CLUE Learning Objectives

# General Chemistry I Learning Outcomes

Students in this course will learn to:

* Explain and model how the existence of atoms leads to the conservation of matter.
* Use appropriate models and theories to describe chemical and physical phenomena.
* Construct representations of chemical species and use them to predict chemical and physical properties.
* Explain how and why the atomic-molecular structure affects the properties of a substance, and vice versa.
* Predict and explain the energy changes associated with interactions of atoms, molecules, and ions.
* How to apply systems thinking to both molecular level and macroscopic systems.

# General Chemistry II Learning Outcomes

Students in this course will learn to:

* Explain and model how the existence of atoms leads to the conservation of matter.
* Use appropriate models and theories to describe chemical and physical phenomena.
* Construct representations of chemical species and use them to predict chemical and physical properties.
* Explain how and why the atomic-molecular structure affects the properties of a substance, and vice versa.
* Predict and explain the energy changes associated with interactions of atoms, molecules, and ions.
* How to apply systems thinking to both molecular level and macroscopic systems.
* Identify common types of reactions and predict the products.
* Understand the factors that affect the rate and extent of reactions.
* Predict and explain the outcome of coupled reactions.

# Gen Chem I & II Learning Objectives

## Chapter 1:

These goals are divided up into “big idea” concepts that we will return to time and time again, and specific learning goals – things you should be able to do.

**Big ideas:**

* Matter is made of atoms. Atoms are the smallest distinguishable part of an element.
* We use models (mental and physical) to represent many chemical entities. We use different models for different purposes.
* Theories (such as atomic theory) change over time according to the evidence available
* All atoms/molecules attract each other because of the attractive electromagnetic forces between them.
* Attractions lower the potential energy of a system, and repulsions tend to raise the potential energy.
* Stable systems form where the attractive forces and the repulsive forces are equal.
* The temperature of a phase change depends on the strength of the intermolecular forces.

**Learning Objectives – (you should be able to do these things):**

1. Make an argument (i.e. make a claim support it with evidence and reasoning) for the existence of:
	1. Atoms
	2. Electrons
	3. The existence of a small, massive, and positively charged nucleus
2. Draw diagrams/pictures of the various models of the atom as they changed over time.
3. Use the models to explain how and why the model of the atom changed over time as new evidence arose.
4. Explain how a scientific theory differs from everyday use.
5. Compare the tenets of the various atomic theories that stayed the same over time and those that changed.
6. Develop a scientific question, a scientific explanation, and use evidence and data to make an argument.
7. Compare and contrast gravitational and electrostatic forces.
8. Construct an atomic level explanation for why two isolated atoms would attract each other as they approach, and why they would repel if they get too close together.
9. Predict and explain the changes in the potential energy, the kinetic energy and the total energy as two isolated helium atoms approach each other.
10. Draw an energy diagram showing potential energy as a function of inter-nuclear distance
11. Construct a diagram/picture and use it to explain how energy is transferred at the atomic level (by collisions that can either add energy to the system or remove it).
12. Predict/rank the relative London Dispersion Forces between atoms and molecules of different sizes.
13. Relate the strength of London Dispersion Forces to relative melting and boiling points for the noble gases, or simple diatomic molecules like H2, N2 O2.
14. Contrast the energy change that occurs when two He atoms and two H atoms combine.
15. Differentiate between London Dispersion Forces and covalent bonding without providing an explanation.

## Stoichiometry: (not in text – use web materials)

**Big Idea 1: Atoms are conserved, and therefore, mass is conserved during a chemical reaction.**

1. Draw a molecular level picture showing how atoms rearrange during a chemical reaction.

2. Predict how many molecules of product will be formed and which reactant is the limiting reagent using molecular level drawings.

**Big Idea 2:** **The mole is the number (of particles) that allows us to convert from the molecular to the macroscopic level.**

1. Use mole-mass conversions to calculate how many moles of a substance are in a given mass (and vice versa).

2. Calculate the mass of any reactant or product given the mass of any other reactant or product.

3. Determine the limiting reagent when given the masses of two or more reactants, and calculate the amount of product produced.

4. Calculate the percent yield for a reaction, when given the experimental yield.

## Chapter 2:

**Big Idea 1: Electromagnetic radiation has wave properties.**

1. Describe the relationship between the frequency and wavelength and velocity (speed) of a wave.
2. Draw and compare two waves of different, frequency, or wavelength, or amplitude.
3. Calculate any one (frequency, wavelength or velocity) given the other two – using appropriate units.
4. Rank e/m radiation in in terms of energy (or wavelength, or frequency) (gamma, X, UV, visible, IR, radio)
5. Describe and identify experimental evidence for why light is a wave (diffraction, interference)
6. Make an argument (claim, evidence and reasoning) about why we can consider electromagnetic radiation as a wave.

