Course Title	Content Area:	Grade Level:	Credit (if applicable)
AP Chemistry	Science	Gr. 10-12	1.0

Course Description

The AP Chemistry course provides students with a college-level foundation to support future advanced coursework in chemistry. Students cultivate their understanding of chemistry through inquiry-based investigations, as they explore content such as:

- Atomic Structure and Properties
- Compound Structure and Properties
- Properties of Substances and Mixtures
- Chemical Reactions
- Kinetics
- Thermochemistry
- Equilibrium
- Acids and Bases
- Thermodynamics and Electrochemistry

This course requires that 25% of instructional time be spent in hands-on laboratory work, with an emphasis on inquiry-based investigations that provide students with opportunities to demonstrate the foundational chemistry principles and apply the science practices. This includes a minimum of 16 hands-on labs (at least six of which are inquiry based).

Inquiry-based laboratory experiences support the AP Chemistry course and AP Course Audit curricular requirements by providing opportunities for students to engage in the science practices as they design plans for experiments, make predictions, collect and analyze data, apply mathematical routines, develop explanations, and communicate about their work.

Colleges may require students to present their laboratory materials from AP science courses before granting college credit for laboratory work, so students should be encouraged to retain their laboratory notebooks, reports, and other materials.

Aligned Core Resources	Connection to the *BPS Vision of the Graduate*
CollegeBoard Course and Exam Description Students must learn the skills necessary to use the following ordinary equipment: Beakers, flasks, test tubes, crucibles, evaporating dishes, watch glasses, burners, plastic and glass tubing, stoppers, valves, spot plates, funnels, reagent bottles, wash bottles, droppers, and measuring equipment. The list of measuring equipment includes: Balances (single pan, double pan, triple beam), thermometers (°C), barometers, graduated cylinders, burets, volumetric pipets, graduated pipets, volumetric flasks, ammeters, voltmeters, pH meters, and spectrophotometers. A more comprehensive list of laboratory equipment can be found in Appendix C of AP Chemistry Guided Inquiry Experiments: Applying the Science Practices. Microscale Experiments One important change in chemistry lab instruction in recent years has been the introduction of microscale experiments. Teachers can determine whether microscale experiments are appropriate for their students and classroom/lab space. Here are some benefits to consider: • Decreased cost of chemicals acquisition and	

disposal
- Reduced storage space requirements and safer storage
- Less need for elaborate lab facilities in schools
- Greater care needed by students to obtain and observe results
- Shorter experiment times as well as easier and faster cleanup
- Ability to carry out some experiments that were once restricted to demonstration because of their hazards in macroscale

Additional Course Information: ***Knowledge/Skill Dependent Courses/Prerequisites***	**Link to *Completed Equity Audit***
Students should have successfully completed an introductory high school chemistry course and Algebra 2. Grade 10 students may take concurrently with Biology ACC with teacher recommendation and "83" average in Physical Science ACC.	

Standard Matrix

Science Practices & Skills	**Unit 1**	**Unit 2**	**Unit 3**	**Unit 4**	**Unit 5**	**Unit 6**	**Unit 7**	**Unit 8**	**Unit 9**
Science Practice 1: Models and Representations *(Describe models and representations, including across scales.)*									
1.A Describe the components of and quantitative information from models and representations that illustrate particulate-level properties only.	X								
1.B Describe the components of and quantitative information from models and representations that illustrate both particulate-level and macroscopic level properties.				X	X	X			
Science Practice 2: Question and Method *(Determine scientific questions and methods.)*									
2.A Identify a testable scientific question based on an observation, data, or model.	X								
2.B Formulate a hypothesis or predict the results of an experiment.				X					
2.C Identify experimental procedures that are aligned to a scientific question (which may include a sketch of a lab setup).			X						
2.D Make observations or collect data from representations of laboratory setups or results, while attending to precision where appropriate.						X		X	
2.E Identify or describe potential sources of experimental error.			X						
2.F Explain how modifications to an experimental procedure will alter results.							X		X
Science Practice 3: Representing Data and Phenomenon *(Create representations or models of chemical phenomena.)*									
3.A Represent chemical phenomena using appropriate graphing techniques, including correct scale and units.		X				X	X		
3.B Represent chemical substances or phenomena with appropriate diagrams ormodels(e.g.,electronconfiguration).		X		X	X				
3.C Represent visually the relationship			X				X		

between the structures and interactions across multiple levels or scales (e.g., particulate to macroscopic).									
Science Practice 4: Model Analysis *(Analyze and interpret models and representations on a single scale or across multiple scales.)*									
4.A Predict and/or explain chemical properties or phenomena (e.g., of atoms or molecules) using given chemical theories, models, and representations.	X		X						
4.B Explain whether a model is consistent with chemical theories.	X								
4.C Explain the connection between particulate-level and macroscopic properties of a substance using models and representations.	X	X	X						
4.D Explain the degree to which a model or representation describes the connection between particulate-level properties and macroscopic properties.			X				X		X
Science Practice 5: Mathematical Routines *(Solve problems using mathematical relationships)*									
5.A Identify quantities needed to solve a problem from given information (e.g., text, mathematical expressions, graphs, or tables).	X					X	X		
5.B Identify an appropriate theory, definition, or mathematical relationship to solve a problem. when one variable changes.	X				X		X	X	
5.C Explain the relationship between variables within an equation			X	X	X		X	X	
5.D Identify information presented graphically to solve a problem.	X							X	
5.E Determine a balanced chemical equation for a given chemical phenomenon.				X	X				
5.F Calculate, estimate, or predict an unknown quantity from known quantities by selecting and following a logical computational pathway and attending to precision (e.g., performing dimensional analysis and attending to significant figures).			X			X	X	X	X
Science Practice 6: Argumentation *(Develop an explanation or scientific argument.)*									
6.A Make a scientific claim.		X							
6.B Support a claim with evidence from experimental data.				X					
6.C Support a claim with evidence from representations or models at the particulate level, such as the structure of atoms and/or molecules.		X						X	X
6.D Provide reasoning to justify a claim using chemical principles or laws, or using mathematical justification.						X	X	X	X
6.E Provide reasoning to justify a claim using connections between particulate and macroscopic scales or levels.			X		X	X			X
6.F Explain the connection between experimental results and chemical concepts, processes, or theories.							X		
6.G Explain how potential sources of experimental error may affect the experimental results.								X	

Unit Links
Unit 1: Atomic Structure and Properties Unit 2 : Compound Structure and Properties Unit 3: Properties of Substances and Mixtures Unit 4: Chemical Reactions Unit 5: Kinetics Unit 6: Thermochemistry Unit 7: Equilibrium Unit 8: Acids and Bases Unit 9: Thermodynamics and Electrochemistry

Unit Title:

Unit 1: Atomic Structure and Properties

Relevant Standards: Bold indicates priority

1.1 Moles and Molar Mass: Students can convert between mass, moles, and number of particles using Avogadro's number, and determine molar mass from chemical formulas to support stoichiometric calculations.

1.2 Mass Spectroscopy of Elements: Students can interpret mass spectra to determine isotopic composition and calculate average atomic mass using weighted isotopic abundances.

1.3 Elemental Composition of Pure Substances: Students can calculate percent composition and determine empirical formulas from experimental or compositional data.

1.4 Composition of Mixtures: Students can determine the composition of mixtures using mass or percent data and distinguish mixtures from pure substances based on quantitative particle-level information.

1.5 Atomic Structure and Electron Configuration: Students can determine subatomic particles and represent isotopes using nuclear notation, and write electron configurations and orbital diagrams using the Aufbau principle, Hund's rule, and the Pauli exclusion principle.

1.6 Photoelectron Spectroscopy (PES): Students can interpret PES data to determine electron configurations and relate peak positions to electron energy levels and sublevels.

1.7 Periodic Trends: Students can predict and explain periodic trends in atomic radius, ionization energy, and electronegativity using effective nuclear charge, shielding, and electron configuration.

Essential Question(s):	**Enduring Understanding(s):**
1. How can the same element be used in nuclear fuel rods and fake diamonds? 2. How can large quantities of objects be counted by weighing? 3. If atoms are too small to be observed directly, how do we know how they're structured? 4. Why does the periodic table have the shape that it does?	1. The arrangement of electrons in atoms determines the structure of the periodic table and influences the chemical behavior of elements. 2. Coulombic attractions between charged particles explain trends in atomic size, ionization energy, and electronegativity. 3. Atomic models are supported by experimental evidence and are used to explain the behavior of matter at the subatomic level. 4. The composition of matter can be represented and quantified using atomic theory, isotopes, and the mole concept. 5. Periodic trends arise from predictable patterns in electron configuration and effective nuclear charge.
Demonstration of Learning:	**Pacing for Unit**
• Success with Unit 1 MCQs and FRQs via AP Classroom • Analysis of a Hydrate Lab • Constant Composition of SG Lab • Separation of a Mixture Lab • Quizzes (AP Classroom MCQs and FRQs) every few topics • Electron POGIL • Progress Check 1 from AP Classroom Unit 1 is 7–9% AP Exam Weighting See AP Exam Description	15 days of instruction 3 lab days 2 days of review / assessment
Family Overview (link below)	**Integration of Technology:**
AP Chemistry FAMILY MATERIALS	*N/A*

Unit-specific Vocabulary:	Aligned Unit Materials, Resources, and Technology (beyond core resources):
element, atom, molecule, compound, chemical formula, formula unit, mixture, homogeneous mixture (solution), heterogeneous mixture, pure substance, atomic number, mass number, isotope, isotopic abundance, average atomic mass, mole, molar mass, avogadro's number (6.022×10^{23}), empirical formula, molecular formula, hydrate, anhydrous, percent composition, electron, proton, neutron, nuclear notation, atomic orbital, s, p, d, f orbitals, electron configuration, aufbau principle, pauli exclusion principle, hund's rule, photoelectron spectroscopy (PES), ionization energy, electronegativity, atomic radius, ionic radius, effective nuclear charge, shielding effect, periodic trends	• Bunsen burners • General lab equipment
Opportunities for Interdisciplinary Connections:	**Anticipated misconceptions:**
• Physics: electromagnetic radiation, photoelectron spectroscopy, Coulombic attraction • Mathematics: scientific notation, dimensional analysis, logarithmic relationships • Engineering: materials design based on atomic properties • Earth Science: isotopes and radioactive dating • History of Science: development of atomic models and quantum theory	Ability to identify periodic trends but difficulty explaining why these trends occur
Connections to Prior Units:	**Connections to Future Units:**
• Structure of the atom introduced in middle school, physical science, and biology • Build upon initial model (likely Bohr model) by introducing quantum view of the atom (orbitals)	• Moles and molar mass is used throughout the curriculum (stoichiometry) • Periodic trends explain reactivity (Unit 4) • Valence electrons dictate bonding (Unit 2)

Differentiation through *Universal Design for Learning*

	Engagement	Representation	Action & Expression
LT1	"Mole Day Estimation Challenge" estimating particles in everyday objects (e.g., drops of water, aluminum cans).	Color-coded dimensional analysis "train tracks" showing unit cancellation and Avogadro's constant	Solve multi-step mole conversions using a choice of digital spreadsheets, handwritten factor-label steps, or video walkthroughs.
LT2	Interactive Mass Spectrometry simulation (e.g., PhET) adjusting isotope abundances to observe shifts in average atomic mass.	Annotated mass spectrum graphs showing peak heights (relative abundance) vs. m/z ratios alongside weighted average formulas.	Calculate average atomic mass from raw spectrum data and present a written or video breakdown identifying unknown elements.
LT3	"Chemical Forensics Mystery" analyzing unknown powder samples from a simulated crime scene to deduce chemical formulas.	Flowchart guide detailing the empirical formula algorithm	Submit empirical/molecular formula calculations via step-by-step problem sets or record a screencast explaining the mole ratio logic.
LT4	"Salad Dressing / Alloy Analysis" determining mass percent of active components in commercial mixtures (e.g., brass, hydrated salts).	Visual pie charts and part-over-whole mass fraction diagrams	Write a laboratory analysis report or construct a visual infographic detailing the separation and mass percentage determination of a mixture.
LT5	"Orbital Hotel" filling game where electrons are guests following hotel rules (Aufbau, Pauli exclusion, Hund's rule).	Color-coded Periodic Table mapping s, p, d, f blocks alongside arrow energy level diagrams and noble gas shorthand notation.	Write full and condensed electron configurations using a choice of interactive orbital board applets, whiteboards, or audio explanations.
LT6	"PES Peak Detective" matching simulated Photoelectron Spectroscopy (PES) graphs to mystery elements based on binding energy.	Dual-display visual showing PES energy plots (X-axis: binding energy decreasing; Y-axis: peak height/number of electrons) aligned to shell models.	Annotate a PES spectrum identifying shell configurations and record a video explaining how nuclear charge shifts binding energy.

LT7	Interactive periodic trend heatmap exploration analyzing atomic radius, ionization energy, electron affinity, and electronegativity.	Graphic organizers comparing nuclear charge, shielding, and effective nuclear charge across periods and down groups.	Draft a Claim-Evidence-Reasoning (CER) justification or record a presentation explaining periodic trend anomalies (e.g., N vs. O first IE).
LT8	Video analysis of alkali metals reacting with water, contrasting reactivity rates down Group 1.	Comparative visual chart mapping valence electron loss/gain ease to ionization energy and electron affinity across families.	Write a comparative reactivity report or record a podcast explaining why reactivity trends differ between metals and nonmetals.

Supporting Multilingual/English Learners (*CELP standards*)

	Emerging	Expanding	Bridging
LT1	Complete dimensional analysis setups using visual cards and color-coded conversion factor templates (grams -> moles -> particles).	Calculate moles, mass, and particle counts using multi-step calculation frames and conversion equation starters.	Perform complex dimensional analysis calculations independently, writing full justifications for unit cancellation and significant figures.
LT2	Identify peak height as "abundance" and peak position as "mass" on a labeled mass spectrum diagram using a word bank.	Explain how average atomic mass is calculated from peak heights and mass positions using cause-and-effect sentence starters.	Calculate fractional abundances and average atomic masses from raw mass spectra data, providing complete written justifications.
LT3	Fill in a structured 4-column empirical formula table (Element -> Mass -> Moles -> Ratio) with formula hints.	Determine empirical formulas using guided calculation steps and sentence frames ("*The mole ratio of X to Y is...*").	Calculate empirical and molecular formulas from percent composition or combustion analysis data with full written explanations.
LT4	Identify pure substances versus mixtures using particulate diagrams and color-coded element/compound cards.	Calculate mass percent of a compound in a mixture using guided calculation frames and explanatory prompts.	Analyze indirect mass composition data (e.g., gravimetric analysis) to determine mixture purity and justify results in writing.
LT5	Fill in orbital energy diagrams (boxes with arrows) using visual rules (Aufbau, Pauli, Hund) and electron configuration word banks.	Write noble gas shorthand electron configurations for atoms and ions using an annotated periodic table guide.	Predict and explain ground-state electron configurations and anomalous exceptions (e.g., Cr, Cu) using formal orbital theory.
LT6	Label PES axes (Binding Energy vs. Relative Number of Electrons) and match peaks to specific subshells using color-coded orbital diagrams.	Explain how peak height relates to electron count and peak position relates to nuclear attraction using comparative sentence frames.	Interpret complex PES data to identify unknown elements, justify subshell binding energy shifts, and explain shielding effects.
LT7	Circle atomic radius or ionization energy trends on a periodic table map using simple directional arrows and visual cues.	Explain trends (atomic radius, ionization energy, electronegativity) using claim-evidence-reasoning (CER) sentence frames ("*As you go across a period...*").	Justify periodic trends and exceptions (e.g., N vs. O ionization energy) using effective nuclear charge (Zeff) and Coulomb's law.
LT8	Sort metals and nonmetals into high/low reactivity categories using a color-coded periodic table.	Explain reactivity trends of alkali metals and halogens using guided paragraph templates incorporating valence electron structure.	Synthesize periodic trends, Zeff, and ionization energy/electron affinity to explain complex reactivity patterns across groups.

Unit Overview

In Unit 1, students will practice identifying components of commonly used models and representations to illustrate chemical phenomena. They will construct models and representations and explain whether they are consistent with chemical theories. Students will also practice translating between data and various representations (e.g., photoelectron spectroscopy data and electron configurations). Students should then be able to use representations (e.g., PES graphs, electron configurations, periodic table, drawings) to explain atomic structure, which is the foundation for all subsequent units. Many of the most useful concepts in chemistry relate to patterns in the behavior of chemical systems, such as periodic trends in atomic and molecular properties. In this unit and all subsequent units, students should learn to analyze data presented graphically to identify patterns and relationships. Once a pattern is identified, students should be able to examine evidence to determine if it supports the pattern or hypothesis pertaining to a testable question.
hypothesis pertaining to a testable question.

Lesson	Learning Target	Success Criteria	Suggested Activities
1.1	**Learning Target 1**	• I can convert between the	**Think-Pair-Share**

Moles and Molar Mass	I can calculate quantities of a substance or its relative number of particles using dimensional analysis and the mole concept.	masses of substances and the actual number of particles using avogadro's number • I can express the mass of an individual atom or molecule in atomic mass units (amu) and understand how it will always be numerically equal to the molar mass of that substance in grams.	Ask students to individually rank three samples in order of increasing number of particles, increasing mass, and increasing mole amounts (Sample A: 1.0 mole of Carbon, SampleB: 18grams of carbon monoxide, SampleC:3.0× 1023 molecules of water). Then have them compare and defend their choices with a partner
1.2 Mass Spectra of Elements	**Learning Target 2** I can explain the quantitative relationship between the mass spectrum of an element and the masses of the element's isotopes.	• I can use the mass spectrum of an element to determine the identity of the isotopes of that element and the relative abundance of each isotope in the sample. • I can determine the average atomic mass of an element from the weighted average of the isotopic masses using the mass of each isotope and its relative abundance.	**Simulations** Conduct a simulation of a mass spectrometer, using a strong magnet and steel ball bearings of various masses, to show students how mass can be used to separate particles based on their ability to be manipulated in an electromagnetic field. Present samples of mass spectra for students to analyze and have them calculate the average atomic mass of an element. Discuss how mass spectrometry could be used to identify the presence of an element within a mixture and the isotopic abundance within an element. Forensic science applications and other modern uses of the technology can be discussed to give relevant context to the concepts.
1.3 Elemental Composition of Pure Substances	**Learning Target 3** I can explain the quantitative relationship between the elemental composition by mass and the empirical formula of a pure substance.	• I can describe the law of definite proportions • I can determine the empirical and molecular formula of a compound given mass data.	**Think-Pair-Share** Have students design an experiment to determine the percent composition of a mixture of sodium carbonate (inert) and sodium bicarbonate. After carrying out the Experiment, provide them with a mock student report to analyze and critique. Then Have them get into pairs and reflect on their particular approach and come up with additional approaches to this problem.
1.4 Composition of Mixtures	**Learning Target 4** I can explain the quantitative relationship between the elemental composition by mass and the composition of substances in a mixture	• I can describe the variable proportion nature of mixtures • I can use data to determine the percent impurity and percent composition of a mixture	**Explore Representations** Translate PES data into an electron configuration and/or predict a PES spectrum based on an element's electron configuration or location in the periodic table. Have students compare their predictions to the actual electron configuration and discuss discrepancies.
1.5 Atomic Structure and Electron Configuration	**Learning Target 5** I can represent the ground-state electron configuration of an atom of an element or its ions using the Aufbau principle	• I can describe the structure of the atom including charge, mass and location of protons, neutrons and electrons • I can use Coulomb's law to describe and compare the force between two charged particles, and therefore the	**Explore Representations** (See 1.4)

		relative amounts of energy to remove different electrons. (related to the distance of the electron from the nucleus and the effective charge of the nucleus ● I can appropriately use the terms "shells (energy levels)" and "subshells (sublevels)," as described by the ground-state electron configuration.Inner electrons are called core electrons, and outer electrons are called valence electrons.	
1.6 Photoelectron Spectroscopy	**Learning Target 6** I can explain the relationship between the photoelectron spectrum of an atom or ion and: i. The Ground-state electron configuration of the species ii. The interactions between the electrons and the nucleus.	I can relate the photoelectron spectroscopy (PES) graph to ● ...the energy required to remove an electron from the corresponding subshell ● ...the number of electrons in that subshell.	**Process Oriented Guided Inquiry Learning (POGIL)** Given ionization energy data from various elements, guide students through a series of questions to help them rationalize the relationship of the charge of the ion to its position on the periodic table, its electronic structure, and reactivity.
1.7 Periodic Trends	**Learning Target 7** I can explain the relationship between trends in atomic properties of elements and electronic structure and periodicity	● I can describe the patterns of recurring properties of the elements, which are explained by patterns of ground-state electron configurations and presence of completely or partially filled shells(and subshells) of electrons in atoms. ● I can follow trends in atomic properties within the periodic table (periodicity) which can be predicted by the position of the element on the periodic table and qualitatively understood using Coulomb's law, the shell model, and the concepts of shielding and effective nuclear charge.	
1.8 Valence Electrons and Ionic Compounds	**Learning Target 8** I can explain the relationship between trends in the reactivity of elements and periodicity	● I can describe the likelihood that two elements will form a chemical bond based on the interactions between the valence electrons and nuclei of the elements. ● I can describe analogous compounds. ● I can describe typical charges of atoms in ionic compounds	
Progress Check 1 ● Multiple-choice: ~20 questions ● Free-response: 2 questions ○ Short ○ Short			

Preparing for the AP Exam
On the AP Exam, students must be able to justify claims with evidence. This starts when students can identify the evidence needed to solve a problem or support a claim and then connect that evidence to known chemical theories. However, many students consistently demonstrated difficulty with this skill. For example, while students can memorize periodic trends, they struggle to explain the electrostatic interactions within an atom that produces period trends as well as exceptions to these trends. Further, students often have difficulty connecting periodic trends to the shell model, Coulomb's law, and elements of quantum theory. To combat these challenges, teachers can ensure that students have a strong foundation in identifying mathematical relationships or patterns from graphical or tabular information and that they can explain how those patterns are consistent with chemical theories and models.

Unit Title:

Unit 2 : Compound Structure and Properties

Relevant Standards: Bold indicates priority

2.1 Types of Chemical Bonds: Students can classify bonds as ionic, polar covalent, or nonpolar covalent based on electronegativity differences and particle-level interactions.

2.2 Intramolecular Force and Potential Energy: Students can explain how bond formation and breaking involve changes in potential energy and stability at the particle level.

2.3 Structure of Ionic Solids: Students can describe the structure of ionic solids as repeating lattices and relate structure to formula ratios and properties.

2.4 Structure of Metals and Alloys: Students can explain metallic bonding using a sea of electrons model and relate structure to conductivity, malleability, and alloy formation.

2.5 Lewis Diagrams: Students can draw Lewis structures for molecules and ions to represent valence electrons and bonding patterns.

2.6 Resonance and Formal Charge: Students can use resonance structures and formal charge calculations to evaluate and compare valid Lewis structures.

2.7 VSEPR and Hybridization: Students can predict molecular geometry and bond angles using VSEPR theory based on electron domain repulsion.

