CHGN 121 Learning Outcomes

Overall Course Learning Outcomes:

After fully participating in this course, you should be able to…

1. **Design and conduct experiments to predict and explain simple chemical and physical processes.**
2. **Calculate or measure quantities using the correct precision and units.**
3. **Predict the products and calculate amounts of substances in a chemical reaction.**
4. **Describe the trends in periodic properties of elements and explain why they occur.**
5. **Identify the primary types of bonds in a substance, and types of intermolecular forces, if present.**
6. **Explain molecular-level differences between solids, liquids, gases for pure ionic and covalent substances.**
7. **Calculate amount of energy flow during a chemical reaction or phase change, classifying it as heat or work.**
8. **Determine 3D shape and polarity of small molecules and polyatomic ions to predict trends in properties.**
9. **Explain in basic terms how each concept above applies to modern science or engineering issues.**

Detailed Learning Outcomes:

Working toward the Overall Course Learning Outcomes I can…

- rearrange algebraic expressions.
- write values in correct scientific notation.
- determine the number of significant figures (estimated uncertainty) in any numerical value.
- determine the number of significant figures (estimate uncertainty) on a balance or marked glassware (buret, graduated cylinder, beaker).
- calculate new quantities to the correct significant figures after a combination of both addition/subtraction and multiplication/division steps.
- use dimensional analysis to arrive at the correct units in a lengthy calculation, if given the conversion factors.
- list (from memory) the SI base units for length, mass, time and temperature.
- calculate volume and density from SI base units.
- convert between any common metric prefixes from memory (factors of 1000 from pico up to Giga).
- convert between Celcius, Kelvin, and Fahrenheit temperature values.
- define and differentiate precision and accuracy.
- determine the accuracy of a measurement.
- compare the precision of various measurements or sets of measurements.
- write a clear and useful hypothesis or prediction statement.
- determine the number of protons and electrons in any neutral atom using only the periodic table
- determine the number of neutrons in any given isotope from its mass number using only the periodic table.
- determine the number of electrons in any cation or anion using only the periodic table.
- calculate the average atomic mass of an element from the natural abundance of its isotopes.
- convert between numbers of particles and numbers of moles.
- convert between number of moles and mass.
- distinguish and explain the differences between the three main states of matter: solid, liquid, gas.
- distinguish between physical and chemical changes at the molecular level.
- distinguish between intensive and extensive properties.
- define mixture, pure substance, element, compound, heterogeneous and homogeneous.

- differentiate between mixtures and pure substances; elements and compounds; and heterogeneous and homogeneous mixtures.
- explain how compounds, e.g. NaCl, are different from the elements, e.g. Na and Cl_2, from which they are composed.
- identify elements as atomic or molecular.
- calculate the molar mass (molecular weight) for an ionic or covalent compound.
- convert between mass, moles and molecules of a compound based on its formula.
- count the atoms in a molecular or structural formula.
- determine whether a substance has ionic, covalent, or metallic bonds based only on the types of atoms present.
- determine if a covalent bond is polar or non-polar based on the electronegativity of each atom.
- explain how covalent compounds tend to exist as discrete (individual) molecules held together by intermolecular forces.
- explain how ionic compounds consist of a repeating pattern of ionic bonds and not discrete molecules.
- explain that bonds are made of matter (electron density).
- describe how the area of the electron density is different and similar in a non-polar covalent, polar covalent or ionic bond.
- explain the atomic-level changes in bonding during a physical or chemical process.
- explain how ions are formed from elements.
- determine the number of electrons gained or lost to form an anion or cation.
- identify the charges formed by the common ions (Groups I, II, III, V, VI and VII)
- write formulas for binary ionic compounds based on the charges of commonly formed ions.
- name binary molecular (covalent) and ionic compounds.
- write molecular formulas from names of binary molecular (covalent) and ionic compounds.
- write names and formulas for common binary acids and oxoacids.
- convert between empirical and molecular formula.
- calculate mass percent of any element from a chemical formula.
- use mass percent as a conversion factor.
- determine an empirical formula from combustion analysis data or elemental percent data.
- write chemical equations with the proper state symbols (s), (l), (g), (aq).
- balance chemical equations according to the Law of Conservation of Mass.
- balance chemical equations with more than two reactants and two products.
- predict the products of a combustion reaction.
- convert between moles of reactants and moles of products.
- convert between mass of reactants and mass of products.
- identify and calculate the amount of limiting or excess reactant.
- determine limiting reactant and calculate amount of limiting reactant for reactions with more than two reactants.
- calculate amount of excess reactant remaining.
- calculate the theoretical yield and percent yield for a given reaction.
- Calculate concentration of solute in an aqueous solution in units of molarity (M).
- Convert between numbers of moles, molecular weight, and volume of solution using molarity.
- Calculate the concentration after a solution has been diluted ($C_1V_1=C_2V_2$).
- Predict whether an ionic compound will be soluble in water based on the solubility guidelines.
- Explain the solubility trends for compounds made from common anions and cations.
- Use solubility trends to create a separation scheme for mixtures of ionic compounds or ions.
- Predict whether a compound (ionic or covalent) will dissociate after it dissolves in solution.

