Course Information Form (CIF)

Course Code: CHE 100 (IAI P1 902L)
Course Title: Introduction to Chemistry

Department: Natural Sciences
Effective Date: Summer 2026
PCS Code: 1.1 - Baccalaureate/Transfer
CIP Code: 40.0501
Repeatability: 0

Credit Hours

Catalog Notation: 3-2-4
Credit Hour Distribution:

Lecture: 3	Lab: 2	Clinical: 0	**Total: 4**

General Course Information

Catalog Description

Introduction to chemical concepts, including chemical composition, the metric system, atomic structure, moles, reactions, bonding, gases, solutions, acids and bases. Designed primarily for those with little or no high school chemistry who expect to continue with CHE 141.

General Course Objectives

To introduce students to concepts and problem solving techniques that would enable the student to succeed in CHE 141 and CHE 142.

Minimum Placement Levels

English	Reading	Math
None	Placement out of CCS 098	Placement out of MAT 072

Prerequisites

None

Recommended before enrolling (not required):
Placement out of MAT 098

Methods of Evaluation

10-20 quizzes, 2-4 one-hour exams, 8-11 lab write-ups, and a comprehensive final exam.

Instructional Materials and Additional Supplies

Introductory Chemistry, 2nd ed. Kevin Revell. Macmillan with Achieve access.
Students need a scientific calculator and lab safety goggles.
The college will provide lab equipment and materials for on-campus students. Online and hybrid students must buy a lab kit.

Course Content

General Learning Outcomes (GLOs)

- Reasoning and Inquiry: Students will demonstrate the ability to solve problems using deductive reasoning and logic, quantitative reasoning, or the scientific method.

Course Segments and Student Learning Outcomes

Course Segment	Learning Outcomes	Lecture Hours	Lab Hours	Clinical Hours
Foundations	1. Define and describe chemistry and describe the impact of chemistry on a variety of other fields. 2. Define matter. Differentiate between elements, compounds, homogeneous mixtures, and heterogeneous mixtures. 3. Describe and state the properties of the three phases of matter and identify examples of phase transitions. 4. Compare and contrast physical and chemical properties and physical and chemical changes. 5. Define heat energy in terms of the motion of particles. 6. Describe the relationship between the potential energy of a system and its potential for change. 7. Identify the difference between an endothermic process and an exothermic process. 8. Explain the differences between a hypothesis, a theory, and a scientific law. 9. Describe the key components of, articulate, and employ the scientific method.	4	2	0
Measurement	1. Convert between standard and scientific notation and solve multiplication and division problems involving scientific notation. 2. State the base metric units for mass, volume, and length and predict which units are best to use when given items. Use the metric system when measuring mass, length, volume, and temperature. 3. Describe the quality of measurements using the terms accuracy and precision. 4. Make measurements using scientific techniques, including how to estimate the last number in a measurement. 5. Identify significant digits in a measured number, and report measurements to an appropriate number of significant digits. 6. Round a calculation to the appropriate number of significant digits. 7. Identify exact numbers. 8. Apply the rules for significant digits to simple calculations, distinguishing between addition/subtraction rules and multiplication/division rules. 9. Convert values between the following metric prefixes: kilo, hecta, deca, deci, centi, milli, and micro. 10. Perform unit conversions using the factor-label (dimensional analysis) method. 11. Calculate density and use it as a conversion factor. 12. Convert between Celsius, Fahrenheit, and Kelvin temperature scales.	4	3	0