Essential Question(s):	**Enduring Understanding(s):**
1. How are molecular compounds arranged? 2. Why are some bonds easier to break than others? 3. In what ways does a diagram drawn on paper accurately reflect the structure of a molecule? In what ways does it not accurately reflect the structure?	1. The type and strength of chemical bonds (ionic, covalent, metallic) arise from electrostatic interactions between charged particles and determine the structure and properties of substances. 2. Lewis structures and molecular geometry models represent the arrangement of valence electrons and atoms, which can be used to predict bonding, polarity, and molecular shape. 3. The three-dimensional structure of molecules, determined by electron domain repulsion and orbital interactions, directly influences physical and chemical properties. 4. Formal charge and resonance structures provide a framework for evaluating and representing the most accurate depiction of electron distribution in molecules and ions. 5. The polarity of bonds and molecular geometry together determine intermolecular interactions, which influence observable properties such as boiling point, solubility, and phase behavior.
Demonstration of Learning:	**Pacing for Unit**
• Success with Unit 2 MCQs and FRQs via AP Classroom • Quizzes (AP Classroom MCQs and FRQs) every few topics • Using VSEPR to build molecules PhET / with modeling kits • Drawing Lewis Structures practice • Progress Check 2 from AP Classroom Unit 2 is 7–9% of AP Exam Weighting See AP Exam Description	10 instructional days 2 lab days 2 review / assess

Family Overview (link below)	Integration of Technology:
AP Chemistry FAMILY MATERIALS	PhET simulations for Molecular Structure
Unit-specific Vocabulary:	**Aligned Unit Materials, Resources, and Technology (beyond core resources):**
atom, element, compound, mixture, pure substance, law of conservation of mass, law of definite proportions, law of multiple proportions, Dalton's atomic theory, subatomic particle, proton, neutron, electron, atomic number, mass number, isotope, average atomic mass, ion, cation, anion, ionic compound, formula unit, molecular compound, molecule, empirical formula, molecular formula, structural formula, Lewis dot diagram, valence electron, octet rule, exception to the octet rule, expanded octet, incomplete octet, resonance, formal charge, electronegativity, electronegativity trend, polar bond, nonpolar bond, ionic character, Coulombic attraction, crystal lattice, ionic solid, network covalent solid, metallic solid, molecular solid, electron sea model	Molecule modeling kits
Opportunities for Interdisciplinary Connections:	**Anticipated misconceptions:**
• Biology: molecular polarity and hydrogen bonding in DNA and proteins • Physics: electrostatic forces and energy interactions • Materials Science: crystal lattices, semiconductors, metallic properties • Engineering: molecular geometry and material strength • Environmental Science: intermolecular forces in atmospheric chemistry and water systems	• Not accounting for lone pairs when predicting molecular geometry • Not drawing Lewis diagrams correctly– incorrect number of total valence electrons, violating octet rule, not accounting for formal charge when determining most likely structure
Connections to Prior Units:	**Connections to Future Units:**
• Location and number of electrons determines type of bond formed • Number of electrons dictates if cation or anion is formed in an ionic compound	• Type of chemical bond dictates behavior in chemical reaction (unit 4), solubility (3.7, 3.10) • Structure of ionic solids relates to net ionic equations (4.2) • Lewis structures are used throughout the year (i.e. when determining most acidic hydrogen in Unit 8)

Differentiation through *Universal Design for Learning*

	Engagement	Representation	Action & Expression
LT1	Physical testing lab comparing conductivity, melting point, and solubility of unknown solids (ionic, covalent, metallic).	Property comparison matrix mapping electronegativity difference (ΔEN) to bond type and macroscopic physical traits.	Create a flowchart or interactive quiz categorizing unknown compounds into bonding types based on experimental property data.
LT2	"Atomic Spring Simulation" adjusting internuclear distance to plot potential energy wells for single vs. double bonds.	Annotated potential energy vs. internuclear distance graph highlighting equilibrium bond length (well minimum) and bond energy (well depth).	Sketch and defend potential energy curves comparing two diatomics (e.g., H2 vs. O2 or HCl vs. HBr) on whiteboards or digital tools.
LT3	"Crystal Builder Challenge" arranging colored sphere manipulatives to construct 3D ionic lattices while accounting for ionic radii.	Coulomb's Law visual scaffold highlighting how ion charge and size affect lattice energy and cleavage planes.	Build a 3D physical or digital particulate model of an ionic crystal lattice and write a justification referencing Coulombic attraction.
LT4	"Sea of Electrons" simulation exploring why metals conduct electricity and deform without shattering compared to rigid ionic crystals.	Diagrammatic models contrasting pure metals, substitutional alloys (similar radii), and interstitial alloys (smaller atoms in holes).	Draw particulate diagrams comparing substitutional vs. interstitial alloys and present an oral argument detailing their mechanical properties.
LT5	"Lewis Structure Speed Challenge" drawing electron dot structures for	Step-by-step Lewis structure protocol card [Valence count -> single bonds ->	Submit hand-drawn or digital Lewis diagrams incorporating resonance

	molecules with expanded octets and resonance.	octets -> formal charge optimization: FC = V - N - (B/2)].	structures and formal charge calculations for assigned species.
LT6	3D molecular modeling lab using physical ball-and-stick kits or digital molecular viewers (e.g., MolView) to inspect molecular geometries.	VSEPR geometry chart connecting electron domains, molecular geometry, bond angles, dipoles, and hybridization.	Construct a digital portfolio or present a 3D model explaining geometry, net dipole moment, and bond order for a complex molecule.

Supporting Multilingual/English Learners (*CELP standards*)

	Emerging	Expanding	Bridging
LT1	Classify substances as ionic, covalent, or metallic using a properties chart (melting point, conductivity) and word banks.	Compare properties of ionic, covalent, and metallic substances using comparative graphic organizers and sentence starters.	Predict and explain bulk substance properties based on bond type, electronegativity differences, and valence electron behavior.
LT2	Identify equilibrium bond length (well minimum) and repulsion/attraction regions on a labeled potential energy curve.	Explain how bond order and atomic radius shift the potential energy minimum depth and position using guided prompts.	Sketch and compare potential energy curves for single, double, and triple bonds, justifying shifts in bond energy and length.
LT3	Arrange positive and negative ion cards into an alternating lattice grid, matching larger/smaller sizes to ion charges.	Explain lattice energy variations using Coulomb's law sentence frames ("*Higher charge leads to stronger attraction because...*").	Construct particulate models of ionic solids and quantitatively/qualitatively justify lattice energy trends using F proportional to (q1 * q2) / r^2.
LT4	Label "electron sea" and "cation cores" on a metallic solid diagram, sorting interstitial vs. substitutional alloy pictures.	Explain malleability and conductivity of metals and alloy structures using structured paragraph frames.	Compare interstitial and substitutional alloys, explaining how lattice alterations impact density, rigidity, and conductivity.
LT5	Draw Lewis structures for simple molecules by following a step-by-step visual checklist (octet rule, valence electron counting).	Construct Lewis diagrams including resonance structures and formal charge calculations using guided templates.	Draw correct Lewis structures for complex or expanded-octet species, evaluating resonance contributors using formal charge analysis.
LT6	Match Lewis diagrams to 3D VSEPR shape names and bond angles using a visual geometry chart.	Predict molecular geometry, bond angles, and net dipole moments using VSEPR sentence frames ("*The shape is ____ because there are ____ electron domains*").	Explain complex molecular geometry, hybridization (sp, sp2, sp3), bond polarity vectors, and overall dipole moments in written detail.

Unit Overview

In this unit, students will learn how to interpret simple graphical representations of changes in potential energy as two atoms approach each other to explain optimal bond length as well as why bonds may or may not occur. Students should also practice constructing representations and models for chemical phenomena (e.g., ionic and metallic solids) and using representations to make claims or predictions. For example, students can use VSEPR theory to draw Lewis structures of molecules and predict their three-dimensional geometry and polarity. Instead of simply connecting chemical theories to phenomena occurring at the atomic level, it is important to provide explanations across scales. For example, teachers can ask students to explain the connection between electronegativity and ionization energy with the type of bond formed and the macroscopic properties of a particular substance. Students should also work with several chemical concepts (Coulomb's law, formal charge, and resonance) to evaluate the accuracy of a model in representing both the particulate-level structure and macroscopic observations. In future units, students will use the practice of constructing and understanding molecular representations to make predictions and claims about interparticle interactions, intermolecular forces, and their connections to macroscopic observations.

Lesson	Learning Target	Success Criteria	Suggested Activities
2.1 Types of Chemical Bonds	**Learning Target 1** I can explain the relationship between the type of bonding and the properties of the elements participating in the bond.	• I can classify a compound as ionic, polar covalent or nonpolar covalent, network ionic or metallic, I can describe the Trends on the periodic table electronegativity. • I can describe the qualitatively relationship between the trends and the electronic structure of the atoms, using the shell model, and Coulomb's law. • I can describe how valence electrons shared between atoms of similar electronegativity constitute a nonpolar covalent bond. For example, bonds between carbon and hydrogen are effectively nonpolar.	

		• I can describe polar covalent bond including how: i. The atom with a higher electronegativity will develop a partial negative charge relative to the other atom in the bond. ii. In single bonds,greater differences in electronegativity lead to greater bond dipoles iii. The difference between ionic and covalent bonds is not distinct but rather a continuum. • I can explain that the best way to characterize the type of bonding in the compound is to examine the properties of a compound • I understand that in a metallic solid, the valence electrons from the metal atoms are considered to be delocalized and not associated with any individual atom.	
2.2 Intramolecular Force and Potential Energy	**Learning Target 2** I can represent the relationship between potential energy and distance between atoms, based on factors that influence the interaction strength.	• I can interpret a graph of potential energy versus the distance between atoms (internuclear distance) as a useful representation for describing the interactions between atoms. Such graphs illustrate both the equilibrium bond length (the separation between atoms at which the potential energy is lowest) and the bond energy (the energy required to separate the atoms). • In a covalent bond, the bond length is influenced by both the size of the atom's core and the bond order (i.e., single, double, triple). Bonds with a higher order are shorter and have larger bond energies. • I can use Coulomb's law to understand the strength of interactions between cations and anions. i. Because the interaction strength is proportional to the charge on each ion, larger charges lead to stronger interactions ii. Because the interaction strength increases as the distance between the centers of the ions (nuclei) decreases, smaller ions lead to stronger interactions.	**Think-Pair-Share** After a review of the graph of potential energy versus internuclear distance in a hydrogen molecule, have students pair up and describe what they believe the graph would look like for various other molecules.
2.3 Structure of Ionic Solids	**Learning Target 3** I can represent an ionic solid with a particulate model that is consistent with Coulomb's law and the properties of the constituent ions	• The cations and anions in an ionic crystal are arranged in a systematic,periodic 3-D array that maximizes the attractive forces among cations and anions while minimizing the repulsive forces.	**Explore Representations** Demonstrate a model of ionic bonding. Put opaque adhesive tape on top of disk magnets to make "+" and "–" signs. Be sure to affix the tape on opposite sides for the differing charges (so that opposite ions have opposite magnetic polarity when arranged on a flat surface). Arrange the ions in an alternating array on the overhead

			projector to show the structure of an ionic crystal. Engage students in a discussion about malleability/ brittleness, and ask why distorting an ionic crystal causes shattering. This also introduces Coulombic forces in a visual and memorable way. Then have students predict and identify the bonding in binary compounds using periodic trends.
2.4 Structure of Metals and Alloys	**Learning Target 4** I can represent a metallic solid and/or alloy using a model to show essential characteristics of the structure and interactions present in the substance.	• Metallic bonding can be represented as an array of positive metal ions surrounded by delocalized valence electrons (i.e.,a"sea of electrons"). • I can describe the differences between types of alloys. Interstitial alloys form between atoms of significantly different radii,where the smaller atoms fill the interstitial space between the larger atoms (e.g., with steel in which carbon occupies the interstices in iron). • Substitutional alloys form between atoms of comparable radius, where one atom substitutes for the other in the lattice. (e.g., in certain brass alloys, other elements, usually zinc, substitute for copper.)	**Manipulatives** Have students use various sized/colored paper plates to illustrate a particular type of alloy (interstitial and/or substitutional). Then have them engage in a gallery walk around the room to listen to others explain the connection between the structure of the different alloys and the properties of each.
2.5 Lewis Diagrams & 2.6 Resonance and Formal Charge	**Learning Target 5** I can represent a molecule with a Lewis diagram.	• I can draw lewis diagrams for compounds including those with expanded octets. • I can use the concept of resonance structures to describe compounds • I can use the octet rules and formal charge as criteria for determining which of several possible valid Lewis diagrams is the best model	**Simulations** Construct various VSEPR shapes using balloons to show the three-dimensional arrangement of atoms in various bonding arrangements. Then use a PhET simulation to help students see the effects of lone pairs and bonding pairs on molecular shape. Students can work on this individually after being shown how to use the interface, or it can be projected and examined as a class. Have students work with the simulation to add/remove bonds and add/remove lone pairs to determine the most likely three-dimensional shape and bond angles in a molecule.
2.7 VSEPR and Hybridization	**Learning Target 6** Based on the relationship between Lewis diagrams, VSEPR theory, bond orders, and bond polarities: I can … i. Explain structural properties of molecules	• I can use VSEPR theory(Coulombic repulsion between electrons) as a basis for predicting the arrangement of electron pairs around a central atom. • I can use both Lewis diagrams and VSEPR theory for predicting electronic and structural properties of	**Simulations** (See 2.5 & 2.6)

	ii. Explain electron properties of molecules.	many covalently bonded molecules and polyatomic ions, including the following: i. Molecular geometry (linear, trigonal planar, tetrahedral, trigonal pyramidal, bent, trigonal bipyramidal,seesaw,T-shaped,octahedral, square pyramidal, square planar) ii. Bond angles iii. Relative bond energies based on bond order iv. Relative bond lengths (multiple bonds, effects of atomic radius) v. Presence of a dipole moment vi. Hybridization Of Valence Orbitals Of Atoms within a molecule or polyatomic ion Vii. sigma and pi bonds. The overlap is stronger in sigma than pi bonds,which is reflected sigma bonds having greater bond energy than pi bonds. The presence of a pi bond also prevents the rotation of the bond and leads to geometric isomers.	

Progress Check 2

- Multiple-choice: ~15 questions
- Free-response: 1 question
    - Long

Preparing for the AP Exam

On the AP Exam, students must be able to construct Lewis structures and make predictions or claims based on them. However, students often struggle to predict the correct molecular shape or bond angle based on VSEPR and the use of formal charge. Mistakes include: using the incorrect number of valence electrons, violating the octet rule, or confusing molecular geometry with bond angles. Teachers can provide students with multiple opportunities to practice drawing Lewis electron-dot diagrams, including resonance structures. Students should also practice predicting and describing molecular shapes, bond angles, and polarities from Lewis structures, and calculating and connecting formal charges in Lewis structures to the predicted structure of a molecule.

Unit Title:

Unit 3: Properties of Substances and Mixtures

Relevant Standards: Bold indicates priority

3.1 Intermolecular & Interparticle Forces: Students can identify and compare intermolecular forces, including dispersion, dipole-dipole, hydrogen bonding, and ion-dipole interactions, and relate them to molecular structure and particle interactions in solution.

3.2 Types of Solids & Properties: Students can classify solids as molecular, ionic, metallic, or network covalent and relate solid type and bonding to macroscopic properties.

3.3 States of Matter: Students can describe solids, liquids, and gases in terms of particle motion, spacing, and intermolecular forces.

3.4 Ideal Gases: Students can use the ideal gas law to relate pressure, volume, temperature, and moles of gas under ideal conditions.

3.5 Gases & Kinetic Molecular Theory: Students can explain gas behavior using kinetic molecular theory assumptions and relate molecular motion to measurable gas properties.

3.6 Deviations from Ideal Behavior: Students can explain how real gases deviate from ideal behavior based on particle volume and intermolecular forces, especially under high pressure and low temperature.

3.7 Solutions & Mixtures: Students can calculate and interpret solution concentration using molarity and related quantitative relationships.

3.8 Particulate Diagrams of Solutions: Students can represent solutions at the particle level to show solute-solvent interactions and relative concentrations.

3.9 Separation of Solutions and Mixtures: Students can describe and select appropriate separation techniques based on differences in physical and chemical properties of mixture components.

3.10 Solubility: Students can predict solubility based on intermolecular forces and explain factors that affect solubility of solutes in solvents.

3.11 Spectroscopy and Electromagnetic Spectrum: Students can describe the electromagnetic spectrum and relate wavelength, frequency, and energy of electromagnetic radiation.

3.12 Photons: Students can explain the particle nature of light and relate photon energy to frequency and electronic transitions.

3.13 Beer's Law: Students can use Beer's Law to relate absorbance to concentration and apply spectrophotometric data to determine solution concentration.

Essential Question(s):	**Enduring Understanding(s):**
1. How do interactions between particles influence the properties of pure substances and mixtures? 2. Why does the smell of perfume only last for a short time? 3. Why can you swim in water, but you can't walk through a wall? 4. How does the spacing and motion of particles relate to a substance's state of matter and the properties of gases? 5. How can you determine the structure and concentration of a chemical species in a mixture?	1. The nature and strength of intermolecular forces determine the physical properties of substances, including phase, boiling point, and solubility. 2. The arrangement and motion of particles in different states of matter are governed by intermolecular forces and kinetic energy. 3. The interactions between solute and solvent particles determine the formation, stability, and concentration of solutions. 4. Interactions between matter and electromagnetic radiation provide experimental evidence for atomic and molecular structure and allow for quantitative analysis of chemical systems.

Demonstration of Learning:	Pacing for Unit
• Success with Unit 3 MCQs and FRQs via AP Classroom • Quizzes (AP Classroom MCQs and FRQs) every few topics • Practice calculations with ideal gas law and combined gas law equations • Molar Mass of a Gas Lab • Particulate view of solids, liquids, gases, solutions • Making Solutions / Beer's Law Lab • Progress Check 3 from AP Classroom Unit 3 is 18–22% AP Exam Weighting See AP Exam Description	18 days of instruction 3 days of lab 3 days review / assess
Family Overview (link below)	**Integration of Technology:**
AP Chemistry FAMILY MATERIALS	• Use of LabQuests / colorimeters to make and test solutions for Beer's Law • Google Suite to graph digitally
Unit-specific Vocabulary:	**Aligned Unit Materials, Resources, and Technology (beyond core resources):**
state of matter, solid, liquid, gas, intermolecular forces, intramolecular forces, chemical bond, ionic bond, covalent bond, metallic bond, London dispersion forces, polarizability, dipole–dipole forces, hydrogen bonding, polar molecule, nonpolar molecule, electronegativity, electronegativity difference, bond polarity, molecular polarity, electron density, partial charge, molecular geometry, Lewis structure, VSEPR theory, electron domain, bonding pair, lone pair, bond angle, linear, bent, trigonal planar, trigonal pyramidal, tetrahedral, square planar, square pyramidal, octahedral, trigonal bipyramidal, hybridization, sigma bond, pi bond, single bond, double bond, triple bond, bond length, bond energy, lattice energy, Coulomb's law, melting point, boiling point, vapor pressure, volatility, surface tension, viscosity, capillary action	Colorimeter (AP suggests using a spectrophotometer, as of 8/2026 this equipment is not available for use in BPS)
Opportunities for Interdisciplinary Connections:	**Anticipated misconceptions:**
• Biology: membrane structure, protein folding, transport of substances • Earth Science: water cycle, atmospheric gases, climate systems • Physics: kinetic molecular theory and gas behavior • Environmental Science: pollution, solubility, water treatment • Medical Science: spectroscopy and diagnostic imaging	• With practice, students can identify the type of IMFs present within a sample, but may struggle to identify which explains specific observations such as boiling or melting point • When given data that contradicts typical behavior, students may struggle to find reasonable explanation as to why • Confusing intramolecular and intermolecular • Difficulty graph reading with Beer's Law • Difficulty picking appropriate equation for gas law calculations • Deviations from ideal gas: identifying which factor causes a gas to deviate from ideal behavior, but difficulty in explaining why that factor causes a deviation.
Connections to Prior Units:	**Connections to Future Units:**

• Type of bond / valence electrons dictates how a substance dissolves
• Type of bond is determined by electronegativity difference, which also dictates intermolecular forces
• Properties of photons, spectroscopy, and the electromagnetic spectrum relate to excitation of electrons (introduced in unit 1)
	• Making solutions comes up throughout the remainder of the curriculum
	  ○ Making solutions in lab
	  ○ Calculating the molarity (i.e. of an acidic or basic solution (Unit 8)
	• How a substance dissolves (i.e. ionic dissociates vs molecular compounds do not) is relevant for net ionic equations (4.2) and acids and bases (unit 8)
	• Kinetic molecular theory returns in unit 5 to explain rate of reactions

Differentiation through *Universal Design for Learning*

	Engagement	Representation	Action & Expression
LT1	Surface tension and evaporative cooling race measuring temperature drops of organic liquids on paper strips.	Structural comparison chart detailing London Dispersion Forces (polarizability/electron cloud size), Dipole-Dipole, and Hydrogen Bonding.	Write a comparative analysis or record a video explaining boiling point trends across homologous series based on IMF strength.
LT2	"Solid Classification Mystery" sorting mystery materials into ionic, metallic, covalent network, or molecular solid classes.	Visual taxonomy matrix linking solid type (e.g., Diamond vs. Ice vs. Cu vs. NaCl) to melting point, hardness, and conductivity.	Design a digital infographic or deliver an oral presentation mapping microscopic solid interactions to macroscopic material properties.
LT3	Phase change animation challenge illustrating particle motion, spacing, and intermolecular interaction breaks.	Side-by-side particulate box diagrams depicting particle arrangement, freedom of motion, and density across solid, liquid, and gas states.	Draw or animate particulate representations of a substance undergoing phase changes and defend particle density and organization.
LT4	"Gas Law Simulator" adjusting volume, temperature, and moles to observe real-time pressure responses on a digital gauge.	Formula organizer highlighting PV = nRT, Dalton's Law of Partial Pressures, and mole fraction relationships.	Solve gas stoichiometry and partial pressure problems via handwritten calculation steps, calculator logs, or video tutorials.
LT5	Maxwell-Boltzmann distribution interactive exploration observing how temperature shifts speed profiles of light vs. heavy gas molecules.	Annotated Maxwell-Boltzmann curve graphs (X-axis: molecular speed; Y-axis: number of molecules) showing temperature and molar mass shifts.	Explain the Kinetic Molecular Theory assumptions and interpret a Maxwell-Boltzmann distribution graph via a written essay or oral report.
LT6	High-pressure/low-temperature real gas simulation demonstrating deviations from ideal PV/RT = 1 baselines.	Comparative visual diagram contrasting ideal point-mass particles with real gas particles possessing finite volume and IMFs.	Write a technical memo evaluating why a specific real gas deviates from ideal behavior under extreme pressure and temperature conditions.
LT7	Solution preparation lab making precise molarity solutions and executing serial dilution procedures.	Volumetric calculation guide highlighting M = n / V and dilution formula M1V1 = M2V2 with step-by-step unit checks.	Complete solution preparation calculations and submit a recorded laboratory demonstration of accurate volumetric technique.
LT8	"Microscopic Mixer Challenge" drawing solute and solvent particles to accurately reflect molar concentration and ion-dipole orientation.	Particulate diagrams showing ion-dipole orientation (e.g., water oxygen facing Na+, hydrogen facing Cl-) and solute particle counts.	Create particulate drawings representing solutions of varying concentrations and defend the orientation of solvent molecules around ions.
LT9	Paper or thin-layer chromatography (TLC) lab separating plant pigments or dye mixtures using polar vs. nonpolar solvents.	Annotated chromatogram diagrams mapping retention factor to stationary vs. mobile phase IMF affinities.	Calculate R_f values from experimental TLC plates and write an explanation linking component travel distance to relative polarity.
LT10	"Like Dissolves Like" lab testing miscibility of polar and nonpolar solute-solvent combinations.	Energetic solubility diagram illustrating solute-solute breaking, solvent-solvent breaking, and solute-solvent forming interactions.	Write a CER justification explaining why certain vitamins (e.g., Vitamin A vs. Vitamin C) are fat-soluble or water-soluble based on IMFs.
LT11	EM spectrum matching activity linking photon wavelengths (UV/Vis, IR, Microwave) to molecular motion types.	EM spectrum chart displaying energy transitions: Microwave -> rotation, Infrared -> vibration, UV/Vis -> electronic transitions.	Construct an explanatory anchor chart or record a video mapping light frequency to specific molecular/atomic excitation modes.