**Big Idea 2: Electromagnetic radiation has particle properties.**

1. Describe and identify experimental evidence for why light is a particle (photoelectric effect)
2. Draw graphs that show the electrical current/# electrons vs. frequency, or wavelength of radiation.
3. Explain **why** the existence of photons (quantized light energy) explains the photoelectric effect.
4. Calculate the energy of photons of a given frequency or wavelength, and vice versa.
5. Make an argument (claim, evidence and reasoning) about why we can consider electromagnetic radiation as a particle.

**Big Idea 3: Electrons in atoms have quantized energy levels.**

1. Explain how (and why) different atoms emit different wavelengths of light.
2. Compare and contrast atomic emission and absorption spectra and how they arise.
3. Use spectra to identify the presence of elements (by comparison, not calculation).
4. Make an argument for why spectra are direct evidence for the existence of quantized energy levels in an atom.

**Big Idea 4: Electrons have wave properties.**

1. Describe experimental evidence for the wave nature of electrons (diffraction).
2. Calculate the wavelength of particles (given mass, velocity).
3. Describe how the wave properties of the electron are taken into account in the current model of the atom.
4. Draw and/or recognize pictures of s and p orbitals.
5. Describe an atomic orbital and what it represents.
6. Write electron configurations for the first 36 elements.
7. Identify core and valence electrons for the first 36 elements.

**Big Idea 5: Atomic theories change over time as new experimental evidence is produced.**

1. Describe how the model of the atom changed from Dalton through Thompson, Rutherford, and Bohr to Schrodinger. Explain why each model changed and point out the problems with the previous model.
2. Choose an atomic model to explain a phenomenon, and explain why you chose that model.

**Big Idea 6: Periodic trends result from the quantized arrangement of electrons in atoms.**

1. Identify valence and core electrons.
2. Describe how ionization energies support the idea of quantized energy levels in atoms.
3. Predict trends in ionization energies based on placement in the periodic table.
4. Explain the concept of effective nuclear charge and how it affects the size of atomic and ionic radii.
5. Apply Coulomb’s law to explain periodic trends in atomic radii and ionization energies.
6. Predict trends in ionization energy based on effective nuclear charge.
7. Predict the relative sizes of isoelectronic atoms and ions.

## Chapter 3:

**Big Ideas: Atoms are formed in stars.**

**The macroscopic properties of a substance can be explained by the interactions at the molecular level.**

**We can explain bonding using a number of different models (e.g. valence bond, molecular orbital, metallic).**

**Chemical bonds determine the identity of the substance.**

**To explain bonding we need to include both electrostatic ideas and quantum mechanical ideas.**

**Learning Objectives:**

1. Describe the chain of events that led from the big bang to the atoms that make up your body (i.e., how were atoms like hydrogen, carbon, oxygen, and iron formed, and how did they come to be part of you?).
2. Differentiate between nuclear fusion, fission and radioactive decay.
3. Complete and balance nuclear equations.
4. Differentiate between properties of atoms, nano-particles, and larger macroscale materials. That is: explain why is it that atoms don’t have macroscopic properties like melting point, boiling point, and color, while macroscopic materials do (and can be identified by them).
5. Predict, draw models (pictures), and explain why relative melting points and boiling points for substances that exist as molecules (like H2) differ from those that exist as continuous extended networks (like diamond or metals).
6. Explain the energy changes that occur when two atoms approach each other and form a bond. Explain why a third body (atom or molecule) is (almost always) needed to form a stable bond.
7. What determines whether a bond/interaction is stable? Explain how temperature influences the stability of bonds/interactions.
8. Explain what forces and energy changes take place when a bond is formed.
9. Compare and contrast the molecular orbital bond model and the valence bond model. What are the similarities? What are the differences?
10. Describe the bonding in diamond and graphite in terms of the hybrid orbital model (valence bond theory). Explain how the macroscopic properties of diamond and graphite (appearance, melting point, strength, electricity conductance, etc.) can be explained in terms of the kind of bonding that occurs in the molecules.
11. Explain on the atomic/molecular level why metals are malleable (you can make them into different shapes), ductile (you can draw them out into wires), shiny, and conduct electricity.