Course Segment	Learning Outcomes	Lecture Hours	Lab Hours	Clinical Hours
Atoms	1. Apply the law of conservation of mass to solve mass problems related to chemical reactions. 2. Describe the development of atomic theory and its key observations about atoms. (optional) 3. Describe the organization of the Periodic Table, being able to locate and identify elements in the groups/families: metals, nonmetals, metalloids, main group elements, alkali metals, alkaline earth metals, transition metals, halogens, and noble gases. 4. Relate an element's symbol to its name. 5. Describe the behavior of charged particles. 6. Describe the relative mass and charge of protons, neutrons, and electrons and their arrangement within an atom, in relation to the nucleus. 7. Relate the number of protons to atomic number. Relate the sum of nuclear particles to mass number. 8. Define isotopes, describe the nuclear structure of isotopes, and write the chemical symbol with a mass number for a given isotope. 9. Calculate average atomic mass from a distribution of isotopes and relative abundances. Differentiate between mass number and average atomic mass. 10. Contrast the description of electrons in the Bohr model and the quantum mechanical model. 11. Identify the overall charge of an atom or ion based on the number of protons and electrons, and conversely, use the charge of an ion to determine its number of protons and electrons.	4	1	0
Light and Electronic Structure	1. Qualitatively describe and calculate the relationships between the wavelength, frequency, and energy of electromagnetic radiation. 2. Describe line spectra, the Bohr model, and how they are related. 3. Draw Bohr/Shell diagrams for atoms. 4. Describe the absorption or emission of light as a function of electron transitions. 5. Describe Heisenberg's uncertainty principle and the wave nature of electrons. 6. Identify the number of orbitals and the maximum electron capacity of the s, p, and d sublevels/subshells. Correlate each primary energy level with the available sublevels. 7. Draw energy diagrams (i.e. orbital diagrams) for atoms and ions 8. Write electron configurations for neutral atoms, using either full notation or noble gas shorthand, for the first 56 elements. 9. Write electron configurations for ions, using either full notation or noble gas shorthand, for the first 56 elements. Define isoelectronic. 10. Identify the inner, outer, and valence electrons in an atom or ion. Apply the octet rule to explain the exceptional stability of the noble gases. 11. Use the periodic table to identify the highest-occupied energy level and sublevel of an element. 12. Use the periodic table to identify the number of valence electrons for main-group elements.	4	3	0

Course Segment	Learning Outcomes	Lecture Hours	Lab Hours	Clinical Hours
Chemical Bonds and Compounds	1. Use the periodic table to identify the number of valence electrons in an atom. 2. Represent valence electrons using Lewis dot symbols. 3. Describe and predict the formation of main-group ions using the octet rule. 4. Identify if an ion is a cation or anion. 5. Name and write formulas for cations, including group 1 & 2 and aluminum monatomic cations, metals that can form more than one ion, and the polyatomic cation, ammonium. 6. Name and write formulas for anions, including monatomic anions and polyatomic anions. 7. Broadly describe the arrangement of ions in an ionic solid. 8. Name ionic compounds. 9. Predict ionic formulas based on cation and anion charges. 10. Describe how nonmetals fulfill the octet rule through covalent bonds. 11. Differentiate between empirical and molecular formulas. 12. Name binary covalent compounds. Write formulas for binary covalent compounds. 13. Classify elements as either atomic or molecular and list the 7 diatomic elements. 14. Distinguish ionic and covalent compounds based on their chemical formulas. 15. Contrast the behavior of ionic compounds and covalent compounds in aqueous solutions. 16. Describe the ionization of acids in aqueous solution. 17. Name binary acids. Write formulas for binary acids 18. Name oxyacids. Write formulas for oxyacids.	4	3	0
Chemical Reactions	1. List evidence of chemical reactions. 2. Identify reactants, products, and phase notations in a chemical equation. 3. Write chemical equations to express the identity and ratio of species in a chemical change. Use a balanced equation to describe the ratio in which atoms or compounds react. 4. Correctly balance a chemical equation. 5. Classify synthesis, decomposition, single-displacement, double-displacement (both precipitation and acid-base), and combustion reactions. 6. Predict the products formed from the reaction of metals and nonmetals in synthesis/combination reactions. Predict the products of a decomposition reaction from a compound to its elements. 7. Write ionic/dissociation/dissolution equations for the dissociation of ionic compounds. 8. Apply the solubility rules to determine whether common ionic compounds are water soluble. 9. Predict the products of precipitation reactions. 10. Describe precipitation and neutralization reactions using molecular, complete ionic, and net ionic equations. 11. Predict the products of acid-base neutralization reactions.	4	3	0

