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Introduction to the
Fundamentals of
Chemistry

by
Rose C. Knowles

A Thesis

Submitted in partial fulfillment
of the requirements of the
Master of Arts Degree in the
Graduate Division of Rowan University
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ABSTRACT

Rose C. Knowles

Introduction to the Fundamentals of Chemistry
May 1998
Dr. Richard Meagher, Community College Education

This thesis designs a curriculum to fulfill the charter of the Community college in providing an opportunity for an individual student whether traditional or non-traditional to prepare to become more proficient in chemical understanding and be able to apply the concepts and knowledge taught to a career or advanced study. The curriculum should benefit those individuals who either lack or have a poor background in chemistry and strengthen their reasoning ability. This paper presents a comprehensive comparison of text books currently being published for introductory chemistry. Texts reviewed were readily found on library shelves and supplied through major distributors; chemical journals were also included in the review.

The curricula that this project presents is aimed toward a first semester freshman college level chemistry class. The topics to be discussed and compared in this curricula include a broad overview of science, chemistry, the five subfields of chemistry, measurement, matter, energy, atomic structure, the periodic table, inorganic and organic nomenclature, balancing of equations, stoichiometry, the states of matter, an introduction to organic classification and reaction mechanisms, and biochemistry.

MINIABSTRACT

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This thesis designs a curriculum to fulfill the charter of the Community College in providing the opportunity for an individual student to prepare to become more proficient in chemical understanding. The curriculum should benefit those individuals who have a poor or no background in chemistry and strengthen their reasoning ability. This paper presents a comprehensive review of the literature currently available for introductory college chemistry.

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Introduction

The Community College was established to provide students and the local community with programs consistent with changing times, abilities, achievements, interests and goals. These colleges are made accessible to all of the community by low tuition, open access to those who are uncertain about educational interests or abilities and wish to explore a variety of occupational and academic programs. The seven main categorical programs of the Community college institution are academic transfer, occupational programs, remediation or basic skills, continuing education, career counseling, community service, and teaching based instruction, not individual research. One of the many purposes of the Community college, not unlike any educational institution, is to develop a student's sense of responsibility, the ability to communicate effectively and obtain a greater facility to think both clearly and critically.

Students that attend the Community college come from a variety of educational experiences, motivations, interests, attitudes, self-concepts, learning styles, cultures, and disciplines. The mission is to ensure excellence in performance and academic competency. A majority of students

are of the "traditional college" student age where as they have recently graduated from high school, plan to attend college with twelve or more credits, and reside at home with parents or guardians. However, the "non-traditional college" student population has increasingly become apparent in the past few years. The age range for beginning students is usually between 28 and 35 years old. These students attend classes for reasons such as change in career or opportunity, are mostly enrolled part time, and commonly are working at least a 40 hour week while maintaining households.

The curriculum's general education component introduces a body of knowledge that increases the interest in intellectual matters, enriches personal lives, develops power of judgment, and develops a commitment to mathematical, scientific, and technological concerns in society and individual professions. Courses offered by Community colleges should provide essential knowledge in order to enter careers and provide the foundation for those that plan on pursuing baccalaureate and graduate degrees.

Chemistry is an integral part of the scientific phase of the curriculum and is concerned with the composition of substances and changes that they may undergo. It is an area

of knowledge that is remarkable for its breadth and depth. Chemistry was one of the earliest sciences to be extensively and intensively studied. To this day, the field and understanding of chemistry is growing and is revolutionizing our culture and the way in which we think.

This paper presents a curricula proposal for an introductory level of the fundamentals of chemistry. Central concepts and basic principles of chemistry need to be explored and enhanced. An introduction to the many subfields of chemistry, including measurement, matter and energy, structure of the atom, periodic trends and classification, structure of compounds, inorganic nomenclature, balancing chemical equations, calculations involving stoichiometry, and all three physical states should be explored. Also, a brief introduction to organic and biochemistry should be covered in an atmosphere conducive to an understanding and comprehension of the aforementioned concepts.

Statement of the Problem

This thesis designs a curriculum to fulfill the charter of the Community college in providing an opportunity for

the individual student whether traditional or returning to prepare to become more proficient in chemical understanding and be able to apply the concepts and knowledge taught to a career or advanced study. This course should benefit those who have a poor background or no background in chemistry and strengthen their ability to reason through problems, either real or hypothetical. The introductory study attempts to aid the returning student who is unfamiliar with this material or simply may wish a comprehensive review course for academic or employment opportunity. Therefore, the applicability of this course may be practical to students who are pursuing curricula that includes additional courses in chemistry or in subjects that have college chemistry as a prerequisite.

In order for an introductory level chemistry course to be complete, laboratory procedures must also be developed. Therefore, basic laboratory work will support lecture topics. Students should learn to rationalize by making both qualitative and quantitative measurements with an outcome in support of theory.

Purpose of the Study

The raison d'être of this study is to set forth several goals. These are: (1.) to offer students curricula current in knowledge and technology in their chosen career path; (2.) to interface with the goals of the Community college; (3.) to provide a solid foundation for students to further their education at the baccalaureate level; (3.) to foster the development of cognitive learning and problem solving through critical thinking and analytical reasoning; (4.) to enhance creativity and develop organizational skills that promote effectiveness in personal and professional lives; (5.) to present the basic and fundamental concepts of chemistry; (6.) to provide the opportunity of individual growth through functional understanding of facts, principles, and concepts with the integration of laboratory experience; (7.) to develop the ability to see what is logical and how to derive quantitative/ qualitative relationships; (8.) to provide individuals with an understanding of the methodology and role that science plays through everyday life; (9.) to develop individuals that are comfortable with reading and integrating scientific literature into everyday life; (10.) to stimulate an

interest in science; and lastly, (11.) to provide a firm foundation for further scientific studies.

Significance of this Study

The significance of this paper is to present a chemistry curriculum enhances student's lives and to prepare individuals with a fundamental understanding of science that could be applied in further studies and/or career development. Throughout this research it is important to keep in mind the objectives as to why this curriculum is being developed. This course should develop an individual's analytical and research skills through integration of laboratory experience and lectured theory. It should remediate those with academic deficiencies in math and science ensuring that basic college level skills are obtained in order for continuation of scientific learning. There should also be a continuous interaction between instructor and student in professional development. Students can easily benefit from personal interactions with an instructor who is highly qualified in their scientific field.

The specific objectives that should conceivably be obtained by the end of this course are as follows:

(1.) The student should understand the concepts of the many subfields of chemistry and be able to properly identify specific work in each.

(2.) The student should be able to properly measure characteristics of matter and perform appropriate calculations.

(3.) The student should be able to explain the structure and states of matter, atomic structure, the basic forms of chemical bonding, and discuss the forms of energy.

(4.) The student should be able to identify the elements and know periodic trends that exist as to date.

(5.) The student should be literate in and able to complete or predict chemical equations and calculate stoichiometric problems.

(6.) The student should acquire inorganic nomenclature and be introduced to naming simple organic compounds.

(7.) The student should develop an understanding that there is a direct link between chemistry and biological functions through biochemistry.

(8.) The student should be able to differentiate between qualitative and quantitative analysis.

(9.) The student should be able to demonstrate proficiency in the use of the metric system, explain International Unit

System, perform calculations in scientific notation, be able to manipulate individual variables algebraically, and perform temperature conversions along with the use of factor analysis.

(10.) The student should be able to experimentally measure mass, volume, and length by all forms of analysis; for example, proper use of a balance both analytical and top-load; to deliver or contain pipettes, graduate cylinders, or volumetric flasks; metric rulers or calipers.

(11.) The student should be able to perform and properly follow experimental procedures with the capability of presenting data in a clear fashion and be able to explain observed results.

(12.) The student should have knowledge of and be able to demonstrate the proper technique in handling chemicals.

(13.) The student should have knowledge of and be able to identify and correctly use the appropriate apparatus for specific experiments.

(14.) The student should have knowledge of and be able to perform proper safety precautions and know to report all incidents of injury no matter how minor.

(15.) The student should have knowledge of and be able to identify appropriate Material Safety Data Sheets and be able to explain health hazards of chemicals.

Review of the Literature

Many individuals, professors and students alike, see chemistry as a vast accumulation of knowledge. In this introductory course, students should learn to manage information and procedures that which have traditionally been identified with an introductory chemistry course. This curricula may be the first step to make "sense" of the physical world. Each student should build a basis of understanding of atoms, molecules, and much more for a little knowledge of chemistry will open the doors to a depth of what is either recently been discovered or yet to be.

The main topics to be discussed in this curricula should include a broad overview of science, chemistry, the five subfields of chemistry, measurement, matter, energy, atomic structure, the periodic table, inorganic and organic nomenclature, balancing of equations, stoichiometry, the states of matter, an introduction to organic classification and reaction mechanisms, and biochemistry.

The first topic to be discussed is an overview of the definition of science and chemistry. The five main subfields of chemistry will also be presented and defined. Many texts present science and the scientific method

similarly. Sherman (1984) presents an abbreviated version of the chemical history time line and discusses chemistry in today's world. The brief chapter ends with career sketches with the application of their field of study. Peter's (1997) presents a short section on chemical history and discusses common terms such as hypothesis, theory and law. This chapter ends with today's chemistry and presents the countries top ten chemical producers. Rogers (1987) presents a cohesive introductory chapter. He defines matter, different forms of matter, and properties governing it. The three Laws of Conservation of Mass, Energy, and Mass/ Energy are clearly defined. The scientific method and processes of hypothesis formation, experimentation, and theory follow. The chapter ends with the presentation of four subfields of chemistry; analytical, biochemical, organic, and inorganic. Williams (1970) has a short introduction to chemistry and scientific approach. Boikess, Breslauer and Edelson (1986) present the definition of chemistry, the nature of matter and energy in separate sections, and discuss forms of measurement. Significant figures along with exponential notation and the International System are introduced. The chapter ends with factor analysis. Stoker and Walker (1988) do not present the divisions of chemistry. They do define chemistry and matter along with composition and structure. The material

discussed in their first chapter will be presented in the second and third topics of this curricula. Biochemistry, chemistry, and matter are defined in a text by McMurry and Castellion (1992). Physical and chemical properties are then discussed. Classification of matter and a brief introduction to the elements are presented in the first chapter. The brief discussion introduces the student to what will be presented in depth in the remaining chapters at a later time. This chapter concludes with a discussion on the difference between potential and kinetic energy. Zumdahl(1996) presents the importance of understanding chemistry, defines chemistry, discusses solving problems using a scientific approach, and the scientific method. Daub and Seese (1996)present all five subfields of chemistry and the scientific approach .

The second topic of this curricula should be to discuss mathematical operations and measurements. An instructor should assume that since this is an introductory course in chemistry that all of the students have not been exposed to scientific notation, exponential notation, significant figures and the laws that govern the operations with proper usage of significant digits, and factor analysis. Corwin (1994) presents the above criteria and he discusses the use of a scientific calculator and percentages. He defines

and discusses uncertainty in the measurements of length, mass, and volume. There are an abundance of mathematical examples throughout the presentation. He continues into a second chapter which goes in depth on measurement in the metric system and conversion factors from English to metric. Volume, volume by displacement, density, specific gravity, temperature conversions, heat, specific heat and the International System are defined and illustrated. Rogers (1987) chapter contains short explanations of the criteria that must be covered in addition with accuracy and precision. There are not many examples within the body of the text. In the newest edition of Introductory Chemistry by Peters and Cracolice (1997), scientific notation is presented similarly as other chemistry texts but dimensional analysis is explained clearly along with examples. Length, volume, mass, significant figures, rounding, mathematical operations, English/metric conversions, temperature conversions, density, and strategies for solving these problems are covered. The chapter concludes with an advisement for pitfalls to avoid in problem solving. This text does not discuss calculator use, it makes an assumption that students already understand how to use a calculator properly and in scientific mode. Area, volume, metric conversions, mass and weight, significant figures, scientific notation and temperature conversions are covered

in Sherman, Sherman, and Russikoff (1984). The material that needs to be covered is redundant; however, this text does not supply many examples in comparison to the others. Williams (1970) discusses methods of measurement of time, and distance, English/ metric conversions, mass, weight, density, specific gravity, temperature conversions, and significant figures in a short chapter. He does not supply enough examples and students may have a difficult time mastering the foundation of unit analysis from this book. Boikess, Breslauer, and Edelson (1986) define measurement, units, accuracy, and precision in their first chapter. They then discuss significant figures and exponential notation much as other texts do. A short section on calculations involving proper usage of significant figures is presented. Mass, volume, density, specific heat, and temperature conversions are presented and defined. Callewaert and Genyea (1980) introduce the processes of measurement in their opening chapter as the previously discussed authors. Measurement, mass, weight, length, volume, and exponential notation are briefly presented. The Law of Conservation of Mass is presented in this section. Density, specific gravity, uncertainty in measurement, significant figures, and conversion factors are presented without many examples. This chapter then covers chemical versus physical changes which will be presented later in the curricula. Stoker and

Walker (1988) begin their discussion by defining the differences between qualitative and quantitative measurement. Precision, accuracy, significant figures along with exponents are illustrated. Length, mass, volume, and dimensional analysis are presented. Density, temperature scales, and energy close this chapter. Many examples can be found throughout this chapter. McMurry and Castellion (1992) begin a discussion in the second chapter and dedicate an entire chapter to measurement. Physical quantities, scientific notation, mass, length, volume, significant figures, and rounding are presented as numerous authors have. Specific gravity, heat, energy, density, and temperature conversions are presented. Dosage calculations are discussed. The chapter revolves around chemical measurements that would be applicable for preparatory medicine. Zumdahl (1996) also dedicates an individual chapter to measurement. He also discusses the differences between qualitative and quantitative measurements. Scientific notation, metric versus English units, length, volume, mass, and uncertainty are discussed. The rules governing significant digits, rounding, and computational calculations are illustrated with many examples. Dimensional analysis is given consideration. Temperature conversions and density are reviewed in a similar fashion to other authors. Daub and Seese (1996) dedicate their second

chapter to measurement. Matter and its characteristics, the metric system, temperature, measuring devices, significant figures, and rules governing them with the inclusion on rounding, and exponential notation are clearly presented. The chapter concludes with factor analysis, density conversions, and specific gravity. There are many examples given within the body of the text that illustrate the concepts.

