick red. This color change is exploited in chemical tests to detect the presence of glucose in the urine. Home tests such as chemical sticks and tablets use Cu^{+2} and the color change from blue to red is calibrated to indicate the quantity of glucose present. Fructose is also a reducing sugar but sucrose (table sugar) is not and will not react with Cu^{+2}. (The equation below is not balanced).

$$\text{Glucose(aq)} + Cu^{+2}\ \text{(blue)} + NaOH \rightarrow \text{Sodium Gluconate} + Cu_2O\ \text{(red)}$$

Commercially prepared copper reduction tablets contain copper(II) sulfate, sodium hydroxide, sodium carbonate and citric acid for pH control. The powder you will use in this experiment is similar to the commercially prepared tablets. The citric acid has been replaced with sodium chloride.

Test glucose to observe a positive result (color change from blue to red) and test sucrose to observe a negative result (color will remain blue).

1. Place a small amount (the size of a pea) of the sample to be tested in a test tube. Add 5 mL of distilled water, stopper and shake.

2. Add a spatula tip of copper reduction powder. Stopper and shake test tube. Wait 1-2 minutes for color to develop. If you do not observe a color change with glucose, add a little more copper reduction powder. Record observations on the data sheet.

20% HCl Test
This test is used to distinguish monosodium glutamate (MSG), a water soluble compound, from two other water soluble compounds, NaCl and sucrose. MSG is soluble in water but when treated with HCl is converted to glutamic acid which is only slightly soluble in water. The glutamic acid should precipitate out of the solution, forming a white solid and indicating a positive test for MSG. The solubility of NaCl and sucrose are unaffected.

1. Place enough MSG to cover the bottom of the test tube. Add 5 mL of distilled water, stopper and shake.

2. Add 1 mL of 20% HCl. Record observations on the data sheet. If a precipitate fails to form, repeat with 32% HCl (muriatic acid).

Evaporation Test
This last test is used to distinguish sucrose and sodium chloride, both of which are water soluble. If a small amount of solution is placed in a spoon spatula and heated until most of the water boils away, the two compounds can be distinguished. Sucrose can be identified by the odor of burnt sugar and salt gives a yellow sodium flame. Test both sucrose and sodium chloride.

1. Place a small amount (the size of a pea) of the sample to be tested in a test tube. Add 5 mL of water and shake.

2. Place a small amount of the solution in a spoon spatula (obtained from the hood). Heat gently until most of the water boils away. Record observations on the data sheet.

3. Once you have completed your data sheet, obtain your instructor's signature and request an unknown.

Part II: Identifying Single Unknowns and Mixtures
Single Unknowns
The first unknown will be a single compound. Begin with the water solubility test and follow the flow chart. If the unknown is insoluble, continue with the iodine test. If the compound is soluble, perform the pH (anthocyanin) test. Once you have tried either the iodine test or the pH test, you will need a fresh sample. Most of the tests rely on a color change and the results of the previous test will mask the results of the next test. Continue to perform as many tests as needed until you are able to identify your unknown. Record all observations on data sheet. When you have identified your unknown, obtain your instructor's signature and request a second unknown.

Mixtures

The second unknown will be a mixture of two compounds. Place 2-3 mL (think creatively, How can you use the calibration test tube to measure this quantity?) of the solid mixture in a test tube. Add 10 mL of distilled water, stopper and shake. Filter to separate the insoluble component. Divide the insoluble component into portions for analysis. You will need to perform 1-3 tests to identify the unknown component. Divide the remaining solution into several test tubes to be used to analyze the soluble component. You should plan on multiple tests to identify the soluble component. Record all observations on data sheet.

Name________________

QUALITATIVE ANALYSIS OF HOUSEHOLD CHEMICALS
DATA SHEET FOR KNOWNS

You do not need to perform every test on every known. All of the boxes do not need to be filled in. You may use any symbol to indicate a positive or negative test or simply use (+) for positive and (-) for negative. It may be helpful to write the solution color for the pH test.

KNOWNS	Formula	Sol. in H_2O	Iodine Test	pH Test	Vinegar Test	NaOH Test	Cu(II) Test	10% HCl Test	20% HCl Test	Evap. Test
Boric acid	H_3BO_3									
Calcium carbonate	$CaCO_3$									
Calcium sulfate	$CaSO_4$									
Cornstarch	Polymer of glucose									
Glucose	$C_6H_{12}O_6$									
Magnesium sulfate	$MgSO_4 \cdot 7H_2O$									
Monosodium glutamate (MSG)	$C_5H_8NNaO_4 \cdot H_2O$									
Potassium bitartrate	$KHC_4H_4O_6$									
Sodium bicarbonate	$NaHCO_3$									
Sodium borate	$Na_2B_4O_7$									
Sodium carbonate	Na_2CO_3									
Sodium chloride	NaCl									
Sodium hydroxide	NaOH									
Sucrose	$C_{12}H_{22}O_{11}$									

Additional Observations:

Instructor's Signature____

Name________________

QUALITATIVE ANALYSIS OF HOUSEHOLD CHEMICALS
DATA SHEET FOR UNKNOWNS

Perform only those tests that you determine are necessary for identifying your unknowns.

