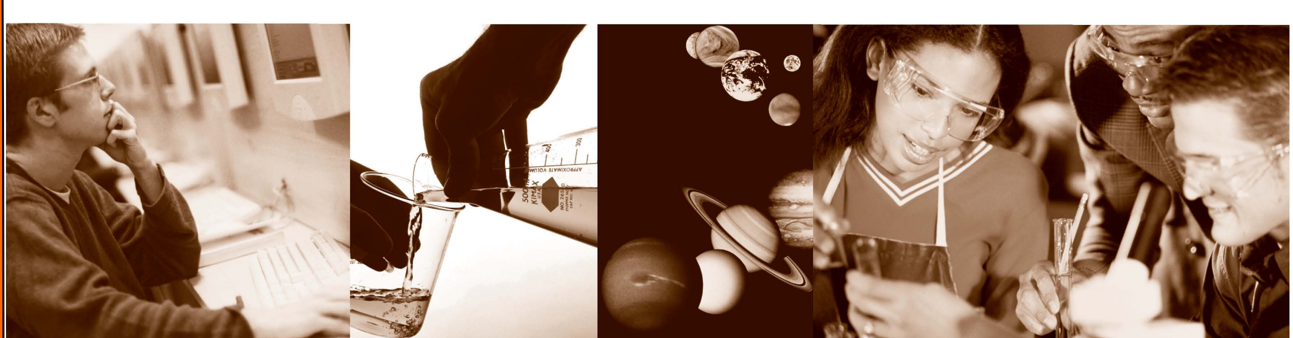
High School Content Expectations

Companion Document



SCIENCE

- Biology
- **Chemistry**
- Earth Science
- Physics

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MICHIGAN Department of Education

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HSSCE Companion Document

The Michigan High School Science Content Expectations establish what every student is expected to know and be able to do by the end of high school and outline the parameters for receiving high school credit as recently mandated by the Merit Curriculum legislation in the state of Michigan. The Science Content Expectations Documents and the Michigan Merit Curriculum Document have raised the bar for our students, teachers and educational systems.

In an effort to support these standards and help our science teachers develop rigorous and relevant curricula to assist students in mastery, the Michigan Science Leadership Academy, in collaboration with the Michigan Mathematics and Science Center Network and the Michigan Science Teachers Association, worked in partnership with Michigan Department of Education to develop this companion document. Our goal is for each student to master the science content expectations as outlined in the merit curriculum.

This companion document is an effort to clarify and support the High School Science Content Expectations and the Michigan Merit Curriculum. The Merit Curriculum has been organized into twelve teachable units – organized around the big ideas and conceptual themes in each of the four discipline areas. The document is similar in format to the Science Assessment and Item Specifications for the 2009 National Assessment for Educational Progress (NAEP). The companion document is intended to provide boundaries to the content expectations. These boundaries are presented as "notes to teachers", not comprehensive descriptions of the full range of science content; they do not stand alone, but rather, work in conjunction with the content expectations. The boundaries use five categories of parameters:

a. **Real World Context** refers to breadth and depth of topic coverage and includes those ideas that are "common" or "familiar" to students and appear frequently in curriculum materials and in most students' experiences outside of school. This section is not intended to guide assessment, but rather, may be used as a context for assessment.
b. **Instruments, measurements, and representations** refer to instruments students are expected to use and the level of precision expected to measure, classify, and interpret phenomena or measurement. This section contains assessable information.
c. **Technical vocabulary** refers to the vocabulary for use and application of the science topics and principles that appear in the content statements and expectations. The words in this section along with those presented within the standard, content statement and content expectation comprise the assessable vocabulary.
d. **Clarification** refers to the restatement of a "key idea" or specific intent or elaboration of the content statements. It is not intended to denote a sense of content priority. The clarifications guide assessment.
e. **Instructional Examples** are included as exemplars of five different modes of instruction appropriate to the unit in which they are listed. These examples include inquiry, reflection, general instruction, enrichment and intervention strategies. These examples are intended for instructional guidance only and are not assessable.