Course Segment	Learning Outcomes	Lecture Hours	Lab Hours	Clinical Hours
Mass Stoichiometry	1. Calculate the formula mass/weight of a compound. 2. Calculate the percent composition (by mass) of elements in a compound. 3. Broadly describe how chemists measure formula mass and percent composition. 4. Use the mole concept to relate masses on the atomic scale to masses on the laboratory scale. Use Avogadro's number (6.022 x 1023) in simple conversions for macroscopic and microscopic items 5. Convert between grams and moles for elements and compounds, convert between moles and number of atoms or molecules/formula units, and convert between grams and number of atoms or molecules/formula units. 6. Use number of atoms in a chemical formula as conversion factors. 7. Apply the mole concept to solve stoichiometry problems, relating the amounts of reactants and products in a chemical change. State and compute mole to mole conversions using a balanced chemical equation. 8. Compute gram to gram stoichiometry conversions using a balanced chemical equation and molar masses. 9. Identify the limiting and excess reagents in chemical reactions and use the limiting reagent to compute how much product will be produced (i.e. the theoretical yield). 10. Differentiate between theoretical, actual, and percent yield. Using the theoretical and actual yields, correctly calculate the percent yield for a chemical reaction. 11. Describe chemical and physical occurrences that can lead to an actual yield that is less than the theoretical yield.	5	5	0
Covalent Bonding and Molecules	1. Define valence electron, predict how many valence electrons a given Main Group element has, and explain the significance of valence electrons. 2. Describe the electronic arrangements of covalent structures, including single, double, and triple covalent bonds, lone pairs, and filled octets (and filled duet - only for hydrogen). 3. Draw Lewis structures for simple covalent molecules. 4. Draw Lewis structures for polyatomic ions. 5. Apply the VSEPR model to predict the electronic and molecular geometry for molecules having two, three, or four electron sets/groups. 6. Describe the trends in electronegativity across the periodic table. 7. Use differences in electronegativity to differentiate between nonpolar covalent, polar covalent, and ionic bonds. 8. Estimate if a molecule is polar or nonpolar overall (i.e. has a molecular dipole) through the combination of polar covalent bonds and molecular shape.	4	3	0
Solids, Liquids, and Gases	1. Describe the motion of particles in a solid, liquid, or gas. 2. Describe the bonding and arrangement of particles in ionic metallic and molecular/covalent substances, and use that to explain their properties. 3. Describe the different types of intermolecular forces, which compounds exhibit each, and relate these differences to relative melting or boiling temperature. 4. Describe the key features of an ideal gas. 5. Describe how to use a liquid barometer to determine pressure. 6. Convert between different units for pressure. 7. Apply the combined gas laws to relate changes in the pressure, volume, and temperature of a gas. 8. Relate the pressure, volume, number of moles, and temperature of a gas using the ideal gas law. 9. Relate the temperature, volume, and pressure of a gas to atomic or molecular motion.	4	3	0

Course Segment	Learning Outcomes	Lecture Hours	Lab Hours	Clinical Hours
Solutions	1. Define and give examples of solutions, solutes, and solvents. 2. Calculate solution concentrations by percent (m/m, v/v, and m/v) and use mass percent as a conversion factor. 3. Calculate solution concentrations by parts per million and parts per billion. 4. Convert between moles, volume, and molarity. 5. Quantitatively describe the preparation of solutions of a known molarity. 6. Perform dilution calculations. 7. Describe the behavior of electrolytes and nonelectrolytes in aqueous solution. 8. Describe reactions that take place in aqueous solution using dissolution equations and molecular and ionic equations. 9. Apply the principles of stoichiometry to solve problems involving solutions.	4	2	0
Acids and Bases	1. Identify properties of acids and bases. Give examples of acids and bases. 2. Name and write formulas for acids, strong bases, and ammonia. 3. Define acids and bases using both the Arrhenius and the Bronsted-Lowry definitions. 4. Describe the process of acid ionization and base dissociation in aqueous solution. 5. Explain the difference between strong and weak acids. Identify strong versus weak acids. 6. Identify the acid, base, conjugate acid, and conjugate base in a chemical equilibrium. 7. Predict the products from acid-base neutralization reactions. 8. Recognize products when metal oxides react with acids and when non-metal oxides react with water to form acids. 9. Explain how the addition of acid or base to an aqueous solution affects the concentration of H+ and OH- in an aqueous solution. Use the pH scale to determine if a solution is acidic, basic, or neutral. Relate the values of pH, [H+] and [OH-] in an aqueous solution. 10. Identify common indicators for acid and base solutions. Identify methods for determining pH or pH ranges of acidic, neutral, or basic solutions. 11. Apply data from a titration experiment to find the concentration of an unknown acid or base. 12. Describe how buffers can stabilize the pH of solutions. Identify a good buffer.	4	2	0

Total Contact Hours

Lecture Hours	Lab Hours	Clinical Hours
45	30	0