UNKNOWN NUMBER	Sol. in H_2O	Iodine Test	pH Test	Vinegar Test	NaOH Test	Cu(II) Test	10% HCl Test	20% HCl Test	Evap. Test
Single									
Mixture									

Additional Observations:

Conclusion: Identify your single unknown compound and unknown mixture. Explain your reasoning.

Instructor's Signature____

Name________________

QUALITATIVE ANALYSIS OF HOUSEHOLD CHEMICALS PRELABORATORY QUESTIONS

1. In this experiment, you will test 14 known compounds before beginning the analysis of an unknown compound. Do you need to start with a fresh sample for each test? Explain.

2. A student, Qual T. Analysis, performing this laboratory experiment added distilled water to his unknown sample. The sample dissolved in distilled water. Next, Qual added 2 drops of tincture of iodine and reported a positive result for starch. To the same test tube, Qual added anthocyanin. Qual reported no change. Qual concluded the unknown contained only one component, cornstarch. Identify at least two errors that Qual made.

3. A student performing this laboratory experiment, Joe Chemistry, obtained an unknown sample. Joe added 5 mL of distilled water to his sample. The sample dissolved. Next, Joe added 5 drops of anthocyanin. The solution sample turned violet. Joe began with a fresh sample of his unknown, added water and then added NaOH. There was no change. Using a fresh sample of his unknown again, Joe added water. To this solution Joe added a spatula tip of the copper(II) reduction powder. A red solid ***did not*** form. With a fresh sample, Joe added water and 20% HCl. There was no change. Finally, Joe dissolved some of his sample in 5 mL of water and placed some of the solution in a spoon spatula and gently heated the solution over a Bunsen burner until it boiled. Joe noticed a distinct odor when he evaporated his solution. What was Joe's unknown? Explain your answer.

4. State all safety concerns specific to this particular laboratory experiment.

THE COLOR OF THE CANDY COATING

INTRODUCTION

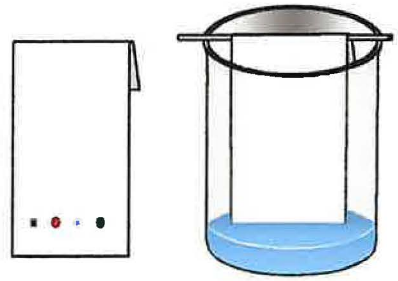
Why are apples red? What would happen to the color of the red apple if you shined a blue light on it instead of a white light? Would you be surprised to find that the red apple will appear black under the blue light? The apple appears red under the white light because it **reflects** the red. The other colors are **absorbed**. A blue light does not contain the color red and so there is no red to reflect. Under a blue light the apple will appear black.

The color of a transparent object depends on the color of the light it **transmits** rather than the color reflected. A piece of red glass appears red because it absorbs the other colors of white light and transmits the red color.

In this experiment you are going to examine the candy coating of M&M's®. Like apples, the candy coating is **opaque** (you can't see through it) and the color depends on which colors of white light are absorbed and which are reflected. You will determine if the color is pure or a mixture by using the experimental technique of **paper chromatography**. If the candy coating is extracted by water, a transparent solution will result. The color of the transparent solution will depend on the color of light it transmits. You will use **spectrophotometry** to determine the nature of the light transmitted from a solution of the candy coating of M&M's®.

PAPER CHROMATOGRAPHY

Chromatography is a separation technique used to separate two or more constituents (parts). In paper chromatography, the substance to be separated (candy coating) is dissolved in a liquid solvent (water). A small drop of the solution is applied to a strip of chromatography paper (similar to filter paper). The paper is then suspended in a beaker (or other container) so that the bottom edge of the paper is in contact with liquid in the bottom of the beaker. See figure on the following page. This liquid is called the **developing solvent**. The developing solvent will move up the paper by capillary action (the same way water moves from the roots of a tree to the top branches) and this will separate the constituents. The process of the solvent moving through the paper is called **elution**.



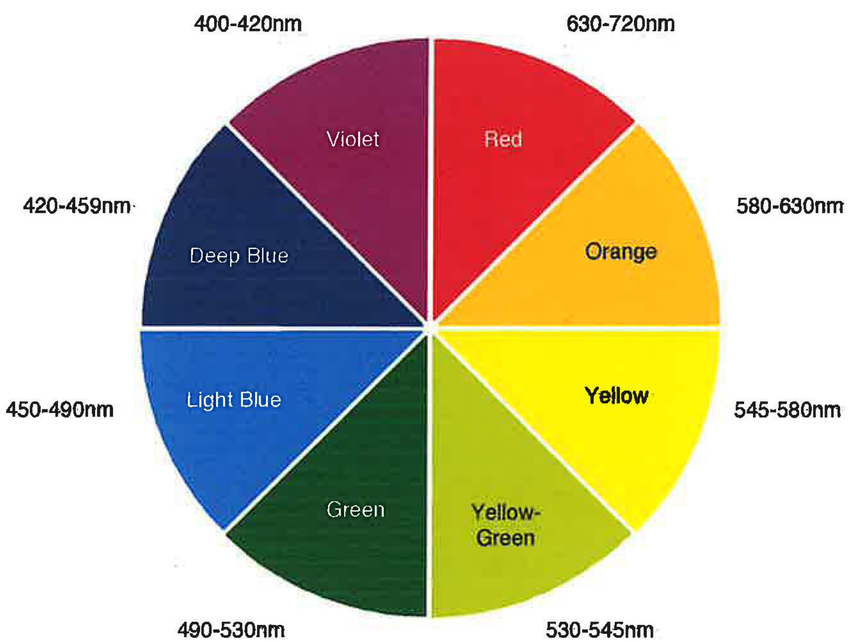
How does this process work? It all depends on the interactions of the substance with the liquid solvent and with the paper. If a substance interacts strongly with the paper, it will tend to stick to the paper and not move along with the moving solvent. Substances that do not interact as strongly with the paper will move with the solvent and the different substances can be separated.