CHEMISTRY
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Chemistry Cross Reference Guide

HSCE Code	Expectation	Clarification Unit	Instructional Example
Standard C1	**INQUIRY, REFLECTION, AND SOCIAL IMPLICATIONS**		
Statement C1.1	**Scientific Inquiry**		
C1.1A	Generate new questions that can be investigated in the laboratory or field.		Lesson 2 i Lesson 10 i Lesson 12 i
C1.1B	Evaluate the uncertainties or validity of scientific conclusions using an understanding of sources of measurement error, the challenges of controlling variables, accuracy of data analysis, logic of argument, logic of experimental design, and/or the dependence on underlying assumptions.		Lesson 5 i Lesson 6 iii Lesson 10 iii
C1.1C	Conduct scientific investigations using appropriate tools and techniques (e.g., selecting an instrument that measures the desired quantity—length, volume, weight, time interval, temperature—with the appropriate level of precision).		Lesson 1 i Lesson 6 i Lesson 6 ii Lesson 6 iii Lesson 10 i Lesson 12 i
C1.1D	Identify patterns in data and relate them to theoretical models.		Lesson 2 i Lesson 3 i Lesson 7 i Lesson 9 i Lesson 11 i
C1.1E	Describe a reason for a given conclusion using evidence from an investigation.		Lesson 1i
C1.1f	Predict what would happen if the variables, methods, or timing of an investigation were changed.		Lesson 1i Lesson 5 i Lesson 6 i Lesson 8 i
C1.1g	Based on empirical evidence, explain and critique the reasoning used to draw a scientific conclusion or explanation.		Lesson 1 i Lesson 2 iii Lesson 2 v Lesson 4 i
C1.1h	Design and conduct a systematic scientific investigation that tests a hypothesis. Draw conclusions from data presented in charts or tables.		Lesson 2 iii Lesson 5 i Lesson 5 ii Lesson 6 i Lesson 10 iii
C1.1i	Distinguish between scientific explanations that are regarded as current scientific consensus and the emerging questions that active researchers investigate.		
Statement C1.2	**Scientific Reflection and Social Implications**		
C1.2A	Critique whether or not specific questions can be answered through scientific investigations.		
C1.2B	Identify and critique arguments about personal or societal issues based on scientific evidence.		

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HSCE Code	Expectation	Clarification Unit	Instructional Example
C1.2C	Develop an understanding of a scientific concept by accessing information from multiple sources. Evaluate the scientific accuracy and significance of the information.		Lesson 2 ii Lesson 2 iii Lesson 4 ii Lesson 6 ii Lesson 7 ii Lesson 8 ii Lesson 9 ii Lesson 10 ii
C1.2D	Evaluate scientific explanations in a peer review process or discussion format.		Lesson 2 v Lesson 6 v
C1.2E	Evaluate the future career and occupational prospects of science fields.		Lesson 12ii
C1.2f	Critique solutions to problems, given criteria and scientific constraints.		
C1.2g	Identify scientific tradeoffs in design decisions and choose among alternative solutions.		
C1.2h	Describe the distinctions between scientific theories, laws, hypotheses, and observations.		Lesson 2 ii
C1.2i	Explain the progression of ideas and explanations that lead to science theories that are part of the current scientific consensus or core knowledge.		Lesson 1 ii Lesson 2 i Lesson 3 ii
C1.2j	Apply science principles or scientific data to anticipate effects of technological design decisions.		Lesson 11ii
C1.2k	Analyze how science and society interact from a historical, political, economic, or social perspective.		
Standard C2	**FORMS OF ENERGY**		
Statement C2.1x	**Chemical Potential Energy**		
C2.1a	Explain the changes in potential energy (due to electrostatic interactions) as a chemical bond forms and use this to explain why bond breaking always requires energy.	Unit 4	Lesson 4 iv
C2.1b	Describe energy changes associated with chemical reactions in terms of bonds broken and formed (including intermolecular forces).	Unit 4	Lesson 4 iv
C2.1c	Compare qualitatively the energy changes associated with melting various types of solids in terms of the types of forces between the particles in the solid.	Unit 9	Lesson 9 iv
Statement C2.2	**Molecules in Motion**		
C2.2A	Describe conduction in terms of molecules bumping into each other to transfer energy. Explain why there is better conduction in solids and liquids than gases.	Unit 7	
C2.2B	Describe the various states of matter in terms of the motion and arrangement of the molecules (atoms) making up the substance.	Unit 7	