S.C. Roy (1997) presents formulas for temperature conversions. A quick conversion of Fahrenheit to Celsius is $F \sim 2C + 30$. However, the traditional formula presented in any chemical text is $F = (99/5 C) + 32$. The newest formula presented in this article is $F = 2(C - 0.1 C + 16)$. Converting from Fahrenheit to Celsius would be $C = (10/9) [(F/2) - 16] + 0.1[(F/2) - 16] + 0.01 [(F/2) - 16] + \text{etc.}$ One can continue this equation until the desired accuracy has been acquired. The error for each consecutive addition is as follows; 10%, 1%, 0.1%, etc. respectively.

It is assumed that a number of students enrolled in this course have not been exposed to exponential notation. Peckham (1997) introduces a newer method for typing and writing exponential notation. Older scientists are reluctant to this change; however, introducing this method

at a beginning level may be helpful. For instance, $[(2.654 \times 10^4)(6.022 \times 10^{23})]/(9.9 \times 10^{-16})$ can be simplified to $[(2.654p4)(6.022p23)]/(9.9n16)$. Advantages to this change are ease of use, saving space, legibility, association and the simple aesthetics involved. The only disadvantages foreseen are the unfamiliarity with the p/n notation and the diversion from established exponential notation traditions.

The next topics that need to be discussed are matter and energy. Both potential and kinetic energies should be fully explained and examples given. The four physical states of matter should be introduced and it will be here that the Laws of Conservation of Mass, Energy, and Mass and Energy are covered. The physical and chemical properties should be introduced along with changes representing both. The laws of definite and constant composition should also be discussed. Texts written by Corwin (1994) and Peters and Cracolice (1997) discuss the aforementioned topics in detail. Corwin (1994) begins with the three physical states and the changes they undergo; this is also the only text so far researched that introduces the fourth state of matter, plasma. Both texts discuss the forms of energy with mathematical expressions. There is a full classification of matter, names and symbols of elements along with the introduction of the periodic table, and the compounds with

chemical formulas explained. Physical and chemical properties and the changes that represent each are given. The three laws of conservation are discussed. The only other differences between these two chapters are Peters and Cracolice (1997) open the chapter with macroscopic, microscopic, and particulate matter and also explains the electrical character of matter. Rogers (1987) discusses the Laws of Conservation in the first chapter along with matter and its forms. Elements, compounds, and mixtures are in a separate chapter with the atomic theory and elements. In this chapter an introduction to chemical equations is given. Sherman, Sherman, and Russikoff (1984) cover the needed information in a brief and understandable manner, but introduce the mole and empirical formula. There are few examples throughout the chapter and material is not covered to the same depth as Corwin (1994) or Peters and Cracolice (1997) texts have. In the earliest published text written by Williams (1970), a chapter and part of another discuss the information needed for this curricula. Exothermic and endothermic reactions are introduced. Dalton's theory leads into structure of the atom which should be covered in a separate section. Boikess, Breslauer, and Edelson (1986) text discuss matter in a chapter with atomic structure and then dedicates two separate chapters on physical properties and chemical properties of matter. Energy is discussed in

the first chapter. Callewaert and Genyea (1980) classify matter in an entire chapter. They speak of the three states, the changes they undergo, and begin an introduction to atomic theory and chemical equations. Stoker and Walker (1988) begin their text with a discussion on the composition and structure of matter. Both physical and chemical properties along with their changes are covered. In a separate chapter, pure substances versus mixtures are presented. Energy and its types are briefly discussed in another chapter. McMurry and Castellion (1992) open their text with matter, energy, and life. The three states of matter are defined and classified. An introduction to the elements is presented and the chapter ends with energy. The chapter is brief. Zumdahl (1996) discusses matter, physical and chemical properties with changes, introduces elements and compounds, classifies matter, and ends the chapter with energy. He illustrates the material and supplies many mathematical examples for energy conversions. Daub and Seese (1996) also dedicate an entire chapter to matter and energy. They open the chapter with defining the three states and classifying them. The Law of Definite Proportions is discussed. Properties and changes of matter are defined. Potential and kinetic energy, specific heat, and the Laws of Mass, Energy and Mass and Energy are

covered. The chapter closes with the division of the elements.

It is important to introduce and discuss in detail the structure of the atom, subatomic particles, and past models which molded our initial understanding of the atom. The periodic table should be discussed as its own topic, one that will follow directly after the atom percievably has been mastered. Corwin (1994) begins with the differences between Dalton, Thomson and Rutherford theories. An in depth discussion on each is presented. The atomic mass unit and isotopes are explained. This text explains Bohr's atom with emission spectrum and the Balmer formula. Principal energy levels and sublevels are discussed with several practice problems to illustrate the material. In another chapter presented much later in the text, the Bohr atom is rediscussed in greater detail and quantum theory introduced. The quantum numbers are also explained and represented in a modern atomic theory. Peters and Cracolice (1997) use two separate chapters to cover the aforementioned material. They begin with Dalton's theory followed by subatomic particles, isotopes, atomic mass and finishing the chapter with an introduction to the periodic table. The second chapter, presented much later in the text, covers the Bohr model, the quantum mechanical model, electronic

configuration, Lewis dot structures and periodic trends. Williams (1970) supplies an explanation of Avagadro's number in two paragraphs. Rogers (1987) text discusses radioactivity, nuclear reactions, and half lives. The energy of the electron and atomic model with electronic information is presented with the periodic table and trends. The chapter ends with ionization energy, formation of ions both monoatomic and polyatomic. Sherman, Sherman, and Russikoff (1984) present the required information in two consecutive chapters. In the first, the atom is defined, theories are presented and subatomic particles are explained. The second chapter discusses the energy levels and configurations. Boikess, Breslauer, and Edelson (1986) dedicate an entire chapter to atoms and electrons. The structure of the atom is briefly discussed, atomic number, mass number, isotopes, calculations involving atomic weight along with natural abundance, and the mole with Avagadro's number are all presented. Energy levels and spectra are covered with a discussion on quantum numbers. Electronic configurations and valance electrons are presented. The periodic table is introduced. This chapter does not contain many examples within the body of the text. Callewaert and Genyea (1980) discuss Dalton's atomic model, atomic structure, ions, and isotopes in one chapter. The next chapter contains the periodic table and electronic

structure. Trends found from the table and classifications open the second chapter. They separate the families and discuss chemical versus physical properties. The Bohr model versus Quantum mechanical model is presented. The chapter concludes with electron configuration and dot formulas. Stoker and Walker (1988) discuss the aforementioned material in two chapters. The first iterates the scientific method, the atomic theory of matter, and the Law of Conservation of Matter. Subatomic particles, atomic number versus mass number, and isotopes are briefly explained without examples. The chapter closes with the periodic law. Electron structure of atoms is presented in a second chapter. Orbitals or shells are discussed along with subshells. The electronic configuration and Auf bau principle follow. Classification systems of the elements close this chapter and not many examples are provided. McMurry and Castellion (1992) provide a single chapter for atoms and the periodic chart. They similarly present atomic theory, subatomic particles, isotopes, orbitals, and configurations. Their chapter ends with periodic trends. Zumdahl (1996) covers elements, symbols, Dalton's theory, Rutherford's model, isotopes, ions, and an introduction to the periodic table in a well defined manner. He discusses in a second chapter, much later in the text, electromagnetic radiation and energy, energy levels, the wave mechanical model, orbitals,

electronic configuration, and the trends found on the periodic table dealing with valence electrons and generalized trends. Daub and Seese (1996) present one chapter on the atomic structure. Dalton's theory, subatomic particles, isotopes, energy levels, electron dot formulas, and orbital geometry are explained. Rutherford's model and calculations involving elemental abundance are also covered. Within the body of the text, there are many examples that provide insight to understanding these concepts.

Richard Pendarvis (1997) describes the electron's probability of being found in certain orbitals. After the students have been exposed to textbook learning of the atomic structure with the orbital being defined, the old definition of the orbital will be refined from 'simply where an electron can be found' to the sophisticated 'a mathematical function describing the wavelike function of an electron, or simply a probability'. Pendarvis (197) describes an analogy of the probability of finding students within a city, 100 miles, or 200 miles from the school's campus during a vacation. As the ratios increase, an anomaly can be made to the probability of where an electron can be found. That location is simply somewhere around the nucleus.

Garofalo (1997) has proposed a method for instructing students on the electronic configuration which includes the principal energy levels, sublevels, and the number of electrons in a sublevel. Since students have a difficult time conceptualizing abstract ideas, the writer has constructed four model houses from foam divided into levels and rooms. Each house represents a primary quantum number, each floor represents the sublevel, while each room is equivalent to the orbital. He uses two different colored beads to represent paired electrons with opposite spin.

Each house is mounted on a two by four foot peg board at different levels. When the levels are arranged properly, the rooms of the houses corresponding to the orbitals, should align as do the boxes on an Auf bau diagram.

Students should be reminded that electrons enter orbitals of lowest energy first. Garofalo (1997) states that he has had success with this model and brings concrete and abstract ideas together.

Tsaparles (1997) describes atomic and molecular structures in chemical education. He initially describes the structure of the atom. The Piagetian developmental perspective, the Ausubelian theory of meaningful learning,

the information processing theory, and the alternative conceptions movement are then discussed and considered as conflicting theories as to why students have difficulty grasping chemical concepts.

The Piagetian developmental perspective considers the four stages of developmental growth from the sensor-motor stage (birth to two years), the preoperational stage (2-4 years), the concrete-operational stage (7-11 years), and the formal operations stage (11-15 years). The majority of students in elementary and junior high levels, even senior high levels, operate at the concrete-operational stage. The question to be considered is what is so different between a senior in high school and freshman in college. Both are operating on concrete or physical attributes in order to conceptually understand. For these students, understanding an atom or molecule is difficult if not impossible.

The Ausubel's theory states that teaching should be based on that which students already know. Therefore, any meaningful learning occurs when preexisting knowledge interacts with new learning. This theory postulates that it is difficult for students to meaningfully learn the intricacies of atomic and molecular structures. Tsparles

contemplates how this difference can be overcome; however, this has not yet been determined.

The information processing theory draws an analogy between long term memory and working memory to a computer. The long term memory allows us to recognize familiar from unfamiliar, while working memory is where we process data for either long term memory or to be forgotten. The manner in which chemistry is taught comes into conflict with this model. Students who do not have a familiarity with chemistry at an introductory level, as should be expected; therefore, they cannot draw upon long term memory. Concepts and ideas presented to them in the working memory will preferentially be processed for long term memory and not to be forgotten. This new kind of concept takes awhile to culture, but once embedded in long term memory, it is a powerful way to learn.

The last theory to be postulated is the Alternative conceptions movement. Students enter a course with their own preformed conceptions, whether correct or not. These misconceptions or students ideas hold them back from learning what has been proven to be true. An overexaggerated example is that an atom of water is just a really small drop

of water. Students bias' must be overcome in order for science to be learned and truly understood.

Although these models pose a threat to those instructing beginner chemistry students, there are central concepts which must be presented. The introduction of the molecule as the smallest bit of matter that retains certain properties of a pure substance comprising two or more atoms strongly bonded together and of course the atom must also be presented in a clear, conceptual manner. Ions, bonds, electropositive and electronegative atoms, activity series, and oxidation numbers should follow.

Tsaparles (1997) proposes a historical approach to scientific discoveries to show the natural development of thought. The subject of atomic and molecular structure can be better understood if we consider evidence as well as evolution of ideas on the chemical bond.

Hernandez and Rodriguez (1996) remind readers of the difficulty that students have visualizing atomic structures. Molecular models are the best tools for students in both general and organic chemistry to see actual spatial arrangements. Models may be divided into a few types; such

as length and shape by ball and stick or space filling which represents Van der Waal radii.