LT12	Flame test or gas discharge tube line emission spectroscopy lab identifying mystery gases.	Energy level diagram showing photon absorption (arrow up) and emission (arrow down) linked to E = hν = hc / λ.	Calculate photon energy, frequency, or wavelength for an electronic transition and present a worked solution set.
LT13	Spectrophotometry / Beer-Lambert Law lab constructing a calibration curve to find the concentration of an unknown dye solution.	Beer-Lambert Law formula scaffold (A = εbc) alongside an absorbance vs. concentration calibration graph.	Construct a Beer-Lambert calibration plot using spreadsheet software and calculate unknown sample concentration with error analysis.

Supporting Multilingual/English Learners (*CELP standards*)

	Emerging	Expanding	Bridging
LT1	Identify IMF types (London dispersion, dipole-dipole, hydrogen bonding) using a flowchart and visual structural formulas.	Compare IMF strengths between pure species or mixtures using comparative writing frames (*"Species A has stronger IMFs because..."*).	Predict and justify relative IMF strengths based on polarizability, dipole moments, and hydrogen bonding capacity in detail.
LT2	Match solid types (ionic, network covalent, metallic, molecular) to visual structural models and physical property cards.	Describe relationships between solid types and properties (vapor pressure, melting point) using graphic organizers.	Explain differences in physical properties among network covalent, molecular, ionic, and metallic solids based on particulate forces.
LT3	Select correct particulate representations for solids, liquids, and gases based on particle spacing and disorder.	Describe phase changes and particle motion differences across states using structured comparison frames.	Draw and critique particulate models representing phase transitions, incorporating density anomalies and IMF disruption.
LT4	Identify variables in PV = nRT and calculate missing values using structured algebra templates and unit guides.	Solve gas law problems (PV = nRT, Dalton's Law) and explain variable relationships using proportional reasoning frames.	Derive and apply gas law relationships (P, V, T, n) in multi-step stoichiometry and partial pressure scenarios with written justifications.
LT5	Match Maxwell-Boltzmann distribution curves to temperature/mass labels using a visual guide.	Explain relationships between temperature, molar mass, and particle speed using KMT sentence starters.	Interpret Maxwell-Boltzmann distributions and particle collision dynamics to explain pressure, temperature, and effusion rate differences.
LT6	Identify high pressure and low temperature conditions as "non-ideal" on a gas behavior diagram.	Explain why real gases deviate from ideal behavior at low T or high P using sentence frames mentioning IMF and particle volume.	Analyze deviations from ideal gas behavior (Z = (P * V) / (R * T) != 1) quantitatively and qualitatively using real gas particle interactions.
LT7	Calculate molarity (M = moles / Liters) using a formula triangle and unit conversion cards.	Perform solution concentration and dilution calculations (M1 * V1 = M2 * V2) using step-by-step problem templates.	Solve complex solution stoichiometry and ion concentration calculations, writing out full conversion steps and justifications.
LT8	Draw solute particles dispersed in solvent molecules using provided ratio boxes and particle icons.	Describe solute-solvent interactions (e.g., ion-dipole) and relative concentration changes using particulate model frames.	Construct precise particulate representations showing ion orientation in solvation shells and relative stoichiometry in mixtures.
LT9	Match chromatography components (stationary/mobile phase) or distillation fractions to polarity labels.	Explain paper chromatography or distillation separation results using polar/nonpolar interaction sentence stems.	Analyze retention factors (Rf) or boiling points to evaluate mixture separation efficiency based on specific IMFs.
LT10	Classify combinations as "soluble" or "insoluble" using the "like dissolves like" rule and visual polarity cards.	Explain solubility differences in water vs. nonpolar solvents using ion-dipole and dipole-dipole interaction frames.	Justify solubility trends of organic and ionic compounds in polar/nonpolar solvents using enthalpy of solution and IMF analysis.
LT11	Match EM regions (UV/Vis, IR, Microwave) to transition types (electronic, vibrational, rotational) using a color chart.	Explain which EM radiation type causes specific molecular motions using guided matching prompts and sentence starters.	Articulate energy relationships (Delta E = h * nu) connecting photon frequencies to electronic, vibrational, and rotational transitions.
LT12	Label ground state, excited state, absorption, and emission on an atomic energy level diagram.	Calculate photon energy, frequency, or wavelength (Delta E = h * nu = (h * c) / lambda) using formula reference cards.	Relate energy level gaps to emission/absorption spectra lines, quantitatively connecting energy levels to photon properties.

LT13	Identify Beer's Law variables (A = e * b * c) and read absorbance values from a calibration curve graph.	Explain how absorbance changes with concentration or path length using direct-proportionality sentence frames.	Analyze Beer's Law data (A = e * b * c) to determine unknown concentrations and evaluate experimental errors (e.g., cuvette smudges).

Unit Overview

This unit requires students to draw upon claims made in Unit 2 about molecular geometry and polarity to support claims about intermolecular forces between molecules. Further, students will practice illustrating such claims by constructing particle representations of pure solids, liquids, gases, and solutions. This unit also requires students to build proficiency with mathematical reasoning skills, essential for success in the remainder of the course. Students should be able to explain relationships between variables in an equation (e.g., the ideal gas law) and then estimate the approximate value of one variable within an equation when the value of another variable changes. Students will practice these skills when choosing and implementing experimental procedures, making observations, and/or collecting data to address a question. Students can then determine the accuracy and precision of the data as well as manipulate it with known mathematical equations to support their claims (e.g., concentration of a substance, properties of substances in a mixture).

Lesson	Learning Target	Success Criteria	Suggested Activities
3.1 Intermolecular and Interparticle Forces	**Learning Target 1** I can explain the relationship between the chemical structures of molecules and the relative strength of their intermolecular forces when: i. The molecules are of the same chemical species. ii. The molecules are of two different chemical species.	• I can describe how London dispersion forces are a result of the Coulombic interactions between temporary, fluctuating dipoles.London Dispersion Forces are often the strongest net intermolecular force between large molecules. i. Dispersion Forces Increase With Increasing contact area between molecules and with increasing polarizability the molecules. ii. The Polarizability Of Molecule Increases with an increasing number of electrons in the molecule and size of the electron cloud. It is enhanced by the presence of pi bonding. iii. The term “London dispersion forces” should not be used interchangeably with “van der waals “ forces • I can explain how the dipole moment of a polar molecule leads to additional interactions with other chemical species. i. Dipole-induced dipole interactions are between a polar and nonpolar molecule. These forces are always attractive. The strength of these forces increases with the magnitude of the dipole of the polar molecule and with the polarizability of the nonpolar molecule. ii. Dipole dipole interactions occur between polar molecules. Iii Ion-dipole forces of attraction are present between ions and polar molecules.	**Demo with Q&A** Fill a long glass tube halfway with water and then layer ethanol over the top and fill the tube, leaving one inch at the top. Have a student mark the liquid level with a permanent marker and invert the tube (with thumb pressed firmly over the top) several times. A noticeable volume decrease occurs, and students should hypothesize why. Introduce a model showing the interparticle spacing between ethanol molecules and water molecules. The model takes into account the spacing between molecules and why volume is not a conserved quantity (unlike mass). Review hydrogen bonding as a relevant interparticle force for this demonstration.

		• Hydrogen bonding is a strong type of intermolecular interaction that exists when hydrogen atoms covalently bonded to the highly electronegative atoms (N, O, and F) are attracted to the negative end of a dipole formed by the electronegative atom (N, O, and F) in a different part of the same molecule. • I can describe hydrogen bonding as a strong type of intermolecular interaction that exists when hydrogen atoms covalently bonded to the highly electronegative atoms (N, O, and F) are attracted to the negative end of a dipole formed by the electronegative atom (N, O, and F)indifferent molecule,or different part of the same molecule.	
3.2 Properties of Solids	**Learning Target 2** I can describe the properties of solids and explain the relationship among the macroscopic properties of a substance, the particulate-level structure of the substance, and the interactions between these particles.	• I can describe the relationship between the properties of liquids and solids and the strengths and types of intermolecular forces present. Because intermolecular interactions are overcome completely when a substance vaporizes,the vapor pressure and boiling point are directly related to the strength of those interactions. Melting points also tend to correlate with interaction strength, but can be more subtle. • I can use Particulate-level representations, to communicate or understand how intermolecular interactions help to establish macroscopic properties. • I can describe ionic solids and that due to strong interactions between ions, ionic solids tend to have low vapor pressures, high melting points, and high boiling points. They tend to be brittle due to the repulsion of like charges caused when one layer slides across another layer. They conduct electricity only when the ions are mobile, as when the ionic solid is melted (i.e., in a molten state) or dissolved in water or another solvent. • I can describe covalent network solids, the atoms are covalently bonded together into a three dimensional network (e.g., diamond) or layers of two-dimensional networks(e.g.,graphite). These are only formed from nonmetals and metalloids: elemental (e.g., diamond, graphite) or binary compounds (e.g., silicon dioxide and silicon carbide). Due to the strong covalent interactions, covalent solids have high melting points.Three-dimensional network solids are also rigid and hard, because the covalent bond angles are fixed. However ,graphite is soft because adjacent layers can slide past each other relatively easily. 3.2.A.5 Molecular solids are composed of distinct, individual units of covalently-bonded molecules attracted to each other through relatively weak intermolecular forces. Molecular solids generally have a low melting point because of the relatively weak intermolecular forces present between the molecules. They do not conduct electricity because their valence electrons are tightly held within the covalent bonds and the lone pairs of each constituent molecule. Molecular solids are sometimes composed of very large molecules or polymers. • I can describe Metallic solids as good conductors of electricity and heat, due to the presence of free valence electrons. They also tend to be malleable and ductile, due to the ease with which the metal cores can rearrange their structure. In an interstitial alloy, interstitial atoms tend to make the lattice more rigid, decreasing malleability and ductility. Alloys typically retain a sea of mobile electrons and so remain conducting. • I can explain how large biomolecules (or polymers), noncovalent	

		interactions may occur between different molecules or between different regions of the same large biomolecule. The functionality and properties of such molecules depend strongly on the shape of the molecule, which is largely dictated by noncovalent interactions	
3.3 Solids, Liquids, and Gases	**Learning Target 3** I can represent the differences between solid, liquid, and gas phases using a particulate level model.	• I can explain how solids can be crystalline, where the particles are arranged in regular three-dimensional structure, or they can be amorphous, where the particles do not have a regular, orderly arrangement. In both cases, the motion of the individual particles is limited, and the particles do not undergo overall translation with respect to each other. The structure of the solid is influenced by interparticle interaction and the ability of the particles to pack together • I can explain how the constituent particles in liquids are in close contact with each other, and they are continually moving and colliding. The arrangement and movement of particles are influenced by the nature and strength of the forces (e.g., polarity, hydrogen bonding, and temperature) between the particles. • I can explain that solid and liquid phases for a particular substance typically have similar molar volume because, in both phases, the constituent particles are in close contact at all times. • I can explain that in the gas phase, the particles are in constant motion. Their frequencies of collision and the average spacing between them are dependent on temperature, pressure, and volume. Because of this constant motion,and minimal effects of forces between particles, a gas has neither a definite volume nor definite shape.	**Explore Representations** Have students create particle representations for samples of solid, liquid, and gaseous H2O. Each diagram should contain 10 molecules, and students should show how the placement and motion of the particles varies in each phase.
3.4 Ideal Gas Law	**Learning Target 4** Explain the relationship between the macroscopic properties of a sample of gas or mixture of gases using the ideal gas law. EQN: PV = nRT.	• I can describe gas mixtures such that in a sample containing a mixture of ideal gases, the pressure exerted by each component (the partial pressure) is independent of the other components. Therefore, the partial pressure of a gas within the mixture is proportional to its mole fraction (X), and the total pressure of the sample is the sum of the partial pressures.	
3.5 Kinetic Molecular Theory	**Learning Target 5** I can explain the relationship between	• I can explain the kinetic molecular theory (KMT) that relates the macroscopic properties of gases to motions of the particles in the gas. The Maxwell-Boltzmann Distribution Describes The distribution of the	

	the motion of particles and the macroscopic properties of gases with: i. The kinetic molecular theory (KMT). ii. A particulate model. iii. A graphical representation.	kinetic energies of particles at a given temperature. Also that all the particles in a sample of matter are in continuous, random motion. The average kinetic energy of a particle is related to its average velocity by the equation: EQN: KE = ½ mv2 . • The Kelvin temperature of a sample of matter is proportional to the average kinetic energy of the particles in the sample. TheMaxwell-Boltzmanndistributionprovides a graphical representation of the energies/ velocities of particles at a given temperature.	
3.6 Deviation from Ideal Gas Law	**Learning Target 6** I can explain the relationship among non-ideal behaviour of gases, interparticle forces, and/or volumes.	• The ideal gas law does not explain the actual behavior of real gases. Deviations from the ideal gas law may result from interparticle attractions among gas molecules, particularly at conditions that are close to those resulting in condensation. Deviations may also arise from particle volumes, particularly at extremely high pressures.	
3.7 Solutions and Mixtures	**Learning Target 7** I can calculate the number of solute particles, volume, or molarity of solutions	• I can describe solutions as sometimes called homogeneous mixtures, which can be solids, liquids, or gases. In a solution, the macroscopic properties do not vary throughout the sample. In a heterogeneous mixture, the macroscopic properties depend on location in the mixture. • Solution composition can be expressed in a variety of ways; molarity is the most common method used in the laboratory. EQN: M = nsolute /Lsolution	**Explore Representations** Begin by telling students that hexane does not mix with water, but ethanol does. Then have them create a particulate representation of each of the mixtures (which illustrate the interactions between the molecules that allow/disallow the solubility).
3.8 Representations of Solutions	**Learning Target 8** I can use particulate models for mixtures to : i. Represent interactions between components. ii. Represent concentrations of components.	• I can use particulate representations of solutions to communicate the structure and properties of solutions, by illustration of the relative concentrations of the components in the solution and/or drawings that show interactions among the components.	**Explore Representations** (See 3.7)
3.9 Separation of Solutions and Mixtures	**Learning Target 9** I can explain the results of a separation experiment based on intermolecular interactions.	• I can explain how the components of a liquid solution cannot be separated by filtration.They Can,however, be separated using processes that take advantage of differences in intermolecular interactions of the components. • i. Chromatography(paper,thin-layer,and column) separates chemical species by taking advantage of the differential strength of intermolecular interactions between	**Post-Lab Discussion** After investigating three different dyes using chromatography, have students determine which of the three dyes is the most polar based on macroscopic observations and an understanding of the interactions between the dyes and the solvent, or between the dyes and the paper. Then have them discuss their answers (based on evidence) and evaluate the strengths of each other's

		and among the components of the solution (the mobile phase) and with the surface components of the stationary phase. The resulting chromatogram can be used to infer the relative polarities of components in a mixture. • ii. Distillation separates chemical species by taking advantage of the differential strength of intermolecular interactions between and among the components and among the components and the effects these interactions have on the vapor pressures of the components in the mixture.	claims using both the evidence and understanding of intermolecular forces.
3.10 Solubility	**Learning Target 10** I can explain the relationship between the solubility of ionic and molecular compounds in aqueous and nonaqueous solvents, and the intermolecular interactions between particles.	• Substances with similar intermolecular interactions tend to be miscible or soluble in one another.	
3.11 Spectroscopy and the Electromagnetic Spectrum	**Learning Target 11** I can explain the relationship between a region of the electromagnetic spectrum and the types of molecular or electronic transitions associated with that region.	• I can describe the difference in absorption or emission of photons in different spectral regions is related to the different types of molecular motion or electronic transition: i. Microwave radiation is associated with transitions in molecular rotational levels. ii. Infrared radiation is associated with transitions in molecular vibrational levels. iii. Ultraviolet/visible radiation is associated with transitions in electronic energy levels.	
3.12 Properties of Photons	**Learning Target 12** I can explain the properties of an absorbed or emitted photon in relationship to an electronic transition in an atom or molecule.	• When a photon is absorbed (or emitted) by an atom or molecule, the energy of the species is increased (or decreased) by an amount equal to the energy of the photon. The wavelength of the electromagnetic wave is related to its frequency and the speed of light by the equation: EQN: $c = \lambda\nu$. • The energy of a photon is related to the frequency of the electromagnetic wave through Planck's equation: EQN: $E = \hbar\nu$.	
3.13 Beer-lambert law	**Learning Target 13** I can explain the amount of light absorbed by a solution of molecules or ions in relation to the concentration, path length, and molar absorptivity.	• I can use TheBeer-Lambert law to relate the absorption of light by a solution to three variables according to the equation: EQN: $A = \varepsilon bc$. • The molar absorptivity, ε, describes how intensely a chemical species absorbs light of a specific wavelength. The Pathlength,b, and concentration, c, are proportional to	**Predict and Confirm** Have students use a Sep-Pak C18 Cartridge (Flinn Scientific AP8917) to separate Grape Kool-Aid into its component red and blue dyes. Then have them compare the separated dyes to reference solutions of common food dyes using a spectrophotometer and measure

		the number light-absorbing particles in the lit path. In most experiments the path length and wavelength of light are held constant. In such cases,the absorbance is proportional only to the concentration of absorbing molecules or ions. The Spectrophotometer Is typically set to the wavelength of maximum absorbance (optimum wavelength) for the species being analyzed to ensure the maximum sensitivity of measurement.	the percent transmittance at 25 nm intervals across the range of 400 nm–750 nm.

Progress Check 3

- Multiple-choice: ~30 questions
- Free-response: 2 questions
    - Short
    - Short

Preparing for the AP Exam

On the AP Exam (in both the multiple-choice and the free-response section), students are required to compare the physical properties of substances and relate them to the attractive forces between particles. Students often struggle with questions that require them to determine the forces of attraction that are present between molecules. Moreover, it can be challenging for them to determine which forces are most important in explaining the differences in physical properties, such as melting points, boiling points, and vapor pressures of molecules in the solid and/or liquid state. Students also confuse the terms intramolecular and intermolecular forces. Another common mistake students make is to simplify their explanations about governing intermolecular forces in a substance by using terms such as "strong" and "weak." Teachers can ensure that students can identify an actual intermolecular force and explain its strength in relation to other forces at play.

Unit Title:

Unit 4: Chemical Reactions

Relevant Standards: Bold indicates priority

4.1 Introduction for Reactions: Students can represent chemical reactions using balanced chemical equations and identify physical states and conservation of matter in reactions.

4.2 Net Ionic Equations: Students can write complete and net ionic equations and identify spectator ions in aqueous reactions.

4.3 Representations of Reactions: Students can represent chemical reactions at the macroscopic, particulate, and symbolic levels to demonstrate conservation of matter.

4.4 Physical and Chemical Changes: Students can distinguish between physical and chemical changes based on particle rearrangement and observable evidence.

4.5 Stoichiometric Relationships: Students can use mole ratios from balanced equations to relate quantities of reactants and products.

4.6 Introduction to Titration: Students can analyze titration data to determine the concentration of an unknown solution using stoichiometric relationships.

4.7 Types of Chemical Reactions: Students can classify chemical reactions as precipitation, acid-base, or redox based on reactants, products, and observable changes.

4.8 Introduction to Acid-Base Reactions: Students can identify acid-base reactions and represent proton transfer processes using chemical equations.

4.9 Oxidation-Reduction Reactions: Students can identify redox reactions, assign oxidation states, and determine electron transfer between species.

Essential Question(s):	**Enduring Understanding(s):**
1. What makes fireworks explode? 2. In what ways can a chemical change be described and documented? 3. How can you predict that a chemical reaction will generate enough product?	1. Chemical reactions involve the rearrangement of atoms through the breaking and forming of bonds, resulting in the conservation of matter. 2. Balanced chemical equations and net ionic equations provide multiple representations of chemical reactions that describe changes at the macroscopic, particulate, and symbolic levels. 3. Stoichiometric relationships in balanced reactions allow chemists to predict the quantities of reactants and products and analyze experimental data such as titrations. 4. Chemical reactions can be classified by patterns of behavior, including acid-base, precipitation, and redox reactions, based on particle-level interactions and electron transfer.
Demonstration of Learning:	**Pacing for Unit**

• Success with unit MCQs and FRQs via AP Classroom
• Quizzes (AP Classroom MCQs and FRQs) every few topics
• Balancing chemical reaction practice
• Stoichiometry calculation practice
  ○ Limiting and excess calculations
  ○ Solution stoichiometry
• Chemical vs Physical Change Lab
• Precipitation Lab
	9 days of instruction
	2 days of lab
	2 days review / assess

• Practice with determining oxidation/ reduction numbers and writing out half reactions • Progress Check 4 from AP Classroom Unit 4 is 7-9% of AP Exam Weighting See AP Exam Description	
Family Overview (link below)	**Integration of Technology:**
AP Chemistry FAMILY MATERIALS	N/A
Unit-specific Vocabulary:	**Aligned Unit Materials, Resources, and Technology (beyond core resources):**
chemical reaction, reactant, product, chemical equation, balanced equation, stoichiometric coefficient, law of conservation of mass, combination reaction, decomposition reaction, single-replacement reaction, double-replacement reaction, combustion reaction, precipitation reaction, acid–base reaction, oxidation–reduction reaction, oxidation, reduction, oxidizing agent, reducing agent, oxidation number, redox reaction, spectator ion, total ionic equation, net ionic equation, solubility, solubility rules, aqueous solution, electrolyte, strong electrolyte, weak electrolyte, nonelectrolyte, limiting reactant, excess reactant, theoretical yield, actual yield, percent yield, mole ratio, stoichiometry, mass–mass calculation, mass–mole calculation, mole–mole calculation	N/A
Opportunities for Interdisciplinary Connections:	**Anticipated misconceptions:**

• Environmental Science: combustion, acid rain, water chemistry
• Biology: cellular respiration and metabolic reactions
• Engineering: industrial synthesis and reaction efficiency
• Mathematics: stoichiometry and proportional reasoning
• Earth Science: geochemical cycling and mineral formation
	• Not following the law of conservation of matter when balancing
	• Difficulty with stoichiometry calculations
	  ○ Comparing mass to mass when determining limiting / excess
	  ○ Not using dimensional analysis correctly when doing stoich problems
	• Difficulty with net ionic equations
	  ○ Including spectator ions
	  ○ Not including correct charge (must memorize main group charges, common transition metals, and common polyatomic ions)
	• Confusing percent yield and percent error calculations

Connections to Prior Units:	**Connections to Future Units:**
• Electrons determine charge in ionic compounds– need to know in order to write out proper net ionic equation • Type of substance (ionic vs molecular (unit 2)) determines how to write out net ionic equation • Type of substance (unit 2) determines type of reaction	• Net ionic equations show up throughout the curriculum – in unit 7 for equilibrium expressions, unit 8 for acid/ base equilibrium and reactions, unit 9 for oxidation / reduction half reactions • Oxidation reduction returns in unit 9 for electrochemistry

Differentiation through *Universal Design for Learning*

	Engagement	Representation	Action & Expression
LT1	"Chemical vs. Physical Station Lab" observing color changes, gas evolution, precipitation, temperature shifts, and phase changes.	T-chart matrix contrasting chemical changes (bond breaking/forming, new substances) with physical changes (interparticle distance shifts).	Record a video walkthrough of demonstration stations classifying each event as physical or chemical based on observational evidence.
LT2	"Net Ionic Equation Puzzles" balancing complete molecular, total ionic, and net	Step-by-step transformation guide showing how to split strong	Write balanced molecular, complete ionic, and net ionic equations for

	ionic equations by removing spectator ions.	electrolytes into ions while keeping solids, liquids, and weak acids intact.	assigned reactions using digital whiteboards or problem sets.
LT3	LEGO or atomic sphere building activity modeling conservation of atoms across chemical reactions.	Before-and-after reaction box diagrams displaying conservation of matter and correct stoichiometric atom counts.	Draw particulate representations of a chemical reaction before and after completion, correctly depicting limiting vs. excess reactants.
LT4	Thermite vs. ice melting comparison analyzing intramolecular covalent/ionic bond breaking vs. intermolecular force disruption.	Energy profile diagrams contrasting high-energy intramolecular bond changes (chemical) with lower-energy IMF changes (physical).	Write a comparative essay or record a video explaining why chemical changes generally involve greater energy changes than physical processes.
LT5	"S'mores Stoichiometry" lab deriving limiting reactants, theoretical yield, and excess reactant remaining.	BCA (Before-Change-After) table scaffold for organizing mole-based stoichiometric calculations.	Solve limiting reactant stoichiometry problems by submitting BCA tables, handwritten calculations, or video explanations.
LT6	Microscale acid-base titration lab using indicators and pH meters to determine an unknown acid concentration.	Titration apparatus schematic paired with stoichiometric equivalence calculation workflow (n_acid = n_base at equivalence point).	Write a formal laboratory report detailing titration execution, equivalence point calculation, and experimental error sources.
LT7	"Reaction Classification Challenge" sorting unlabelled chemical equations into acid-base, redox, or precipitation categories.	Diagnostic decision tree evaluating proton transfer (acid-base), electron transfer/oxidation number change (redox), or solid formation.	Classify mystery chemical reactions and defend decisions using oxidation state changes, proton transfers, or solubility rules.
LT8	Brønsted-Lowry acid-base matching game identifying conjugate acid-base pairs in solution.	Structural reaction diagram highlighting proton (H+) transfer from donor (acid) to acceptor (base) forming conjugate pairs.	Write balanced Brønsted-Lowry acid-base net ionic equations and identify conjugate pairs for given reaction prompts.
LT9	"Electrochemistry Half-Reaction Balancing Lab" balancing complex redox reactions in acidic or basic solutions.	Half-reaction balancing method chart (Balance key elements -> O with H2O -> H with H+ -> e- for charge -> equalize e-).	Balance complex acidic/basic redox equations on whiteboards or submit step-by-step half-reaction calculation sheets.