- Identify an ionic or covalent substance as a strong electrolyte, weak electrolyte, or non-electrolyte.
- Write chemical equations for simple dissolution or dissociation processes with the proper state symbols (s), (l), (g), (aq).
- Explain which attractions (bonds/IMFs) are broken and formed when a ionic or covalent substance dissolves or dissociates.
- Identify substances as strong or weak acids or bases (Bronsted-Lowry definition) based on name, formula and hints given on Equation Sheet.
- Explain behavior of strong or weak acids or bases in water alone.
- Recognize diprotic and triprotic acids from chemical formulas.
- Write Molecular, Complete, and Net Ionic equations for any ionic or acid/base reaction including weak acids/bases.
- Determine oxidation numbers for all elements in a compound.
- Identify and differentiate precipitation, acid/base, gas evolution and redox reactions.
- Use the Metals Activity Series to predict spontaneous redox reactions between metals.
- Explain what pressure is from a molecular point of view.
- Convert between the various units of absolute *and* relative pressure, including psig.
- Explain the difference between absolute and relative pressure and which to use in a given calculation.
- Explain the difference between absolute and gauge pressure and calculate one from the other.
- Explain how the ideal gas law combines the three simple gas laws into one equation.
- Calculate an unknown state variable (P, V, n, or T) using the ideal gas law and the gas constant, *R*, with the appropriate value and units.
- Define standard temperature and pressure and molar volume of an ideal gas.
- Convert between molar volume, molar mass, and density of a gas.
- Calculate the partial pressure of a gaseous component in a mixture.
- Calculate the mole fraction of a gaseous component in a mixture.
- Perform calculations involving the collection of gases over a solution.
- Calculate moles or mass of gas in a chemical reaction involving gases using proper gas-phase stoichiometry.
- Define & explain each of the five postulates/assumptions of kinetic molecular theory of gases.
- Describe the derivation of the ideal gas law from the kinetic molecular theory.
- Explain why all gases at the same temperature have the same average kinetic energy.
- Explain the relationship between gaseous molecular speed and molar mass.
- Interpret the graphical representation of the distribution of molecular speeds.
- Define diffusion and effusion and explain how they are related to the kinetic molecular theory.
- Explain the conditions under which the ideal gas law approximation is appropriate to use.
- Explain why the ideal gas law does not hold true at low temperature and/or high pressure.
- Explain how non-ideal behavior arises from the finite volume of gas particles and the intermolecular forces between particles.
- Identify the components of the van der Waals equation.
- explain how physical or chemical changes involve the flow of energy as heat.
- define system vs surroundings for any physical or chemical change.
- explain the first law of thermodynamics and how it is used.
- identify sign (+ or -) for heat of the system or surroundings.
- convert between energy units (J, L*atm, calories, etc.).
- convert between heat flow and temperature changes using specific heat capacity.
- define enthalpy change as the heat flow under constant pressure.
- identify a process as endothermic or exothermic.

- explain the molecular-level behavior of endothermic & exothermic reactions.
- explain that breaking a chemical bond always requires energy (and thus raises the potential energy of the atoms).
- explain that the formation of a chemical bond always releases energy (and thus lowers the potential energy of the atoms).
- estimate reaction enthalpies using average bond energies for all bonds broken and formed in a chemical reaction.
- write thermochemical equations for the formation of compounds from their elements.
- calculate the enthalpy of reaction using enthalpies of formation of products and reactants.
- use Hess's Law to calculate reaction enthalpy (heat of reaction) from known heats of formation or heats of reaction by multiplying chemical equations by a factor, reversing chemical equations, and summing a series of chemical equations.
- identify the correct method to calculate the enthalpy of reaction from a variety of data types: calorimetry data, bond dissociation energies, known heats of reaction, or known heats of formation.
- define and appropriately use standard state and standard enthalpy of formation.
- identify the *s*, *p*, *d*, and *f* blocks of the periodic table and list the sub-shells within a given shell from lowest to highest energy ($s < p < d < f$).
- identify the shell, sub-shell and orbital of an electron based on the shorthand orbital notation.
- define valence and core (inner-shell) electrons.
- explain why chemical reactivity of element is due to the number and location of electrons that occupy valence orbitals.
- explain how electron configurations lead to periodic trends and characteristic reactions of the main group elements.
- write electron configurations for any element using only the periodic table by following the Aufbau principle (NOT including anomalous configurations).
- define and use the Pauli exclusion principle when writing electron configurations.
- explain the term "degenerate orbitals".
- define and use Hund's Rule when writing electron configurations.
- define and use the concepts of shielding (repulsion) and Z_{eff} (effective nuclear charge) - to predict & explain the following trends:
  - atomic radius gets larger down a group and smaller across a period.
  - first ionization energy gets smaller down a group and smaller across a period.
  - higher ionization energies are larger than the first IE and are especially large when you reach the noble gas core.
  - electron affinity tends to increase across a period and decrease down a group.
  - predict and explain anomalies to this trend.
  - effective nuclear charge gets smaller down a group and larger to the right across a period.
- write electron configurations for anions & cations.
- explain why cations are always smaller than their neutral atom, while anions are always larger than their neutral atom.
- predict the size of atoms/ions in a group or an isoelectronic series.
- explain how periodic trends in ionization energy and electron affinity lead to periodic trends in electronegativity.
- explain that more electronegative atoms will pull electron density away from less electronegative atoms during a chemical reaction.
- explain differences between structure & properties of ionic vs. covalent vs. metallic substances.