SPECTROPHOTOMETRY

A **spectrophotometer** is an instrument that measures the amount of **ultraviolet** (UV), visible or **infrared** (IR) **radiation** absorbed by solutions of chemical substances. **Radiation** is a process of energy transmission (including heat energy). Energy that is transmitted by radiation is in the form of **electromagnetic waves**. Violet light has the shortest visible wavelength. Ultraviolet (beyond the violet) radiation has shorter wavelengths. You may know about UV radiation from discussions about the harmful affects of the sun. This kind of radiation is not visible – you cannot see it. Red light has the longest visible wavelength. Infrared (below the red) radiation has longer wavelengths. You are also familiar with IR radiation although you may not have known it by its name. If you have felt the warmth from a flame, you have experienced IR radiation. Spectrophotometers measure the amount of radiation absorbed in the UV, visible and IR range. A **colorimeter** is similar to a spectrophotometer but it measures radiation in the visible region only.

In a spectrophotometer, a source generates UV, visible or IR radiation. Consider the visible spectrum. The source (a light bulb) generates white light that is separated into specific wavelengths by a diffraction grating (similar to a prism). The spectrophotometer has a wavelength selector dial that can be used to control or select specific wavelengths. Unabsorbed light passes through the solution and reaches a detector. The amount of light absorbed is measured. (The amount of light transmitted can also be measured and converted to light absorbed by a mathematical relationship).

Consider a solution that absorbs light at 670 nm. This is the red region of the visible spectrum and the solution will appear green. This is the **complementary color** of red. The figure below shows the wheel of complementary colors. The red light has been absorbed, no red light reaches your eyes and the solution appears green. If you want to measure the intensity of light absorbed for this solution, use the wavelength dial of the spectrophotometer to select 670 nm. This is the wavelength that will allow for maximum absorption.



One more thing about spectrophotometry: you will need to use a **blank sample**. The need for a blank arises from the reflection and refraction (bending) of light as it passes from one medium into another. If you have ever hung your legs into a swimming pool, you may have noticed a difference between how your legs look in and out of the pool. This is refraction. Light is bent differently in water than it is in air which is why your leg looks different as the light passes from air to water. Likewise, when light in a spectrometer passes from air to glass to solution and out again, light can be reflected or refracted. This light will not reach the detector and will be measured as absorbed light. To avoid this you can make a blank sample that will contain the reagents used in the sample to be measured **except** the one reagent that is absorbing the light. You must use the same kind of glass container for the blank sample that you use for the sample to be studied. This is called a **cuvette**. Cuvettes look like test tubes and are used to contain the solution to be tested. These must be checked out from your instructor. You will use this blank to "zero" your instrument between readings.

Your lab will have two types of spectrophotometers: analog and digital. You may have an analog watch – a watch with a dial or meter that you interpret or read. You may have a digital watch where the reading is displayed in numbers. Likewise, analog instruments have meters that need to be read or interpreted and digital instruments display the readings in numbers. Digital instruments are easier to read but they are not more accurate.

In this experiment, you will extract the color from the candy coating from an M&M®. Using a spectrophotometer, you will vary the wavelength from 340 to 720 nm and measure the absorbance for the candy coating solution. A plot of absorbance vs. wavelength will reveal the maximum absorbance for the solution.

SAFETY AND WASTE DISPOSAL

The reagents used in this experiment are not toxic. Do not eat any of the reagents unless directed by your instructor. This experiment is conducted in a chemical laboratory. Wash hands with soap and water before leaving and especially before eating or drinking.

There is no special disposal required in this experiment.

THE COLOR OF THE CANDY COATING

PROCEDURE

PAPER CHROMATOGRAPHY

1. Place a brown M&M's® candy (take care to avoid broken candy) in a 50 mL beaker. Add 10 mL of distilled water. Swish the contents of the beaker for about 20 seconds. Pour off the solution (liquid only) into a test tube.

2. Concentrate the solution by evaporation. Pour the solution into an evaporating dish. Place the evaporating dish on a hot plate and gently evaporate until the volume of solution is about half.

