COURSE OUTLINE: Prepared by: Dr. Victor Huang

September 2016

COURSE TITLE: General Chemistry Principles I

COURSE CODE: CHM 152

CREDITS: 4

CONTACT HOURS: Lecture: 3 Laboratory: 3

CATALOG DESCRIPTION: The first part of a two-semester sequence in General Chemistry with laboratory. This course covers the qualitative and quantitative aspects of scientific measurement, the nature of matter, mole concept, gases, liquids and solids, energy, atomic theory, properties of elements, chemical bonding, molecular structure and properties, stoichiometry, thermochemistry and solutions.

PREREQUISITE: High School Algebra 2 (Regents level); CHM 124 (or equivalent) or Regents High School Chemistry or equivalent.

IMPORTANT NOTE: BOTH THEORY AND LABORATORY PARTS OF THIS COURSE MUST BE TAKEN CONCURRENTLY IN ORDER TO RECEIVE CREDIT.

REQUIRED FOR: Liberal Arts and Sciences, Pre-Health and Life Sciences, Bioscience, Computer Science, Medical Laboratory Technology and Manufacturing Engineering Technology.

GENERAL EDUCATION: This course satisfies 4 credits of the Natural Sciences competency area of the General Education requirements at Farmingdale State College.

ELECTIVE FOR: Liberal Arts and Sciences – Non-science majors

*Laboratory Manual for General Chemistry Principles, Part I* by Giannotti, and Mark et al.

OPTIONAL TEXT: None

REQUIRED SUPPLIES: Calculator, laboratory coat and safety glasses or goggles. Other items as mandated by instructor.
FARMINGDALE STATE COLLEGE
DEPARTMENT OF CHEMISTRY

CHM 152 - General Chemistry Principles I

Course Outline

The objectives given below are keyed to the lecture sections:

I. Fundamental Concepts and Units of Measurements.

In this unit students are introduced to chemistry as a central science. They explore the different branches of chemistry and its importance in our life. Students are exposed to the following topics including: Scientific Method, Matter and Its Classification, Properties of Matter, Measurements and SI System, Uncertainty in Measurements, Units Conversion, and Density.

Section I: At the end of this section, the student should be able to:

1. Know the System Internationale S.I. (Metric System) units and observations and perform conversions of length, mass and volume.

2. Calculate density from appropriate data and or interconvert to mass or volume from density data.

3. Recognize states of matter, both homogeneous and heterogeneous; know the difference between chemical and physical properties and relate it to elements and compounds.

4. Dalton’s Atomic Theory should be understood as it relates to the Laws of Conservation of Mass, Definite Proportions and Multiple Proportions.

5. From subatomic particle data calculate the Atomic Number, Neutron Number and Mass Number; from % isotopic abundance data calculate Atomic Weights/Mass. From Atomic Weights/Mass data calculate Molecular Weight.

6. Calculate % Composition from both a Molecular Formula and Experimental Data.

7. Understand the mole concept; calculate moles of elements and compounds and numbers of atoms and molecules.
II. Elements, Compounds and Chemical Reactions.

In this unit students are introduced to the concept of elements, atoms and atomic theory. A deeper discussion on atoms and subatomic particles is also cover. The periodic table as well as how it can be used to arrange elements as metals, nonmetals and metalloids is also cover in this chapter. How to write chemical formulas, equations and their use in writing chemical reactions is also introduced here. Finally, the definitions of compounds including molecular and ionic compounds are discussed in this section.

Section II: At the end of this section, the student should be able to:


2. Become familiar with Dalton’s atomic theory and the evidence that supports such theory.

3. Students should be familiar with the periodic table and how elements are arranged in groups and periods.

4. Be able to write chemical equations and define and recognize a molecular and ionic compound.

5. Be able to name ionic and molecular compounds.

III. The Mole: Relating the microscopic world of atoms to laboratory measurements.

In unit chapter students are introduced to the concept of the mole and how it conveniently links mass to number of atoms or molecules. How chemical formulas relates to amounts of substances in a compound. How chemical formulas can be determined from experimental mass measurements. How chemical equations link amounts of substances in a reaction. Finally, students should have a clear understanding of the concept of percent and theoretical yield in a chemical reaction.
Section III: At the end of this section, the student should be able to:

1. The mole, Avogadro’s number and how mole relates to mass, number of atoms or molecules.

2. The concept of Stoichiometry, Stoichiometric equivalencies and how a chemical formula relates to amounts of substance.

