Course Outline For:

General Chemistry Principles (CHM 152)

Credits: 4  Contact Hours: Lecture: 3  Lab: 3

NOTE on Laboratory: Both Lecture and Laboratory must be taken simultaneously; separate grades will not be given for either. Students must pass the laboratory section to receive a passing grade in the entire course.

Semesters Offered: Fall, Spring, & Summer

Prerequisites: Regents Algebra 2 or MTH 116 or equivalent AND Regents High School Chemistry or CHM 124 or equivalent

Catalog Course Description:

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

Required Course for: Bioscience; Medical Laboratory Technology; Certificate for Health Professions

Elective Course for: Science, Technology, & Society; Liberal Arts & Sciences; Mechanical Engineering

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

Course Texts:


CHM 152 Laboratory Manual for General Chemistry Principles (Giannotti, Mark, et al., FSC Chemistry Dept.)

Other Required Course Materials

Calculator, laboratory coat and safety glasses or goggles.

NOTE: This course outline supersedes any course syllabus provided by a professor.
Course Learning Objectives:

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.

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

1. Know the System Internationale (S.I., i.e. 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.

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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 covered. The periodic table as well as how it can be used to arrange elements as metals, nonmetals and metalloids is also covered 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.

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.

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.

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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.

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: An object has energy if it 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.

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.

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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.

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 Mod and the Modern Concept of an atom.
7. Understand and use quantum numbers to describe electrons.
8. Interconvert between nlx notation and quantum numbers.
9. Recognize and know which elements in the Periodic Table are classified as representative, transition and inner transition elements.
10. Recognize trends in the Periodic Table as they relate to Atomic Size, Ionization Energy and Electron Affinity.
11. 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.

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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.

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 used 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.

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

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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.

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.
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.

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## Laboratory Schedule

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