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HSCE Code	Expectation	Clarification Unit	Instructional Example
Statement C2.2x	**Molecular Entropy**		
C2.2c	Explain changes in pressure, volume, and temperature for gases using the kinetic molecular model.	Unit 7	
C2.2d	Explain convection and the difference in transfer of thermal energy for solids, liquids, and gases using evidence that molecules are in constant motion.	Unit 9	Lesson 9 v
C2.2e	Compare the entropy of solids, liquids, and gases.	Unit 12	Lesson 12 v
C2.2f	Compare the average kinetic energy of the molecules in a metal object and a wood object at room temperature.	Unit 7	
Statement C2.3x	**Breaking Chemical Bond**		
C2.3a	Explain how the rate of a given chemical reaction is dependent on the temperature and the activation energy.	Unit 12	
C2.3b	Draw and analyze a diagram to show the activation energy for an exothermic reaction that is very slow at room temperature.	Unit 12	
Statement C2.4x	**Electron Movement**		
C2.4a	Describe energy changes in flame tests of common elements in terms of the (characteristic) electron transitions.	Unit 3	Lesson 3 i
C2.4b	Contrast the mechanism of energy changes and the appearance of absorption and emission spectra.	Unit 3	Lesson 3 i
C2.4c	Explain why an atom can absorb only certain wavelengths of light.	Unit 3	Lesson 3 ii
C2.4d	Compare various wavelengths of light (visible and nonvisible) in terms of frequency and relative energy.	Unit 3	Lesson 3 ii Lesson 3 iv
Statement C2.5x	**Nuclear Stability**		
C2.5a	Determine the age of materials using the ratio of stable and unstable isotopes of a particular type.	Unit 1	
C2.r5b	Illustrate how elements can change in nuclear reactions using balanced equations. *(recommended)*	R	
C2.r5c	Describe the potential energy changes as two protons approach each other. *(recommended)*	R	
C2.r5d	Describe how and where all the elements on earth were formed. *(recommended)*	R	
Standard C3	**ENERGY TRANSFER AND CONSERVATION**		
Statement C3.1x	**Hess's Law**		
C3.1a	Calculate the ΔH for a given reaction using Hess's Law.	Unit 12	Lesson 12 iv
C3.1b	Draw enthalpy diagrams for exothermic and endothermic reactions.	Unit 12	
C3.1c	Calculate the ΔH for a chemical reaction using simple coffee cup calorimetry.	Unit 9	Lesson 9 iii
C3.1d	Calculate the amount of heat produced for a given mass of reactant from a balanced chemical equation.	Unit 9	Lesson iii

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HSCE Code	Expectation	Clarification Unit	Instructional Example
Statement C3.2x	**Enthalpy**		
C3.2a	Describe the energy changes in photosynthesis and in the combustion of sugar in terms of bond breaking and bond making.	Unit 12	
C3.2b	Describe the relative strength of single, double, and triple covalent bonds between nitrogen atoms.	Unit 4	
Statement C3.3	**Heating Impacts**		
C3.3A	Describe how heat is conducted in a solid.	Unit 7	
C3.3B	Describe melting on a molecular level.	Unit 7	
Statement C3.3x	**Bond Energy**		
C3.3c	Explain why it is necessary for a molecule to absorb energy in order to break a chemical bond.	Unit 4	Lesson 4 iv
Statement C3.4	**Endothermic and Exothermic Reactions**		
C3.4A	Use the terms endothermic and exothermic correctly to describe chemical reactions in the laboratory.	Unit 6	Lesson 6 iv
C3.4B	Explain why chemical reactions will either release or absorb energy.	Unit 12	
Statement C3.4x	**Enthalpy and Entropy**		
C3.4c	Write chemical equations including the heat term as a part of equation or using □*H* notation.	Unit 6	Lesson 6 iv
C3.4d	Draw enthalpy diagrams for reactants and products in endothermic and exothermic reactions.	Unit 12	
C3.4e	Predict if a chemical reaction is spontaneous given the enthalpy (ΔH) and entropy (ΔS) changes for the reaction using Gibb's Free Energy, $\Delta G = \Delta H - T\Delta S$ (Note: mathematical computation of ΔG is not required.)	Unit 12	Lesson 12 iii
C3.4f	Explain why some endothermic reactions are spontaneous at room temperature.	Unit 12	
C3.4g	Explain why gases are less soluble in warm water than cold water.	Unit 9	Lesson 9 ii
Statement C3.5x	**Mass Defect**		
C3.5a	Explain why matter is not conserved in nuclear reactions.	Unit 1	
Standard C4	**PROPERTIES OF MATTER**		
Statement C4.1x	**Molecular and Empirical Formulae**		
C4.1a	Calculate the percent by weight of each element in a compound based on the compound formula.	Unit 5	Lesson 5 i Lesson 5v
C4.1b	Calculate the empirical formula of a compound based on the percent by weight of each element in the compound.	Unit 5	Lesson 5 iii
C4.1c	Use the empirical formula and molecular weight of a compound to determine the molecular formula.	Unit 5	Lesson 5 iii
Statement C4.2	**Nomenclature**		
C4.2A	Name simple binary compounds using their formulae.	Unit 5	