Supporting Multilingual/English Learners (*CELP standards*)

	Emerging	Expanding	Bridging
LT1	Sort real-life observation cards (color change, gas produced, melting, dissolving) into Physical vs. Chemical bins.	Describe observations indicating chemical vs. physical changes using evidence-based sentence starters.	Differentiate physical vs. chemical changes at macroscopic and microscopic levels, justifying edge cases (e.g., salt dissolving).
LT2	Balance simple chemical equations by counting atom icons on reactant and product sides.	Write balanced molecular, complete ionic, and net ionic equations using solubility rules and step-by-step guides.	Predict reaction products, write net ionic equations for complex systems, and assign spectator ions with full written justifications.
LT3	Match reactant particle diagrams to correct product particle diagrams ensuring conservation of mass.	Draw particulate models showing reactants before and products after a reaction, accounting for limiting reactants.	Construct accurate particulate diagrams representing stoichiometry, limiting reactants, physical state changes, and ion conservation.
LT4	Classify processes as bond breaking (requires energy) or bond forming (releases energy) using visual cards.	Explain macroscopically observed temperature changes during physical/chemical processes using bond-breaking/forming frames.	Analyze energy changes in physical and chemical processes by comparing the relative strengths of bonds/IMFs broken vs. formed.
LT5	Complete BCA (Before-Change-After) tables for stoichiometry using simple whole-number particle counts.	Solve stoichiometry, limiting reactant, and percent yield problems using structured BCA tables and calculation steps.	Execute complex limiting reactant, theoretical yield, and excess reactant stoichiometry problems with complete written work.
LT6	Identify equivalence point volume on a titration curve graph and label burette/flask equipment.	Calculate analyte concentration at the equivalence point using titration stoichiometry templates (MA * VA = MB * VB for 1:1).	Design a titration procedure, analyze experimental titration data, and calculate analyte concentrations for non-1:1 stoichiometry.

LT7	Sort chemical reaction equations into Acid-Base, Redox, or Precipitation categories using a visual key.	Classify reaction types by identifying proton transfer, electron transfer, or precipitate formation using sentence frames.	Categorize complex reactions and justify classifications by tracking oxidation numbers, H+ transfer, or precipitate solubility.
LT8	Identify Bronsted-Lowry acids (H+ donor) and bases (H+ acceptor) in simple equations using color coding.	Write acid-base neutralization equations and identify conjugate acid-base pairs using guided templates.	Predict products of strong/weak acid-base reactions and write net ionic equations for various neutralization combinations.
LT9	Assign oxidation numbers to elements in simple binary compounds using oxidation rule reference cards.	Balance simple redox reactions by splitting them into oxidation and reduction half-reactions using guided steps.	Balance complex redox reactions in acidic or basic solutions using the half-reaction method, tracking electron conservation.

Unit Overview

In Unit 4, students will describe and construct equations of chemical systems and learn to balance those equations. Students should be able to identify and effectively represent types of reactions (e.g., acid-base, redox, precipitation) and then use that knowledge to make hypotheses or predictions about the outcome of a reaction. Additionally, students should be able to support their claims about the identity and amount of product yield through evidence gained with both experimentation and the principles of stoichiometry. Further, students should be able to determine the output of a reaction when the number of moles of reactants change or are in limited/excess supply. This practice of effectively representing balanced chemical equations and using stoichiometry to calculate outcomes of such reactions is critical to student success in the remainder of the course.

Lesson	Learning Target	Success Criteria	Suggested Activities
4.1 Introduction for Reactions	**Learning Target 1** I can identify evidence of chemical and physical changes in matter.	• I can define a physical change as occurring when a substance undergoes a change in properties but not a change in composition. Changes in the phase of a substance (solid, liquid, gas) or formation/ separation of mixtures of substances are common physical changes. • I can define a chemical change as occurring when substances are transformed into new substances, typically with different compositions.Production Of heat or light, formation of a gas, formation of a precipitate, and/or color change provide possible evidence that a chemical change has occurred.	
4.2 Net Ionic Equations	**Learning Target 2** I can represent changes in matter with a balanced chemical or net ionic equation: i. For physical changes. ii. For given information about the identity of the reactants and/or product iii. For ions in a given chemical reaction.	• I can represent physical and chemical processes symbolically by balanced equations. These processes are the result of a rearrangement of atoms into new combinations; thus, any representation of a chemical change must contain equal numbers of atoms of every element before and after the change occurred. Equations thus demonstrate that mass and charge are conserved in chemical reactions. • Balanced molecular, complete ionic, and net ionic equations are different symbolic forms used to represent a chemical reaction. The form used to represent the reaction depends on the context in which it is to be used.	**Explore Representations** Have students work through an online simulation of particulate-level representations of various single-displacement reactions. Then have them translate these particle level views into net ionic equations
4.3 Representations of	**Learning Target 3** I can represent a given chemical reaction or	• I can balance chemical equations in their various forms and translate them into	**Explore Representations** (See 4.2)

Reactions	physical process with a consistent particulate model.	symbolic particulate representations.	
4.4 Physical and Chemical Changes	**Learning Target 4** I can explain the relationship between macroscopic characteristics and bond interactions for: i. Chemical processes ii. Physical processes.	• I can describe the processes that involve the breaking and/or formation of chemical bonds as chemical processes. Processes that involve only changes in intermolecular interactions, such as phase changes, are typically classified as physical processes. Sometimes physical processes involve the breaking of chemical bonds. For example, plausible arguments could be made for the dissolution of a salt in water, as either a physical or chemical process, involves breaking of ionic bonds, and the formation of ion-dipole interactions between ions and solvent.	
4.5 Stoichiometry	**Learning Target 5** I can explain changes in the amounts of reactants and products based on the balanced reaction equation for a chemical process.	• I can explain that because atoms must be conserved during a chemical process, it is possible to calculate product amounts by using known reactant amounts, or to calculate reactant amounts given known product amounts. Coefficients Of Balanced Chemical equations contain information regarding the proportionality of the amounts of substances involved in the reaction. These values can be used in chemical calculations involving the mole concept. Stoichiometric calculations can be combined with the ideal gas law and calculations involving molarity to quantitatively study gases and solutions.	**Simulations** Have students view a simulated reaction pertaining to a limiting reagent problem. Each iteration of the simulation provides students with different unknown concentrations of the reactants from which students calculate the amount of product that is dissolved. Then have them check their answers upon completion of the simulation.
4.6 Introduction to Titration	**Learning Target 6** I can perform titrations and identify the equivalence point in a titration based on the amounts of the titrant and analyte, assuming the titration reaction goes to completion.	• I can use the method of Titrations to determine the amount of an analyte in solution. The titrant has a known concentration of a species that reacts specifically and quantitatively with the analyte. The equivalence point of the titration occurs when the analyte is totally consumed by the reacting species in the titrant. The equivalence point is often indicated by a change in a property (such as color) that occurs when the equivalence point is reached. This observable event is called the endpoint of the titration.	**Think-Pair-Share** Ask students to connect four different particulate representations with a strong acid-strong base titration curve between HCl + NaOH. The representations depict the acid before base has been added, the half equivalence point of the titration, the equivalence point of the titration, and some point beyond the equivalence point (excess base). Have students defend their choices with a partner.
4.7 Types of	**Learning Target 7**	• I can describe Acid-base reactions as involving transfer	**Critique Reasoning** After a review of different types of chemical reactions

Chemical Reactions	I can identify a reaction as acid base, oxidation-reduction, or precipitation.	of one or more protons (H+ ions) between chemical species. And Oxidation-reduction(redox)reactions as involving the transfer of one or more electrons between chemical species, as indicated by changes in oxidation numbers of the involved species. Also,combustion is an important subclass of oxidation-reduction reactions,which species reacts with oxygen gas. In the case of hydrocarbons, carbon dioxide and water are products of complete combustion. In a redox reaction, electrons are transferred from species that are oxidized to the species that is reduced. • I can assign Oxidation numbers to each of the atoms in the reactants and products; this is often an effective way to identify the oxidized and reduced species in a redox reaction. • I can describe how precipitation reactions frequently involve mixing ions in aqueous solution to produce an insoluble or sparingly soluble ionic compound. All sodium, potassium, ammonium, and nitrate salts are soluble in water	(acid-base, redox, precipitation), give students a series of 10 reactions (both the equation and a short demo of the reaction taking place). Have them identify what type of reaction is taking place and justify that claim with evidence. Then have them pair up and evaluate the strength of each other's claims.
4.8 Introduction to Acid-Base Reactions	**Learning Target 8** I can describe and write acid base reactions	• I can identify species as Brønsted Lowry acids, bases, and/or conjugate acid-base pairs, based on proton-transfer involving those species. • I can describe amphoteric species. • When an acid or base ionizes in water, I can identify the conjugate acid-base pairs and compare their relative strengths.	
4.9 Oxidation-Reduction (Redox) Reactions	**Learning Target 9** I can represent a balanced redox reaction equation using half-reactions	• I can construct balanced chemical equations for redox reactions from half-reactions	**Simulations** After viewing a simulation on metal/metal ion reactions, provide students with several 1 molar solutions and a piece of aluminum and ask them to select a solution that would react to coat the Al. Students who select incorrect solutions should go back and revisit the simulation.
Progress Check 4 • Multiple-choice: ~20 questions • Free-response: 1 question			

- Long

Preparing for the AP Exam
On the AP Exam, students must be able to demonstrate proficiency in writing and balancing chemical equations (molecular, complete, net ionic) and calculating quantities in multiple contexts using more than just 1:1 stoichiometric ratios. Students often struggle with questions that require them to justify their identification of a particular type of reaction using an equation. They also encounter difficulty with determining the limiting reactant using stoichiometry. For example, with stoichiometric calculations, students often make the mistake of comparing mass to mass instead of mole to mole when determining the limiting reactant. Teachers can ensure that students practice writing balanced equations (for net ionic and molecular) and that they develop a strong understanding of the mole concept and gain proficiency with dimensional analysis. This will help them correctly calculate required quantities using stoichiometric ratios.

Unit Title:

Unit 5: Kinetics

Relevant Standards: Bold indicates priority

5.1 Reaction Rates: Students can describe reaction rate in terms of changes in concentration over time and interpret graphical representations of rate data.

5.2 Introduction to Rate Law: Students can determine the rate law from experimental data and relate reaction rate to reactant concentrations.

5.3 Concentration Changes Over Time: Students can use integrated rate laws to determine concentration, time, or rate constants for zero-, first-, and second-order reactions.

5.4 Elementary Reactions: Students can identify elementary steps in a reaction mechanism and relate molecularity to rate law expressions.

5.5 Collision Model: Students can explain how effective collisions, orientation, and activation energy determine whether a reaction occurs.

5.6 Reaction Energy Profile: Students can interpret potential energy diagrams to identify activation energy, reaction intermediates, and catalysts.

5.7 Introduction to Reaction Mechanisms: Students can describe how reaction mechanisms consist of a sequence of elementary steps that lead to an overall reaction.

5.8 Reaction Mechanisms and Rate Law: Students can determine rate laws from proposed reaction mechanisms and identify rate-determining steps.

5.9 Pre-Equilibrium Approximation: Students can apply the pre-equilibrium approximation to derive rate laws for multi-step reactions involving fast equilibria.

5.10 Multistep Reaction Energy Profiles: Students can interpret complex energy diagrams for multi-step reactions, including intermediates and multiple activation energies.

5.11 Catalysts: Students can explain how catalysts increase reaction rate by providing an alternative reaction pathway with lower activation energy without being consumed.

Essential Question(s):	**Enduring Understanding(s):**
1. * Why are some reactions faster than other reactions? 2. * Why are some medications taken every 8 hours and others once a day? 3. * Why is food stored in a refrigerator, but bread dough is kept in a warm place to rise? 4. * How can the speed of a reaction be controlled by understanding the collisions that occur on the particle level?	1. Reaction rates are determined by the frequency and effectiveness of particle collisions, which depend on concentration, temperature, and molecular orientation. 2. The rate law describes the relationship between reactant concentrations and reaction rate and must be determined experimentally, not from the balanced equation. 3. Reaction mechanisms consist of a series of elementary steps, and the slowest step (rate-determining step) controls the overall reaction rate. 4. Energy changes during reactions, including activation energy barriers, govern whether collisions result in successful product formation. 5. Catalysts increase reaction rates by providing an alternative pathway with lower activation energy without being consumed in the reaction.
Demonstration of Learning:	**Pacing for Unit**
• Success with unit MCQs and FRQs via AP Classroom • Quizzes (AP Classroom MCQs and FRQs) every few topics	10 instructional days 3 days of lab 2 days review / assess

• Practice identify reaction order based on data, graphs, equations, elementary steps, energy profiles • Practice calculating rate constant k using data, graphs, equations • Practice drawing energy profiles for single and multistep reaction mechanism • Progress Check 5 from AP Classroom Unit 5 is 7-9% of AP Exam Weighting See AP Exam Description	
Family Overview (link below)	**Integration of Technology:**
AP Chemistry FAMILY MATERIALS	N/A
Unit-specific Vocabulary:	**Aligned Unit Materials, Resources, and Technology (beyond core resources):**
reaction rate, rate of reaction, average rate, instantaneous rate, rate law, rate constant, reaction order, overall reaction order, zero-order reaction, first-order reaction, second-order reaction, integrated rate law, half-life, collision theory, activation energy, activated complex, transition state, energy diagram, reaction mechanism, elementary step, rate-determining step, intermediate, catalyst, homogeneous catalyst, heterogeneous catalyst, enzyme, temperature, concentration, surface area, reaction pathway, Arrhenius equation, frequency factor	N/A
Opportunities for Interdisciplinary Connections:	**Anticipated misconceptions:**
• Biology: enzyme activity and metabolic reaction rates • Medicine: drug dosage timing and reaction mechanisms • Environmental Science: decomposition rates and atmospheric reactions • Physics: collision theory and energy transfer • Engineering: catalysts in industrial manufacturing and automotive systems	• Difficulty with graph reading ○ Not noticing differences in y-axis labels for 0th, 1st, and 2nd order graphs • Difficulty determining units for k depending on the order of reaction • Assuming that the stoichiometric coefficients of the balanced chemical reaction determines order of the chemical reaction • Including products in rate expression • Difficulty determining k for multistep reaction mechanisms (especially when first step is fast and at equilibrium)
Connections to Prior Units:	**Connections to Future Units:**
• Kinetic molecular theory was introduced in unit 3 but also helps explain collision theory and activation energy • Balanced equations (unit 4) are needed to determine the order of elementary steps • Concentration (unit 3) is needed to perform kinetic calculations	• Energy profiles are seen again in unit 6 to determine if a reaction is endothermic or exothermic and in unit 9 • Rate constant k must be contrasted with equilibrium constant K in units 7, 8, 9 as they are different concepts.

Differentiation through *Universal Design for Learning*

	Engagement	Representation	Action & Expression
LT1	"Alka-Seltzer Reaction Rate Lab" measuring gas evolution speed across varying temperatures, particle sizes, and concentrations.	Graphical overlay showing reaction rate curves (Concentration vs. Time) steepening with temperature, surface area, or catalyst addition.	Design an experiment testing a kinetic parameter and write a lab report summarizing rate dependencies.
LT2	"Differential Rate Law Speed Challenge" deducing reaction orders (0, 1, 2) from initial rate method data tables.	Method of Initial Rates formula scaffold [Rate2 / Rate1 = ([A]2 / [A]1)^m] with rate constant k unit decoder.	Solve initial rate problems to determine rate laws and calculate k with correct units via handwritten or video solutions.

LT3	Integrated rate law graphing lab using Excel/Google Sheets to plot [A], ln[A], and 1/[A] vs. time to find linearity.	Linear graph comparison chart (0th order: [A] vs. t; 1st order: ln[A] vs. t; 2nd order: 1/[A] vs. t) with half-life formulas.	Determine reaction order from time-series concentration data by submitting linearized plots and calculating rate constant k.
LT4	Molecular collision simulation observing unimolecular vs. bimolecular collisions forming products.	Conversion chart translating elementary step stoichiometry (e.g., A + B -> Products) directly into rate law [Rate = k[A][B]].	Write rate laws for elementary reaction steps and justify why molecularity dictates order for elementary steps but not overall reactions.
	Interactive Collision Theory simulation adjusting collision speed and orientation angle to observe successful reaction percentage.	Visual collision diagrams displaying effective collisions (correct orientation + E ≥ E_a) vs. ineffective collisions (incorrect orientation or low E).	Explain Collision Theory principles and defend why increasing temperature increases effective collision percentage using a video or written report.
	Potential energy profile sketching challenge drawing activation energy barriers (E_a) and enthalpy changes (ΔH).	Annotated potential energy profile diagram labeling reactants, transition state/activated complex, products, E_a (forward/reverse), and ΔH.	Draw and label potential energy profile diagrams for endothermic and exothermic elementary reactions on whiteboards or digital tools.
	"Mechanism Detective" tracking species produced and consumed across multistep elementary mechanisms.	Color-coded mechanism table distinguishing Intermediates (produced then consumed) from Catalysts (consumed then regenerated).	Identify intermediates and catalysts in a multistep mechanism and write the overall balanced chemical equation.
	Multistep mechanism puzzle deriving the overall rate law directly from a slow first elementary step.	Flowchart showing that when Step 1 is slow, Overall Rate Law = Rate Law of Step 1 (Slow Step = Rate-Determining Step).	Derive overall rate laws for mechanisms with a slow first step and explain why step 1 governs the overall rate.
	Fast-equilibrium mechanism problem set substituting intermediate terms using reversible first-step equilibrium expressions.	Derivation scaffold showing intermediate substitution [k_f[A][B] = k_r[Intermediate]] into the slow step rate law.	Derive overall rate laws for mechanisms with a fast reversible first step using written algebraic derivations or screencast walkthroughs.
	Multistep energy profile drawing challenge modeling multi-peak reaction landscapes corresponding to mechanism steps.	Multi-peak potential energy diagram highlighting intermediate valleys, activation energy peaks for each step, and the rate-limiting peak.	Sketch a multi-peak reaction profile for a two-step mechanism, identifying the rate-determining step based on peak height.
LT5	Catalyzed reaction demonstration (e.g., decomposition of H2O2 with MnO2 or yeast) observing reaction speedup and catalyst recovery.	Side-by-side reaction energy profile showing uncatalyzed (high E_a) vs. catalyzed (lower E_a alternative pathway) curves.	Write a CER justification or record a video explaining how a catalyst accelerates a reaction without being consumed or altering ΔH.

Supporting Multilingual/English Learners (*CELP standards*)

	Emerging	Expanding	Bridging
LT1	Identify factors that speed up reactions (temperature, concentration, surface area, catalyst) using visual icon cards.	Explain how changing parameters (concentration, temperature, surface area) affects rate using cause-and-effect sentence frames.	Analyze rate data under varying experimental parameters, explaining rate changes using collision theory and kinetic energy distributions.
LT2	Identify reaction order (0th, 1st, 2nd) from simple rate law expressions (Rate = k[A]^n).	Write rate law expressions and determine units of k using structured problem-solving frames.	Derive rate law expressions, calculate rate constant k with correct units, and evaluate reaction orders from experimental data.
LT3	Match linear integrated rate law graphs ([A], ln[A], 1/[A] vs. time) to 0th, 1st, and 2nd order labels.	Determine reaction order from concentration vs. time data or graphs using guided decision templates.	Determine reaction orders, rate constants, and half-lives using graphical analysis, integrated rate laws, and initial rate data.
LT4	Write rate laws for simple unimolecular or bimolecular elementary steps (A -> B => Rate = k[A]).	Determine molecularity and write rate law expressions for elementary steps using reaction stoichiometry prompts.	Formulate rate laws for elementary steps and explain why molecularity directly dictates reaction order for elementary processes.
LT5	Circle "effective collisions" in diagrams showing correct vs. incorrect particle orientation and minimum energy.	Explain collision theory requirements (frequency, energy, orientation) using structured writing frames.	Evaluate collision dynamics and explain how temperature and orientation factors quantitatively/qualitatively affect reaction rates.

LT6	Label reactants, products, activation energy (Ea), and Delta H on a reaction coordinate diagram.	Sketch exothermic and endothermic potential energy diagrams and identify Ea and Delta H values using guided prompts.	Construct and interpret potential energy profiles, calculating Ea and Delta H for forward and reverse elementary reactions.
LT7	Identify catalysts (consumed first, produced later) and intermediates (produced first, consumed later) using color coding in a mechanism.	Categorize species in a multi-step mechanism as reactants, products, intermediates, or catalysts using structured tables.	Analyze multi-step reaction mechanisms to identify all chemical species and write the overall balanced reaction equation.
LT8	Write the overall rate law directly from the first step of a mechanism using a visual slow-step rule card.	Explain why the rate-determining step dictates the overall rate law using guided paragraph starters.	Derive the overall rate law for a mechanism with a slow initial step, justifying why subsequent steps do not affect the rate law.
LT9	Identify the slow step in a 2-step mechanism and circle the intermediate species to be substituted.	Express overall rate laws for mechanisms with fast equilibrium pre-steps using step-by-step substitution guides.	Derive overall rate laws for complex mechanisms with fast reversible initial steps using the steady-state or pre-equilibrium approximation.
LT10	Count the number of peaks (steps) and valleys (intermediates) on a multi-step potential energy diagram.	Draw multi-step energy profile diagrams identifying rate-limiting steps (Ea peak height) using guided prompts.	Construct multi-step reaction energy profiles, accurately representing relative Ea heights, intermediates, transition states, and overall Delta H.
LT11	Identify catalyzed vs. uncatalyzed curves on a reaction coordinate diagram (catalyzed peak is lower).	Explain how catalysts lower Ea and increase effective collisions using comparative sentence frames.	Analyze catalyzed vs. uncatalyzed energy profiles and collision models, explaining mechanism changes and activation energy reduction.