They propose an inexpensive way to construct models without the need to purchase physical models. Thin flexible wires such as telephone hook up wires or even pipe cleaners would easily do the trick. Electronic organic orbitals designated as sp^3 and sp^3d^2 can be pictorialled with two to five ten centimeter long wires twisted together. Transition state geometries can also be constructed using this method.

The next topic which should be explored is the periodic table and the trends it predicts. Students should be taught the significance of what our chemical forefathers discovered and further predicted on constructing this table. The chapter should include an explanation of family and period definition, valance shell and electronic trends. Corwin (1994) has a chapter dedicated solely to the purpose of instructing the reader of the periodic table. Peters and Cracolice (1997) have only one small section devoted to the periodic table and trends. It does not discuss the names of the groups or electronic trends. Rogers (1987) and Sherman, Sherman and Russikoff (1984) present this information in a similar fashion as Corwin (1994). One can conclude that not much has changed on the periodic table. Williams (1970)

text written almost thirty years ago presents much of the needed information. He has three chapters dedicated to the halogens, nonmetals, and metals. He discusses properties and basic chemical reactions in which they undergo. Boikess, Breslauer, and Edelson (1986) summarize families, periods, and classification of the elements in two pages. Callewaert and Genyea (1980) present the periodic table in half of a chapter. They cover all of the intended information and how to decipher the table. Stoker and Walker (1988) introduce the periodic law in a chapter on atomic structure and in a later chapter discuss trends in electronic configuration. They omit many of the physical and chemical trends that the table predicts. McMurry and Castellion (1992) present the periodic table, group characteristics, and electronic configuration trends in detail. In a section of another chapter, they dedicate the discussion of periodic properties on ion formation. Zumdahl (1996) introduces the periodic table, groups, families, physical properties and ion trends about one third into his text. He then continues his discussion later in the text on electron arrangement, configuration, orbital filling, and atomic properties which include ionization energy and radii. Daub and Seese (1996) dedicate an entire chapter to the periodic table. The law, periods, groups, and characteristics are all defined. Daub and Seese (1996) and

Corwin (1994) present the needed information in one cohesive chapter; whereas, McMurry and Castellion (1992) and Zumdahl (1996) introduce the table early in the text and continue the finer details later.

Several articles published in the Journal of Chemical Education in 1997 discuss elements of the periodic table. Jensen (1997) discusses Group VI A elements. He defines what is meant by chalcogen and the history of when it was derived. Modern texts claim that this term means chalk former, glass former, and even brass giver or copper producing. The first two meanings are unambiguously incorrect and the other two imply that the ores formed are the actual chalcogens rather than the elements of oxygen and sulfur.

Originally, the chalcogens were proposed to be amphigens. This term denotes the ability of these elements to form both acidic and basic compounds.

Although the article was brief Jensen (1997) presents background on Group VI A elements. He clearly discusses more information on this group in the periodic table than what can be found in many texts.

Venkataraman, Wilson, Hirsch, Zhang, and Moore (1997) present coordination geometry of the d-block (d denotes diffuse) elements and their ions is presented. The article is written such that an introductory chemistry student would have difficulty in reading. They present a graph of molecular geometries, totaling twelve shapes with corresponding IUPAC (International Union of Applied and Pure Chemistry) symbols and ideal angles is presented.

After the periodic table and predictable trends have been presented, the structures of compounds need to be explored. Chemical bonds need to be discussed and the types of bonds formed from ionic, polar, nonpolar, to covalent, shape and bond angles delved into. Electronegativity, the octet rule, polyatomic ions, Lewis dot structures, multiple bonds and geometry need to be covered. Corwin (1994) begins a discussion on monoatomic and polyatomic ions in a chapter on nomenclature. He discusses the material in a chapter much later in the text. There is a discussion on bond types, formation of ions, structure and electron dot formulas. He presents many examples within the chapter that illustrate main topics on bonding characteristics and Lewis dot structures. Peters and Cracolice (1997) present chemical bonding in a chapter that covers explanations on ions, all bond types, multiple bonds and octet rule

exceptions. In a following chapter, structure and shape are covered. Lewis dot, electron pair repulsion, molecular geometry and an introduction to organic structure are presented. Rogers (1987) covers all of the basic definitions, Lewis dot, shapes of molecules, polarity and electrolytes. The chapter does not contain as many examples as other texts. Sherman, Sherman, and Russikoff (1984) cover the required topics with the exclusion of geometrical shapes and they introduce a little of nomenclature. Williams (1970) gives a basic introduction to the understanding of ionic bonds, writing formulas, covalent bonds, valance shell, oxidation-reduction, and multiple valances. Boikess, Breslauer, and Edelson (1986) discuss the difference between ionic and covalent structure and geometrical shapes. They do not supply many examples. Callewaert and Genyea (1980) dedicate an entire chapter to chemical compounds and bonds. They discuss the various kinds of bonds, electronegativity, Lewis structures, geometry, and resonance structures. The chapter contains few problems. Stoker and Walker (1988) divide a chapter into two sections, ionic and covalent. They define and discuss electron dot formulas, structures, and nomenclature. They have drawn very definite lines between the two bonds and supply few examples. McMurry and Castellion (1992) divide ionic compounds and covalent compounds into two separate consecutive chapters. In the

first chapter, they define ionic bonding, Lewis structures, properties, ions, nomenclature, polyatomic ions, and formula units. There are many examples within the body of the text. In the other chapter, covalent bonds, multiple bonds, structural formulas, Lewis structures, shapes, electronegativity, polarity, and properties are illustrated. They also supply many examples problems. Zumdahl (1996) dedicates an entire chapter to chemical bonding. He begins with the definition of kinds of bonds, then continues with electronegativity and determination of polarity. Dipole moments, stable electron configurations, and ionic charges lead into ionic bonding and structure. Lewis dot formulas of both ionic and covalent compounds with multiple bonds are presented. Resonance is also introduced. The VSEPR (Valance Shell Electron-Pair Repulsion) model is discussed. The geometrical shapes are illustrated. The structures of compounds is discussed in one chapter by Daub and Seese (1996). Chemical bonds are defined, ions, oxidation numbers, ionization, electronegativity, polarity, shapes, and formula writing are covered.

It is imperative that the student be able to name a chemical correctly. Various forms of nomenclature should be discussed. One should be literate in writing and reading chemical compounds interchangeably by the completion of this

topic. Corwin (1994) covers the aforementioned material in detail. The International Stock System and Latin systems are discussed, and writing and predicting chemical formulas are presented. Peters and Cracolice (1997) cover the topics in the same fashion as Corwin (1994); however, many examples are given throughout the chapter. Many graphs and charts are presented. Rogers (1987) dedicates only a few pages toward nomenclature. Sherman, Sherman, and Russikoff (1984) discuss in a stepwise fashion how to name chemicals by the IUPAC system in two pages. Williams (1970) does not have a section discussing nomenclature. Boikess, Breslauer, and Edelson (1986) also do not discuss nomenclature. Callewaert and Genyea (1980) discuss nomenclature for some simple inorganic compounds in an abbreviated fashion in an appendix. Both Stoker and Walker (1988) and McMurry and Castellion (1992) discuss nomenclature within chapters on ionic and covalent compounds. Zumdahl (1996) dedicates an entire chapter to nomenclature. He describes how to name ionic compounds, types I and II, and binary compounds, type III (only nonmetals). Polyatomic compounds and acids are also developed. He concludes the chapter with writing formulas. Daub and Seese (1996) also dedicate an entire chapter to nomenclature. They discuss systematic naming of metals with and without fixed oxidation numbers, ternary and

higher compounds, acids, bases, and salts. Their chapter is brief and does not supply many example problems.

It is important that students understand that chemistry is not a stagnant study. Additional elements are being discovered and experimentally processed. Many periodic tables account to element 106; whereas, 101-106 are unnamed.

In three issues published in Chemical and Engineering News, heavy elements atomic masses 101-109 have officially been named. February 24 and March 17, 1997, issues present elements 104-109 as Rutherfordium, Dubnium, Seaborgium, Bohrium, Hassium, and Meitnerium, respectively. The latest issue in September, 1997, elements 101-103 are named Mendeleevium, Nobelium, and Lawrencium. Elements 110-112 remain unnamed as of this date.

Software can be purchased from the Journal of Chemical Education (1997) which aids students with inorganic nomenclature. Either formula to name or vice versa can be practiced. This is strictly a drill and practice program in which a question is presented and five multiple choices follow. The student is to choose the answer which best fits. The advantage of this program is that when a student selects an incorrect answer, the program explains why the

answer is incorrect and supplies the correct statement. This software could enhance this portion of the curricula. Students need to be quizzed on their individual study habits and this program seems to be beneficial for learning nomenclature.

The next topic to be discussed in the curricula is the many kinds of chemical reactions; such as single and double displacement, combination, decomposition, neutralization, and balancing chemical equations. Corwin (1994) explains evidence for chemical reactions, writing chemical equations, stepwise balancing and classifying chemical reactions. He illustrates each of the reaction types with many examples provided along with practice problems. Peters and Cracolice (1997) cover the information in a similar manner as Corwin (1994); however, there are not as many practice problems. Rogers (1987) discusses the many kinds of reactions and evidence confirming that a reaction has taken place. He intertwines balancing amidst the discussion on chemical reactions. Oxidation-reduction reactions are discussed. Other topics covered in the chapter are mass relationships, percent yield, limiting reagents, and enthalpy. The mathematical operations will be discussed in the following section and enthalpy will not be developed within the content of an introductory course. Sherman, Sherman, and

Russikoff (1984) cover the required information and provide examples within the chapter. Williams (1970) does discuss the kinds of chemical reactions but only briefly presents the concept of balancing. Boikess, Breslauer, and Edelson (1986) cover chemical equations in one section and general types of reactions in the following section. Callewaert and Genyea (1980) do not discuss balancing of equations, or types of reactions. They dedicate an entire chapter to intermolecular attractions which will be concisely discussed in the introduction to organic chemistry. Stoker and Walker (1988) present balancing within a chapter on calculations. They do not cover the forms of reactions. McMurry and Castellion (1992) discuss balancing of equations in a stepwise fashion in one section of a chapter on calculations and later in the same chapter the types of chemical reactions are discussed. Zumdahl (1996) presents the evidence for chemical reactions and balancing in one chapter. In another chapter, he classifies reactions. Daub and Seese (1996) dedicate one chapter to balancing and types of equations. They provide many examples within the body of the text and present many stepwise instructions on how to work through these problems.

The topic to be explored next in this curriculum should be chemical calculations. By the completion of this

section, the student should understand the following concepts: the mole, Avagadro's number, conversion into molecules, ions, atoms, molar mass, percent composition, empirical versus molecular formulas, interpretation of balanced equations, mole/mole, mass/ mass, mass/volume, volume/volume at STP, accuracy and precision, and lastly limiting reagent problems. Corwin (1994) covers the aforementioned material in depth and provides many illustrations demonstrating each. The material is presented in two chapters. However, he does not discuss the limiting reagent. Peters and Cracolice (1997) present the above information in two chapters. The first chapter covers information from Avagadro's number to molecular formulas. The second chapter covers the remainder of the material with discussions on thermochemical equations and stoichiometry. This text supplies many mathematical examples. In the text presented by Rogers (1987), a small section in one chapter is dedicated to stoichiometry. It covers mass relationships, percent yield, and limiting reagents. Sherman, Sherman, and Russikoff (1984) cover mole concept, a brief section on quantities of reactants and products, and limiting reagent problems. Williams (1970) does not cover stoichiometry. In the text written by Boikess, Breslauer, and Edelson(1986), atomic weights, percent natural abundance, the mole, and Avagadro's number are

presented very early in the text. In fact, these topics are discussed in chapter two. They present molecular weight and composition in chemical formulas in a small section in another chapter. Callewaert and Genyca (1980) cover atomic and molecular weights, Avagadro's number, the mole, mass conversions, and discuss the difference between exothermic and endothermic reactions in a single chapter with few mathematical examples. Stoker and Walker(1988) dedicate a chapter to chemical calculations. They begin with formula weight and then introduce the mole along with Avagadro's number. Molar mass, mass, mole and volume calculations are explained. There are many more examples in this text in comparison to the previous two texts. McMurry and Castellion (1992) also present chemical calculations as part of one chapter. They begin with a discussion on balancing chemical equations, Avagadro's number, the mole, gram-mole conversions, mole-mole conversions, and percent yield. Although they supply many examples, Stoker and Walker's (1988) and McMurry and Castellion's (1992) texts are written more for a medical perspective. Zumdahl(1996) dedicates two chapters to calculations. In the first chapter, atomic mass, the mole, Avagadro's number, molar mass, percent composition, calculations involving empirical and molecular formulas are presented and illustrated. He uses many examples and flow charts. In his second chapter, the

information given by chemical equations is presented along with mole-mole, mass-mole, and mass-mass calculations. He presents calculations involving limiting reagents. He dedicates many pages and examples to this topic. The chapter concludes with percent yield. Zumdahl(1996) covers all of the requirements to be presented on stoichiometry for this curricula. Daub and Seese(1996) also present this material in two chapters. They begin with calculations involving elements and compounds. Formula and molecular mass, Avagadro's number, molar volume at STP, percent composition, empirical and molecular formula are also explained with many examples. Their second chapter begins with methods for solving stoichiometry problems. Mass-mass, mass-volume, and volume-volume problems are defined and illustrated. Percent yield and limiting reagent problems are discussed for each type of conversion. Endothermic and exothermic reactions conclude this chapter.