## Chapter 4:

1. Translate between different representations of simple molecules (i.e., space filling models, ball and stick models, perspective formulae, and Lewis structures).
2. Draw Lewis structures when given the molecular formula of hydrocarbons (including multiple bonds), and compounds containing C, H, N, O, F, Cl, Br, S, P.
3. Draw structural isomers when given a molecular formula. Distinguish between structural isomers and alternative representations of the same molecule.
4. Explain why rotation is (generally) easy around single (sigma) bond, but more difficult around pi bonds. Predict relative potential energies of different rotations around C–C bonds in hydrocarbons.
5. Use Lewis structures and VSEPR to deduce **electron pair geometry** and **molecular shape** of molecules.
6. Use Lewis structures to deduce the hybridization of the atoms in a molecule.
7. Convert between Lewis structures and 3-dimensional representations of structure – and vice versa.
8. Predict the polarity of bonds using electronegativities and the polarities of molecules using bond polarity and molecular shape.
9. Use molecular structure and polarity to predict the types of intermolecular forces present in molecules, including London dispersion forces, dipole-dipole interactions and hydrogen bonding.
10. Use intermolecular forces to explain properties of molecular compounds. Predict/rank relative melting and boiling points of given compounds
11. Explain the difference between covalent and ionic bonding using the idea that there is a continuum of bonding ranging from pure covalent to more or less ionic.
12. Explain why metals tend to form positive ions and non-metals tend to form negative ions. Predict the charge on ion for common ions.
13. Describe the structure of ionic compounds and use it to explain the properties of ionic compounds such as high melting point and boiling point, hard, and brittle.

## Chapter 5:

**Big Ideas**

* The temperature at which a phase change occurs depends upon the molecular structure of the compound
* As a molecular substance changes phase, intermolecular forces are overcome (not chemical bonds)
* The direction of change is determined by an increase in the total entropy change, ΔStotal or the Gibbs energy change, ΔG
* Measuring temperature changes can be related to molecular level changes in interaction strength by the thermodynamic function ΔH, and bond energies

**Learning Objectives**

1. Explain the difference/relationship between temperature, thermal energy, and kinetic energy.
2. Explain why particles in gases move with a range of different velocities at a given temperature. Identify the Boltzmann distributions of particles at different temperatures, or for particles of different molecular weights at a specified temperature.
3. Explain how temperature and kinetic energy are related, including the energy associated with vibration, rotation, and translation in different phases.
4. Explain the causes of water’s anomalous properties (high melting point, boiling point, lower density of ice relative to liquid water, specific heat).
5. Explain why the heat capacity of a substance is affected by the molecular-level structure.
6. Draw heating or cooling curves showing how the temperature changes when thermal energy is added to a substance (including a phase change). Explain why the temperature changes except during the phase change.
7. Define and give examples of open, closed and isolated systems.
8. Explain the difference between state and path functions and give examples.
9. For exothermic and endothermic processes, identify the direction of thermal energy change and the sign of q or ΔH.
10. For phase changes, identify the direction of thermal energy change and the sign of q or ΔH.
11. Explain the role of probability in entropy changes.
12. Predict the sign of the entropy change for simple systems.
13. Explain the second law of thermodynamics in terms of the system and surroundings
14. Explain why we usually use ΔG instead of the total entropy change to predict whether a process is thermodynamically favorable.

## Chapter 6:

**Big Ideas**

* A solution is a stable heterogeneous molecular mixture; it will not become “unmixed” over time
* The factors that affect solubility are the enthalpy change (ΔH), and the entropy change (ΔS)
* A rule of thumb for predicting solubility is “like dissolves like” – but from a molecular perspective, this is not the reason why a solute dissolves in a solute. Also, it does not explain quantitative differences in solubility (for example, unlike dissolves at low levels into like).
* Molecules that have both polar and non-polar parts will often assemble to form larger structures, through a process driven by a decrease in entropy (involving the water molecules, primarily.)

**Learning Objectives**

1. Understand and explain the terms: solute, solvent, solution, solvation, molarity, concentration, dilute, concentrated, micelle, colloid, emulsion.
2. Perform calculations using molarity (M) and volume of solution. Be able to calculate mass of solute needed, and determine volume of a dilute or concentrated solution.
3. Interconvert between solution units such as % by mass, ppm and ppb.
4. Relate temperature changes to the interactions that are broken and formed during the solution process.
5. Describe the solution process in terms of entropy changes for both solute, and solvent.
6. Explain why free energy change is the thermodynamic function that allows us to predict solubility. (Why can’t we just use entropy or energy changes?)
7. Predict which materials might be soluble in water given the interactions that hold their particles together. (“like dissolves like” – but remember this is not the explanation for why something dissolves).
8. Draw molecular level diagrams showing how solute and solvent interact, including how water molecules solvate ionic compounds.
9. Predict which molecular compounds can form hydrogen bonds with water.
10. Understand that solubility is not an all or nothing property – and be able to estimate the **relative** solubilities of a range of given compounds.
11. Explain why non-polar molecules are not soluble in water, being sure to include the role of entropy in your discussion.
12. Discuss how large molecules with both polar and non-polar “parts” behave, including the role of the hydrophobic effect on how biomolecules fold, and micelles and bilayers form.
13. Understand and explain the terms colloids and emulsions. Explain how they differ from solutions.
14. Explain the role of temperature in solubility.
15. Explain how different inclusions in metals can alter the properties of the metal. (consider steel, brass, bronze)
16. Predict and explain the effect of temperature on the solubility of gases and solids.