Unit Overview

In prior units, students developed their ability to describe symbolic and quantitative information from representations (e.g., Lewis structures, chemical reactions) that illustrate both the particulate and macroscopic level of a chemical phenomenon. In Unit 5, students will build on these explanations and representations by constructing and describing rate laws consistent with experimental evidence. To that end, students will collect data by spectrophotometry and choose an appropriate mathematical routine to determine how concentration varies with time during the course of a reaction. In addition, students will examine proposed reaction mechanisms to determine if there is a match between observed experimental data and constructed rate law expressions. Students will learn to identify any intermediates or catalysts that are included in the reaction mechanism, as well as the rate-determining step ,and be able to justify their claims. To do so, students must learn to construct and analyze energy profiles for chemical reactions and identify how such profiles may change with the addition of a catalyst.

Lesson	Learning Target	Success Criteria	Suggested Activities
5.1 Reaction Rates	**Learning Target 1** I can explain the relationship between the rate of a chemical reaction and experimental parameters.	• I can use experimental methods to monitor the amounts of reactants and/or products of a reaction over time and determine the rate of the reaction. • I can express the rate of a reaction using a rate law that shows how it depends on the concentration of each reactant raised to a power. • I can determine the order of a reaction with respect to each reactant and use this to find the overall reaction order as the sum of the individual orders. • I can describe the rate constant and explain how its value depends on temperature	**Post-Lab Discussion** As an introduction to kinetics, have students form small groups to design an experiment to establish a relationship between the rate and a specific reaction parameter of Alka-Seltzer tablets in water. Have them select varying temperature, concentration, mass, or surface area and decide which data to collect. Groups use whiteboards to present their data and major findings to the rest of the class.

		and how its units reflect the overall reaction order. ● I can use initial rate data to determine the rate law and reaction orders for a chemical reaction.	
5.2 Introduction to Rate Law	**Learning Target 2** I can represent experimental data with a consistent rate law expression	● I can use experimental data to determine a consistent rate law expression that describes how reactant or product concentrations change over time. ● I can explain how reaction rates are determined by monitoring changes in reactant or product concentrations over time. ● I can write a rate law that expresses reaction rate as proportional to the concentration of each reactant raised to a power. ● I can determine the order of a reaction with respect to each reactant and explain how the sum of these orders gives the overall reaction order. ● I can describe the rate constant, including how it depends on temperature and how its units are determined by the overall reaction order. ● I can use the method of initial rates to determine the order of a reaction with respect to each reactant.	**Post-Lab Discussion** Using a spectrophotometer, have students measure the absorbance of a solution of green food coloring after bleach has been added. Have them use Excel to prepare different graphs of the data, such as absorbance vs. time, and 1/(absorbance) vs. time. Students should use a linear regression analysis to determine the most linear fit, the order of the reaction, and the effect on the value of k when the concentration of bleach is increased. Have student groups share and compare their results.
5.3 Concentration Changes Over Time	**Learning Target 3** I can identify the rate law expression of a chemical reaction using data that show how the concentrations of reaction species change over time.	● I can determine the rate law expression of a chemical reaction using data that show how the concentrations of reactants change over time. ● I can infer the order of a reaction by analyzing a graph of reactant concentration versus time. ● I can determine if a reaction is first order by recognizing that a plot of ln(concentration) versus time is linear. ● I can determine if a reaction is second order by recognizing that a plot of 1/concentration versus time is linear. ● I can use integrated rate law equations and the slopes of linearized graphs to determine the rate constant for zeroth-, first-, and second-order	**Critique Reasoning** Using a balance and a stopwatch, have students determine the rate order of a burning birthday candle by preparing graphs in Excel, and use a linear regression analysis to determine the most linear fit and the value of the rate constant, k. Have students justify why the rate of mass disappearance of the candle does not change as the candle burns down. Then have them compare their results with other groups to see if their results are consistent.

		reactions. • I can explain that half-life is a key feature of first-order reactions because it remains constant regardless of concentration. • I can use the relationship t{1/2} = 0.693/k to calculate half-life or the rate constant for a first-order reaction. • I can describe how radioactive decay is an example of first-order kinetics.	
5.4 Elementary Reactions	**Learning Target 4** I can represent an elementary reaction as a rate law expression using the stoichiometry of the reacting particles.	• I can infer the rate law of an elementary reaction directly from its balanced equation and collision stoichiometry. • I can explain that the molecularity of an elementary step determines its rate law expression. • I can describe why elementary reactions involving the simultaneous collision of three or more particles are rare.	
5.5 Collision Model	**Learning Target 5** I can explain that the rate of an elementary reaction depends on the frequency of particle collisions, the energy of those collisions, and the orientation of the reacting particles.	• I can describe how successful reactions require collisions with sufficient energy to overcome activation energy and the correct orientation for bonds to break and form. • I can explain why only a small fraction of collisions result in a reaction, even when particles are colliding frequently. • I can use the Maxwell-Boltzmann distribution to describe the range of particle energies in a system. • I can use the Maxwell-Boltzmann distribution to explain how temperature affects the fraction of particles with enough energy to react.	
5.6 Reaction Energy Profile	**Learning Target 6** I can represent the activation energy and overall energy change of an elementary reaction using a reaction energy profile diagram.	• I can explain that elementary reactions involve the breaking of bonds in reactants and the formation of new bonds in products. • I can describe the reaction coordinate as the pathway that represents the progression of reactants through the transition state to products. • I can identify activation energy as the energy difference between the reactants and the transition state on a reaction energy profile. • I can explain how temperature affects reaction rate by changing the fraction of collisions that have enough energy to reach the transition state. • I can describe how the Arrhenius equation relates temperature and activation energy to the rate of an elementary reaction.	
5.7 Introduction to Reaction Mechanisms	**Learning Target 7** I can identify the components of a reaction mechanism, including reactants, products, intermediates, and catalysts.	• I can describe a reaction mechanism as a series of elementary steps that occur in sequence. • I can explain that the sum of all elementary steps in a mechanism must match the overall balanced chemical equation. • I can define a reaction intermediate as a species that is produced in one step and consumed in a later step, so it does not appear in the overall reaction. • I can explain how the	**Critique Reasoning** Working in small groups, have students evaluate the appropriateness of reaction mechanisms for a given reaction for which the rate law is established. Have groups share their conclusions with the rest of the class and then discuss why certain choices must be eliminated and why there might be more than one possible mechanism that is valid. Have classmates provide feedback to the groups on the validity of their conclusions.

		experimental detection of intermediates can provide evidence for or against a proposed reaction mechanism.	
5.8 Reaction Mechanism and Rate Law	**Learning Target 8** I can identify the rate law for a reaction from a mechanism in which the first step is rate limiting.	• I can explain that when each elementary step is irreversible or when the first step is slow, the rate law is determined by that slow (rate-limiting) step. • I can determine the rate law of a reaction by using the molecularity of the rate-determining step.	**Critique Reasoning** (see 5.7)
5.9 Pre-Equilibrium Approximation	**Learning Target 9** I can identify the rate law for a reaction from a mechanism in which the first step is not rate limiting.	• I can explain that when the first elementary step is not rate limiting, additional assumptions or approximations are needed to determine the rate law. • I can apply approaches such as the pre-equilibrium approximation to derive a rate law expression from a multi-step mechanism.	**Manipulatives** Give students a blank multistep reaction energy profile with a series of labels on the side. Have them work with a partner to correctly place the labels next to the blanks indicated on the profile and then share/evaluate their diagrams with another pair of students.
5.10 Multistep Reaction Energy Profile	**Learning Target 10** I can represent the activation energy and overall energy change in a multistep reaction using a reaction energy profile diagram.	• I can explain how each elementary step in a reaction mechanism contributes to the overall energy profile. • I can construct an energy profile for a multistep reaction using knowledge of the energetics of each step in the mechanism. • I can identify multiple transition states and intermediates on a multistep reaction energy diagram.	
5.11 Catalysis	**Learning Target 11** I can explain how a catalyst increases the rate of a reaction by increasing the number of effective collisions and/or providing an alternative reaction pathway with lower activation energy.	• I can describe how a catalyst participates in a reaction mechanism by being consumed in one step and regenerated in a later step, resulting in no net change in catalyst concentration. • I can explain how catalysts can form temporary bonds with reactants to improve orientation or lower activation energy, often creating new intermediates (as in enzyme catalysis). • I can describe how acid–base catalysis involves proton transfer steps that create new intermediates and modify the reaction mechanism. • I can explain how surface catalysis involves reactants binding to a surface, forming intermediates and introducing new elementary steps in the mechanism.	

Progress Check 5

- Multiple-choice: ~25 questions
- Free-response: 2 questions
    - Short
    - Long

Preparing for the AP Exam

On the AP Exam, students must be able to navigate between experimental data (tabular or graphed), a given or constructed rate law, and a proposed mechanism. Students generally struggle with reading a graph of reactant concentration versus time and drawing appropriate conclusions (i.e., order and rate constant) from the graphed data. Specifically, students confuse the units of the graphs with the units represented in the chemical equation. Teachers can ensure that students have multiple opportunities to graph concentration versus time or concentration versus rate data (using appropriate increments and units for the axes). Once students learn how to graph this data, teachers can help them analyze the graphs to determine the order of a reaction.

Unit Title:

Unit 6: Thermochemistry

Relevant Standards: Bold indicates priority

6.1 Endothermic and Exothermic Processes: Students can distinguish between endothermic and exothermic processes based on energy flow between system and surroundings.

6.2 Energy Diagrams: Students can interpret potential energy diagrams to identify enthalpy change and activation energy for reactions.

6.3 Heat Transfer and Calorimetry: Students can use calorimetry data to calculate heat absorbed or released during chemical and physical processes.

6.4 Heat Capacity and Specific Heat: Students can relate heat transfer to temperature change using specific heat capacity and mass.

6.5 Enthalpy of Phase Changes: Students can calculate energy changes associated with phase transitions using molar enthalpies of fusion, vaporization, and other phase changes, and relate these energy changes to particle interactions and changes in state.

6.6 Introduction to Enthalpy of Reaction: Students can calculate enthalpy changes using experimental data, bond energies, or Hess's Law.

6.7 Bond Enthalpies: Students can estimate enthalpy changes of reactions using bond dissociation energies.

6.8 Enthalpies of Formation: Students can calculate standard enthalpy changes using standard enthalpies of formation.

6.9 Hess's Law: Students can determine enthalpy changes for chemical and physical processes by applying Hess's Law to combine, reverse, and scale multi-step reactions to obtain an overall reaction enthalpy.

Essential Question(s):	**Enduring Understanding(s):**
1. Why is energy released when liquid water becomes an ice cube? 2. Why does your skin feel cold when water evaporates off of it? 3. How does a thermal energy transfer affect temperature, states of matter, and chemical bonds? 4. How can energy changes be tracked and measured when energy can't be seen? 5. Why do combustion reactions that form carbon dioxide release energy?	1. Energy is transferred between a system and its surroundings through heat and work, and these transfers can change the internal energy of a system. 2. Enthalpy is a state function that represents heat transfer at constant pressure, and enthalpy changes can be used to describe endothermic and exothermic processes. 3. Calorimetry allows experimental determination of heat transfer using temperature changes, heat capacity, and specific heat relationships. 4. Energy changes in chemical reactions can be calculated using Hess's Law, standard enthalpies of formation, or average bond enthalpies. 5. Phase changes and chemical processes involve energy changes associated with breaking and forming intermolecular or intramolecular interactions.
Demonstration of Learning:	**Pacing for Unit**
• Success with unit MCQs and FRQs via AP Classroom • Quizzes (AP Classroom MCQs and FRQs) every few topics • Practice with the 4 ways to calculate the enthalpy of a reaction • Practice calculating the energy involved in heating / cooling a substance ($q = mc\Delta T$) • Practice calculating energy required to change phase	9 instructional days 2 days lab 2 days review / assess

• Calorimetry/ Specific Heat Lab • Combustion of a Cheeto Lab • Practice drawing energy profiles • Progress Check 6 from AP Classroom Unit 6 is 7-9% of AP Exam Weighting See AP Exam Description	
Family Overview (link below)	**Integration of Technology:**
AP Chemistry FAMILY MATERIALS	N/A
Unit-specific Vocabulary:	**Aligned Unit Materials, Resources, and Technology (beyond core resources):**
Energy, system, surroundings, boundary, thermal energy, kinetic energy, potential energy, chemical potential energy, heat, work, temperature, internal energy, enthalpy, enthalpy change, endothermic, exothermic, calorimetry, calorimeter, coffee-cup calorimeter, bomb calorimeter, heat capacity, specific heat capacity, molar heat capacity, $q = mc\Delta T$, standard state, standard enthalpy change, standard enthalpy of formation, standard enthalpy of combustion, Hess's law, bond enthalpy, bond dissociation energy, state function, phase change, fusion, vaporization, sublimation, condensation, deposition, heating curve, cooling curve, extensive property, intensive property, first law of thermodynamics, energy transfer, reaction enthalpy, thermochemical equation, average bond enthalpy	• Digital thermometers • Bunsen burners • Calorimetry set up (double cup calorimeter)
Opportunities for Interdisciplinary Connections:	**Anticipated misconceptions:**
• Physics: conservation of energy and thermodynamics • Earth Science: phase changes, climate systems, geothermal processes • Biology: metabolism and energy transfer in living organisms • Engineering: energy efficiency, fuels, calorimetry, heat transfer systems • Environmental Science: combustion and energy resources	• Difficulty determining system vs surroundings and their change in temperature when considering an endo or exothermic reaction • Not considering stoichiometric coefficients in enthalpy calculations • Difficulty picking correct method to calculate enthalpy using the information provided in the question ○ Using the wrong equation when using bond enthalpy vs enthalpy of formation • Confusing enthalpy (ΔH) and heat (q) • When calculating the amount of energy to heat a sample and change phase (for example), only using one equation to calculate one step of the process (versus energy required to change temp + energy required to change phase) • Difficulty manipulating given chemical reactions **and** the enthalpy values in order to use Hess's Law to calculate enthalpy of overall process
Connections to Prior Units:	**Connections to Future Units:**
• Students learned about heat transfer, specific heat and specific heat calculations, and phase changes in Physical Science • Energy reaction profiles were introduced in Unit 5. Now students can compare reactants and products to determine if overall reaction is endothermic or exothermic	• Hess's law manipulations are similar but different to how you can manipulate equilibrium constant K in Units 7, 8, and 9 • ΔH returns in Unit 9 to determine if a reaction is thermodynamically favorable or not. • Activation energy can help determine why a thermodynamically favorable reaction may not occur at appreciable rates (9.4)
Differentiation through *Universal Design for Learning*	

	Engagement	Representation	Action & Expression
LT1	Hand warmer (exothermic) and cold pack (endothermic) exploration observing temperature changes during transformations.	Classification chart linking temperature increase (ΔT > 0 -> Exothermic, ΔH < 0) and temperature decrease (ΔT < 0 -> Endothermic, ΔH > 0).	Classify physical or chemical processes as endothermic or exothermic based on lab observations and write an energy flow explanation.
LT2	"Energy Curve Builder" drawing system energy changes for dissolution, phase changes, and reactions.	Reaction coordinate energy diagrams labeling system initial energy state, final energy state, and net heat enthalpy change (ΔH).	Draw energy diagrams for physical phase changes or chemical reactions and label system vs. surroundings heat flow.
LT3	Thermal equilibrium simulation observing heat transfer from hot metal blocks to cold water via kinetic particle collisions.	Particulate collision diagrams showing fast-moving high-T particles transferring kinetic energy to slower-moving low-T particles until thermal equilibrium.	Explain zeroth law/thermal equilibrium principles using particulate collision models via written descriptions or oral presentations.
LT4	Coffee-cup calorimetry lab calculating heat capacity (c) of metal washers or enthalpy of neutralization.	Calorimetry formula scaffold ($q = mc\Delta T$) with sign conventions ($q_{sys} = -q_{surr}$) and unit cancellation guides.	Calculate heat exchange and specific heat values from experimental calorimetry data using handwritten solutions or spreadsheets.
LT5	Heating curve lab for water or lauric acid analyzing plateau regions where energy breaks intermolecular forces.	Annotated heating/cooling curve graph (X-axis: heat added; Y-axis: temperature) paired with $q = n \times \Delta H_{phase}$ calculations.	Calculate heat required for multi-step phase transitions (e.g., ice at -10°C to steam at 110°C) and present worked steps.
LT6	Stoichiometric heat calculation challenge converting grams of limiting reactant to kilojoules of heat produced or absorbed.	Stoichiometric conversion bridge: Mass -> Moles -> Heat ($q = n \times \Delta H_{rxn}$).	Solve reaction enthalpy stoichiometry problems by submitting BCA tables, dimensional analysis steps, or video walkthroughs.
LT7	"Bond Energy Accounting" breaking Lewis structure bonds (endothermic input) and assembling new product bonds (exothermic output).	Bond energy formula scaffold: $\Delta H_{rxn} = sum(BE_{broken}) - sum(BE_{formed})$ with structural Lewis model visual cues.	Calculate ΔH_{rxn} using average bond energy tables by drawing Lewis structures for all species and tracking broken/formed bonds.
LT8	Heat of formation calculation challenge comparing theoretical reaction enthalpies across fuel combustion reactions.	Standard enthalpy formula scaffold: $\Delta H^\circ_{rxn} = sum(n \times \Delta H^\circ_f(products)) - sum(m \times \Delta H^\circ_f(reactants))$ with element baseline rules ($\Delta H^\circ_f = 0$).	Calculate ΔH°_{rxn} from standard enthalpy of formation tables for chemical reactions and present worked solution sets.
LT9	"Hess's Law Puzzle" rearranging and flipping elementary chemical equations to align with a target overall reaction.	Step-by-step Hess's Law rules card (Flip reaction -> change sign of ΔH; Multiply coefficients -> multiply ΔH).	Manipulate and sum target chemical equations and their corresponding ΔH values to build overall reaction pathways on whiteboards.
LT10	Hess's Law experimental verification lab combining two reactions (e.g., NaOH dissolution and neutralization) to yield a third.	Algebraic visual showing state function properties where $\Delta H_{overall} = \Delta H1 + \Delta H2 + \Delta H3$ regardless of path.	Write a formal lab report or record a video verifying Hess's Law experimentally by summing multi-step reaction enthalpies.

Supporting Multilingual/English Learners (*CELP standards*)

	Emerging	Expanding	Bridging
LT1	Classify reactions as exothermic (Delta H < 0, feels hot) or endothermic (Delta H > 0, feels cold) using visual cards.	Describe energy changes during transformations using structured paragraph frames ("*Energy is absorbed/released because...*").	Relate macroscopic temperature observations to microscopic enthalpy changes (Delta H) and system-surrounding heat flow.
LT2	Draw simple enthalpy diagrams showing reactant and product energy levels for exothermic and endothermic processes.	Sketch energy diagrams for phase changes and chemical reactions, labeling Delta H using guided prompts.	Construct accurate enthalpy level diagrams for physical and chemical transformations, justifying relative energy levels.
LT3	Draw arrows showing heat transfer from hot objects (fast particles) to cold objects (slow particles).	Explain thermal equilibrium and heat transfer using particle collision sentence starters ("*Faster particles collide with...*").	Explain the zeroth law and kinetic thermal energy transfer at the particulate level, detailing how thermal equilibrium is reached.

LT4	Calculate heat using q = m * c * Delta T with formula variable cards and unit templates.	Solve calorimetry problems (q_system = -q_surroundings) using step-by-step calculation frames.	Analyze coffee-cup or bomb calorimetry data to calculate q, molar enthalpy changes, and specific heat capacities.
LT5	Identify heating curve plateaus as phase changes where temperature remains constant.	Calculate heat required for phase changes (q = n * Delta H_fusion/vap) using step-by-step calculation templates.	Perform multi-step heating/cooling curve calculations combining q = m * c * Delta T and q = n * Delta H across phase transitions.
LT6	Calculate heat released using dimensional analysis templates (moles * Delta H_rxn).	Calculate reaction heat for given reactant amounts using stoichiometry and enthalpy conversion frames.	Solve stoichiometry-enthalpy problems involving limiting reactants and calculate molar enthalpy of reaction (Delta H_rxn).
LT7	List bonds broken and bonds formed in a reaction table using Lewis structure representations.	Calculate Delta H_rxn using sum(Bonds Broken) - sum(Bonds Formed) formula templates.	Calculate reaction enthalpies from bond energies, explaining why bond-energy calculations are estimates compared to Delta Hf°.
LT8	Locate Delta Hf° values in a reference table and plug into sum(n * Delta Hf°(products)) - sum(m * Delta Hf°(reactants)).	Calculate Delta H_rxn° using standard formation values and structured equation templates.	Compute standard reaction enthalpies using formation data, explaining why Delta Hf° for pure elements in standard states is zero.
LT9	Rearrange chemical equation cards to show how individual steps sum to an overall target equation.	Combine intermediate reaction steps to yield an overall reaction equation using guided step-by-step rules.	Formulate logical reaction sequences that represent complex transformations, verifying stoichiometry and species cancellation.
LT10	Calculate total Delta H by adding given Delta H values of intermediate steps on a visual diagram.	Apply Hess's Law to flip or multiply reaction steps and calculate overall Delta H using guided templates.	Execute multi-step Hess's Law calculations, manipulating reaction equations and their corresponding Delta H values with full justifications.

Unit Overview

The ability to link atomic- and particulate level phenomena and models to macroscopic phenomena is central to the study of chemistry. In previous units, students used representations, equations, and reasoning to demonstrate this ability. In Unit 6, students will develop justifications for claims made about the direction of thermal energy transfer of a system in relation to its surroundings when a temperature change, physical change, or a chemical reaction occurs. Students will construct representations of energy using appropriate diagrams with arrows showing the direction of energy transfer between the system and the surroundings. They will continue to develop their explanations of chemical phenomena by explaining how the change in energy of a system is balanced by transfer of energy by either heat or work into or out of the system.