Carla Krieger(1997) has discovered a methodology to overcome students confusion on moles. Instead of blindly memorizing dimensional analysis, she invented "Moe's Mall" which is a motivating, humorous, and memorable way to understand the concept of the mole.

There are a few steps that must be imposed before Moe's mall can be introduced: (1.) material must be presented in small pieces, (2.) students must be questioned on new information after it is presented, and (3.) periodically review all steps when adding new additions. The mall's design is as follows: there is a central square in which the only entrance and exit and to three corners of this square are three halls that lead to three individual stores of this mall. The store connected to the northeast corner is a car parts store whose motto is "Parts is Parts", the northwest store sells bulk foods, and the southwest is a record store.

As students enter the fictitious mall on the south face of the central square, they are greeted by Moe to take them to see Shemp who is selling avocados. As they pass by, Moe eats 6.022×10^{23} avocados, They continue up to the parts store, where no matter what part you may need any old part will do. When coming back down the hall, Moe stops again to eat 6.022×10^{23} avocados. The students soon realize that the mole or Moe contains 6.022×10^{23} avocados or Avagadro's number.

Moe then takes students to see Curly in the bulk food store, where mass quantities are visited. Then they walk

down the southwest corridor to meet Larry at the music store. The motto there is "Pump up the Volume". Here the louder the volume, the higher the price. The northeast hallway is labeled by avocados or Avagadro's number, the northwest hallway is the Moe-Curly mass or molecular mass, and lastly the southwest hall is Moe-Larry or molarity.

Kriegar (1997) does not supply a visual; students must draw their own mall, thus promoting long term memory. Each time a student goes from one store to another a conversion factor must be applied. Soon students are using dimensional analysis easily; for instance, to get from moles to particles means traveling from Moe's square to the Parts store while stopping to eat 6.022×10^{23} avocados. This technique is also presentable with the concept of molarity; however, in this course molarity is not taught due to restraint of time.

She uses a reward system based on the Three Stooges theme for quizzes or tests. Moe's Mall can also be expanded to solve mole-mole, mole-mass, mass-mass, mass-volume, volume-volume, and solution stoichiometry by adding mirror images with bridges.

This is a unique way of introducing stoichiometry. The students learn dimensional analysis while traveling from store to store.

Kashmar (1997) introduces a new method for illustration of molecules and limiting reagents. He uses a variety of colored acetate transparencies in which circles are cut out that are approximately two centimeters in diameter. Elemental symbols can be written directly on these circles with a felt tip pen. Different elements are represented by different colors. An illustration of two molecules of hydrogen and one molecule of oxygen would require four hydrogen circles and two oxygen circles. Now, to combine them to form water, an oxygen circle will have two hydrogen circles attached to it; therefore, two molecules of water would be formed. The physical rearrangement of the molecules on the overhead projector illustrates the fact that molecules actually do separate and recombine during chemical reactions.

These circles can also be used to explain limiting reagents. Six molecules of hydrogen and two molecules of oxygen can be placed on the overhead. The idea is to combine as many molecules of water as possible. Students will visually see that only four molecules of water can be

formed with the original amount of reactants. Two hydrogen molecules are left untouched. Therefore, since all of the oxygen circles were used, it limits the quantity of product to be formed.

According to Kashmar, this is an easy and inexpensive way to illustrate reactions and limiting reagents that should prove successful when introduced in an introductory chemistry course.

The next topic that needs to be explored should be the physical states of matter. Since this is an introductory level chemistry course, plasma should be defined but not expanded on. It is imperative that the student learns gaseous properties, variables that affect them, and the many laws governing them. Once this task is accomplished, liquids and solids should be covered. State properties and intermolecular forces along with physical changes should be explained. Corwin (1994) breaks the above material into two chapters. He states the physical properties of gases and then defines atmospheric pressure, the barometer, vapor pressure and boiling point. A few examples are given within this text that helps illustrate these concepts. The laws covered in this chapter are Dalton's partial pressure, Boyle's, Charles', Gay-Lussac's, Combined gas and Ideal gas.

Variables affecting gas pressure are explored. There are many self test questions given to solve the various laws. Peters and Cracolice(1997) present gases in three and a half chapters. They separate the Ideal gas from combined gases and the other laws. In the first chapter, the Ideal gas law is given. Stoichiometry is also revisited, now only to the application of gases. Many examples can be found within the body of the chapter. The second chapter deals with the combined gas law and specific mathematical problems that illustrate it. Dalton's law is presented briefly in the chapter with solids and liquids. Properties, the barometer, Kelvin scale, Charles', Boyle's, and Combined laws are explained with many examples in the third chapter. Rogers (1987) presents the above information in one chapter with a short explanation of stoichiometry. There are few examples within the chapter. Sherman, Sherman, and Russikoff(1984) present the gas laws and properties in one chapter. Williams (1970) explains all of the laws in one chapter also. His text does not have individual chapters dedicated to solids and liquids. He discusses the properties and physical states briefly in the beginning of the text. Corwin (1994) discusses the properties of liquids, vapor pressure, viscosity, surface tension and intermolecular forces. He then explains the structures they form along with heats of fusion and vaporization. The

chapter finishes with water. Peters and Cracolice (1997) discuss the properties of liquids, intermolecular forces, liquid-vapor equilibrium, and boiling in an expanded explanation. They distinguish the forms of solids, types of crystallization, energy change with state, specific heat, and the heats of vaporization and fusion. Rogers(1987) begins his chapter with kinetic and potential energies. He then discusses intermolecular forces, physical properties of solids and liquids and the forces which govern them. There are not many examples within the content of the material. Sherman, Sherman, and Russikoff(1984) cover the physical changes of matter and the properties of both liquids and solids. Their chapter does not contain many mathematical relationships. Boikess, Breslauer, and Edelson (1986) discuss the physical properties of matter in a single chapter. They define the differences between the three states of matter. Boyle's, Charles', Gay-Lussac's, the Ideal gas laws and Dalton's partial pressure are presented in the first third of the chapter. Solids are defined next along with changes of state. Liquids are briefly covered. Callewaert and Genyea (1980) also place gases, solids, and liquids in a single chapter. All of the aforementioned gas laws are presented without many mathematical examples. Solids and liquids are concisely discussed. The chapter concludes with enthalpy, entropy, free energy, and the

direction of change. Stoker and Walker (1988) begin a chapter on states by defining property differences between physical states. The kinetic molecular theory of matter and electrostatic interactions are defined. Each state is individually presented and all of the gas laws illustrated. The text supplies many biological examples. McMurry and Castellion (1992) also present this material in much the same manner as the previous authors. They discuss the kinetic theory, pressure, all of the gas laws, liquids, intermolecular forces, solids, and changes of state very similiarly to other text books. This chapter contains many examples. The previous four texts have covered the intended information in a single chapter. Both Zumdahl(1996) and Daub and Seese(1996) use two chapters to present this information. Zumdahl (1996) discusses gases in his first chapter. Pressure is defined, Boyle's, Charles', the Ideal gas and Dalton's laws are presented and given numerous examples. Gas stoichiometry is revisited and the kinetic molecular theory established. In the following chapter, liquids and solids are presented. Water and its phase changes open the chapter. Energy requirements, intermolecular forces, evaporation, and vapor pressure follow. The types of solids are bonding concludes the chapter. Daub and Seese(1996) present gases in one chapter. Characteristics of ideal gases and the kinetic

theory are presented. Pressure is defined followed by the gas laws. They conclude their chapter with calculations related to the gas laws.

There are two experiments introduced in different issues of the Journal of Chemical Education in 1997 to illustrate Boyle's Law. Lewis(1997) demonstrates that a volume of fixed mass of confined gas maintained at constant pressure is inversely proportional to the pressure of the gas.

The experiment requires a bathroom scale, a 60 milliliter syringe, and a wooden block. The syringe is attached to the block and the plunger is placed into the top of the syringe. When applying force to the top of the syringe, the whole apparatus needs to be on the scale.

Results can be graphed as the change in pressure to the inverse of the volume of gas inside the syringe by any excel spreadsheet to yield a straight line. This is a quick and easy way to present Boyle's Law.

Richmond and Parr (1997) use a U-tube filled with fluid is used to demonstrate this law. A ten milliliter graduated pipet is sealed at one end and heated in order to bend it

into a U shape. Air is trapped near the closed end by pipetting a liquid into the open end. The U-tube is attached to a pressure gauge by flexible tubing. By varying the pressure, changes in trapped air volume will be apparent. Students should plot pressure versus volume to see what curve results. The graph can be made linear by plotting pressure against the inverse of the volume.

In order for the curriculum to be a complete introduction to the fundamentals of chemistry, the student should be exposed to organic chemistry. A definition of organic chemistry should have been introduced in the beginning of the course. Hydrocarbons should be discussed in depth along with structural isomers. Nomenclature of alkanes, alkenes, alkynes, and aromatics also to be included. Simple reactions such as addition, substitution, and halogenation need to be introduced. Properties of each hydrocarbon should be presented. The student should be able to recognize and name simple functional groups; such as alcohols, ethers, esters, carboxylic acids, amines, amides, halides, ketones, and aldehydes. The texts written by Corwin (1994) and Peters and Cracolice(1997) present the mentioned material in much of the same fashion. Peters and Cracolice (1997) define and supply examples of polymers with their uses at the end of the chapter. Rogers (1987)

discusses general characteristics, nomenclature, and functional groups. He does not discuss basic organic reactions but does end the chapter with a brief paragraph on polymers. Sherman, Sherman, and Russikoff (1984) separate organic chemistry into two chapters. The first chapter deals with hydrocarbons, structures, alkanes, alkenes, alkynes, cyclics, and aromatics. Examples are given throughout the chapter. Their second chapter is dedicated to functional groups, nomenclature, and examples. Williams (1970) has limiting material in the introduction to organic chemistry. He explains isomers, bond angles and bond structures. Alkanes, alkenes, alkynes, and aromatics are presented along with substitution and addition reactions. The only functional groups discussed are alcohols, aldehydes, ketones, acids, and amines. The following four texts by Boikess, Breslauer, and Edelson (1986), Callewaert and Genyey (1980), Stoker and Walker (1988), and McMurry and Castellion (1992) present an introduction to organic chemistry similarly but in many chapters. Boikess, Breslauer, and Edelson (1986) use six chapters to illustrate the introduction, hydrocarbons, oxygen containing compounds, carbonyl compounds, carboxylic acids and nitrogen, sulfur, and phosphorus containing compounds. Callewaert and Genyey (1980) similarly present their topic on organic chemistry which spans six chapters. Stoker and Walker (1988) go much

more in depth with their discussions, as can easily be seen in a total of seven chapters. McMurry and Castellion (1992) follow much the same presentation as Stoker and Walker. Zumdahl (1996) presents the topic of organic chemistry in a single chapter. He introduces carbon bonding, alkanes, isomerism, simple reactions, alkenes, alkynes, aromatic compounds, and nomenclature of each. Functional groups are also discussed along with properties of each. Alcohols, aldehydes/ ketones, carboxylic acids, and esters are explained. The chapter concludes with an introduction to polymers. Daub and Seese (1996) also present organic chemistry in a single chapter. They begin with defining organic chemistry and hydrocarbons. The geometry of organic molecules is concisely discussed followed by the traditional presentation of alkanes, alkenes, alkynes, isomers, aromatic compounds, halides, alcohols, ketones, aldehydes, acids, amines, amides, nomenclature of each, and simple reactions.