- define bond as nonpolar covalent bond, polar covalent, or ionic based on differences in electronegativity (ΔEN) between atoms.
- define dipole moment and explain how it can predict polarity of covalent molecules.
- draw atoms & ionic compounds of main-group elements with their valence electrons represented as dots.
- explain how covalent and ionic substances share or transfer electrons according to the octet rule.
- explain why most nonmetal atoms prefer to be surrounded by eight valence electrons, but hydrogen is stable with only two.
- use Lewis theory to show how pairs of electrons can form a single bond between two atoms, giving each constituent atom an octet.
- explain that a double bond is the sharing of two pairs of electrons between two atoms and a triple bond is the sharing of three pairs of electrons between two atoms.
- identify and draw covalent compounds with single, double and triple bonds between constituent atoms.
- draw Lewis structures for any molecular compound and polyatomic ion.
- explain how the same chemical formula can give different molecular structures - isomers.
- define resonance structures and explain how Lewis structures represent individual and hybrid structures.
- define formal charge and use it to determine the best Lewis structure for a given molecule.
- draw Lewis structures for odd-electron species – free radicals.
- draw Lewis structures for molecules containing atoms with incomplete & expanded octets.
- explain why first and second-period elements cannot have expanded octets.
- define bond energy.
- explain the inverse relationship between bond length and bond strength.
- explain the importance of distinguishing a 2D from a 3D molecular structure.
- define VSEPR (valence shell electron pair repulsion theory).
- identify the five basic shapes of VSEPR according to the number of electron groups surrounding a central atom.
- identify bond angles & molecular shape for each basic VSEPR shape, *using the class equation sheet.*
- identify the various molecular geometries for all electron domain geometries.
- predict electron geometry & molecular geometry for molecules with one central atom.
- explain how lone pair electrons can distort the bond angles in the molecular geometry.
- identify electron & molecular geometry around each atom in a molecule with more than one central atom.
- predict polarity of bonds based on electronegativity.
- predict polarity of molecules based on Lewis structure and VSEPR shapes.
- create a Born Haber cycle for the formation of an ionic compound.
- calculate any portion of the Born Haber cycle from known values for the other portions, including multiple ionization energy atoms.
- explain how the Born Haber cycle is really just an example of Hess's Law.
- determine polarity of molecules and use to predict IMFs for that molecule.
- explain how intermolecular forces (IMFs) are responsible for liquid and solid states of matter, and originate from interactions between ions and/or partial charges on molecules, atoms, and ions.
- explain how molecular polarity determines strength of IMFs in covalently bonded substances.
    - explain that Coulomb's Law predicts energy of attraction in IMF or chemical bond based on magnitude of charges and distance between particles.

  - explain how dispersion (London) forces result from fluctuations of electron distribution within neighboring atoms as they move closer together.
  - predict how the shapes and sizes of molecules or atoms affect the magnitude of dispersion forces.
  - explain that polar molecules have permanent dipoles that attract each other through dipole-dipole interactions.
  - explain the phenomenon of hydrogen bonding and predict the ability of molecules to exhibit hydrogen bonding.
  - recognize hydrogen bonding as a force that holds together important substances, such as liquid water, solid ice and double-stranded DNA.
  - Explain how ion-dipole forces result from the attraction of cations/anions to the permanent dipole of polar molecules in mixtures of ionic and polar covalent compounds.
- rank a series of molecular compounds with respect to melting or boiling point based on type and magnitude of IMFs present, molecular/atomic weight and surface area.
- identify how the net strength of IMFs differs between phases of matter.
- give examples of surface tension, viscosity, & capillary action, and explain how these are caused by cohesive & adhesive IMFs and how temperature and surface area can affect them.
- explain the process of vaporization and how it changes with temperature, surface area and type of intermolecular forces.
- define heat of vaporization, ΔH_{vap}, as quantity of energy required to vaporize molecules of a substance.
- calculate and interconvert mass, moles, and energy using heat of vaporization.
- explain how the vapor pressure of a liquid depends on temperature but the boiling point of liquid depends on external pressure.
- define all phase change processes and explain how they appear at the molecular level: fusion and freezing, boiling and vaporization, sublimation and deposition.
- use heat of fusion, ΔH_{fus} to calculate and convert between energy, masses and moles.
- identify the macroscopic and molecular level changes that occur in different segments of the heating curve for H_2O, ranging from below melting point to above boiling point.
- calculate the energy changes associated with heating substances (like H_2O) through a series of temperature changes and phase changes – heating curves.