Lesson	Learning Target	Success Criteria	Suggested Activities
6.1 Endothermic and Exothermic Processes	**Learning Target 1** I can explain the relationship between experimental observations and energy changes associated with chemical and physical transformations.	• I can use temperature changes in a system to infer whether energy has been absorbed or released. • I can describe energy changes in terms of endothermic and exothermic processes, including heating/cooling, phase changes, and chemical reactions. • I can explain that in exothermic processes energy is transferred from the system to the surroundings, and in endothermic processes energy is transferred from the surroundings to the system. • I can explain how the formation of a solution can be either endothermic or exothermic depending on the relative strengths of interactions	**Think-Pair-Share** Have student pairs generate a list of exothermic and endothermic processes that occur in their everyday life. Have them share their lists with other pairs to determine if they have correctly identified these common processes in terms of endo- or exothermicity.

		broken and formed during dissolution.	
6.2 Energy Diagrams	**Learning Target 2** I can represent a chemical or physical transformation using an energy diagram.	• I can use an energy diagram to show whether a process is endothermic or exothermic. • I can interpret energy diagrams to describe changes in energy during a physical or chemical process.	**Think-Pair-Share** (see 6.1)
6.3 Heat Transfer and Thermal Equilibrium	**Learning Target 3** I can explain the relationship between the transfer of thermal energy and molecular collisions.	• I can describe how particles in a warmer substance have greater average kinetic energy than particles in a cooler substance. • I can explain how collisions between particles in thermal contact result in the transfer of energy as heat. • I can describe how energy transfer continues through particle collisions until thermal equilibrium is reached. • I can explain that at thermal equilibrium, the average kinetic energy and temperature of both systems are equal.	**Demo with Q&A** After working a few practice problems in groups with the q = mcΔT equation, demonstrate that heating 40 g of copper pellets to 80°C and placing them into 40 g of 20°C water does not result in 50°C as a final temperature. Have students reason why and then record the final temperature of the copper/water mixture. Then have them calculate the specific heat capacity of copper and compare it to published values. As a class discussion, account for deviations from the expected results.
6.4 Heat Capacity and Calorimetry	**Learning Target 4** I can calculate the heat (q) absorbed or released by a system using the equation q = mcΔ T.	• I can explain how calorimetry experiments are used to measure heat transfer between a system and its surroundings. • I can apply the first law of thermodynamics to explain that energy is conserved during heating and cooling processes. • I can explain why equal masses of different substances experience different temperature changes when the same amount of heat is transferred. • I can describe how specific heat capacity and molar heat capacity are used in calculations involving thermal energy transfer. • I can explain that chemical systems change energy through heating/cooling, phase transitions, and chemical reactions. • I can describe how calorimetry can be used to determine the direction of energy flow during a process. • I can explain that an increase in temperature of a solution indicates that the process is exothermic and releases thermal energy. • I can explain that a decrease in temperature of a solution indicates that the process is endothermic and absorbs thermal energy.	**Process Oriented Guided Inquiry Learning (POGIL)** Have students wet one finger with water and keep one finger dry then wave them in the air to see which feels cooler. Have them respond to a series of guided questions about the energy transfers involved in the evaporation process. Next, two beakers are heated side by side on a hot plate. Heating a beaker with 100 g of water on the same hot plate alongside a beaker with 100 g of 1-propanol results in very different changes in temperature. Through guided inquiry, students derive the concept of specific heat. As a class, compare whether the two liquids have been treated “fairly,” and the concept of molar heat capacity is established and compared to specific heat capacity.

6.5 Energy of Phase Changes	**Learning Target 5** I can explain changes in heat (q) absorbed or released during a phase transition using the amount of substance in moles and the molar enthalpy of the phase change.	• I can describe how energy must be absorbed for melting and vaporization, increasing the system's energy, and how energy is released during freezing and condensation, decreasing the system's energy. • I can explain why the temperature of a pure substance remains constant during a phase change even though energy is being transferred. • I can calculate the heat involved in a phase change using molar enthalpy and moles of substance. • I can explain that the energy absorbed in one direction of a phase change is equal in magnitude and opposite in sign to the energy released in the reverse process (e.g., fusion vs. freezing, vaporization vs. condensation).	
6.6 Introduction to Enthalpy of Reaction	**Learning Target 6** I can calculate the heat (q) absorbed or released by a system undergoing a chemical reaction using the amount of reacting substance in moles and the molar enthalpy of reaction.	• I can explain that the enthalpy change of a reaction represents the heat released (negative ΔH) or absorbed (positive ΔH) at constant pressure. • I can describe how energy is transferred between a system and its surroundings during exothermic and endothermic reactions to reach thermal equilibrium. • I can explain that changes in chemical potential energy during reactions result from breaking and forming bonds, which leads to a change in the kinetic energy of particles.	
6.7 Bond Enthalpies	**Learning Target 7** I can calculate the enthalpy change of a reaction using average bond energies of bonds broken and bonds formed.	• I can explain that chemical reactions involve breaking bonds in reactants and forming new bonds in products, which changes the system's potential energy. • I can estimate the total energy required to break reactant bonds by summing the average bond energies of all bonds broken. • I can estimate the total energy released when product bonds form using average bond energies. • I can determine whether a reaction is exothermic or endothermic by comparing the energy required to break bonds with the energy released when bonds form.	**Think-Pair-Share** Have pairs of students examine tables of average bond enthalpy and establish patterns with regard to bond order, atomic radius, and bond length. Similar patterns are examined for the standard enthalpies of formation. Have student pairs work through several practice problems using bond energies and enthalpies of formation to determine the enthalpy of a chemical reaction and compare their calculations.
6.8 Enthalpy of Formation	**Learning Target 8** I can calculate the enthalpy change for a chemical or physical process using standard enthalpies of formation.	• I can use tabulated standard enthalpies of formation to determine the overall enthalpy change of a reaction. • I can apply the equation $\Delta H^\circ_{reaction} = \Sigma \Delta Hf^\circ_{products} - \Sigma \Delta Hf^\circ_{reactants}$ to calculate reaction enthalpy.	**Think-Pair-Share** (see 6.7)
6.9 Hess's Law	**Learning Target 9** I can represent a chemical or physical process as a sequence of steps.	• I can explain that many processes can be broken down into a series of steps, each with its own energy change. • I can describe how the overall energy change of a process is related to the sum of the energy changes of its individual steps.	**Post-Lab Discussion** Have students apply Hess's law by reacting magnesium metal and magnesium oxide with hydrochloric acid to determine the enthalpy change of the following reaction: Mg + O 2 → MgO. Then have them evaluate their results and discuss sources of error.
	Learning Target 10 I can explain that the enthalpy of a chemical	• I can describe how energy is conserved in a multi-step process, so the total enthalpy change equals the net energy transferred to or from the surroundings.	

	or physical process is equal to the sum of the enthalpies of its individual steps.	● I can explain that at constant pressure, the overall enthalpy change is the sum of the enthalpy changes of each step in a reaction sequence. ● I can apply Hess's Law to calculate enthalpy changes by reversing reactions (changing the sign of ΔH), multiplying reactions (scaling ΔH accordingly), and adding reactions together to obtain an overall ΔH.

Progress Check 6

- Multiple-choice: ~20 questions
- Free-response: 2 questions
    - Short
    - Short

Preparing for the AP Exam

On the AP Exam, students must be able to translate between a balanced chemical reaction and a calculation involving the energies of bonds broken and bonds formed within the reaction. In addition, students will be required to analyze calorimetry data and apply mathematical routines to calculate or estimate the heat transferred and the overall enthalpy of a reaction. In a question that asks students to apply mathematical routines to estimate or calculate the overall enthalpy of a reaction, students often struggle to determine the number of bonds that were broken and made in the reaction. Teachers can ensure that students are able to identify the bonds broken and formed in the reaction and use the enthalpies for such to determine the overall enthalpy for the reaction, in addition to their ability to represent a chemical reaction with its associated equation.

Unit Title:

Unit 7: Equilibrium

Relevant Standards: Bold indicates priority

7.1 Introduction to Equilibrium: Students can use chemical principles and mathematical reasoning to justify claims about dynamic equilibrium in reversible reactions, including the relationship between forward and reverse reaction rates.

7.2 Direction of Reversible Reactions: Students can explain how particulate-level interactions connect to macroscopic behavior in reversible reactions and use representations to determine the direction a reaction will proceed.

7.3 Reaction Quotient and Equilibrium Constant: Students can represent chemical equilibrium systems using mathematical expressions and graphs, and interpret the reaction quotient (Q) in relation to the equilibrium constant (K).

7.4 Calculating the Equilibrium Constant: Students can calculate equilibrium constants from experimental data and explain how changes in system variables affect equilibrium relationships.

7.5 Magnitude of the Equilibrium Constant: Students can use chemical principles and mathematical reasoning to justify how the magnitude of K relates to the extent of a reaction at equilibrium.

7.6 Properties of the Equilibrium Constant: Students can identify and apply the mathematical relationships and definitions required to determine equilibrium constants from given chemical information.

7.7 Calculating Equilibrium Concentrations: Students can calculate equilibrium concentrations or partial pressures using appropriate mathematical models, including ICE tables, and represent results clearly using correct units and precision.

7.8 Representations of Equilibrium: Students can construct and interpret particulate, graphical, and symbolic representations of equilibrium systems across multiple scales to show the relationship between structure and observable properties.

7.9 Introduction to Le Châtelier's Principle: Students can explain how experimental observations of equilibrium systems support Le Châtelier's Principle and relate shifts in equilibrium to changes in conditions.

7.10 Reaction Quotient and Le Châtelier's Principle: Students can predict the direction of a system's shift toward equilibrium by comparing Q and K and applying quantitative reasoning to equilibrium changes.

7.11 Introduction to Solubility Equilibria: Students can select and apply appropriate equilibrium relationships to solve problems involving solubility equilibria.

7.12 Common-Ion Effect: Students can explain how modifications to a system (such as addition of a common ion) alter equilibrium position and solubility based on equilibrium principles.

Essential Question(s):	**Enduring Understanding(s):**
1. How are the rates of forward and reverse reactions related to the direction that a reversible reaction proceeds? 2. How can the composition of a mixture at equilibrium be predicted? 3. How can an equilibrium system be manipulated to maximize product yield? 4. Why do paramedics administer pure oxygen to people with carbon monoxide poisoning? 5. What factors influence the degree to which a salt will dissolve?	1. Some reactions can occur in both forward and reverse directions, sometimes resulting in a dynamic equilibrium state. 2. A system at equilibrium depends on the relationships between concentrations, partial pressures of chemical species, and the equilibrium constant K. 3. Systems at equilibrium respond to external stresses to offset the effect of the stress. 4. The dissolution of a salt is a reversible process that can be influenced by environmental factors such as pH or other dissolved ions.
Demonstration of Learning:	**Pacing for Unit**
• Success with unit MCQs and FRQs via AP Classroom	10 instructional days

• Quizzes (AP Classroom MCQs and FRQs) every few topics • K vs Q particle modeling • Practice RICE calculations • Practice writing out equilibrium expressions • Practice calculating the effect of a common ion and responding to FRQ questions about common ion effect • Solubility of Calcium Chloride Lab (Ksp) • Progress Check 7 from AP Classroom Unit 7 is 7-9% of AP Exam Weighting See AP Exam Description	2 days for lab 2 days for review / assess
Family Overview (link below)	**Integration of Technology:**
AP Chemistry FAMILY MATERIALS	N/A
Unit-specific Vocabulary:	**Aligned Unit Materials, Resources, and Technology (beyond core resources):**
Equilibrium, chemical equilibrium, dynamic equilibrium, Le Chatelier's principle, equilibrium constant (K), reaction quotient (Q), heterogeneous equilibrium, homogeneous equilibrium, partial pressure, Kp, Kc, relationship between Kp and Kc, concentration changes, pressure changes, temperature changes, catalyst effect on equilibrium, solubility equilibrium, solubility product constant (Ksp), ion product (Qsp), common ion effect, precipitation, dissolution, saturated solution, unsaturated solution, supersaturated solution	• Stir plates • Milk of magnesia • Application of spectrophotometer if we had one • Pop beads and petri dishes to model K and Q
Opportunities for Interdisciplinary Connections:	**Anticipated misconceptions:**

• Biology: homeostasis and dynamic equilibrium in body systems
• Environmental Science: ocean acidification and atmospheric equilibrium
• Medicine: oxygen transport and blood chemistry equilibria
• Engineering: industrial equilibrium processes such as ammonia synthesis
• Mathematics: modeling dynamic systems and equilibrium calculations
	• Confusing rate constant k and equilibrium constant K
	• Writing out equilibrium expression incorrectly
	  ○ Including solids or liquids in equilibrium expression
	  ○ Putting products on the bottom
	  ○ Not including stoichiometry coefficients as superscripts
	  ○ Using K_c vs K_p when appropriate. Writing out the two equilibrium expressions with appropriate notation
	• Inappropriate or insufficient language to describe impact of stressor
	  ○ Shift left or right vs towards reactants or products vs proceed forwards or in reverse direction
	  ○ Using comparison of reaction quotient Q and equilibrium constant K to justify answer for Le Chatelier's or common ion effect
	• Using incorrect units in RICE table
	• Misuse of 5% approximation

Connections to Prior Units:	**Connections to Future Units:**
• Students must be able to determine how an ionic compound will dissolve and write out net ionic equation (4.2) • Students must be able to calculate molarity of solutions (3.7) • Students must be able calculate partial pressures to use in a K_p expression (3.4)	• Equilibrium expressions and RICE tables are used in unit 8 with equilibrium of weak acid or weak base in water • Effect of pH on solubility is included in Unit 8 • Equilibrium constant K can be used with electrochemical calculations to determine if a cell will be thermodynamically favorable or not (unit 9)

• Students must be able to reference reaction rates and catalysts to determine how a catalyst impacts equilibrium	

Differentiation through *Universal Design for Learning*

	Engagement	Representation	Action & Expression
LT1	Reversible reaction lab observing color shifts in the CoCl4^2- / [Co(H2O)6]^2+ system upon heating and cooling.	Dynamic equilibrium visual showing non-zero forward and reverse rates maintaining constant macroscopic concentrations.	Write a description or record a video explaining how dynamic equilibrium differs from a static reaction ending point.
LT2	Rate vs. Time graph exploration showing forward rate decreasing and reverse rate increasing until r_forward = r_reverse.	Side-by-side graph comparison displaying Rate vs. Time and Concentration vs. Time (concentrations plateau).	Interpret Rate vs. Time and Concentration vs. Time graphs to identify when equilibrium is established in a reversible system.
LT3	Writing equilibrium expressions for heterogeneous reactions containing pure solids and liquids.	Mass action expression guide: Q_c = ([C]^c [D]^d) / ([A]^a [B]^b), emphasizing exclusion of pure solids (s) and liquids (l).	Write correct Q_c and Q_p expressions for homogeneous and heterogeneous chemical systems on problem sets or digital whiteboards.
LT4	Equilibrium constant calculation lab analyzing the formation of FeSCN^2+ using spectrophotometry data.	Equilibrium calculation formula card converting equilibrium concentrations/pressures into numerical K_c or K_p values with K_p = K_c(RT)^Δn.	Calculate K_c or K_p values from equilibrium measurement data and convert between K_c and K_p using worked solutions.
LT5	"Product-Favored vs. Reactant-Favored Sorting" matching K values (K >> 1, K ≈ 1, K << 1) to mixture composition images.	Visual scale mapping K > 10^3 (product-dominated at equilibrium) vs. K < 10^-3 (reactant-dominated at equilibrium).	Analyze given K values and state whether reactants or products dominate at equilibrium, supporting claims with mathematical justifications.
LT6	Multi-step equilibrium puzzle calculating K_overall by multiplying step constants (K_overall = K1 × K2).	Summary rules card for manipulating K: Reverse reaction -> 1/K; Multiply coefficients by n -> K^n; Add reactions -> K1 × K2.	Calculate K for modified or multi-step equilibrium equations using mathematical rules and present step-by-step work.
LT7	"RICE Table Challenge" solving for equilibrium concentrations given initial values and K (including 5% approximation rule).	RICE table template (Reaction, Initial, Change, Equilibrium) with algebraic variable guides.	Solve equilibrium concentrations using RICE tables and validate or refute 5% approximation assumptions via written problem sets.
LT8	Particulate box counting activity verifying whether a particle diagram matches the ratio specified by K.	Particulate box diagrams displaying molecular counts before and at equilibrium, confirming constant particle ratios.	Draw particulate representations of an equilibrium mixture matching a specific K value and defend particle count ratios.
LT9	Le Châtelier's principle lab applying stress (concentration change, temperature shift, volume/pressure change) to equilibrium systems.	Stress-Response summary matrix detailing system shifts: Add reactant -> shift right; Increase T (exothermic) -> shift left; Decrease V -> shift to fewer gas moles.	Predict shift direction for equilibrium systems subjected to external stresses and defend predictions using Le Châtelier's principle.
LT10	"Q vs. K Number Line" comparing calculated Q values against K to determine net reaction direction.	Number line visual comparison: Q < K (shift right ->), Q = K (at equilibrium), Q > K (shift left <-).	Calculate Q, compare it to K, and write a justification stating which direction the system must shift to reach equilibrium.
LT11	Molar solubility calculation challenge computing dissolved ion concentrations for sparingly soluble salts (e.g., AgCl, PbI2).	K_sp expression scaffold [K_sp = [M^n+]^m [X^m-]^n] paired with molar solubility (x) substitution equations.	Calculate molar solubility (x) and ion concentrations from K_sp values for 1:1, 1:2, and 1:3 sparingly soluble salts.
LT12	"Common Ion Effect Lab" adding NaCl to a saturated PbCl2 solution to observe increased precipitation.	Le Châtelier K_sp visual diagram showing how adding a common ion shifts solubility equilibrium left, decreasing molar solubility.	Calculate reduced molar solubility of a salt in the presence of a common ion and explain the shift using K_sp expressions.

Supporting Multilingual/English Learners (*CELP standards*)

	Emerging	Expanding	Bridging

LT1	Identify equilibrium on a concentration vs. time graph as the point where concentrations become constant.	Explain dynamic equilibrium using sentence starters (*"At equilibrium, the forward rate equals..."*).	Describe dynamic chemical equilibrium, emphasizing equal forward/reverse rates and constant macroscopic properties.
LT2	Compare forward and reverse rate arrows on a visual diagram before and at equilibrium.	Explain how unequal forward and reverse rates shift species concentrations toward equilibrium using guided prompts.	Analyze rate vs. time graphs to evaluate how non-equilibrium systems adjust forward and reverse rates to achieve dynamic equilibrium.
LT3	Write equilibrium expressions (K = [Products] / [Reactants]) using visual formula templates (excluding solids/liquids).	Write Qc and Qp expressions for balanced reversible reactions using structured graphic organizers.	Formulate correct Qc and Qp expressions for heterogeneous and homogeneous equilibria, justifying why pure solids/liquids are omitted.
LT4	Calculate K by plugging given equilibrium concentrations directly into a provided K expression.	Calculate Kc or Kp from equilibrium data using structured calculation templates and ICE tables.	Calculate equilibrium constants (Kc, Kp) from initial conditions and partial equilibrium data using ICE tables and unit analysis.
LT5	Classify reactions as "product-favored" (K > 1) or "reactant-favored" (K < 1) using visual scale cards.	Explain what K >> 1 or K << 1 implies about equilibrium composition using structured sentence frames.	Interpret the physical significance of K magnitudes, relating K values to reaction extent and thermodynamic favorability.
LT6	Combine K values of combined steps using the rule: K_overall = K1 * K2 on a visual template card.	Calculate overall K for combined reactions (reversing, multiplying coefficients) using rule-based calculation guides.	Derive overall equilibrium expressions and calculate K_overall for manipulated multi-step equilibrium systems.
LT7	Set up an ICE (Initial, Change, Equilibrium) table grid using provided initial concentrations and +x / -x hints.	Solve ICE table problems for equilibrium concentrations using algebraic templates and the 5% approximation rule.	Calculate equilibrium concentrations or partial pressures using ICE tables, solving quadratic equations or applying the 5% rule when valid.
LT8	Count reactant and product particles in box diagrams to determine if a system is at equilibrium.	Construct particulate diagrams representing systems before, during, and at equilibrium using ratio guidelines.	Draw and evaluate particulate models of equilibrium mixtures, ensuring particle counts reflect K magnitudes and stoichiometry.
LT9	Predict response direction (Shift Left/Right) when adding/removing species using a visual seesaw diagram.	Predict equilibrium shifts due to concentration, pressure, volume, or temperature changes using Le Châtelier sentence frames.	Apply Le Châtelier's principle to predict and explain shifts, species concentration changes, and K value changes (for T changes).
LT10	Compare Q and K on a number line card (Q < K -> shift right; Q > K -> shift left).	Explain reaction direction shifts by comparing Q to K using structured writing templates (*"Since Q < K, the system must..."*).	Calculate Q, compare it to K, and mathematically explain why the system proceeds forward or reverse to restore equilibrium.
LT11	Write Ksp expressions for simple ionic salts using visual dissociation equation templates.	Calculate molar solubility (s) or Ksp using structured ICE tables and formula templates.	Calculate molar solubility, mass solubility, and Ksp for various salts, accounting for stoichiometry and polyatomic ions.
LT12	Identify the "common ion" in a mixture using color-coded chemical formula cards.	Explain why adding a common ion decreases solubility using Le Châtelier sentence starters.	Calculate molar solubility of a salt in the presence of a common ion using ICE tables and Ksp expressions.

Unit Overview

Building on practices from earlier units where students translated between representations of chemical systems, they will now construct equilibrium expressions from reaction equations. Students should also illustrate the dynamic nature of the chemical reaction through particulate level representations, portraying both the forward and reverse rates of the reaction equations. They will construct and describe graphs that represent a chemical system in equilibrium and connect them to their particulate-levelrepresentationsand equilibrium expressions. In conjunction with their constructed equilibrium expressions, students will practice using experimental data to calculate the reaction quotient (Q) and equilibrium constant (K) for a reaction. Using Le Châtelier's principle, they will also support claims made about the dominant direction of a reaction once stresses like changes in concentration, pressure, volume, or temperature are introduced.

Lesson	Learning Target	Success Criteria	Suggested Activities

7.1 Introduction to Equilibrium	**Learning Target 1** I can explain the relationship between reversible chemical or physical processes and the establishment of equilibrium using experimental observations.	• I can identify examples of reversible processes, including phase changes, gas exchange, dissolution/precipitation, acid–base reactions, and redox reactions. • I can describe equilibrium as a state where reactants and products are both present and their concentrations or partial pressures remain constant over time. • I can explain that chemical equilibrium is dynamic, meaning forward and reverse reactions continue to occur at equal rates with no net observable change. • I can interpret graphs of concentration, partial pressure, or reaction rate versus time to determine when equilibrium has been established.	
7.2 Direction of Reversible Reactions	**Learning Target 2** I can explain the relationship between the direction a reversible reaction proceeds and the relative rates of the forward and reverse reactions.	• I can describe how a faster forward reaction than reverse reaction results in a net formation of products. • I can describe how a faster reverse reaction than forward reaction results in a net formation of reactants. • I can explain that equilibrium is reached when the forward and reverse reaction rates are equal, resulting in no net change in concentrations.	
7.3 Reaction Quotient and Equilibrium Constant	**Learning Target 3** I can represent the reaction quotient (Q) for a reversible reaction using the relative concentrations or partial pressures of reactants and products.	• I can explain that the reaction quotient describes the ratio of product species to reactant species at any point in time, based on their concentrations or partial pressures. • I can distinguish between Qc (based on concentration) and Qp (based on partial pressure) for gas-phase reactions. • I can explain that the equilibrium constant (K) has the same mathematical structure as the reaction quotient but applies specifically when the system is at equilibrium. • I can explain that as a reaction approaches equilibrium, the value of Q changes until it becomes equal to K, at which point the system is at equilibrium.	**Manipulatives** Give groups of students containers that hold objects representing particles in an equilibrium mix (beads work well here). Each bead represents a molecule in a reversible synthesis reaction. The law of mass action is introduced, and students are asked to calculate K. Each group should get the same value for K, even though the number of particles in each container is different. Each group of students then gets a new container that represents a mixture not at equilibrium, and they calculate the ratio using the law of mass action. The concept of Q is introduced and then students determine if and how they could get the ratio of reactants and products to be equal to K by attaching or detaching beads.
7.4 Calculating the Equilibrium Constant	**Learning Target 4** I can calculate the equilibrium constant (Kc or Kp) using experimental measurements of concentrations or partial pressures at equilibrium.	• I can determine equilibrium constant values from data collected once a reaction has reached equilibrium. • I can apply equilibrium data to quantify the relative amounts of reactants and products in a system at equilibrium.	**Identify Subtasks** Given a gaseous equilibrium process, have students construct the expression that can ultimately be used to calculate the K_p.