3. Obtain a piece of chromatography paper tapered. Draw a pencil (do not use ink or other markers) line across the bottom of the chromatography paper. Use a capillary tube to draw up some of the solution. Apply a small amount of the solution from the capillary tube to the pencil line. Allow the spot of solution to dry and then reapply. Repeat this until the spot is dark enough to see well. By allowing the spot to dry between applications, you will keep the spot small. Large spots do not separate well.

4. Use a 400 mL beaker as the developing chamber. Add enough developing solvent (water) to the beaker so that the height of the liquid in the beaker is about 1 cm.

5. Suspend the chromatography paper with the spot in the beaker so that the tapered edge of the paper is in contact with liquid in the bottom of the beaker. The spot must be above the level of the liquid. Do not submerse the spot in the liquid.

6. Allow the water to move up the paper. When the water is about 1 cm from the top of the chromatography paper, remove it from the beaker and trace the boundary of the water with a pencil. (This is called the solvent boundary).

7. Allow the paper to air dry. Record your observations on the data sheet.

8. Attach the chromatagram (the developed chromatography paper) to your data sheet.

9. Repeat steps 1-7 with two different colors of M&M's® candies.

THE COLOR OF THE CANDY COATING

PROCEDURE

SPECTROPHOTOMETRY

1. Obtain a set of cuvettes from your instructor.

2. Place a brown M&M® candy (take care to avoid broken candy) in a 50 mL beaker. Add 10 mL of distilled water. Swish the contents of the beaker for about 20 seconds. Pour off the solution (liquid only) into a cuvette. Fill the cuvette about three-quarters full.

3. Prepare a **blank sample** by filling a cuvette three-quarters full of distilled water (the solvent).

4. Set the wavelength to 340 nm for the first reading.

5. Use the on/off dial (front left) to zero the spectrophotometer so that it reads 0% light transmission. There should be nothing in the sample compartment and the cover should be closed.

6. Place the blank sample in the sample compartment and close the cover. Before placing any sample in the compartment, wipe the cuvette off with a Kim-wipe and hold it so that you do not leave fingerprints on the cuvette (these will interfere with the light).

7. With the blank in the sample compartment with the cover closed, set the spectrometer to blank (100% transmittance) using the light adjustment dial (front right).

8. Remove the blank sample (do not make any adjustments).

9. Place the sample to be measured in the sample compartment and close the cover. Remember to wipe the cuvette with a Kim-wipe and avoid fingerprints. Do not use a paper towel or any other material on the cuvette.

10. Read the absorbance of the sample on the meter or digital display. Record absorbance reading.

11. Remove the sample from the compartment.

12. Repeat steps 4-11 changing the wavelength in 20 nm increments until you have measured the absorbance of the solution from 340 to 720 nm. **NOTE:** under the on/off dial, there is a lever for the filter. The lever must be to the left for measurements of 340-599 nm and to the right for measurements 600-950 nm.

13. Construct a plot of absorbance (y-axis) vs. wavelength (x-axis).

14. Clean up. Wash cuvettes with soap and water. Rinse with distilled water. Return to instructor. All solutions may go down the sink. Solid waste (M&M's®) may go in the trash. Wash your hands with soap and water.

Name________________

THE COLOR OF THE CANDY COATING
DATA SHEET

PAPER CHROMATOGRAPHY

1. Attach the chromatograms here.

2. Which, if any, of the candy coatings consists of more than one colored component? In each case, identify the color of the coating and the color of the components.

3. Are these results you would have predicted based on your experience? (i.e., would you predict that those colors would have more than one component)

Name_______________

THE COLOR OF THE CANDY COATING
REPORT SHEET

SPECTROPHOTOMETRY

WAVELENGTH	ABSORBANCE
340	
360	
380	
400	
420	
440	
460	
480	
500	
520	
540	
560	
580	
600 **CHANGE FILTER**	
620	
640	
660	
680	
700	
720	

1. Plot absorbance vs. wavelength. From your graph, what is the maximum absorbance of the M&M® candy solution?

2. What color is being absorbed? What is the complementary color that reaches your eyes?

3. There is a second, very small peak following the maximum absorption peak. Locate this peak on your graph. What is the wavelength of this peak? What color is being absorbed? What is the complementary color that reaches your eyes?

4. Is there any correlation between the complementary colors that reach your eyes (your answers to questions 2 and 3) and the colors you observed in the paper chromatography of the brown M&M® candy? (i.e., do any of the colors not absorbed in the spectrophotometry also appear in the paper chromatography?)