The next topic of this curricula is biochemistry. It is important that the student be presented with a firm foundation in chemistry and that the biological realm intertwine with the chemical principles governing it. This section should include a reiteration of the definition of biochemistry and a brief introduction as to why this study is important. Amino acids, proteins, enzymes,

carbohydrates, lipids and nucleic acids will be defined and classified. Bonds that make up each and terminology that describes each should be presented. Williams (1970) describes the composition of lipids, carbohydrates, proteins, nucleic acids, and vitamins. He defines metabolism and applies this to proteins, lipids, and carbohydrates. Sherman, Sherman, and Russikoff (1984) also present metabolism, anabolism, and catabolism. They classify carbohydrates, lipids, proteins, and vitamins and minerals. The chapter provides explanations and illustrations of each. Peters and Cracolice (1997) not only define each term but describe macromolecules. Proteins are described by their amino acids, peptide bonds, alpha amino and carboxy termini, and primary to quaternary structures. The terms and models of enzymes are covered. Carbohydrates are broken down into monosaccharides, disaccharides, and polysaccharides each being given many examples. Fats, oils, waxes, phospholipids, and steroids are discussed in a lipid section. The chapter ends with nucleic acids. DNA and RNA and the process from transcription to translation of protein is covered. Boikess, Breslauer, and Edelson (1986) present biochemistry in eleven chapters. They discuss in detail, stereoisomerism, carbohydrates, lipids, proteins, enzymes, nucleic acids, energetics, lipid and protein metabolism, body

fluids and health in the environment. Callewaert and Genyee (1980) similarly discuss the material also in eleven chapters. Stoker and Walker (1988) discuss stereoisomerism, carbohydrates, proteins, lipids, enzymes, nucleic acids, vitamins, minerals, hormones, metabolism, mitochondrial oxidation, and extracellular fluids all in thirteen chapters. McMurry and Castellion (1992) fill ten chapters to explore this topic. The above four texts provide a vast amount of information on this single topic; therefore, they can not be used in this presentation of this course. These texts provide enough information to be discussed in a two semester course. McMurry and Castellion (1992) and Stoker and Walker (1988) text books are both excellently written and could have been considered if (1.) the curricula being developed were a two semester course, and (2.) the course were biologically based. Zumdahl (1996) dedicates one chapter to biochemistry. He discusses proteins, amino acids, structure of proteins, function of protein, enzymes, carbohydrates, nucleic acids, DNA and protein synthesis, and concludes with lipids. He intermittedly throughout the text presents a cumulative review every two to three chapters.

The topic which should conclude this curricula is safety and the correct handling of chemicals. Although

individuals may assume that the laboratory is safe possibly due to the minute amount of chemicals being handled, it may not be. Many issues of reports in journals account for accidents and in some cases fatalities can readily be found.

A few examples are (1.) Chemistry Lab Fire at University of Texas (1996). A total destruction of three laboratories occurred. The fire was caused by a disposal of sodium into a drain after treated with alcohol. Apparently all of the sodium was not decomposed and the residue metal caught fire igniting solvent mixtures stored near the sink, (2.) Plant Explosion at Arkansas Plant Kills Three Firefighters and Injures 17 (1996). This particular fire blazed for six days and the cause was an insecticide known as azinphos-methyl, (3.) Dartmouth College Chemistry Professor Dies of Mercury Poisoning (1997). Karen Wetterhahn was preparing a mercury standard for the NMR when she was accidentally killed. Her death was an accident that could happen to any chemist even experienced ones.

Safety must always be the number one priority in the laboratory. Chemists often work in dangerous situations where they are exposed to a variety of hazards.

Lois Kramer (1997) reviews McMurry and Castellion's Fundamentals of General, Organic, and Biological Chemistry 2 ed.. This text is intended as an introduction to chemistry for students who are preparing for allied health or life sciences. The authors use many medicinal mathematical problems to assist those who are math phobic. This text is written in a clear fashion that provides countless examples to illustrate concepts. The text also provides numerous colorful pictures and graphs which aid a beginning student. Goals begin each chapter to guide the student to exactly what will be obtained at the completion of the chapter. The text also describes the relevance of chemistry in our lives without losing the content needed to be covered.

Nakhleh, Lowrey, and Mitchell (1996) tried an innovative method for presenting General Chemistry II. They tried narrowing the gap between conceptual and algorithmic processing. There was a distinct metamorphosis from heavy emphasis on mathematical problem solving to a mixture of conceptual questions.

The intent of this course is to become more interactive and stimulate motivation by linking concepts to everyday life. Initial surveys from students were acquired followed by a final survey. Overall, students were pleased by the

change in presentation and felt that they received much more from the course.

The class met three times a week, each session was 50 minutes. In the first 20 minutes, a topic would be developed while the remainder of the class would be used for illustrations of problems. The change occurred by using one of the three sessions as an interactive problem solving time. Students were divided into groups and each group was given a particular problem to solve in 25 minutes. When the time had elapsed, each individual group presented the method by which they solved it. The professor interjected only when incorrect information was given. Students responses' were that they were never required to think conceptually and asked to decipher problems before.

Each exam contained approximately eleven free response questions, which were an equal mixture of mathematical and conceptual based. The first exam set a guideline and basis, results were not wonderful. Students felt that the special session was a worthwhile part of this course and the instructor felt it was helpful in spotting student's weakness' of fundamental concepts. Subsequent exams proved to be much better in scores once the students became aware of the exams style.

In a final interview, the professor was extremely enthusiastic about the changes. He notes that the students seemed more interested, alert, and active.

Laing (1996) iterates that students only know the "plug and chug" method in solving equations. Once again, students seldom face a problem in which they do not readily apply conceptual knowledge. They need to be presented with historical facts and encouraged to actually think.

Weiner and Peters (1998) covers a wide variety of experiments that can be used to supplement lectures. Many of the laboratories that should be performed as to accompany lecture materials are as follows: (1.) Properties and Changes of Matter, (2.) Measured Quantities and Derived Quantities: Significant Figures, (3.) Densities of Liquids and Solids, (4.) Simplest Formula of a Compound, (5.) Hydrates, (6.) Percentage of Oxygen in Potassium Chlorate, (7.) Calorimetry, (8.) Chemical Names and Formulas, (9.) Chemical Equations, (10.) Molecular Models, (11.) Boyle's Law, (12.) Hydrocarbons and Alcohols, (13.) Aldehydes, Ketones, and Carboxylic Acids, (14.) Carbohydrates, and (15.) Amino Acids and Proteins.

Design of the Process

The curricula that this project is presenting is aimed toward a first semester freshman college level chemistry class. Previously in this paper a diagnosis of need, formulated objectives, and a selected content to be covered in the curricula was given; which can be found in the statement of the problem, the purpose of the study, and the significance of the study. The methodology in reviewing the literature presents the organization of the material to be covered, the preferred sequence of topics, and a comprehensive comparison of text books currently being published for introductory chemistry.

The following chapters attempt to present an introduction to topics essential for the students knowledge in chemistry. In the first chapter, a few questions will be posed for the students to ponder. For example, what is chemistry and how many substances can one name as not being a chemical? One may possibly promote water not a chemical or maybe an inanimate object such as a desk. However, each molecule of water has three components. The water molecule is composed of two atoms of hydrogen to every one atom of oxygen. The desk is composed of many organic and inorganic substances from in the wood, to the binding substances, to

the stain. These substances range in their chemical composition from being organic in nature to inorganic manmade molecules. Therefore, one should soon conclude that everything both synthetic and natural is composed of some form of chemical. Once the students are made aware of this concept, many definitions need to be covered. Science in general must be defined followed by a definition of chemistry. The five subfields of chemistry need to be presented and defined. These subfields of chemistry, analytical, biochemical, inorganic, organic, and physical, should be given examples and related employment opportunities explored. The chemical field ranges not only to industry but also to environmental testing and quantities of contaminants found in drinking water. The chapter will conclude with a presentation of the scientific method, which is comprised of experimentation, hypothesis formation, further experimentation and/or theory formation. Although this chapter is brief, many definitions need to be presented along with the concept of scientific methodology.

The second topic to be addressed are the mathematical operations employed in chemistry. It should be assumed that the class will be composed of many students who are math phobic. This fear needs to be addressed and the professor must put the students at ease by correcting misconcepts and

the lack of understanding of the sequence of operational procedures. The following areas to be discussed in this chapter are measurement and its uncertainty, measuring devices for length, volume, and mass, along with precision and accuracy. Significant digits should be explained and the rules governing their assignment for use in mathematical operations, for instance the rules for addition and subtraction differ from multiplication and division significant digit assignment. Rules for rounding numbers should also be presented in the same section and an explanation to why the last digit taken in a measurement is questionable or insignificant. Exponential notation and scientific notation and the difference between them will be explored following significant figures. The student should then be exposed to the metric systems. Presently there are three accepted metric systems; (1.) the Meter, Second, Kelvin, Ampere (MSKA) used in physics and engineering, (2.) the meter, second, gram, calorie (msgc) metric used in chemistry, and (3.) the International system(SI) accepted throughout the world. It should be explained that this system which is based on the unit of ten is much easier to perform mathematical computations and that the United States is the only country which is not promoting the conversion at an early educational level. A few examples of the differences between these systems are (1.) for volume; the

liter versus the cubic meter is used in the msgc metric versus the SI and MSKA; and (2.) for the unit of energy, the calorie versus the joule differentiates the msgc metric versus the SI and MSKA systems. Since these students are learning new concepts in mathematics, a short review in the usage of calculators as applied to chemical calculations and equations should be presented. The everyday functions such as addition, subtraction, multiplication, and division need not be explained. The use with exponential notation on the calculator, however, does need to be explained. It must be stressed that calculators purchased must have an EE or Exp (exponential expression) button on it, along with functions such as log, square, and square root. Any 'scientific' calculator would be appropriate for this course. However, the students should be presented the mathematical concepts that relate to the operations so that they do not rely solely on the calculator. The chapter should close with the introduction of unit conversion or dimensional analysis of English/English, English/metric, and metric/metric. For example, those students who are pursuing a career in the health field must be able to convert from the English system to metric and vice versa; for instance, the reported daily recommended percentages found on the reverse side of product labels are in metric units. The mathematical conversions should be broached with a basic application so that students

may not feel overcome. An example could be; How many dollars are there in five quarters?

These first two chapters are not what students would consider 'real' chemistry. Students may wish to see chemical reactions take place before understanding theories behind what constitutes chemistry. A basis for chemical definitions is presented and the foundation for all of the mathematical calculations to be used throughout the remainder of this curricula in factor analysis should be completely described and given many illustrations.

Chapter One: Introduction to Chemistry

1.1 Introduction

Look around you, just where you are sitting, and what do you see? Are there any chemicals around you? Try to name something that you would not consider to be classified as a chemical. Chemicals are all around, everything can be broken down into some form of chemical substituent. Our bodies have a vast amount of water in them. You may not consider water as a chemical; however, one molecule of water is composed of two atoms of hydrogen to every one atom of oxygen. Books, desks, pens, paints, soaps, household cleaners, and clothes are only a few items that consist of chemical compounds. Chemicals can be synthetic or manmade such as the leisure suits manufactured in the 1970's which were composed of a common substance known as polyester. Every man that owned one of these suits knew the term polyester; but did they know it was a chemical manufactured for the purpose of producing clothing that was durable and stylish? Polyester is a polymer of ester which has been constructed in the laboratory. The ester is a chemical, which may have many unique properties, which constitutes the monomer or single unit which is bonded to itself to form the

polymer of polyester. Other polymers you may not have been aware of are the HDPE and LDPE, these letters denote 'high' and 'low' density polyethylene. These polymers are plastics. Plastic can be found just about everywhere from soda bottles to children's play-yards.

Bill Nye, the Science Guy, and Beakman's World, although very entertaining and not just written television programs for children I might add, can teach us many different insights to the process of how we can learn about science.

The term chemical has been used many times in the opening paragraph. Do you think that all chemicals are harmful, corrosive, or even toxic? After being presented the above material, your answer to this posed question should be 'no'. However, many chemicals when exposed to above the determined PEL, or permissible exposure limit determined by a government agency OSHA (Occupational Safety and Hazard Agency), can be hazardous to health. But, millions of chemicals we deal with everyday of our lives pose no threat to our well being.

A firm foundation in chemistry can be useful to just about everyone; since, chemistry occurs around us all of the

time. Chemistry can play a key role in assisting doctors, lawyers, business people, paramedics, and firefighters in making key decisions. Doctors and paramedics need to understand how certain chemicals effect our biological systems in emergencies or simply in prescribing medications. Lawyers need to be insightful in this field by understanding the results determined in forensic testing or DNA testing. Firefighters need to know about various chemical reactions in order to combat fires correctly; certain chemicals when on fire should never be extinguished with water. Business people are involved with marketing, both in buying and selling, of chemicals. As consumers, we also should be aware of the chemicals we purchase to use. These are only a few examples as to the numerous professions that rely on a basic understanding of chemistry. Although you may not choose a field which directly relates to chemistry, an introduction to the fundamentals of chemistry will greatly enhance your life.

While studying a course in chemistry, you will inadvertently become a much better problem solver. Many students feel that a course in chemistry is impossible to endure due to the level of perceived difficulty. You need not fall into that trap. This text will illustrate ' how to' solve various types of mathematical functions associated

with chemical equations by employing a systematic and logical approach.

Many companies recruit individuals who are problem solvers and a course in chemistry will strengthen your power of reasoning. Keep this in mind as you continue this course. This text is written to present what and how chemistry works and is aimed to anyone who is willing, ready, patient, and interested in this field of science.

1.2 Science and Chemistry

According to Webster's dictionary (1988), science is defined as " knowledge reduced to a system; the facts pertaining to any department of mind or matter in their due connections; or a skill resulting from training". Science in a less broad a sense is an organized or systematized knowledge gained by an application of the 'scientific' method of researching and collection of data to define and resolve a problem.