3. The concept of mass percent and how we can determine percentage compositions in order to identify unknown compounds.

4. How to determine empirical formulas from: mass percentage and indirect analysis.

5. How to determine molecular formulas using empirical formulas and molecular masses.

6. Writing and balancing chemical equations.

7. How we can use mole in order to calculate the following: density, mass and molecular weight.

8. How to calculate percent yield and percent recovery.

IV. Reactions of Ions and Molecules in aqueous solutions.

In this unit students should become familiar with special terminology that applies to solutions. How ionic solutions conduct electricity when dissolved in water. The concept of Acid and Bases including: naming and special properties of such compounds. Students should learn how to predict ionic reactions as well as determining the concentration of solutions. Learn the concept of molarity and how it is used for problems in solution stoichiometry.
Section IV: At the end of this section, the student should have a clear understanding of:

1. The definition of solution, solvent, concentrated solution, dilute solution and saturated, unsaturated solutions.

2. The concept of electrolytes, weak and strong, as well as writing molecular, ionic and net ionic equation.

3. The Arrhenius definition of Acid and Bases.

4. Capable of classifying Acids and Bases as: strong or weak as well as strong or weak electrolytes.

5. Capable of predicting ionic reactions and identify them as: precipitation, neutralization, and combustion reactions.

6. Molarity, and how it can be calculated from volume and moles of reactions.

7. Students should be able to have a clear understanding of the concept of titration and related terminology.

V. Energy and Chemical Change.

In this unit, students study the nature of energy, how it is measured, and how it relates to chemical change. The following concepts are explained to students: object have energy if is capable of doing work. Internal energy is the total energy of an object’s molecules. Heat can be determined by measuring temperature changes. Energy is absorbed or released during most chemical reactions. Heats of reaction are measured at constant volume or constant pressure. Thermochemical equations are chemical equations that quantitatively include heat. Thermochemical equations can be combined because enthalpy is a state
function. Tabulated standard heats of reaction can be used to predict any heat of reaction using Hess’s law.

Section V: At the end of this section, the student should have a clear understanding of:

1. The concept of Kinetic Energy and Potential Energy as well as factors that affect both.

2. The law of conservation of energy.

3. Understand the difference between temperature in heat. Calculate the heat lost or gained using water as a calorimeter. Be able to extend the results of heat lost and gained to determine the specific heat of a metal.

4. The kinetic molecular theory as well as internal energy as state of function.

5. How heat can be determined by the change as well as the concept of closed, open and isolated systems.

6. Heat and specific heat capacity and how to carry out calculations using such formula.

7. The concept of endothermic, exothermic, pressure and calorimetry.

8. Pressure-volume work or P-V work and how they are related.

10. Enthalpy and thermodynamic equations including how they can be manipulated.

11. Hess’s law and how heats of reaction can be predicted using such law.

VI. The Quantum mechanical atom.

In this unit students should gain a solid understanding of electromagnetic radiation and the clues that it provides to electronic structures of atoms. They should also gain a deeper understanding about electrons properties including: electron spin, electron distribution among orbitals in the atom and how electron configuration explains the structure of the periodic table. Finally students should be introduced to quantum theory and how it predicts the shape of atomic orbitals.

Section VI: At the end of this section the student should be able to:

1. Understand the relationship between wavelength frequency and speed of electromagnetic radiation.

2. Know the contributions of Bohr in determining the structure of the Hydrogen atom.

3. Calculate the energy of Hydrogen atom from the orbit number, n.

4. Interconvert between energy, frequency and wavelength via the Plank equation.

5. Recognize and know the significance of the deBroglie equation and quantum theory.

6. Describe electron distribution via the Bohn Model and the Modern Concept of an atom. Understand and use quantum numbers to describe electrons.