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HSCE Code	Expectation	Clarification Unit	Instructional Example
C4.2B	Given the name, write the formula of simple binary compounds.	Unit 5	
Statement C4.2x	**Nomenclature**		
C4.2c	Given a formula, name the compound.	Unit 5	
C4.2d	Given the name, write the formula of ionic and molecular compounds.	Unit 5	Lesson 5 ii
C4.2e	Given the formula for a simple hydrocarbon, draw and name the isomers.	Unit 5	
Statement C4.3	**Properties of Substances**		
C4.3A	Recognize that substances that are solid at room temperature have stronger attractive forces than liquids at room temperature, which have stronger attractive forces than gases at room temperature.	Unit 7	
C4.3B	Recognize that solids have a more ordered, regular arrangement of their particles than liquids and that liquids are more ordered than gases.	Unit 7	Lesson 7 iii Lesson 7 iv
Statement C4.3x	**Solids**		
C4.3c	Compare the relative strengths of forces between molecules based on the melting point and boiling point of the substances.	Unit 8	
C4.3d	Compare the strength of the forces of attraction between molecules of different elements. (For example, at room temperature, chlorine is a gas and iodine is a solid.)	Unit 8	Lesson 8 i Lesson 8 iii Lesson 8 iv
C4.3e	Predict whether the forces of attraction in a solid are primarily metallic, covalent, network covalent, or ionic based upon the elements' location on the periodic table.	Unit 8	Lesson 8 v
C4.3f	Identify the elements necessary for hydrogen bonding (N, O, F).	Unit 8	Lesson 8 iii
C4.3g	Given the structural formula of a compound, indicate all the intermolecular forces present (dispersion, dipolar, hydrogen bonding).	Unit 8	Lesson 8 iv Lesson 8 v
C4.3h	Explain properties of various solids such as malleability, conductivity, and melting point in terms of the solid's structure and bonding.	Unit 8	Lesson 8 i Lesson 8 ii
C4.3i	Explain why ionic solids have higher melting points than covalent solids. (For example, NaF has a melting point of 995°C while water has a melting point of 0° C.)	Unit 8	Lesson 8 v
Statement C4.4x	**Molecular Polarity**		
C4.4a	Explain why at room temperature different compounds can exist in different phases.	Unit 4	Lesson 4 v
C4.4b	Identify if a molecule is polar or nonpolar given a structural formula for the compound.	Unit 4	Lesson 4 i Lesson 4 ii Lesson 4 iii
Statement C4.5x	**Ideal Gas Law**		
C4.5a	Provide macroscopic examples, atomic and molecular explanations, and mathematical representations (graphs and equations) for the pressure-volume relationship in gases.	Unit 7	Lesson 7 i Lesson 7 ii