Chemistry is the science which deals with the composition of structure and the properties of substances and the changes they undergo when reacting to each other. "Substance" in the preceding definition can be replaced by

the term matter. Matter is anything which has mass and occupies space. Mass is a quantity of matter in a particular body or object. These two definitions may seem a little ambiguous. However, the composition of a substance can be identified by experimental methods and the quantity of each of the components determined. Once the composition of matter is defined, it's structure can be hypothesized. This would be considered the physical arrangement of all of the atoms in a particular substance. Lastly, the property of matter can either be physical or chemical. Physical properties range from color, odor, taste, solubility, density, melting point and boiling point. You can think of a physical property much like the physical appearance of a person's height, weight, and eye color. A chemical property can only be observed when matter undergoes change in its composition. Examples of this property are paper burning, food digesting, or the metal on a car rusting.

Chemistry is an integral part of science. It's study supports other fields of science such as biology, botany, geology, physics, and zoology. In many instances, chemists will work with biologists or members of other sciences on projects.

Within the field of chemistry, there have been traditionally five subfields of study. These are analytical, biochemical, inorganic, organic, and physical chemistries.

Analytical chemistry is concerned with the 'what and how much' of a trace quantity can be found in a sample. This is to say that an analytical chemist will perform qualitative and quantitative tests on samples. Examples of what an analytical chemist would be testing for are: what contaminants and in what quantity these contaminants can be found in drinking water or what pollutant levels are emitted from a motor vehicle into the air.

Biochemistry is by far a most predominant subfield of chemistry today. It is the study of chemical reactions in living systems which range from botany, to zoology, to the human body. Biochemists are responsible for the discovery of many structures that play key roles in our biological systems. It is this field that proposed the structures of DNA, the genetic code, proteins, and enzymes which are the catalysts that aid chemical reactions to take place in our body. Due to the advances in the field, cures for many diseases have been found and a number of vaccinations developed to maintain an individuals health, such as the DTP

(Diphtheria, Tetanus, Pertussis) and MMR (Measles, Mumps, Rubella). The study of how the digestion of food takes place beginning in the mouth and ending in the colon is also a discovery of biochemistry.

Organic chemistry is the study of properties and reactions of compounds which contain the element of carbon. This field of chemistry was originally termed as the chemistry of living things. However, living organisms do contain the element of carbon but many nonliving things also contain carbon. Biochemistry and organic chemistry share many chemical reactions and properties; there seems to be a commonality between the two fields. Many polymers are organic, they contain carbon, but are obviously not alive. The synthesis of pharmaceutical drugs, petroleum, polymers, and many kinds of rubber are examples of what an organic chemist may do.

Inorganic chemistry is the subfield of chemistry concerned with the study of anything which does not contain the element of carbon. Although at first glance everything may seem to be organic or biochemical, many chemicals are inorganic. For example, the chemical which allows a car battery to work is inorganic. That acid is known as

sulfuric acid. Electrochemistry or the flow of electricity is considered inorganic chemistry.

Physical chemistry examines the physics of chemical change. It studies the structures of substances, how fast substances may change, kinetics, and the role that heat plays in the reaction, thermodynamics. A physical chemist would study energy involved in the changes in the physical states between gas, liquid, and solid.

A few other subfields of chemistry are (1.) nuclear chemistry, the study of energy production, (2.) geochemistry, study of the structure of the earth as to its rocks, strata, soil, minerals, organic remains and changes, and (3.) environmental chemistry, the study of how products from chemical manufacturers effect our earth.

There is not always a distinct line between these subfields of chemistry, many times they interact and overlap with each other. If in a drinking water sample trace amounts of pesticide were detected, three areas of chemical study may be involved. The analytical chemist would find the contamination and determine the exact concentration of the contaminants. If the element of carbon were involved, an organic chemist would determine its structure and a

biochemist would study possible side effects or carcinogenicity to living matter.

Chemists are in demand for employment opportunities such as in research and development, quality assurance and quality control, synthesis supervision, sales, and more. The petroleum industry is a large employer of chemists; this field studies organic substances which compose fossil fuels. Environmental chemists study the products and byproducts of industry or manufacturers to determine their effects on our environment. An example would be to assess the effects of CFC's on the upper atmosphere. CFC's stand for chlorofluorocarbons; this chemical destroys the ozone in the atmosphere exposing the earth to high energy radiation. This chemical was used in freon for refrigeration units and packing materials until its hazards were determined. Now, environmentally friendly chemicals manufactured under trade names such as Oz-12, Hc -12 and R-12 replace CFC's.

Table 1.1 The Five Subfields of Chemistry

<u>Subfield of Chemistry</u>	<u>Study</u>
Analytical	Qualitative and quantitative study of samples.
Biochemical	Study of chemical reactions in living matter.
Organic	Study of substances containing carbon.
Inorganic	Study of compounds that do not contain carbon.
Physical	Study of kinetics and thermodynamics of substances.

1.3 The Scientific Method

Previously in the chapter, science was defined as an organized and systematized knowledge gathered by the scientific method. All scientists follow this method when performing an experiment or simply solving a problem.

For example, the ice cream parlor is approximately 2 miles from my house. I would like to determine the fastest route to it. If I drive one way to the ice cream parlor, I need to drive through two traffic lights and turn onto a

main highway. The other route would be to drive through a housing development where the speed limit is 25 miles per hour. I propose that the second route would be faster. Therefore to determine which route is quicker, I drive both ways several times so that I can take an average of both miles and time. Route one took seven minutes and is 2.2 miles from my house. Route two took twelve minutes and is 1.7 miles from my house. While examining the data although route two is closer to my house, it is not the fastest route to the ice cream parlor. My fastest route would actually be the further distance. If, however, I was in dire need of chocolate ice cream and was very low on gasoline in the car, I would have to drive the slightly longer route because the route through the development would take a longer time, therefore, the engine would be consuming more gasoline. I stated a problem to be solved, hypothesized which route is quicker, and collected data. My data did support my initial postulate, so I need not reformulate my hypothesis or collect more data along with more ice cream sundaes.

The scientific method involves three steps:

(1.) Experimentation: a collection of data or facts by observing events under carefully planned conditions.

(2.) Hypothesis formation: after examining data collected by the experiment, a tentative explanation is proposed. A hypothesis is subject to verification or rejection.

(3.) Further experimentation: more experimentation needs to be performed in order to support the hypothesis, reformulation of the hypothesis, or rejection of the hypothesis. If the hypothesis is verified by many experimenters under the same conditions, then a scientific theory can be stated. If the theory proves true over a given time period, it then becomes a scientific law.

All experiments result in some form of hypothesis formation. Very few are verified to theories and even fewer result in scientific law. Hypothesis' can be modified or even discarded after further experimentation. However for a law to be postulated, the experiment must be reproducible by many other scientists under identical laboratory conditions.

Figure 1.1
The Scientific Method

Experimentation

A collection of data
under carefully
controlled conditions.

Hypothesis Formation

A proposal of a possible
explanation for an observed
event by correlating data.

Further Experimentation

Re-experimentation under
identical laboratory
conditions.

**Reformulate or Reject
Hypothesis**

If results do not support
original hypothesis,
reformulate or reject
initial hypothesis.

Extensive Experimentation

Reproducible results give
rise to theory
and law.

Chapter Two: Measurements

2.1 Introduction

In order to truly understand chemistry, you need to feel comfortable and confident with computational mathematics. Although this chapter presents the proper usage of calculators, you must still understand the basic concepts of mathematics which lie behind the calculation.

This chapter presents the basic mathematical principles, thereby laying a foundation for chemical concepts and calculations applied to chemistry presented in future chapters. This chapter will also include measurement and its uncertainty, precision, accuracy, significant figures and rules governing their assignment for use in mathematical operations, rounding numbers, exponential notation, scientific notation, the metric systems, and dimensional analysis.

Once you have mastered the above skills, try some of the 'end of the chapter' questions for practice.

2.2 Uncertainty in Measurement

There is nothing new about taking measurements, you do them everyday without being aware of it. How much gas did you put in the car? How many miles did you drive to get to class today? How much did you eat for breakfast? What is your height or weight?

A chemist observes, measures, and records events which occur during an experiment. Measurements are obtained in the laboratory by the use of scientific instruments. Whenever measurements are performed with an instrument, whether with a scale or a more sophisticated instrument such as a balance, some sort of estimate is required. It is not possible to make an exact measurement because there are no instruments that can be calibrated exactly. All instruments have a certain amount of uncertainty and reported values should account for this by a reported value followed by a +/- value; for instance 6 +/- 0.3 inches. The uncertainty of a measurement depends on the instrument itself and the environment in which the measurement is made, such as temperature or humidity, which effect the accuracy of the instrument. For instance, in measuring the length of a pen with a scale which has divisions indicating tenths versus a ruler which indicates hundredths of a centimeter, the

measurement taken with the ruler that has smaller divisions is more accurate because you can determine the smaller units of the length and do not have to estimate the smaller unit of measure as with the scale with the larger divisions.

Measurements need to be accurately and precisely performed. These two terms are used synonymously in the nonscientific world; however, they do have different definitions and meaning in the scientific world. Precision refers to consistency of measurements, for instance the weight of a man is taken three times consecutively to be 195 pounds, 196 pounds, and 195 pounds. These measurements are said to be precise because they are reproducible within an acceptable degree of tolerance of the +/- value. Accuracy refers to the proximity of repeated measurements to the actual accepted value. An example of this would be an experimental value for the density of water at 20 degrees Celsius to be 0.9998 g/cm^3 , the accepted value is 1.0000 g/cm^3 . Usually high precision will give rise to high accuracy. However, this may not always hold true. It is possible to have high precision but low accuracy. This could arise by the use of poorly calibrated instruments. Although the measurements are easily repeatable, they would not be accurate if taken on an uncalibrated or poorly calibrated instrument.

Different instruments measure different properties of a substance. Typically chemists measure mass in terms of grams, volume in liters, and length in meters (g, l, and m). You probably have used a measuring cup or teaspoon in cooking or taking oral medications. Chemists use a variety of instruments to measure the volume of liquids; a few examples are graduated cylinders, pipets, burets, and beakers. Volume is defined as the amount of space occupied by a solid, liquid, or gas. The most often used graduated cylinders are 10, 25, 50, and 100 milliliter capacity. The accuracy ranges from ± 0.1 to 0.5 milliliters (ml). Pipets can either be classified as to contain "TC" or to deliver "TD". TC pipets have one single etched line above the glass bulb and a liquid is drawn to that calibration line. The TD pipets have divisions much like a graduated cylinder on them. Pipets range in volume from as low as 0.50 ml to as large as 100 ml. The accuracy of pipets ranges from ± 0.01 to 0.1 ml. The smaller the pipet, the smaller the uncertainty would be. A buret is a long narrow piece of calibrated glass tubing. The 50 ml buret is the most commonly used. A value is read from the top of the buret before the liquid has been dispensed and then a final reading is taken after the liquid has been released. The difference between the two readings would be the amount of

fluid drained from the buret. This value is very important and calculations will be done with these values later in this text. Uncertainty ranges from +/- 0.01 to 0.1 ml.

Length is measured by a scale. You have performed many measurements in the past with scales; however, the scales most likely have been calibrated in inches or feet. Chemists measure length in centimeters or meters. Later in this chapter, the conversion between inches and centimeters will be presented.

The mass, as presented in Chapter 1, of an object is a measure of the amount of matter it possesses. Balances measure mass and are designed to null the effect of gravitational pull, which will be discussed shortly. There are three different designs of balances used in the chemical laboratory; (1.) a platform balance in which the object is placed on one pan and you physically place objects in the opposite pan to counterbalance, this balance appears to be a teeter-totter, (2.) a triple beam balance in which the object is placed in the pan and you move counter masses until the pan is in a level position, and (3.) an electronic balance in which the object is placed on the pan and the mass reading will appear on the digital display. These

balances can measure mass as low as 0.0001 grams. The accuracy ranges from +/- 0.0001 to 0.1 grams.

Unfortunately, the terms weight and mass are used interchangeably. However, they should not be and it is important that you understand the difference. Weight is defined as the gravitational force of attraction between an objects mass and the mass of the planet or other object. The mass of an object remains constant from location to location; whereas, an object's weight will change according to location, due to the change in gravitational force. A person who weighs 125 pounds on the earth will weigh 21 pounds on the moon because the moon has less of a gravitational force, due to its smaller mass than that of the earth. Spring scales are typically used to measure the weight of an object. You may have even used one last time you bought produce in the grocery market. In chemistry, the mass of an object is the unit used not its weight.

The smaller the uncertainty, the more accurate the measurement. This holds true regardless of the instrument to be used whether measuring mass, length, or volume.

(This is a note for the presenter of this information: It would be most beneficial for the students to see what the

above instruments physically look like while explaining them and demonstrating this use or application.)

2.3 Significant Digits

As discussed in the previous section, all measurements involve an estimate. Therefore, this gives rise to some sort of uncertainty. Significant digits of significant figures takes into account this uncertainty. The number of significant figures in a quantity is the number of digits that are known accurately plus the digit in question. The questionable digit is always the last digit recorded or reported. For example, a man weighs 194 pounds, the questionable or doubtful digit in this number is 4 because the man could weigh 193 or 195 pounds.

To determine the number of significant digits in a measurement, we must follow certain rules.