7.5 Magnitude of the Equilibrium Constant	**Learning Target 5** I can explain the relationship between the magnitude of the equilibrium constant (K) and the relative amounts of reactants and products at equilibrium.	• I can describe that very large K values indicate that products are favored and the reaction proceeds nearly to completion. • I can describe that very small K values indicate that reactants are favored and very little product is formed. • I can use the value of K to predict whether a reaction mixture at equilibrium contains mostly reactants, mostly products, or a significant amount of both.	**Identify Subtasks** (See 7.4)
7.6 Properties of the Equilibrium Constant	**Learning Target 6** I can represent a multistep process using an overall equilibrium expression derived from the equilibrium expressions of each individual step.	• I can explain that reversing a reaction inverts the value of the equilibrium constant. • I can explain that multiplying a reaction by a factor changes the equilibrium constant by raising it to that same power. • I can explain that adding multiple reactions together results in an overall equilibrium constant equal to the product of the individual equilibrium constants. • I can apply the same algebraic rules used for equilibrium constants to reaction quotients because they share the same mathematical form.	
7.7 Calculating Equilibrium Concentrations	**Learning Target 7** I can identify the concentrations or partial pressures of chemical species at equilibrium using the balanced equation, initial conditions, and the equilibrium constant (K).	• I can predict whether a reaction will shift toward products or reactants by comparing the reaction quotient (Q) to the equilibrium constant (K). • I can explain that when Q < K, the reaction shifts toward products, consuming reactants. • I can explain that when Q > K, the reaction shifts toward reactants, consuming products. • I can explain that when Q = K, the system is at dynamic equilibrium with no net change in concentrations or partial pressures.	
7.8 Representations of Equilibrium	**Learning Target 8** I can represent a system undergoing a reversible reaction using a particulate model.	• I can use particulate diagrams to show the relative numbers of reactant and product particles before equilibrium is reached. • I can use particulate models to represent the composition of a system at equilibrium. • I can connect particulate representations to the magnitude of the equilibrium constant to describe whether reactants or products are favored.	
7.9 Introduction to Le Châtelier's Principle	**Learning Target 9** I can identify how a system at equilibrium responds to external stresses using Le Châtelier's principle.	• I can predict how adding or removing a chemical species shifts the position of equilibrium. • I can predict how changes in temperature affect the direction of a shift in equilibrium. • I can predict how changes in volume or pressure affect gas-phase equilibrium systems. • I can explain how dilution affects the equilibrium position in a solution. • I can use Le Châtelier's principle to predict changes in observable properties such as pH, temperature, and color.	**Demo with Q&A** Prepare a solution of cobalt (II) chloride in dry ethanol. Demonstrate various methods to shift the equilibrium position: adding water, heating, cooling, layering with dry acetone, adding silver nitrate to precipitate chloride ions from solution, and measuring the temperature change of the solution as concentrated hydrochloric acid is added. As a class, have students analyze what each change does to the predominant species in the equilibrium mixture and then generalize patterns for LeChâtelier's

			principle.
7.10 Reaction Quotient and Le Châtelier's Principle	**Learning Target 10** I can explain the relationship between Q, K, and the direction a reversible reaction will proceed to reach equilibrium.	• I can describe how a disturbance to a system at equilibrium causes Q to differ from K, shifting the system out of equilibrium. • I can explain that the system responds by shifting the reaction forward or reverse until Q again equals K. • I can distinguish between changes that affect Q (such as concentration, pressure, or volume changes) and changes that affect K (such as temperature changes). • I can explain how changes in concentration or pressure affect Q, while temperature changes alter the value of K and lead to a new equilibrium state.	**Demo with Q&A** (see 7.9)
7.11 Introduction to Solubility Equilibria	**Learning Target 11** I can calculate the solubility of a salt using its solubility product constant (Ksp).	• I can describe dissolution as a reversible process that reaches equilibrium in saturated solutions. • I can use Ksp values to compare the relative solubility of different salts. • I can relate solubility rules to Ksp values to predict whether a salt is soluble or insoluble. • I can determine Ksp from the molar solubility of a substance in a saturated solution.	**Post-Lab Discussion** After examining the K sp tables for patterns (including ion charge, ionic radius, polyatomic vs. monoatomic ions, etc.), have students investigate the K sp of lead (II) iodide. One drop of 0.1 M potassium iodide is added to 250 mL of 0.01 M lead (II) nitrate. A precipitate forms but then dissolves as it dissipates through the solution. Based on K sp, have students calculate whether the precipitate should have formed and connect this calculation with what was initially observed. Have them determine how many milliliters of the 0.1 M KI solution would need to be added for a lasting precipitate to be formed. Then have them share their calculated values and agree as a class which is the best answer
7.12 Common-Ion Effect	**Learning Target 12** I can identify how the solubility of a salt changes when a common ion is already present in solution.	• I can explain that the presence of a common ion reduces the solubility of a salt by shifting the dissolution equilibrium. • I can use Le Châtelier's principle to predict the direction of shift in a system affected by the common-ion effect. • I can calculate solubility or Ksp values in solutions that already contain a common ion.	

Progress Check 3

- Multiple-choice: ~30 questions
- Free-response: 2 questions
  - Short
  - Long

Preparing for the AP Exam

On the AP Exam, students must be able to connect what is happening at the molecular level to a model for a system at equilibrium. For example, when students are asked to connect the value of the equilibrium constant (K) from the equilibrium expression to the dominant direction of the reaction, they struggle to connect the value of a large K to a

reaction proceeding essentially to completion. This lack of connection leads students to use ineffective mathematical routines and then incorrectly calculate the concentration of the product in solution. To help students avoid this type of misunderstanding, teachers can ensure that students connect the value of the equilibrium constant to the experimental data or observations provided. Additionally, teachers can help students visualize the effects of a large or small equilibrium constant on the concentrations of all species in equilibrium.

Unit Title:

Unit 8: Acids and Bases

Relevant Standards: Bold indicates priority

8.1 Introduction to Acids and Bases: Students can identify and apply appropriate acid–base theories, definitions, and mathematical relationships to solve chemical problems.

8.2 pH and pOH of Strong Acids and Bases: Students can select and apply appropriate mathematical relationships to determine pH and pOH for strong acid and base solutions.

8.3 Weak Acid and Base Equilibria: Students can explain relationships between variables in acid–base equilibrium expressions and predict how changes in one variable affect the system.

8.4 Acid–Base Reactions and Buffers: Students can calculate, estimate, and predict equilibrium quantities in acid–base systems, including buffer solutions, using logical computational pathways and graphical representations.

8.5 Acid–Base Titrations: Students can analyze titration data to determine the concentration of an unknown acid or base and interpret titration curves to identify key equivalence and buffering regions.

8.6 Molecular Structure of Acids and Bases: Students can use particulate-level representations and molecular structure to support claims about acid and base behavior and justify those claims using chemical principles and data.

8.9 Henderson–Hasselbalch Equation: Students can calculate pH, pOH, or related quantities in buffer systems using the Henderson–Hasselbalch equation with appropriate precision.

8.10 Buffer Capacity: Students can explain how buffer systems resist changes in pH and analyze how experimental conditions and sources of error may affect buffer performance.

8.11 pH and Solubility: Students can interpret experimental data and representations to analyze the relationship between pH and solubility in chemical systems.

Essential Question(s):	**Enduring Understanding(s):**
1. How is pH related to the concentration and strength of an acid, a base, or a mixture of them? 2. How does acid or base strength relate to the concentrations of reactants and products when a system reaches equilibrium? 3. Why are some acids stronger than others? 4. How does your body maintain pH balance?	1. Acid–base behavior can be understood in terms of proton transfer and equilibrium, where the relative strengths of acids and bases determine the extent to which reactants and products are favored. 2. pH and pOH are quantitative measures of hydrogen ion and hydroxide ion concentrations, and they provide a consistent way to describe and compare acidic and basic solutions, including mixtures and buffers. 3. Acid and base strength is directly related to equilibrium position: stronger acids and bases dissociate more completely in water, leading to larger concentrations of products at equilibrium. 4. Buffer systems resist changes in pH because they contain conjugate acid–base pairs that can shift equilibrium to consume added acids or bases. 5. The structure and bonding within molecules influence acid–base behavior, including why some substances act as strong acids or bases while others remain weak.
Demonstration of Learning:	**Pacing for Unit**
• Success with unit MCQs and FRQs via AP Classroom • Quizzes (AP Classroom MCQs and FRQs) every few topics • Practice converting using the pH loop (between pH, pOH, [H+], [OH-] • Practice writing out equilibrium expressions of weak	18 instructional days 4 lab days 3 days for review / assess

acid or base in water • Standardization of NaOH Lab • Determining the Ka of Acetic Acid Lab • Strength of Acids POGIL • Practice predicting the pH of different solutions / different points within a titration • Practice with strong acid/ strong base titration and neutralization calculations • Practice with weak / strong titration and neutralization calculations • Progress Check 8 from AP Classroom Unit 8 is 11-15% See AP Exam Description	
Family Overview (link below)	**Integration of Technology:**
AP Chemistry FAMILY MATERIALS	N/A
Unit-specific Vocabulary:	**Aligned Unit Materials, Resources, and Technology (beyond core resources):**
Acid, base, Brønsted–Lowry acid, Brønsted–Lowry base, Lewis acid, proton donor, proton acceptor, acidic hydrogen, Lewis base, strong acid, weak acid, strong base, weak base, pH, pOH, Ka, Kb, pKa, pKb, percent ionization, hydrolysis, buffer, buffer capacity, titration, equivalence point, half equivalence point, Henderson–Hasselbalch equation	• pH probes with Labquests • Stir plates • Burets • Conductivity tester lightbulb
Opportunities for Interdisciplinary Connections:	**Anticipated misconceptions:**
• Biology: blood pH regulation, digestion, enzyme function • Environmental Science: acid rain, soil chemistry, aquatic ecosystems • Medicine: antacids, buffer systems, pharmaceuticals • Agriculture: soil pH and nutrient availability • Engineering: wastewater treatment and corrosion prevention	• Students must memorize the strong acids and bases in order to correctly determine what a substance will do in water (i.e. dissociate fully or only partially and reach equilibrium) • Difficulty with certain math skills that have not yet been taught in math class (logarithms and 10^x) • Difficulty determining when to use a RICE or BCA table • Picking the incorrect Ka or Kb constant • Misidentifying pH at key points during a titration • When do use calculations vs when to predict based on species in solution • Using insufficient explanations to justify claims on FRQs regarding acid or base strength, buffer capacity • Difficulty gathering important info from a titration curve to use in a calculation or justification of FRQ
Connections to Prior Units:	**Connections to Future Units:**
• Students must be able to write the net ionic equations for the dissolution of acids and bases (4.2) • Students must be able to make solutions and do calculations with concentrations (3.7) • Uses equilibrium expressions and RICE tables introduced in unit 7 • Acid/base neutralization reactions and titrations are introduced in Unit 4 • Students must be able to draw and analyze Lewis structures to determine acidic hydrogens • Electrolytes are introduced in Unit 2 with ionic compounds. In unit 8, students can analyze weak and strong acids and bases based on the percent ionization (and thus, how electrolytic these	N/A

substances are).
- 8.11 uses Le Chatelier's and solubility product constant to predict the solubility of salts in acidic or basic solutions.

Differentiation through *Universal Design for Learning*

	Engagement	Representation	Action & Expression
LT1	Logarithmic scale exploration mapping household liquids (lemon juice, bleach) onto a pH scale.	pH mathematical square visual connecting [H3O+], [OH-], pH, and pOH via 10^-14 and pH + pOH = 14.	Convert between [H3O+], [OH-], pH, and pOH using problem sets, spreadsheets, or video walkthroughs.
LT2	"100% Dissociation Challenge" calculating pH for strong monoprotic/diprotic acids and group 1/2 metal hydroxides.	Complete ionization visual chart: [H3O+] = M_acid for strong monoprotic acids; [OH-] = 2 M_base for Ca(OH)2.	Calculate pH and pOH of strong acid and strong base solutions of varying concentrations with worked steps.
LT3	Weak acid ionization lab measuring pH of acetic acid vs. hydrochloric acid at identical molarities.	Weak acid K_a RICE table scaffold and percent ionization formula.	Calculate K_a/K_b, percent ionization, and pH for weak acids and weak bases using RICE tables and handwritten solutions.
LT4	Stoichiometric neutralization lab mixing strong acid/base combinations and calculating excess reactant pH.	ICF (Initial-Change-Final) mole table scaffold for neutralization stoichiometry followed by V_total molarity conversion.	Solve neutralization stoichiometry problems by submitting ICF tables and calculating final pH based on excess reactant.
LT5	Digital titration curve analysis plotting pH vs. volume of titrant for strong vs. weak acid titrations.	Annotated titration curve diagram labeling initial pH, half-equivalence point (pH = pKa), equivalence point, and buffer region.	Interpret titration curves to identify equivalence point volume, pKa, and unknown acid concentration via written or oral reports.
LT6	Molecular structure comparison analyzing acid strength trends across binary acids (HF vs. HCl) and oxyacids (HClO vs. HClO4).	Structural diagram highlighting bond polarity, bond strength, electronegativity, and conjugate base anion stability.	Write a CER argument explaining why acid strength increases with higher central atom electronegativity and oxygen count in oxyacids.
LT7	Speciation curve analysis evaluating the relative abundance of [HA] vs. [A-] as solution pH changes.	Speciation scale visual: pH < pKa -> [HA] > [A-]; pH = pKa -> [HA] = [A-]; pH > pKa -> [A-] > [HA].	Predict the dominant chemical species in a weak acid solution at specified pH values and support predictions mathematically.
LT8	"Buffer Resistance Lab" adding strong acid/base drops to water vs. a buffered solution while monitoring pH.	Particulate reaction diagram showing buffer action: added H+ consumed by A-; added OH- consumed by HA.	Draw particulate diagrams illustrating how a buffer consumes added H+ or OH- ions and write corresponding net ionic equations.
LT9	"Buffer Recipe Lab" mixing calculated volumes of weak acid and conjugate base to hit a target pH.	Henderson-Hasselbalch equation scaffold: pH = pKa + log([A-] / [HA]) with logarithmic ratio rules.	Calculate buffer pH using the Henderson-Hasselbalch equation and design a buffer solution recipe for a specific target pH.
LT10	Buffer capacity stress testing adding large amounts of strong titrant to 0.1 M vs. 1.0 M buffer solutions until capacity fails.	Visual capacity comparison showing high concentration buffers containing more moles of HA and A- to neutralize added stress.	Write an analysis comparing two buffer solutions of identical pH but different concentrations, explaining differences in buffer capacity.
LT11	Dissolution lab adding acid to CaCO3 or Mg(OH)2 precipitates to observe dramatic solubility increases.	Coupled equilibrium visual showing added H+ removing basic anions (CO3^2- or OH-), shifting K_sp right.	Explain why salts containing basic anions become more soluble in acidic solutions using Le Châtelier's principle and coupled equations.

Supporting Multilingual/English Learners (*CELP standards*)

	Emerging	Expanding	Bridging
LT1	Calculate pH from [H3O+] using a pH square chart (pH + pOH = 14) and calculator steps.	Calculate pH, pOH, [H3O+], and [OH-] using formula templates and logarithmic calculation guides.	Convert fluently between pH, pOH, [H3O+], and [OH-] at various temperatures, justifying Kw temperature dependence.

LT2	Identify strong acids/bases from a reference list and set [H3O+] = [Acid]_0 for 1:1 strong acids.	Calculate pH of strong acid and base solutions (accounting for stoichiometry like Ca(OH)2) using calculation templates.	Calculate pH and pOH for single and mixed strong acid/base solutions before reaction, writing complete justifications.
LT3	Identify weak acids/bases as "partially ionized" (<< 100%) using particulate dissociation diagrams.	Calculate pH or Ka/Kb for weak acids/bases using structured ICE tables and percent ionization templates.	Solve weak acid/base equilibrium problems, calculating Ka, Kb, pH, and percent ionization, verifying the 5% approximation.
LT4	Identify excess reactant (acid or base) in a neutralization reaction using BCA table templates.	Calculate pH after mixing strong acids and bases using structured stoichiometry (BCA) tables and excess volume frames.	Calculate equilibrium pH and all major species concentrations following complete or partial neutralization reactions.
LT5	Label key points on a titration curve (initial pH, half-equivalence, equivalence point) using a visual guide.	Explain titration curve features (buffer region, equivalence point pH) using structured comparison templates.	Interpret mono- and polyprotic titration curves, calculating Ka from pH = pKa at half-equivalence and explaining pH at equivalence.
LT6	Match acid strength to electronegativity or bond strength trends using visual structural formula cards.	Explain acid strength trends (binary acids, oxyacids) using bond strength and electronegativity sentence frames.	Justify relative strengths of binary acids and oxyacids based on bond polarity, bond strength, and conjugate base resonance stabilization.
LT7	Select dominant species ([HA] > [A-] or [A-] > [HA]) using a pH vs. pKa number line card.	Determine major species in solution by comparing pH to pKa using structured conditional sentence frames.	Predict and justify dominant conjugate species concentrations at any given pH using the Henderson-Hasselbalch relationship.
LT8	Identify buffer solutions as containing "weak acid + conjugate base" using chemical recipe cards.	Explain how a buffer resists pH change upon adding H+ or OH- using neutralization reaction frames.	Explain buffer action at the particulate level, writing net ionic equations for buffer component reactions with added strong acids/bases.
LT9	Calculate buffer pH using the Henderson-Hasselbalch equation template: pH = pKa + log([A-] / [HA]).	Calculate buffer pH and prepare buffers of specific pH using guided Henderson-Hasselbalch worksheets.	Calculate buffer pH before and after adding strong acid/base titrants, combining stoichiometry and Henderson-Hasselbalch calculations.
LT10	Select the buffer with highest capacity (highest concentration of components) using visual concentration cards.	Explain buffer capacity differences using comparative sentence frames (*"Buffer A has higher capacity because..."*).	Evaluate buffer capacity and optimal buffer range (pH = pKa +/- 1), designing buffers for target pH values.
LT11	Identify salts containing basic anions (e.g., F-, CO3^2-) whose solubility increases in acid using a visual rule card.	Explain how adding acid increases salt solubility by removing basic anions using Le Châtelier sentence starters.	Predict and mathematically/conceptually justify how changing pH affects salt solubility using coupled equilibrium equations.

Unit Overview

In Unit 8, students will apply the explanations and calculations they learned in Unit 7 to the acid-base equilibrium system. Students will collect titration data and develop titration curves to represent a variety of acid-base systems. They will analyze these titration curves to describe the similarities and differences between a strong acid-strong base and a weak acid-strong base or weak base-strong acid titration, identify the equivalence points and the half-equivalence points, and identify the buffering regions of the curves. Students will use the information presented graphically in the titration curves to complete calculations to find the equilibrium constant for the reactions (Ka or Kb), determine the concentration of an unknown, and support claims about how a particular buffer system may work when an acid or base is introduced. From these calculations and what is known about the chemical system, students will then develop explanations for how potential sources of error may have affected experimental results and associated calculations.

Lesson	Learning Target	Success Criteria	Suggested Activities
8.1 Introduction to Acids and Bases	**Learning Target 1** I can calculate pH and pOH using hydronium and hydroxide ion concentrations.	• I can use pH = -log[H3O+] and pOH = -log[OH-] to determine acidity or basicity. • I can relate hydronium and hydroxide ion concentrations using Kw. • I can explain that at 25°C, Kw = 1.0 × 10^-14 and pH + pOH = 14. • I can describe how pure water is neutral at 25°C but may not be neutral at other temperatures.	

8.2 pH and pOH of Strong Acids and Bases	**Learning Target 2** I can calculate pH and pOH for strong acid and strong base solutions.	• I can determine hydronium ion concentration from the initial concentration of a strong acid. • I can determine hydroxide ion concentration from the initial concentration of a strong base. • I can use the fact that strong acids and bases fully dissociate in water. • I can calculate pH and pOH from ion concentrations using logarithmic relationships.	**Post-Lab Discussion** Rainbow Acid Indicator (Flinn Scientific Item U0012) is added to 0.001 M solutions of HCl, H_2SO_4, and $HC_2H_3O_2$. Have students reason out why the pH values are not the same, concept of Ka. Then have them calculate the pH of each solution to explain their earlier observations. Percent ionization is discussed and how ICE charts reflect the percent ionization is explained.
8.3 Weak Acid and Base Equilibria	**Learning Target 3** I can explain how pH, pOH, and species concentrations are related in weak acid and weak base solutions.	• I can describe that weak acids and weak bases only partially ionize in water and establish equilibrium with their conjugates. • I can explain that pH of a weak acid depends on initial concentration and Ka (or pKa), and weak base pH depends on initial concentration and Kb (or pKb). • I can relate Ka and Kb through Kw for conjugate acid-base pairs. • I can calculate percent ionization of weak acids or bases using equilibrium and initial concentration values.	**Post-Lab Discussion** (See 8.2)
8.4 Acid-Base Reactions and Buffers	**Learning Target 4** I can explain how major species concentrations change when strong acids and strong bases are mixed, resulting in complete neutralization and determining pH from excess reactant.	• I can describe how mixing a weak acid with a strong base can produce a buffer, where pH depends on the relative amounts of the weak acid and its conjugate base. • I can explain how mixing a weak base with a strong acid can also produce a buffer, where pH depends on the ratio of the weak base and its conjugate acid. • I can determine pH in acid–base mixtures based on whether one reactant is in excess, equimolar, or forms a buffer system. • I can explain that when weak acids and weak bases are mixed, the system reaches equilibrium with both conjugate acid–base pairs present.	
8.5 Acid-Base Titrations	**Learning Target 5** I can explain how titration curves represent pH changes as a titrant is added to a mono- or polyprotic acid or base solution.	• I can determine the concentration of an unknown solution using the equivalence point where moles of acid and base are equal. • I can explain that at the half-equivalence point of a weak acid or base titration, the pH equals the pKa of the conjugate acid-base pair. • I can describe how the pH at the equivalence point depends on	**Post-Lab Discussion** After collecting data on a weak acid/strong base titration, have students create a titration curve (pH as a function of the volume of base added). Then have them identify relative points on the graph based on group discussion (e.g., equivalence point).

		the strength of the acid or base and the properties of the conjugate species formed. • I can use titration curves to identify the number of acidic protons in a polyprotic acid and determine corresponding pKa values.	
8.6 Molecular Structure of Acids and Bases	**Learning Target 6** I can explain how the strength of an acid or base is related to its molecular or ionic structure.	• I can identify which protons in a molecule are likely to participate in acid–base reactions based on structure. • I can explain that strong acids have weak conjugate bases that are stabilized by electronegativity, resonance, or inductive effects. • I can describe how structural features such as electronegative atoms increase acid strength by stabilizing the conjugate base. • I can compare strong and weak acids and bases by relating their strength to the stability of their conjugate partners.	
8.7 pH & pKa	**Learning Target 7** I can explain how the dominant form of a weak acid or base in solution depends on the relationship between pH and pKa (or pKb).	• I can predict whether the protonated or deprotonated form is more abundant by comparing pH to pKa. • I can explain that when pH is less than pKa, the protonated (acid) form is predominant, and when pH is greater than pKa, the deprotonated (base) form is predominant. • I can describe how acid–base indicators change color based on shifts between protonated and deprotonated forms. • I can explain why appropriate indicators are chosen in titrations based on matching their pKa to the equivalence point pH.	
8.8 Properties of Buffers	**Learning Target 8** I can explain how buffers stabilize pH by containing significant amounts of a conjugate acid-base pair.	• I can describe how the conjugate base in a buffer reacts with added acid to minimize pH change. • I can describe how the conjugate acid in a buffer reacts with added base to minimize pH change. • I can explain that buffer effectiveness depends on the ability of both components to neutralize added acid or base. • I can connect buffer action to the equilibrium between a weak acid and its conjugate base.	**Demo with Q&A** Add an Alka-Seltzer tablet to 200 mL of water and pour the resulting solution into three small beakers. Add deionized water to three more beakers. Add universal indicator to all six beakers and then add strong acids and strong bases to each beaker to demonstrate buffering ability and buffer capacity. Have students develop particulate-level drawings to illustrate what is happening in the beakers in the context of “buffering ability.”
8.9 Henderson-Hasselbalch Equation	**Learning Target 9** I can identify the pH of a buffer solution using the pKa of the weak acid and the ratio of conjugate base to weak acid concentrations.	• I can explain that buffer pH depends on the equilibrium between a weak acid and its conjugate base. • I can describe that adding small amounts of acid or base does not significantly change the ratio of conjugate pairs, so pH changes very little. • I can explain that buffers resist pH change because added acid or base is consumed by the conjugate pair. • I can relate buffer pH to the relative amounts of acid and base in the buffer system.	**Simulations** Using a Chem Collective virtual lab, ask students to develop a buffer that will have a particular pH after an amount of strong acid is added.