Name________________

THE COLOR OF THE CANDY COATING

POSTLABORATORY QUESTIONS

1. Why is pencil used to mark the paper chromatography paper? Why is a pen or other ink marker not used?

2. Why is it important to keep the spot above the water line?

3. What is the color of a solution that absorbs light in the 610 nm region?

4. If a solution absorbs light at 550 nm, the yellow range of the visible spectrum, it will appear deep blue. Explain.

Name________________

THE COLOR OF THE CANDY COATING

PRELABORATORY QUESTIONS

1. A blue object will appear _____ under a red light.
   (a) blue
   (b) red
   (c) black
   (d) white

2. Paper chromatography is a technique used to
   (a) measure the amount of light transmitted through a solution
   (b) separate a substance into its constituents
   (c) filter solutions
   (d) measure the amount of light absorbed by a solution

3. How does paper chromatography work?

4. Spectrophotometry measures
   (a) the amount of light transmitted through a solution
   (b) ultraviolet radiation
   (c) Infrared radiation
   (d) the amount of light absorbed by a solution

5. If a solution absorbs light at 540 nm, what color will it appear?

6. State all safety concerns in this laboratory.

CHEMICAL REACTIONS

INTRODUCTION

How are kidney stones formed? Are there any dietary restrictions for patients with kidney stones? Why is fluoride (notice it's an ion, not an element) used to strengthen teeth? How is glycogen formed from glucose and stored in the body? How is it released? Glycogen in the muscle cannot be used to raise blood sugar (glucose) levels. Why not? Why do our muscles burn during vigorous exercise? Why do we continue to breathe hard even after we have stopped exercising? A cure for hyperventilation is to have the patient breathe into a paper bag. This actually works, but how? If you know something about chemical reactions, you could answer these questions.

CHEMICAL REACTIONS

Sodium chloride, NaCl, is commonly known as table salt. If you put salt in your mouth what happens? The salt dissolves and separates into ions – this ionic break up is called **dissociation** in chemistry. The salt dissociates into the ions that make it up: chloride ions, Cl^-, and sodium ions, Na^+. Sodium chloride dissolved in water is an **aqueous solution** of sodium chloride. There isn't any of the sodium chloride compound in the solution – the compound has dissociated into sodium and chloride ions. Sodium chloride is an ionic compound and many ionic compounds dissociate in water. Those compounds that dissolve in water are called **soluble**. Many other ionic compounds will not dissolve in water and these are called **insoluble**.

What happens when a solution of one ionic compound is mixed with a solution of a second ionic compound? A chemical reaction may or may not occur. Consider what happens when a solution of sodium carbonate (Na_2CO_3) is mixed with a solution of calcium chloride ($CaCl_2$). Both solutions are clear and colorless but when mixed the resulting solution becomes cloudy and a white solid settles at the bottom. The insoluble white solid is called a **precipitate**. A chemical reaction has occurred. What is the precipitate in this reaction?

To answer the question, you must first determine which ions are present in the solutions. Sodium carbonate will dissociate into sodium ions, Na^+, and carbonate ions, CO_3^{-2}. Likewise, calcium chloride will dissociate into calcium ions, Ca^{+2}, and chloride ions, Cl^-. All of the ions are together in solution and can interact with each other. Remember that opposite charges attract and like charges repel. This means Na^+ and Ca^{+2} ions are repelled by each other; likewise, CO_3^{-2} and Cl^- ions are repelled by each other. Which ions can come together and form a precipitate?

Consider the possibilities:

- Na^+ and CO_3^{-2} can combine to form Na_2CO_3
- Na^+ and Cl^- can combine to form NaCl
- Ca^{+2} and Cl^- can combine to form $CaCl_2$
- Ca^{+2} and CO_3^{-2} can combine to form $CaCO_3$

Which of these is most likely the precipitate? You already know that Na_2CO_3 and $CaCl_2$ will dissociate in water. Since both compounds started out as aqueous solutions, they will not form precipitates. One of the other possibilities is NaCl, table salt, which you know will dissociate in water (think about the ocean). The precipitate is most likely $CaCO_3$. This is the main ingredient in seashells. Do you think that sea shells are water-soluble or water insoluble? Does it seem likely that $CaCO_3$ is the precipitate?

HOW DO YOU KNOW A CHEMICAL REACTION HAS OCCURRED?

A chemical reaction has occurred when a chemical change has taken place; in other words, a new substance (this is the **product** of the reaction) has been formed. In this experiment you will mix two chemicals (called **reactants** or **reagents**) together and make observations. What will you look for? From the example above, you know the formation of a precipitate is evidence that a chemical reaction has occurred. Other observable signs include production of bubbles (evolution of a gas), change in solution color (may or may not include the formation of a precipitate) and a change in temperature (the mixture gets hot or cold).

In the first part of this experiment, you will mix two reagents together and make observations. You will interpret your observations and decide if a chemical reaction has occurred. In the second part, you will predict the products, if any, using the logic modeled above and confirm your predictions with appropriate procedures and observations

OBSERVATIONS VS. CONCLUSIONS

What is the difference between an observation and a conclusion?

Observations are things we can detect with our senses – things we can see, hear, smell, taste and touch. **Never** touch or taste any chemicals in the laboratory. Even smelling chemicals can be dangerous. Your laboratory instructor will demonstrate how to safely smell any chemicals, if needed. How can you "hear" a reaction? Some reactions do produce sound: think explosion. Chemists most often rely on their sight to make observations.