(1.) Nonzero digits: 1,2,3,4,5,6,7,8, and 9 are always significant.

1.2	2 significant digits
6.45	3 significant digits
1234.5	5 significant digits

(2.) Leading zeros: Zeros that appear at the beginning of a number are never significant because they serve only to place the position of the decimal point in a number less than one.

0.346 3 significant digits
 0.00378 3 significant digits
 0.0001 1 significant digit

(3.) Confined zeros: Zeros that are sandwiched between two significant digits are always significant.

108 3 significant digits
 10.08 4 significant digits
 2.0004 5 significant digits

(4.) Trailing zeros: Zeros at the end of a number are significant only if the number has a decimal point or contains an overbar.

1500 2 significant digits
 150.0 4 significant digits
 150 3 significant digits

(5.) Significant figures do not apply to exact numbers. An exact number has no uncertainty. For example, there are 60 seconds in 1 hour and 12 eggs in 1 dozen. These are exact by definition as $2.54 \text{ centimeters} = 1 \text{ inch}$.

Example 2.1: Determine the number of significant digits in the following numbers.

<u>Number</u>	<u>Answer</u>	<u>Number</u>	<u>Answer</u>
a. 16	2	f. 0.0600	3
b. 10.01	4	g. 13,460	4
c. 1000	1	h. 3121.00	6
d. 17.60	4	i. 3121.00360	9
e. 0.0056	2	j. 1.0	2

2.4 Rounding off of Nonsignificant Digits

Although all digits taken from instrumental readings are considered significant, mathematical calculations performed with them give rise to both significant and nonsignificant numbers. Nonsignificant digits may exceed the data in the measurement. For instance, in performing a calculation on the calculator, the number of digits displayed is usually greater than the number of significant figures that the result should contain. Therefore, the number needs to be rounded off. You may have already learned this method of rounding off numbers.

Rules for rounding are as follows:

- (1.) If the first number to be dropped is less than five, leave the digit before it unchanged. Thus, 6.92 is rounded to 6.9 for two significant digits, the two is dropped and the nine unchanged.
- (2.) If the first digit to be dropped is five followed by numbers other than zero or is greater than five, increase the last digit to be retained by one. Therefore, both 6.51 and 6.6 round to 7 for one significant digit.

- (3.) If the first nonsignificant digit is five and followed by zeros, drop the five and increase the last digit to be retained by one, if it is odd. If the number is even, leave it unchanged. Hence, 16.50 will round to 16 for two significant digits and 0.350 will round to 0.4 for one significant digit.
- (4.) Nonsignificant digits to the left of the decimal point are not to be discarded but replaced by zeros. Thus, 1546 will round to 1500 for two significant digits.

Example 2.2: Round the following numbers to three significant digits.

<u>Number</u>	<u>Answer</u>	<u>Number</u>	<u>Answer</u>
a. 123.6	124	f. 0.072650	0.0727
b. 435.4	435	g. 0.27056	0.271
c. 447.50	448	h. 93,425	93,400
d. 868.50	869	I. 0.00687	0.00687
e. 1284	1280	j. 1.0034	1.00

2.5 Mathematical Operations Involving Significant Figures

Rounding aids the experimenter in calculating the 'correct' number of significant digits. However, what is meant by the correct number? This question depends on which mathematical operation is being performed.

Rules for Addition and Subtraction:

When adding or subtracting numbers, the answer is limited by the measurement with the most uncertainty. In other words, the number with the least number of decimal places limits how the answer will be reported. The answer must not contain a smaller place unit than the number with the smallest place unit. For example, $5 + 5.0 + 5.00 = 15.00$. However, the first number being added has the least number of decimal places; in fact it has no decimal places, therefore, the answer is reported as 15. Thus, $16.73 - 1.6 = 15.13$, but is reported as 15.1 due to the 1.6 limiting the decimal place unit.

Rules for Multiplication and Division:

In multiplication and division, the answer must not contain any more significant digits than the least number of significant figures in any of the terms used. For example, $25.0 * 6.253 = 156.325$, but is recorded as 156 because the numbers being multiplied contain three and four significant digits. Therefore, the answer must be reported with three significant figures. Also, $1.3/6.24 = 0.20835$ on the calculator but is reported as 0.21 for the reason that 1.3 has two significant digits and is rounded off.

Order of Mathematical Operations:

Many times a calculation may contain more than one mathematical operation to be performed. But, which operation is performed first? When this occurs, there is a list which prioritizes the order of operations. Anything in parentheses is performed first followed by resolving the exponents, then multiplication or division, and ending with addition or subtraction of the remaining values. This can easily be remembered by PEMDAS, the old adage stands for Please Excuse My Dear Aunt Sally. The underlined letters stand for the order of operations.

Example 2.3: Perform the indicated Mathematical operations and express the answer in the correct number of significant digits.

<u>Operation</u>	<u>Answer</u>
a. $16.1 + 5.682 + 4.7$	26.5
b. $5.63 - 2.555$	3.08
c. $62.1 * 0.346$	21.5
d. $13.65 * 2.36$	5.78
e. $1000.65 + 13$	1014
f. $(1.35 + 6.3) / 16.54$	0.5
g. $181.8 / 75$	2.4
h. $0.345 - 0.46$	0.16
I. $94.3 * 5.6$	530
j. $(1.34 * 1.6) / 1.354$	1.6

2.6: Exponential Notation and Scientific Notation

Chemists use very large and extremely small numbers which are associated with scientific measurements. For instance, the distance between the sun and the earth is on the order of 93,000,000 miles. Writing these large or small numbers is cumbersome and tiring not to mention space consuming. Exponential and scientific notation are methods for making these numbers more compact, therefore, easier to write and perform calculations.

When a value is multiplied by itself, the process is indicated by a number written as a superscript. Therefore, an exponent is a whole number or symbol written as a superscript to another whole number or symbol, the base. The exponent actually indicates that the base is multiplied by itself the number of times equal to the superscript. Thus, $6^3 = 6 * 6 * 6 = 216$ and $2^4 = 2 * 2 * 2 * 2 = 16$. Likewise, very small numbers can be represented by negative exponents. For instance, $6^{-3} = 1/6 * 1/6 * 1/6 = 1/216$ and $2^{-4} = 1/2 * 1/2 * 1/2 * 1/2 = 1/16$. The negative exponent indicates that the base is inverted and multiplied the exponents number of times.

When ten is raised to an exponential power, it is known as a power of ten. This takes the general form of 10^n , where 10 is the base and n is the exponent. Therefore, ten to the n^{th} power is equal to ten multiplied by itself n times. Thus $10^3 = 10 * 10 * 10 = 1000$. Chemists and scientists use both positive and negative exponents to write large and small numbers. This form is referred to as exponential notation which contains two portions, one number written as a decimal and the other as a power of ten. Some examples are as follows: $1,600,000 = 1.6 * 10^6$ and $0.00061 = 6.1 * 10^{-4}$. Where the 1.6 in the $1.6 * 10^6$ and the $6.1 * 10^{-4}$ are called the mantissa.

The powers of ten that you must become familiar with are:

$$10^4 = 10000$$

$$10^3 = 1000$$

$$10^2 = 100$$

$$10^1 = 10$$

$$10^0 = 1$$

$$10^{-1} = 0.1$$

$$10^{-2} = 0.01$$

$$10^{-3} = 0.001$$

$$10^{-4} = 0.0001$$

Helpful Hints:

(1.) A positive exponent means that the number is larger than one. A negative exponent indicates that the number is less than one.

(2.) The exponent is equivalent to the number of places the decimal place is moved. Shifting the decimal to the left

will result in a positive exponent and shifting the decimal to the right will result in the exponent being negative.

Example 2.4 : Express the following numbers in exponential notation.

<u>Number</u>	<u>Answer</u>
a. 0.0034	$3.4 * 10^{-3}$
b. 34,000	$3.4 * 10^4$
c. 16,000,000	$1.6 * 10^7$
d. 0.000000078	$7.8 * 10^{-8}$
e. 0.3456	$3.456 * 10^{-1}$

The difference between scientific notation and exponential notation is very slight. In exponential notation, the relevance of where the decimal point is placed does not matter. In scientific notation, the decimal point must be placed immediately to the right of the first nonzero number. Thus, $41.0 * 10^4$ written in scientific notation equals $4.10 * 10^5$. Take note that when you move the decimal point in a number already expressed in exponential notation, you must change the exponent also.

Summary of Procedures for Writing Scientific Notation:

- (1.) Place the decimal immediately after the first nonzero number and write $* 10$.
 $5,600,000 = 5.6 * 10$
- (2.) Count the number of places you moved the decimal point in the original number and make that the exponent. If

you moved the decimal to the left, the exponent will be positive. However, if you move the decimal to the right, the exponent will be negative.

$$5,600,000 = 5.6 * 10^6 \text{ and } 0.00654 = 6.54 * 10^{-3}$$

Example 2.5 : Convert the following numbers into scientific notation.

<u>Number</u>	<u>Answer</u>
a. 63.4 * 10 ⁻³	6.34 * 10 ⁻²
b. 0.000532	5.32 * 10 ⁻⁴
c. 651,000,000	6.51 * 10 ⁸
d. 1300	1.3 * 10 ³
e. 0.000000000012	1.2 * 10 ⁻¹¹

2.7: Calculators and Scientific Notation

In order to make otherwise tedious calculations quickly, You must purchase a scientific calculator. In selecting a calculator, choose one that has an exponent key, EE or Exp, a log key, and a +/- toggle key. Most calculators do not display exponents as a power of ten. Therefore, a special notation is available to write an exponential number from a calculator display. In order to enter $3.31 * 10^5$ into the calculator, 3.31 is entered using the normal number keys followed by the EE or Exp key and enter 5. It will appear on the display screen as 3.31 EE 5. A number such as $1.23 * 10^{-3}$ can be entered on the calculator as 1.23 followed by the EE button and -3. The

key which has the +/- signs is the toggle switch which will change the sign of the number entered.

The following examples illustrate how to perform mathematical calculations on the calculator.

<u>Expression</u>	<u>How to input</u>	<u>Answer</u>
$(3.46 * 10^4) (1.54 * 10^6)$	3.46EE4 * 1.54EE6	$5.33 * 10^{10}$
$(7.73 * 10^{-3}) (9.11 * 10^2)$	7.73EE-3 * 9.11EE2	$7.04 * 10^3$
$(8.94 * 10^6) / (4.35 * 10^4)$	8.94EE6 / 4.35EE4	$2.06 * 10^2$
$(5.08 * 10^{-3}) / (7.23 * 10^{-9})$	5.08EE-3 / 7.23EE-9	$7.03 * 10^5$

There is a newer method for writing scientific notation which has recently been proposed. However, this method has not been adopted by the scientific community. Instead of writing $2.06 * 10^2$, the new notation would be 2.06 p 2. The lower case p stands for positive, the exponent is positive in the above example. If the exponent were negative, a lower case n would be written. This text will continue to use the conventional method in writing scientific notation.

Example 2.6: Perform the following calculations and report the answers in the proper number of significant figures.

<u>Operation</u>	<u>Answer</u>
a. $2.75 * 10^3 + 3.2 * 10^2$	$3.1 * 10^3$
b. $(4.72 * 10^5) (1.83 * 10^{-2})$	$8.64 * 10^3$
c. $(2.40 * 10^5) / (1.30 * 10^{-3})$	$1.85 * 10^8$
d. $4.20 * 10^{-3} - 0.12 * 10^{-3}$	$4.1 * 10^{-3}$
e. $1 * 10^{16} + 3 * 10^4$	$1 * 10^{16}$

2.8 : The Metric Systems

Units are labels that describe something that is being measured or counted. Units can be anything; 3 dozen eggs, 2 quarts, 9 pints, 6 pounds, etc. However from this point on in the text, a unit will always accompany a number. The number 2 is meaningless without a definition of 2 what? You must also follow this rule in order to avoid ambiguity.

Two formal systems of units of measure are in use in the United States today: (1.) English based or Customary system of units and (2.) the metric system of units. In addition physicists and engineers use the Meter, Second, Kilogram, Ampere (MSKA) system, while scientists employ a system of units known as the International System of Units (SI) from the French Systeme International. Both of the aforementioned systems are based on the metric system. The English based system contains units such as inches, feet, pounds, quarts, and gallons. This system is primarily used in the nonscientific community and in some aspects of commerce in the United States. The metric system, developed in the nineteenth century, is used in the scientific world and throughout the rest of the world. Units in this system include grams, meters, and liters.

The United States is now in the process of a voluntary conversion to the metric system. In fact, many metric system units can be found on consumer products. When was the last time you bought a two liter bottle of soda? Did it even occur to you that the liter is metric? Boxes of cereal now have masses listed in ounces and grams.

The metric system is simpler and more coherent than the English based system for several reasons; (1.) it uses a single base unit for each quantity measured, and (2.) conversion from one unit size into another can be accomplished by moving a decimal point to the left or right. Therefore, this system is based on the unit of ten. The base unit for mass is the gram; the base unit for volume is the liter; and the base unit for length is the meter.