## Chapter 7:

**Big Ideas**

* Chemical reactions involve rearrangements of atoms to produce new chemical species.
* Chemical reactions involve only changes in arrangements of atomic cores and valence electrons.
* Chemical reactions can be classified by how the valence electrons behave during the course of the reaction. For example: acid base reactions or redox reactions
* Energy changes in reactions arise from the changes in bond energies as bonds in reactants are broken (which requires an energy input) and new bonds are formed (which releases energy to the surroundings).

**Learning Objectives**

1. Explain, using a molecular level explanation, the difference between a chemical change and a phase change. Describe the evidence you would use to support your explanation.
2. Identify the molecular structural features that lead to acidic and basic properties in a molecule. Explain why these features are important.
3. Know common strong acids (HCl, HBr, HI, HNO3, H2SO4) and bases (NaOH, KOH).
4. Use the words strong, weak, dilute and concentrated appropriately.
5. Predict and explain relative strengths of a range of related acids (or bases) using thermodynamic and electronegativity arguments.
6. Explain how the progression of Arrhenius, Bronsted-Lowry, and Lewis acid-base theories build on each other.
7. Predict the products of acid base reactions, based on the relative strengths of the acids and bases.
8. Predict the extent of acid-base reactions (more products or more reactants) based on relative acid/base strength.
9. Use curved arrow notation to predict the products of Lewis acid-base reactions and simple nucleophilic attacks at electrophilic carbons (N-methylations).
10. Calculate pH from given concentrations of strong (i.e. completely ionized) acids or bases.
11. Use pH to calculate [H3O+] or [–OH].
12. Use pH of solutions of the same concentration to determine relative strengths of acids and bases.
13. Identify and explain why acid base reactions are different from redox reactions.
14. Recognize redox reactions by looking for changes in oxidation numbers.
15. Use bond energies to predict and calculate enthalpy changes in reactions.
16. Construct drawings and graphs to show how energy changes over the course of a reaction can be determined by the bonds that break and forming.
17. Use molarity (mol/L) to calculate amounts of reactants and products for reactions in solution.

## Chapters 8 & 9:

* Rates of reactions depend on:
	+ The probability that molecules will collide (this is increased by increasing concentration)
	+ The energy of the collision (this is increased by increasing the temperature)
	+ The height of the activation energy barrier (this depends upon the type of reaction)
* The extent of a reaction (position of equilibrium) is related to the free energy change from reactants to products and the temperature
* The position of equilibrium (but not the equilibrium constant) can be changed by changing concentrations of reactants or products.
* Reactions can be coupled by common intermediates.

**Learning Objectives:**

1. Explain what rate means when discussing a chemical reaction
2. Discuss the factors that affect rates of reactions and explain how and why each factor affects the rate.
3. Graphically determine the **rate** and the **order** of reaction from concentration v time data
4. Draw graphs that show how concentration changes with time for all reactants
5. Define and use half life for a first order reaction
6. Draw and label reaction energy profiles when given a reaction mechanism, and show the effect of a catalyst.
7. Use rate equation to determine which mechanism is more likely
8. Calculate an equilibrium constant when given concentrations at equilibrium
9. Discuss the concept of dynamic equilibrium and explain how it differs from a steady state
10. Predict which side the equilibrium lies from data such as the equilibrium constant or the Gibbs energy
11. Use acid dissociation constants (Ka) to calculate pH of weak acid. Use the pH of a weak acid to calculate the Ka
12. Predict and explain how a position of equilibrium will shift when conditions are changed.
13. Use the relationship between Q and K to predict how a reaction will shift
14. Explain how and why a buffer can resist changes in pH, and relate this to the buffer capacity and range.
15. Calculate pH of buffer solutions, and the change in pH when a strong acid or base is added
16. Predict the approximate pH of a salt, and explain what processes cause a pH change when it dissolves in water.
17. Explain how and why reactions are coupled by a common intermediate, and give examples.