Conclusions are judgments or interpretations or explanations based on observations. Consider the following: Your lab partner wears the same black shirt every lab day. This is because black is his favorite color. Which of these statements is an observation and which is an interpretation (conclusion)? You see your partner wear the same black shirt – this is an observation. Your interpretation is that black is his favorite color. This is a judgment based on your observation. It might be that the black shirt is his least favorite one and he wears it to lab because he is doesn't care if chemicals are spilled on it. It might possibly be his only shirt. Or it could simply be coincidence. Whatever the reason, you are making a judgment based on an observation and this is a conclusion.

Case Study
You watch a child throwing rocks into a lake. Which of the following statements are observations and which are conclusions?

1. The child is throwing rocks into a lake.
2. The rocks sink to the bottom.
3. Throwing rocks in the lake makes the child happy.
4. Some water from the lake evaporates in the heat.

Statement 1 is an observation. Statement 2 is a conclusion unless you could see the bottom of the lake and see the rocks sink to the bottom. It is a reasonable conclusion, of course, but it is still an interpretation of an observation. Statement 3 is a conclusion. Statement 4 is similar to Statement 2. It is a reasonable conclusion but it is not something you observed. What if you observed a small puddle and not a lake? You might observe the puddle getting smaller in size during the day. Stating that the water evaporated is interpreting your observation. It is a reasonable conclusion but it is not an observation.

What does this have to do with the experiment? You will make observations based on what you see. You will be looking for the formation of a precipitate, change in color, or bubbles. You will interpret your observations and conclude if a reaction has occurred. The fourth sign that a reaction has taken place is a change in temperature. A change in temperature may be detected by touching the reaction container (test tube or beaker) or by using a thermometer. In this experiment, it is not practical to touch the container or use a thermometer so you will be limited to observations of the other three.

CHEMICAL REACTIONS CONTINUED

The examples of chemical reactions discussed above depend on the mixing of two ionic solutions together. Do reactions occur only when two ionic solutions are mixed? Think about some of the chemical reactions with which you are familiar. Rust occurs when oxygen in the air reacts with iron metal (formation of a solid precipitate). A spark will ignite the gas on your stove (change in temperature). Hydrogen peroxide reacts with the cut on your finger (bubbles). Red meat turns brown when exposed to air (change in color). You can conclude that two ionic solutions are not necessary for a chemical reaction to occur.

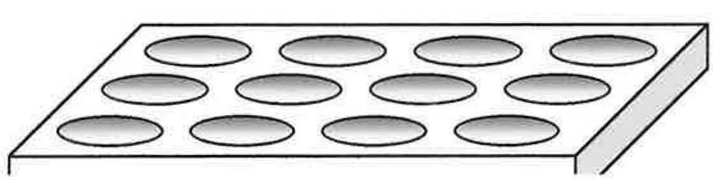
In the first part of this experiment, you will use ionic solutions and two solid metals to test for evidence of a chemical reaction. The ionic solutions are sodium carbonate (Na_2CO_3), sodium hydroxide (NaOH), silver nitrate ($AgNO_3$), hydrochloric acid (HCl) and phenolphthalein, an indicator or dye. (Acid-base ndicators are actually organic compounds that change color when they react with acids and bases. The color change "indicates" the presence of an acid or base depending on the change. Hence, the name "indicator".)

To observe the interaction between two ionic solutions (or indicator) you will first observe the original solutions (note the color) separately and then mix 5 drops of each solution in a spot plate. To observe the interaction between two metals, you will record the appearance of each metal before and after "mixing". When mixing a metal with an ionic solution (or indicator), record the appearance of the metal prior to mixing and observe changes after adding 5 drops of the solution.

In the second part of this experiment, you will use only combinations of ionic solutions to predict the products, if any, of a chemical reaction. The reagents you will use are calcium nitrate, $Ca(NO_3)_2$, lead(II) nitrate, $Pb(NO_3)_2$, mercury(II) nitrate, $Hg(NO_3)_2$, potassium bromide, KBr, potassium iodide, KI, silver nitrate, $AgNO_3$, and sodium chloride, NaCl. Again, you will use the spot plate and mix 5 drops of each solution of the pair of reagents in the spot plate. You will record your observations and determine if a reaction occurred. If there is a chemical reaction, you will predict the products of the reaction using the logic modeled above.

USE OF SPOT PLATES

The type of spot plate used in this experiment has 12 depressions in it. Each depression is about the size of a thumb print (see figure below). Reagents are mixed in this depression and observed. This allows a chemist to perform many experiments on a small scale – much smaller than using test tubes. Spot plates are an efficient way to carry out numerous reactions. You may have seen spot



plates, or something similar, used in commercial laboratories or clinical labs where the same reaction is carried out many times. This is also cost effective because smaller quantities of reagents are used and less waste is produced. Previously in the discussion it was mentioned that temperature change would not be measured because spot plates would be in use. It is not practical to touch the spot plate as you would a test tube to determine if there has been a temperature change. A temperature probe can be used to detect temperature changes using a spot plate but you will not use that in this experiment.

SAFETY AND WASTE DISPOSAL

Take care when handling all reagents. HCl is a strong acid. NaOH is a strong base. Both reagents are highly reactive. Avoid contact with skin and eyes. Do not inhale fumes. $AgNO_3$ will react with the proteins in your skin and discolor them. $Pb(NO_3)_2$ and $Hg(NO_3)_2$ contain heavy metals and are toxic. Take care when handling these reagents. Wear safety goggles at all times. Wash hands with soap and water before leaving and especially before eating or drinking.