The metric system was updated in 1960 by International agreement. A revision of the traditional metric system changed to the SI system in the sciences. A less known subdivision of the metric system mentioned briefly in the beginning of this section is the MKSA used primarily by engineers and physicists. The differences between these three systems are minor, as shown in Table 2.1. This text will present material in the traditional metric system switching to SI units when appropriate.

Table 2.1

<u>Quantity measured</u>	<u>English</u>	<u>Traditional</u>	<u>SI</u>	<u>MSKA</u>
length	inch, feet, yard	centimeter (cm)	meter (m)	(m)
mass	pound, ounce	gram (g)	kilogram (kg)	kg
volume	pint, quart	liter (l)	(l)	(l)
time	second (s)	(s)	(s)	(s)
electric current	volt	ampere (A)	(A)	(A)
temperature	Fahrenheit	Celsius (C)	kelvin (K)	(K)
amount of substance		mole	mole	
intensity	lux		candela	

Now that you have been introduced to the three main base units, gram, liter, and meter, the metric system uses prefixes to express a multiple or fraction of the base unit. All prefixes are related to the power of ten; therefore, the metric system is a decimal system. The prefix increases or decreases the base unit by powers of ten. The kilometer (km) is 1000 times larger than the meter; whereas, the decimeter is one-tenth the size of the meter. Keep in mind that the base unit can be any of the units listed in Table 2.1. The base units most commonly used in chemistry are liter, meter, and gram. Table 2.2 lists all of the prefixes that you must learn in order to work with problems.

Table 2.2 Metric Prefixes

<u>Large Prefixes</u>		
<u>prefix</u>	<u>symbol</u>	<u>multiple</u>
tera-	T	10^{12}
giga-	G	10^9
mega-	M	10^6
kilo-	k	10^3
hecto-	h	10^2
deca-	da	10^1
base unit		10^0

<u>Small Prefixes</u>		
<u>prefix</u>	<u>symbol</u>	<u>multiple</u>
bas unit		10^0
deci-	d	10^{-1}
centi-	c	10^{-2}
milli-	m	10^{-3}
micro-	u	10^{-6}
nano-	n	10^{-9}
pico-	p	10^{-12}

2.9: Conversion Factors and Dimensional Analysis

Many times a need arises to convert units of one quantity of measurement into another quantity with different units, for example volume to mass (liter to kilogram). The new units can either be in the same system (metric or customary) or a different system, from metric to customary or customary to metric. For example, since the United States presently uses both the English based system and the metric system, we must be able to relate and convert measurements from one system to the other. For instance, a typical piece of paper measures 8.5 inches by 11.5 inches;

in the metric system the same piece of paper measures 21.59 centimeters by 27.94 centimeters.

The mathematical operation employed in the above example is known as dimensional analysis or unit analysis. It is a problem solving method used in conversion factoring. One inch is equivalent to 2.54 centimeters. A conversion factor is simply a ratio with a value that specifies a relationship between two units. How many days are in four weeks? You can quickly answer 28 days. In answering this question you employed unit analysis. You knew that one week is equivalent to seven days. You took the 7 days and multiplied by four weeks. Some English/English based and English/metric based unit conversion factors can be found in Tables 2.3 and 2.4.

Table 2.3: English/English conversion factors

<u>Mass</u>	<u>Length</u>	<u>Volume</u>
16 ounces=1 pound	12 inches=1 foot	16 ounces=1 pint
2000 pounds=1 ton	3 feet=1 yard	2 pints= 1 quart
	5280 feet= 1 mile	4 quarts= 1 gallon

Table 2.4: English/metric conversion factors

<u>Mass</u>	<u>Length</u>	<u>Volume</u>
1 pound=454 grams	1 inch=2.54 cm	1.06 quarts=1 liter

The following steps show how to set up a dimensional analysis problem.

(1.) Read the problem carefully. Identify and write all known quantities, both the numerical value and units; followed by the unknown units in the solution to the problem. Set up the problem by placing the known quantities on the left of an equal sign and unknown quantity on the right. In this method, units are treated in the same manner as numbers in that they can be multiplied, divided, or canceled.

(2.) Apply one or more unit factors to convert the units of the given value to the unit of the unknown quantity.

(3.) Multiply the given quantity by the conversion factors so that the original unit cancels out leaving behind the desired unit.

(4.) Check the answer to be sure that both the number and unit 'make sense' and be careful about significant digits.

The following problems use dimensional analysis using the steps outlined above. The first problem involves an English/English unit based conversion, followed by two English/metric unit conversions, and ending with two metric/metric unit based conversions.

Example 2.7: How many feet are in 42 inches?

step (1.) Identify the given quantity and unknown quantity.
42 inches and feet?

step (2.) Apply conversion factor.
12 inches = 1 foot

step (3.) Set up problem.

$$42 \text{ inches} * \frac{1 \text{ foot}}{12 \text{ inches}} = \frac{42 \text{ foot}}{12} = 3.5 \text{ feet}$$

step (4.) Check numerical answer and significant digits.

The inches unit cancels out and the proper number of significant digits should be reported as two.

Example 2.8: A piece of material has the length of 5.21 inches, how many centimeters of material are there?

step (1.) Identify the known quantity and unknown quantity.
5.21 inches and centimeters?

step (2.) Apply conversion factor.
1 inch = 2.54 centimeters

step (3.) Set up the problem and solve.

$$5.21 \text{ inches} * \frac{2.54 \text{ cm}}{1 \text{ inch}} = 13.2 \text{ cm}$$

step (4.) Check mathematical answer and significant figures.

Example 2.9: How many liters are in 2 quarts of milk?

step (1.) Identify known and unknown values.
2 quarts and liters?

step (2.) Apply conversion factors.
1.06 quarts = 1 liter

step (3.) Set up problem and solve.

$$2 \text{ quarts} * \frac{1 \text{ liter}}{1.06 \text{ quarts}} = 1.89 \text{ liters}$$

step (4.) Check numerical value and significant figures.

Example 2.10: Convert 1.64 kilometers to meters.

step (1.) Identify known and unknown quantities.
1.64 kilometers and meters?

step (2.) Apply conversion factors.
1 kilometer = 1000 meters

step (3.) Set up mathematical problem and multiply.

$$1.64 \text{ km} * \frac{1000 \text{ m}}{1 \text{ km}} = 1640 \text{ m}$$

step (4.) Check answer and significant digits.

Example 2.11: Convert $3.82 * 10^4$ milliliters to kiloliters.

step (1.) Identify known and unknown quantities.
 $3.82 * 10^4$ ml and kl?

step (2.) Apply conversion factors.
1000 ml = 1 l and 100 l = 1 kl

step (3.) Set up problem and multiply.

$$3.82 * 10^4 \text{ ml} * \frac{1 \text{ l}}{1000 \text{ ml}} * \frac{1 \text{ kl}}{1000 \text{ l}} = \frac{3.82 * 10^4 \text{ kl}}{1 * 10^6}$$

This expression can be inputted into the calculator as:
3.82 EE 4 / 1 EE 6

The final answer equals $3.82 * 10^{-2}$ kl.

step (4.) Check final answer and significant digits.

The only way to learn how to solve problems is to practice solving them for yourself. A few more examples are illustrated with individual steps omitted and final answers

supplied. Please try the problem first and then check your answer along with the rational for solving these.

Example 2.12: Convert 2.20 pounds to ounces.

$$2.20 \text{ lbs} \quad * \quad \frac{16 \text{ oz}}{1 \text{ lb}} \quad = \quad 35.2 \text{ oz}$$

Example 2.13: Convert 66.0 milliliters to gallons.

$$66.0 \text{ ml} \quad * \quad \frac{1 \text{ l}}{1000 \text{ ml}} \quad * \quad \frac{1.06 \text{ quart}}{1 \text{ l}} \quad * \quad \frac{1 \text{ gallon}}{4 \text{ quart}} \quad =$$

$$\frac{66.0 * 1 * 1.06 * 1 \text{ gallon}}{1000 * 1} \quad = \quad \frac{69.96}{4000} \text{ gallon} =$$

$$0.0175 \text{ gallons or } 1.75 * 10^{-2} \text{ gallons}$$

Example 2.14: Calculate how many decimeters are in $1.56 * 10^{-3}$ kilometers.

$$1.56 * 10^{-3} \text{ km} \quad * \quad \frac{1000 \text{ m}}{1 \text{ km}} \quad * \quad \frac{10 \text{ dm}}{1 \text{ m}} \quad =$$

$$\frac{1.56 * 10^{-2} * 1000 * 10 \text{ dm}}{1 * 1} =$$

$$15.6 \text{ dm or } 1.56 * 10^1 \text{ dm}$$

Example 2.15: Perform the following conversions:

<u>Operation</u>	<u>Answer</u>
a. 16 inches to feet	1.33 feet
b. 1.56 pounds to grams	708 grams
c. $2.3 * 10^{-6}$ kilometers to meters	$2.3 * 10^{-3}$ meters
d. $1.872 * 10^{14}$ milliliters to kilometers	$1.872 * 10^8$ kilometers
e. 164 seconds to milliseconds	$1.64 * 10^6$ milliseconds

The six basic steps to solve mathematical problems are as follows:

- (1.) Read the problem carefully; preferably at least two times.
- (2.) List all quantities that are given with units attached.
- (3.) Identify for what the problem is asking, including units.
- (4.) Determine the relationship between the given quantity and the unknown quantity. Decide how to set up the problem as to what the numerator and denominator will be.
- (5.) Write the calculation step with appropriate units and calculate the answer.
- (6.) Challenge the answer. Make sure that the number seems reasonable and that the units and significant digits are correct.

Chapter 2 Problems:

1. Determine the number of significant figures in the following.

- a. 2001 l
- b. 0.006 ml
- c. 1×10^{-3} cm
- d. 34,100,000 km
- e. 0.002007890 g

2. Round the following numbers to three significant numbers.

- a. 0.146823 mg
- b. 218,946.80 dl
- c. 13.801 cm
- d. 138.150 kg
- e. 224.59 kl

3. Perform the following calculations and report answer in proper significant digits.

- a. 22.10 mm - 10.5 mm
- b. 4 cg + 16.3 cg + 0.95 cg
- c. 12.3 km * 16.876 km
- d. 22.342 cl / 2 cl
- e. (13 kg + 2 kg) / 5.0 kg

4. Express the following in scientific notation.

- a. 1634 cm
- b. 3,703,000 l
- c. 0.0000000346 km
- d. 0.24000 g
- e. 6,800 ml

5. Express the following in decimal form.

- a. 1.6×10^{-2} cm
- b. 1.6×10^4 ml
- c. 3.46×10^{-1} dm
- d. 6.0×10^{-12} kg
- e. 2.2×10^{10} l

6. Calculate the following operations and report in scientific notation with proper amount of significant figures.

- a. $(4 \times 10^2 \text{ km}) (2 \times 10^{-4} \text{ km})$
 b. $\frac{3.62 \times 10^4 \text{ l}}{1.60 \times 10^{-3} \text{ l}}$
 c. $(4.78 \times 10^{-2} \text{ mg}) + (7.3654 \times 10^3 \text{ mg})$
 d. $(4.3 \times 10^{-3} \text{ g}) - (2.98 \times 10^{-2} \text{ g})$
 e. $\frac{(4.36 \times 10^8 \text{ km}) (1.82 \times 10^{-3} \text{ km})}{(0.0856 \text{ km}) (4.7 \times 10^{-6} \text{ km})}$

7. Convert the following.

- a. 2.5 pounds to ounces
 b. 0.500 quarts to milliliters
 c. 3.0×10^8 millimeters to centimeters
 d. 31.9 grams to centigrams
 e. 18,000 liters to kiloliters

Answers for end of chapter problems.

1. (a.) 4 l (b.) 1 ml (c.) 1 cm (d.) 3 km (e.) 7 g
 2. (a.) 0.147 mg (b.) 219,000.00 dl (c.) 13.8 cm
 (d.) 138 kg (e.) 225 kl
 3. (a.) 11.6 mm (b.) 21 cg (c.) 208 km² (d.) 10 (e.) 3
 4. (a.) $1.634 \times 10^3 \text{ cm}$ (b.) $3.703 \times 10^6 \text{ l}$ (c.) $3.46 \times 10^{-8} \text{ km}$
 (d.) $2.4000 \times 10^{-1} \text{ g}$ (e.) $6.8 \times 10^3 \text{ ml}$
 5. (a.) 0.016 cm (b.) 16,000 ml (c.) 0.346 dm
 (d.) 0.0000000000060 kg (e.) 22,000,000,000 l
 6. (a.) $8 \times 10^{-2} \text{ km}^2$ (b.) 2.26×10^7 (c.) $7.37 \times 10^3 \text{ mg}$
 (d.) $-2.5 \times 10^{-2} \text{ g}$ (e.) 2.0×10^{12}
 7. (a.) 40 oz (b.) 472 ml (c.) $3.0 \times 10^3 \text{ cm}$
 (d.) $3.19 \times 10^3 \text{ cg}$ (e.) $1.8 \times 10^1 \text{ kl}$

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