8.10 Buffer Capacity	**Learning Target 10** I can explain how buffer capacity depends on the total concentrations of the conjugate acid and conjugate base.	• I can explain that increasing the concentrations of both buffer components increases buffer capacity without changing pH if the ratio stays the same. • I can describe that buffer capacity is the ability of a buffer to resist changes in pH when acid or base is added. • I can explain that a buffer with more conjugate acid than base resists added base more effectively than added acid. • I can explain that a buffer with more conjugate base than acid resists added acid more effectively than added base.	**Simulations** (See 8.9)
8.11 pH and Solubility	**Learning Target 11** I can explain how changes in pH can affect the solubility of certain salts.	• I can identify that salts containing ions that act as weak acids, weak bases, or hydroxide ions are pH sensitive. • I can describe how adding acid or base can shift equilibrium and change solubility. • I can predict whether solubility increases or decreases based on the type of ion present. • I can relate pH changes to shifts in dissolution equilibrium.	

Progress Check 3

- Multiple-choice: ~30 questions
- Free-response: 1 question
    - Long

Preparing for the AP Exam

On the AP Exam, students must be able to use experimental data to make calculations and support claims. Students often encounter difficulty with questions that require them to use titration curves to identify the equivalence and half-equivalence points or to complete calculations or estimations of either the concentration or pH of an unknown at a particular point on the curve. They also struggle to justify the selection of an appropriate indicator for the end point of the titration. In these situations, students can face challenges with unit conversion, or they can confuse half-equivalence, equivalence, and endpoint. Students may also struggle to understand what is represented in different types of titration curves. Teachers can provide students with multiple opportunities to describe why titration curves have characteristic shapes for certain acid-base equilibrium systems. Teachers can also provide opportunities to choose and implement mathematical routines to manipulate and interpret titration data and connect that interpretation to chemistry concepts. The half-equivalence point is a helpful reference point to visualize ratios of acid/conjugate base, particularly when combined with particulate representations.

Unit Title:

Unit 9: Thermodynamics and Electrochemistry

Relevant Standards: Bold indicates priority

9.1 Introduction to Entropy: Students can use particulate-level models to explain entropy as the number of possible arrangements of particles and energy within a system.

9.2 Absolute Entropy and Entropy Change: Students can calculate and predict entropy changes for chemical and physical processes using thermodynamic data and mathematical relationships.

9.3 Gibbs Free Energy and Thermodynamic Favorability: Students can explain how enthalpy, entropy, and temperature determine whether a process is thermodynamically favorable.

9.4 Thermodynamic and Kinetic Control: Students can distinguish between thermodynamic favorability and reaction rate by connecting particulate-level processes to macroscopic observations.

9.5 Free Energy and Equilibrium: Students can use thermodynamic principles and mathematical relationships to explain the connection between Gibbs free energy and equilibrium.

9.6 Free Energy of Dissolution: Students can explain how intermolecular interactions and entropy changes affect the favorability of dissolution processes.

9.7 Coupled Reactions: Students can explain how energy released from one process can drive another process that is not thermodynamically favorable.

9.8 Galvanic (Voltaic) and Electrolytic Cells: Students can analyze electrochemical cells and predict how modifications to cell conditions or procedures affect cell behavior and experimental outcomes.

9.9 Cell Potential and Free Energy: Students can calculate relationships among cell potential, free energy, and spontaneity using electrochemical data.

9.10 Cell Potential Under Nonstandard Conditions: Students can justify how concentration and nonstandard conditions affect cell potential using chemical principles and mathematical reasoning.

9.11 Electrolysis and Faraday's Law: Students can apply Faraday's Law and electrochemical relationships to calculate quantities involved in electrolysis processes.

Essential Question(s):	**Enduring Understanding(s):**
1. Why do some chemical reactions occur without intervention, but others require the input of energy? 2. How can we determine the conditions under which a chemical or physical transformation is likely to occur? 3. How is electrical energy generated using chemical reactions?	1. The spontaneity of a chemical or physical process depends on the balance between energy changes and entropy changes within a system and its surroundings. 2. Gibbs free energy provides a quantitative way to predict whether a process is thermodynamically favorable under specific conditions. 3. Chemical reactions can transfer energy in the form of electrical work through the movement of electrons in redox reactions. 4. Electrochemical cells use oxidation–reduction reactions to convert chemical energy into electrical energy or use electrical energy to drive nonspontaneous reactions. 5. The direction and favorability of redox processes depend on factors such as cell potential, concentration, and thermodynamic conditions.
Demonstration of Learning:	**Pacing for Unit**
• Success with unit MCQs and FRQs via AP Classroom • Quizzes (AP Classroom MCQs and FRQs) every few topics	10 instructional days 2 lab days 2 days review / assess

• Practice with entropy, enthalpy, and Gibb's free energy calculations • What's the Driving Force Lab • Building a Galvanic Cell worksheet • Practice with cell potential calculations at standard and nonstandard conditions • Practice with Nernst calculations and responses to FRQs • Progress Check 9 from AP Classroom Unit 9 is 7-9% of AP Exam Weighting See AP Exam Description	
Family Overview (link below)	**Integration of Technology:**
AP Chemistry FAMILY MATERIALS	N/A
Unit-specific Vocabulary:	**Aligned Unit Materials, Resources, and Technology (beyond core resources):**

Entropy, Gibbs free energy, microstates, arrangements, positional possibilities, oxidation, reduction, redox reaction, oxidation number, half-reaction, electrochemical cell, galvanic cell, voltaic cell, anode, cathode, electrode, salt bridge, electrolyte, standard reduction potential, cell potential (Ecell), Nernst equation, spontaneous reaction, thermodynamically favorable, nonspontaneous reaction, thermodynamically unfavorable, Faraday's constant, electrolysis, concentration cell, fuel cell, standard hydrogen electrode (SHE)	• Multimeters to build galvanic cells
	  ○ Zinc and copper electrodes
	  ○ Batteries
	• Steel wool

Opportunities for Interdisciplinary Connections:	**Anticipated misconceptions:**
• Physics: entropy, spontaneity, and free energy • Biology: ATP and biochemical energy transformations • Environmental Science: energy efficiency and sustainability • Engineering: fuel cells, batteries, and energy systems • Earth Science: natural energy flow and geochemical processes	• Confusing enthalpy and entropy • Insufficient or inappropriate language to answer Unit 9 FRQs ○ disorder vs microstates, spontaneous vs thermodynamically favorable • Not considering the order of importance for determining change in entropy (change in moles of gas is more important than change in state) and explanation as to why • Assuming no appreciable reaction means it is at equilibrium or will not occur (vs under kinetic control) • Not changing the sign of the oxidized half reaction • Difficulty identifying parts of a galvanic cell • Not considering standard vs nonstandard conditions when calculating cell potential.
Connections to Prior Units:	**Connections to Future Units:**
• Students must know how to identify the oxidation number of all species within a reaction and write out oxidation / reduction half reactions (4.9) • Students must be able to calculate enthalpy using the 4 methods introduced in unit 6 • Students must be able to calculate the reaction quotient and equilibrium constant and determine the direction of reaction based on the magnitude (unit 7) • Students must be able to use dimensional analysis and stoichiometry to calculate change in mass of electrode when a current is applied (pre-work and Unit 4)	N/A
Differentiation through *Universal Design for Learning*	

	Engagement	Representation	Action & Expression
LT1	"Microstate Sorting Game" arranging gas particles across connected flasks to observe probability of distribution/disorder (ΔS).	Conceptual matrix mapping entropy increases (ΔS > 0) to phase changes (s -> l -> g), temperature increases, and gas mole generation.	Predict the sign of ΔS° for physical phase changes and chemical reactions, justifying answers with structural molecular reasoning.
LT2	Molar entropy calculation challenge computing ΔS°_rxn for synthesis vs. decomposition reactions.	Standard entropy formula scaffold: ΔS°_rxn = sum(n × S°(products)) - sum(m × S°(reactants)) (noting S° > 0 for elements).	Calculate ΔS°_rxn from standard molar entropy tables and confirm that the algebraic sign matches structural predictions.
LT3	Gibbs free energy decision lab evaluating temperature dependent spontaneity for ice melting vs. combustion.	ΔG° = ΔH° - TΔS° sign matrix detailing thermodynamic favorability conditions (ΔG° < 0 -> Favored).	Calculate ΔG° using enthalpy and entropy values and state whether a reaction is thermodynamically favored at given temperatures.
LT4	"Thermodynamic vs. Kinetic Control Debate" analyzing why diamond conversion to graphite (ΔG° < 0) does not occur perceptibly.	Energy profile diagram contrasting a large negative ΔG° (thermodynamically favored) with a high activation energy E_a barrier (kinetic control).	Write a CER response explaining why a thermodynamically favored reaction may require a catalyst or high heat to proceed at a measurable rate.
LT5	Mathematical modeling of ΔG° = -RT ln(K) observing how equilibrium constants scale exponentially with free energy changes.	Dual-axis conversion chart connecting ΔG° < 0 <-> K > 1 (product favored) and ΔG° > 0 <-> K < 1 (reactant favored).	Calculate K from ΔG° values using ΔG° = -RT ln(K) and explain the thermodynamic favorability of the system.
LT6	Dissolution enthalpy/entropy audit evaluating why endothermic dissolution processes (e.g., NH4NO3) occur spontaneously due to +ΔS.	Dissolution thermodynamic balance diagram illustrating solute/solvent lattice breaking (ΔH > 0) vs. hydration entropy gain (ΔS > 0).	Write an explanation detailing how the balance between ΔH_soln and ΔS_soln determines salt solubility across temperatures.
LT7	Electrolysis demonstration using an external power source to drive water splitting (ΔG° > 0) into H2 and O2.	Reaction coupling schematic showing an unfavorable process (ΔG1° > 0) coupled to a favorable process or external current (ΔG_overall° < 0).	Design a coupled reaction scheme or explain how electrolysis uses electrical energy to drive a non-spontaneous reaction.
LT8	Construction lab building a Galvanic cell (Cu/Zn) using beakers, wire, voltmeter, and a salt bridge.	Diagram of a galvanic cell labeling Anode (oxidation, -), Cathode (reduction, +), electron flow (A -> C), and ion movement in the salt bridge.	Construct a functional galvanic cell, draw a fully labeled diagram, and write net ionic equations and electron/ion flow directions.
LT9	Standard reduction potential table activity calculating E°_cell = E°_cat - E°_an for various metal electrode combinations.	Formula connection visual linking E°_cell > 0 <-> ΔG° < 0 via ΔG° = -nFE°_cell (where F = 96,485 C/mol e-).	Calculate E°_cell and ΔG° for electrochemical cells and determine whether cell reactions are thermodynamically favored.
LT10	Non-standard voltaic cell lab measuring voltage shifts as ion concentrations in anode/cathode compartments change.	Nernst equation visual scaffold [E_cell = E°_cell - (RT / nF) ln(Q)] linking Q < 1 -> E_cell > E°_cell and Q > 1 -> E_cell < E°_cell.	Predict how altering reactant/product concentrations shifts Q and cell potential E_cell away from standard potential E°_cell.
LT11	Electroplating lab electroplating copper onto a nickel coin using a constant current power supply.	Stoichiometric electrochemistry bridge: Current (I = q / t) -> Charge (q) -> Moles of e- -> Moles/Grams of Metal.	Solve stoichiometry problems for electrolysis and electroplating using I = q / t and Faraday's constant (F = 96,485 C/mol e-).

Supporting Multilingual/English Learners (*CELP standards*)

	Emerging	Expanding	Bridging
LT1			
LT2			
LT3			
LT4			
LT5			

Unit Overview

To achieve success in AP Chemistry and in Unit 9 specifically ,students must connect principles and calculations across the areas of kinetics, thermodynamics, equilibrium, and electrochemistry to explain and support claims about

what is happening in chemical systems. Sometimes support of such claims comes from students being able to estimate an approximate value of a required characteristic of a chemical system rather than use a routine application of an algorithm. Students are introduced to entropy as a factor that is necessary to explain why some endothermic reactions occur in spite of the higher energy that products may have in these changes. Students will use particulate representations and graphical distribution of kinetic energy to describe the increase in entropy with increasing temperature. In addition, students will explore how to use Gibbs free energy for determining the thermodynamic favorability by considering the change in both enthalpy and entropy. Students will use the concepts of thermodynamics to generate more comprehensive claims about what is happening in a galvanic or electrolytic cell.

Lesson	Learning Target	Success Criteria	Suggested Activities
9.1 Introduction to Entropy	**Learning Target 1** I can identify whether entropy increases or decreases during chemical and physical processes.	• I can explain that entropy increases when matter becomes more dispersed, such as during phase changes from solid to liquid or liquid to gas. • I can describe how entropy increases when gases expand or when the number of gas particles increases in a reaction. • I can explain that entropy increases with temperature because energy becomes more dispersed among particles. • I can compare relative entropy changes based on changes in phase, volume, number of gas particles, and temperature.	
9.2 Absolute Entropy and Entropy Change	**Learning Target 2** I can calculate the standard entropy change for a reaction using standard molar entropy values.	• I can determine entropy change by subtracting the total entropy of reactants from the total entropy of products. • I can apply tabulated entropy data to chemical and physical processes. • I can predict whether entropy increases or decreases based on calculated values. • I can use entropy values to analyze disorder changes in a system.	
9.3 Gibbs Free Energy and Thermodynamic Favorability	**Learning Target 3** I can explain whether a process is thermodynamically favored based on the sign of Gibbs free energy change (ΔG°).	• I can determine that a process is thermodynamically favored when ΔG° is less than zero. • I can calculate ΔG° using standard Gibbs free energy of formation values for reactants and products. • I can calculate ΔG° using the relationship between enthalpy, entropy, and temperature. • I can predict thermodynamic favorability by analyzing the signs of ΔH°, ΔS°, and temperature without full calculation.	**Think-Pair-Share** Given a problem pertaining to thermodynamic favorability, have students think through how enthalpy and/or entropy is driving the thermodynamic favorability of the reaction. Have them pair up and explain their reasoning for whether or not the reaction is thermodynamically favorable and what is driving that favorability. After the pairs have discussed their responses, have them share with other pairs to get feedback on their rationale.
9.4 Thermodynamic and Kinetic Control	**Learning Target 4** I can explain that a thermodynamically favored reaction may still not occur at a measurable rate due to kinetic barriers.	• I can describe that high activation energy can prevent a reaction from proceeding quickly even if it is thermodynamically favorable. • I can explain that such reactions are said to be under kinetic control. • I can distinguish between thermodynamic favorability and reaction rate. • I can explain that a lack of observable reaction does not necessarily mean the system is at equilibrium.	
9.5 Free Energy and Equilibrium	**Learning Target 5** I can explain whether a process is thermodynamically favored using the	• I can describe that a process is thermodynamically favored when ΔG° is less than zero and K is greater than 1. • I can relate ΔG° and K using the equations $\Delta G^\circ = -RT \ln K$ and $K = e^{\wedge}(-\Delta G^\circ/RT)$. • I can explain that when ΔG° is close to zero, K is close to 1 and neither reactants nor products are strongly favored.	

	relationship between ΔG°, K, and temperature.	• I can predict whether reactants or products are favored by comparing the sign and magnitude of ΔG° relative to RT.	
9.6 Free Energy of Dissolution	**Learning Target 6** I can explain how the solubility of a salt depends on changes in enthalpy and entropy during dissolution.	• I can describe how breaking the ionic lattice and forming solute–solvent interactions affects the enthalpy change of dissolution. • I can explain how increased disorder from dissolved ions contributes to the entropy change. • I can describe how solvation of ions by water molecules influences both enthalpy and entropy. • I can explain that solubility depends on the overall free energy change, which results from a balance of enthalpy and entropy effects.	
9.7 Coupled Reactions	**Learning Target 7** I can explain how external energy sources can drive thermodynamically unfavorable processes.	• I can describe how electrical energy can be used to force nonspontaneous reactions in electrolysis or battery charging. • I can explain how light energy can drive unfavorable reactions such as photosynthesis. • I can describe how coupling a thermodynamically unfavorable reaction with a favorable one can make the overall process favorable. • I can explain that coupled reactions share intermediates and result in an overall process with negative ΔG°.	
9.8 Galvanic (Voltaic) and Electrolytic Cells	**Learning Target 8** I can explain how each component of an electrochemical cell contributes to the overall operation of the cell.	• I can describe how electrodes, solutions, and the salt bridge work together to allow electron and ion flow. • I can explain the difference between galvanic and electrolytic cells in terms of thermodynamic favorability and energy input. • I can identify that oxidation occurs at the anode and reduction occurs at the cathode in all electrochemical cells. • I can relate macroscopic observations (voltage, mass changes, gas formation) to particle-level electron transfer in half-reactions.	**Demo with Q&A** Construct a simple battery by submerging two electrodes (Mg and Cu) into orange juice and attaching it to the battery compartment of a quartz clock. Instruct students to ask as many questions as they can while the clock is running. Allow the clock to run for as long as possible and then examine the magnesium anode after a day to see if it corrodes away. As a class, examine the table of standard reduction potentials and discuss where the electrons are coming from and going to in order to power the clock.
9.9 Cell Potential and Free Energy	**Learning Target 9** I can explain whether an electrochemical cell is thermodynamically favored based on the sign of its standard cell potential.	• I can determine that a cell is thermodynamically favored when the standard cell potential (E°_{cell}) is positive. • I can explain that a negative E°_{cell} indicates a thermodynamically unfavored reaction that requires external energy. • I can identify oxidation and reduction half-reactions using standard reduction potentials to calculate E°_{cell}. • I can relate cell potential to Gibbs free energy using $\Delta G^\circ = -nFE^\circ$, where positive cell potential corresponds to negative ΔG°.	
9.10 Cell Potential Under Nonstandard Conditions	**Learning Target 10** I can explain how changes in concentration from standard conditions affect cell potential.	• I can describe that cell potential depends on how far a system is from equilibrium. • I can explain that cell potential is zero at equilibrium and decreases in magnitude as equilibrium is approached. • I can describe that deviations from standard conditions can increase or decrease cell potential depending on how they shift the system relative to equilibrium.	

		• I can determine the direction of spontaneous electron flow based on the system's tendency to reach equilibrium. • I can explain how nonstandard concentrations affect the cell potential of an electrochemical cell. • I can describe that cell potential decreases as a system approaches equilibrium. • I can use the Nernst equation conceptually to relate cell potential to reaction conditions. • I can explain that increasing Q lowers the cell potential, while decreasing Q increases it. • I can predict qualitative changes in cell potential without relying solely on calculations.	
9.11 Electrolysis and Faraday's Law	**Learning Target 11** I can calculate the amount of charge transferred in an electrochemical cell using changes in reactants and products.	• I can relate the number of electrons transferred in a redox reaction to the amount of substance produced or consumed. • I can use Faraday's laws to connect current, time, and charge in electrochemical processes. • I can apply the relationship I = q/t to determine current or charge in a system. • I can use stoichiometry to connect ionic charge, electron transfer, and mass changes at electrodes.	**Critique Reasoning** Have students mass a new penny with an analytical balance. They attach the penny to the negative electrode, which is attached to a 9-volt battery. A zinc strip is attached to the positive electrode. The penny is submerged for 10 minutes in a 1.0M NaOH solution with zinc dust and the zinc electrode. Students dry the penny and mass it again. Using Faraday's laws, have them calculate the current that must have been delivered to plate the zinc on to the penny. Then have student pairs share and peer review each other's reasoning.

Progress Check 9

- Multiple-choice: ~30 questions
- Free-response: 2 questions
    - Short
    - Long

Preparing for the AP Exam

On the AP Exam, students must be able to provide an appropriate explanation of the connection between entropy, enthalpy, and Gibbs free energy and the thermodynamic favorability of a chemical reaction. Students often struggle with questions that require them to reason about whether enthalpy, entropy, or both drive a reaction. They will state that both enthalpy and entropy drive the reaction by using the equation for Gibbs free energy instead of reasoning about which might be more of a driving factor than the other. Further, some students fail to connect macroscopic observational data to the concepts of entropy and enthalpy to support claims about which concept was driving the reaction. Teachers can ensure that students understand that the first step in making thermodynamic favorability predictions is to define and interrelate enthalpy, entropy, and Gibbs free energy in relation to driving chemical or physical processes. Students should be shown how the concepts of kinetics, equilibrium, and thermodynamics are connected in order to explain why thermodynamically favorable reactions might produce small concentrations of product and why unfavorable reactions can produce large concentrations of product.

AP Exam Description

The AP Chemistry Exam assesses student application of the science practices and understanding of the learning objectives outlined in the course framework. The exam is 3 hours and 15 minutes long and includes 60 multiple-choice questions (MCQ) and 7 free-response questions (FRQ). A scientific or graphing calculator is recommended for use on both sections of the exam. Students are provided with the periodic table and a formula sheet that lists specific and relevant formulas for use on the exam (see CED Appendixes). The details of the exam, including exam weighting and timing, can be found below:

Section	Question Type	Number of Questions	Exam Weighting	Timing
I	**Multiple-choice questions**	**60**	**50%**	**90 minutes**
II	**Free-response questions**		**50%**	**105 minutes**
	Long questions (10 points each)	3		
	Short questions (4 points each)	4		

Units of Instruction	Exam Weighting
Unit 1: Atomic Structure and Properties	**7–9%**
Unit 2: Compound Structure and Properties	**7–9%**
Unit 3: Properties of Substances and Mixtures	**18–22%**
Unit 4: Chemical Reactions	**7–9%**
Unit 5: Kinetics	**7–9%**
Unit 6: Thermochemistry	**7–9%**
Unit 7: Principles of Equilibrium	**7–9%**
Unit 8: Acids and Bases	**11–15%**
Unit 9: Thermodynamics and Electrochemistry	**7–9%**

Section I: Multiple-Choice (MCQ)

Science Practices 1, 2, 4, 5, and 6 are all assessed in the multiple-choice section with the following exam weighting (Science Practice 3 is not assessed in the multiple-choice section):

Science Practice	Exam Weighting
Practice 1: Models and Representations	**8–12%**
Practice 2: Question and Method	**8–12%**
Practice 4: Model Analysis	**23–30%**
Practice 5: Mathematical Routines	**35–42%**
Practice 6: Argumentation	**8–12%**

Section II: Free-Response (FRQ)

All six science practices are assessed in the free-response section with the following exam weighting:

Science Practice	Exam Weighting
Practice 1: Models and Representations	**2–4%**
Practice 2: Question and Method	**10–16%**
Practice 3: Representing Data and Phenomena	**8–16%**
Practice 4: Model Analysis	**5–9%**
Practice 5: Mathematical Routines	**43–53%**
Practice 6: Argumentation	**15–24%**