Clean up the spot plate by emptying the contents onto a paper towel. Dispose of the paper towel in the waste container. Clean the spot plate with soap and water. Make the final rinse with distilled water. Take care to clean the spot plate thoroughly so as not to contaminate the next experiment.

PROCEDURE

CHEMICAL REACTIONS, PART I

1. Obtain 3 spot plates (each spot plate will have 12 depressions) from the hood.

2. Obtain a set of reagents from the hood. Each reagent in the set will be labeled.

3. Look carefully at the matrix on the data sheet. The reagents will be mixed in the order prescribed in the matrix. For example, Na_2CO_3 is in the first row. The first depression will contain a mix of Na_2CO_3 and NaOH. The second depression will contain Na_2CO_3 and $AgNO_3$. The third will contain Na_2CO_3 and phenolphthalein and so on. There are 31 combinations to be observed. Consider the most efficient way of organizing the 3 spot plates and the reagents before going further.

4. To observe the interaction between a metal and a solution, place one metal strip in the spot plate and record the appearance of the metal. Record observations in the box at the intersection of the two reagents. Before adding the solution, observe and record the appearance of the solution. The most significant observation regarding the solution is its color. Add 5 drops of the solution to the spot plate with the metal strip and record observations. If there is no change after 20 seconds, wait 1-2 minutes and observe. Record all observations. Be brief when recording observations; complete sentences are not required. If there is no observable change, record "no change".

5. To observe the interaction between two solutions, place 5 drops of the first solution in the spot plate and record the appearance of the solution. Use the box at the intersection of the two reagents to record observations. Before adding the second solution, observe and record its appearance. ***The observations of each interaction need to be recorded even if there is no change.*** Add 5 drops of the second solution to the spot plate and record observations.

6. To observe the interaction between a metal and a metal, place one metal strip in the spot plate and record the appearance of the metal in the box at the intersection of the two reagents. Before adding the second metal strip, observe and record its appearance. Add the metal strip and observe the interaction. If there is no change after 20 seconds, wait 1-2 minutes and observe. Record all observations.

7. Show your matrix with observations to your instructor and obtain your instructor's signature.

8. Clean up the spot plate by emptying the contents onto a paper towel. Dispose of the paper towel in the solid waste container. Clean the spot plate with soap and water. Make the final rinse with distilled water. Take care to clean the spot plate thoroughly so as not to contaminate or ruin the results of the next section of the experiment.

9. Return all reagents to the hood.

CHEMICAL REACTIONS, PART II

1. Obtain 4 spot plates (each spot plate will have 12 depressions).

2. Obtain a set of reagents. Each reagent in the set will be labeled.

3. Look carefully at the matrix on the data sheet. The reagents will be mixed in the order prescribed by the matrix. For example, $Pb(NO_3)_2$ is in the first row. The first depression will contain a mix of $Pb(NO_3)_2$ and $Hg(NO_3)_2$. The second depression will contain $Pb(NO_3)_2$ and KBr. The third will contain $Pb(NO_3)_2$ and KI and so on. There are 37 combinations to be observed. Consider the most efficient way of organizing the 4 spot plates and the reagents before going further.

4. To observe the interaction between two ionic solutions, place 5 drops of the first solution in the spot plate and record the appearance of the solution. Use the box at the intersection of the two reagents to record observations. Before adding the second solution, observe and record its appearance. ***The observations of each interaction need to be recorded even if there is no change.*** Add 5 drops of the second solution to the spot plate and record observations.

5. Show your matrix with observations to your instructor and obtain your instructor's signature.

6. Clean up the spot plate by emptying the contents onto a paper towel. Dispose of the paper towel in the solid waste container. Clean the spot plate with soap and water. Make the final rinse with distilled water. Take care to clean the spot plate thoroughly so as not to contaminate or ruin the results of the next section of the experiment.

7. Return all reagents to the hood.

8. Wash your hands with soap and water before leaving the lab and especially before eating or drinking.

Name________________

DATA SHEET
CHEMICAL REACTIONS, PART I

Record observations in pencil in the boxes where the reagents intersect. Observations are not required for boxes marked with X.

	Na_2CO_3	NaOH	$AgNO_3$	Phenolphth.	Zn	HCl
Na_2CO_3	X					
NaOH		X				
$AgNO_3$			X			
Phenol-phthalein				X		
Zn					X	
Mg						

INSTRUCTOR SIGNATURE______________________________

Name_______________

DATA SHEET
CHEMICAL REACTIONS, PART II

Record observations in pencil in the boxes where the reagents intersect.
Observations are not required for boxes marked with X.