7. Interconvert between \( nl^k \) notation and quantum numbers.

8. Recognize and know which elements in the Periodic Table are classified as representative, transition and inner transition elements.
9. Recognize trends in the Periodic Table as they relate to Atomic Size, Ionization Energy and Electron Affinity.

10. Know that differences in Electronegativity lead to the formation of polarized bonds resulting in covalent, hydrogen and ionic bonding as well as it being responsible for differences in solubility.

VII. Chemical Bonding: General Concepts.

In this unit, students are introduced the following concepts: How electron transfer leads to the formation of ionic compounds. How to write Lewis dot structures as well as how covalent bonds are formed. How covalent bonds can have partial charges at opposite ends leading to the formation of polar covalent bonds. How the reactivities of metals and nonmetals can be related to their electronegativity.

Section VII: At the end of this section, the student should have a clear understanding of:

1. Lattice energies and how it is influenced by ionic size and charge.

2. The octet rule and how it can be use to write ionic compounds and covalent molecules.

3. How to write Lewis dot structures of both ionic and covalent compounds.

4. How covalent bonds are formed.

5. How dipole moments can lead to the formation of polar covalent bonds.

VIII. Chemical Bonding and Molecular Structure.

In this unit, students will learn about the different types of shapes that molecules have and the way the electronic structures of atoms influence the chemical bonds that determine molecular geometry. How molecules are arranged in a three dimensional shape and how they are built from basic arrangements. The VSEPR
model and how it is used to predict molecular shapes. How symmetry affects the polarity of molecules and how the valence bond theory is used to explain bonding as an overlap of atomic orbitals. Hybrid orbitals are also used to explain experimental molecular geometries and described multiple bonds. How molecular orbital theories explains bonding as constructive interference of atomic orbitals. Finally how molecular orbital theory uses delocalized orbitals to describe molecules with resonance structures.

Section VIII: At the end of this section, the student should have a clear understanding of:

1. The different types of molecular shapes and bond angles based on Lewis dot structures.

2. How to use and apply the VSEPR model.

3. How molecular symmetry affects the polarity of molecules.

4. How the Valence Bond Theory is used to explain bonding as an overlap of atomic orbitals.

5. Hybridization theory and how it can be used to predict shapes and angles. What a sigma bond and a pi bond consist of.

6. Molecular orbitals and how it is used to described resonance.

IX. Properties of Gases.

In this unit, students should learn how several properties of gases can be explained at the molecular level. How pressure is a measure property of gases and how gas laws summarize experimental observations. How gas volumes can be used in solving stoichiometric problems as well as how the ideal gas law relates: pressure, volume, temperature and the numbers of moles of gas (n). Graham’s law and how it leads to the concept of effusion and diffusion. How kinetic molecular theories explains the gas laws and how real gases don not obey the ideal gas low perfectly.
Section IX: At the end of this section, the student should have a clear understanding of:

1. The concept of pressure and the SI units related to it.

2. Boyle’s law, Gay-Lussac’s law, Avogadro’s principle, combine gas laws and the law of combine volumes.

3. The ideal gas law $PV=nRT$ and how it can be manipulated to solve for $P$, $V$, $n$, and $T$.

4. How molar masses can be calculated from measurements of $P$, $V$, $T$ and mass.

5. The concept of partial pressure, mole fractions and mole percents.

6. Effusion and Diffusion and how it leads to Graham’s Law.

FARMINGDALE STATE COLLEGE  
DEPARTMENT OF CHEMISTRY  

CHM-152  General Chemistry Principles I  

TEXT: Laboratory Manual for General Chemistry Principles – Part I  

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GRADING POLICY

Lecture

The lecture portion of the course constitutes 75% of the final grade. Four unit examinations are given during the semester. No make-up examination will be given for any missed exam. A comprehensive final is optional for students to take, and can be used to replace a missed or poor examination grade. The final lecture grade will be an average of the four best lecture exam grades (comprehensive final counts the same as a unit exam).

Laboratory

The laboratory portion of the course constitutes 25% of the final grade. There are 13 laboratory experiments in total. The lowest laboratory grade will be dropped. The final laboratory grade will be an average of all laboratory report grades.

Final Grade = 75% Lecture + 25% Lab Grade.