	$Pb(NO_3)_2$	$Hg(NO_3)_2$	KBr	KI	$AgNO_3$	NaCl
$Pb(NO_3)_2$	X					
$Hg(NO_3)_2$		X				
KBr			X			
KI				X		
$AgNO_3$					X	
NaCl						X
$Ca(NO_3)_2$						

INSTRUCTOR SIGNATURE______________________________

Name_______________

POSTLABORATORY QUESTIONS

CHEMICAL REACTIONS, PART I

1. Which combinations of reagents produced reactions? Complete the table below by identifying the two reagents and the observations that led you to conclude a reaction had occurred.
   Example: **Na_2CO_3 and $CaCl_2$, white precipitate formed**

REAGENT 1	REAGENT 2	OBSERVATION

2. Which combinations of reagents produced precipitates? Complete the table below by identifying the two reagents and the color of the precipitate.
Example: **Na_2CO_3 and $CaCl_2$, white precipitate**

REAGENT 1	REAGENT 2	COLOR

3. Using the data from the table above, identify any reagents that are associated with particular precipitates. For example, if a white precipitate is formed more than once, are any reagents in common? Develop your own table identifying the relationships, if any, between reagents and the precipitates they produce.

4. Which combinations of reagents produced bubbles? Complete the table below by identifying the two reagents that produced bubbles.

REAGENT 1	REAGENT 2

5. Using the data from the table above, identify any reagents that are associated with the evolution of a gas (bubbles). For example, if bubbles are produced more than once, are any reagents in common? Develop your own table identifying the relationships, if any, between reagents and evolution of a gas.

6. Which combinations of reagents produced a color change (not including formation of a precipitate)? Complete the table below by identifying the two reagents that produced a color change.

REAGENT 1	REAGENT 2	OBSERVATION

7. Using the data from the table above, identify any reagents that are associated with a color change. Develop your own table identifying the relationships, if any, between reagents and a color change.

8. Summarize any trends that you have observed regarding chemical reactions.

Name________________

CHEMICAL REACTIONS, PART II: PREDICTING PRODUCTS

1. Which combinations of reagents produced precipitates? Complete the table below by identifying the two reagents and the color of the precipitate.

REAGENT 1	REAGENT 2	COLOR

2. For each reaction that produced a precipitate, identify the solid using the logic in the discussion. Write out a chemical equation (you do not need to balance the equation but all chemical formulas must be correct). Underline the compound you have concluded is the solid product.

   Example: $Na_2CO_3 + CaCl_2 \rightarrow NaCl + \underline{CaCO_3}$

3. Your observations showed that the interaction of lead(II) nitrate, $Pb(NO_3)_2$ and calcium nitrate, $Ca(NO_3)_2$ produced no change. Explain in your own words why these two reagents will not form a solid precipitate.

4. Which interactions produced no change? List the reaction pairs.

5. Considering your answer to question 3, are there any interactions that you could predict, even before mixing, would not produce a solid?

6. Consider your answer to question 2. The underlined products are insoluble and the other products are soluble. Consider your answer to question 4. The interactions that had no change yielded products that are soluble (which is why you did not observe a change). Using data your answers to both of these questions, construct two tables, one identifying the soluble products and a second identifying the insoluble products.

SOLUBLE	**INSOLUBLE**

7. Of the 6 cations you tested (Ca^{+2}, Pb^{+2}, Hg^{+2}, K^{+}, Ag^{+}, Na^{+}), which form *soluble* compounds when mixed with various anions?

8. Of the 6 cations you tested, which form *insoluble* compounds when mixed with various anions?

9. Of the 4 anions you tested (NO_3^-, Cl^-, Br^-, I^-), which form *soluble* compounds when mixed with various cations?

10. Of the 4 anions you tested, which form *insoluble* compounds when mixed with various cations?

11. Based on your data, complete the following table:

Compounds containing the following cations are generally soluble	**Exceptions**	**Compounds containing the following anions are generally soluble**	**Exceptions**

12. If 5 drops of $Ca(NO_3)_2$ is mixed with 5 drops of Na_2SO_4, a solid white precipitate is formed. Predict all possible products. Using your data, predict which of the possible products is the white precipitate. If you know that $CaSO_4$ is the main ingredient of Plaster of Paris, does this help you determine which product is most likely the white precipitate?

Name________________

CHEMICAL REACTIONS
PRELABORATORY QUESTIONS

1. List the four criteria used as evidence that a chemical reaction has occurred.

2. What is a spot plate?

3. Consider step 3 in the procedure for part I. Briefly describe how you will mix the reagents.

4. State all safety concerns specific to this laboratory experiment.

Appendix B

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LITERATURE CITED

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Appendix C

Chemistry 2A
Faculty Survey

CHEMISTRY 2A FACULTY SURVEY
STUDENT LABORATORY NEEDS ASSESSMENT

The purpose of this survey is to determine the needs of the students, significance of the skills currently emphasized, and relevancy of the laboratory topics for Chemistry 2A students.

1. Please rank the following topics in order of importance, with (1) being the most important:

Observations
Measurement
Graphical Analysis
Density
Specific heat of a metal
Change of state
Properties of gases
Properties of water
Chemical reactions
Acids and bases
Neutralization titration
Cation/Anion analysis
Formula of a hydrate
Empirical formula
Ionic equations
Solutions
Oxidation/reduction

Are there any topics that were not included that you think should have been?