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HSCE Code	Expectation	Clarification Unit	Instructional Example
C4.5b	Provide macroscopic examples, atomic and molecular explanations, and mathematical representations (graphs and equations) for the pressure-temperature relationship in gases.	Unit 7	Lesson 7 ii
C4.5c	Provide macroscopic examples, atomic and molecular explanations, and mathematical representations (graphs and equations) for the temperature-volume relationship in gases.	Unit 7	Lesson 7 ii Lesson 7 v
Statement C4.6x	**Moles**		
C4.6a	Calculate the number of moles of any compound or element given the mass of the substance.	Unit 5	Lesson 5 iv
C4.6b	Calculate the number of particles of any compound or element given the mass of the substance.	Unit 5	
Statement C4.7x	**Solutions**		
C4.7a	Investigate the difference in the boiling point or freezing point of pure water and a salt solution.	Unit 9	Lesson 9 i
C4.7b	Compare the density of pure water to that of a sugar solution.	Unit 1	
Statement C4.8	**Atomic Structure**		
C4.8A	Identify the location, relative mass, and charge for electrons, protons, and neutrons.	Unit 1	Lesson 1 iii
C4.8B	Describe the atom as mostly empty space with an extremely small, dense nucleus consisting of the protons and neutrons and an electron cloud surrounding the nucleus.	Unit 1	
C4.8C	Recognize that protons repel each other and that a strong force needs to be present to keep the nucleus intact.	Unit 1	
C4.8D	Give the number of electrons and protons present if the fluoride ion has a -1 charge.	Unit 1	Lesson 1 iv
Statement C4.8x	**Electron Configuration**		
C4.8e	Write the complete electron configuration of elements in the first four rows of the periodic table.	Unit 3	Lesson 3 v
C4.8f	Write kernel structures for main group elements.	Unit 3	Lesson 3 v
C4.8g	Predict oxidation states and bonding capacity for main group elements using their electron structure.	Unit 3	Lesson 3 v
C4.8h	Describe the shape and orientation of *s* and *p* orbitals.	Unit 3	Lesson 3 iii
C4.8i	Describe the fact that the electron location cannot be exactly determined at any given time.	Unit 3	Lesson 3 iv
Statement C4.9	**Periodic Table**		
C4.9A	Identify elements with similar chemical and physical properties using the periodic table.	Unit 2	Lesson 2 i
Statement C4.9x	**Electron Energy Levels**		
C4.9b	Identify metals, non-metals, and metalloids using the periodic table.	Unit 2	

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HSCE Code	Expectation	Clarification Unit	Instructional Example
C4.9c	Predict general trends in atomic radius, first ionization energy, and electonegativity of the elements using the periodic table.	Unit 2	Lesson 2 iii Lesson 2 v
Statement C4.10	**Neutral Atoms, Ions, and Isotopes**		
C4.10A	List the number of protons, neutrons, and electrons for any given ion or isotope.	Unit 1	Lesson 1 v
C4.10B	Recognize that an element always contains the same number of protons.	Unit 1	Lesson 1 v
Statement C4.10x	**Average Atomic Mass**		
C4.10c	Calculate the average atomic mass of an element given the percent abundance and mass of the individual isotopes.	Unit 2	Lesson 2 iv
C4.10d	Predict which isotope will have the greatest abundance given the possible isotopes for an element and the average atomic mass in the periodic table.	Unit 2	Lesson 2 iv
C4.10e	Write the symbol for an isotope, *X Z A* , where *Z* is the atomic number, *A* is the mass number, and *X* is the symbol for the element.	Unit 1	
Standard C5	**CHANGES IN MATTER**		
Statement C5.r1x	**Rates of Reactions *(recommended)***		
C5.r1a	Predict how the rate of a chemical reaction will be influenced by changes in concentration, temperature, and pressure. *(recommended)*	R	
C5.r1b	Explain how the rate of a reaction will depend on concentration, temperature, pressure, and nature of reactant. *(recommended)*	R	
Statement C5.2	**Chemical Changes**		
C5.2A	Balance simple chemical equations applying the conservation of matter.	Unit 6	
C5.2B	Distinguish between chemical and physical changes in terms of the properties of the reactants and products.	Unit 6	
C5.2C	Draw pictures to distinguish the relationships between atoms in physical and chemical changes.	Unit 1	
Statement C5.2x	**Balancing Equations**		
C5.2d	Calculate the mass of a particular compound formed from the masses of starting materials.	Unit 6	Lesson 6 ii
C5.2e	Identify the limiting reagent when given the masses of more than one reactant.	Unit 6	Lesson 6 i Lesson 6 v
C5.2f	Predict volumes of product gases using initial volumes of gases at the same temperature and pressure.	Unit 6	Lesson 6 iii
C5.2g	Calculate the number of atoms present in a given mass of element.	Unit 2	Lesson 2 ii
Statement C5.3x	**Equilibrium**		
C5.3a	Describe equilibrium shifts in a chemical system caused by changing conditions (Le Chatelier's Principle).	Unit 11	Lesson 11 iii
C5.3b	Predict shifts in a chemical system caused by changing conditions (Le Chatelier's Principle).	Unit 11	Lesson 11 iii

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HSCE Code	Expectation	Clarification Unit	Instructional Example
C5.3c	Predict the extent reactants are converted to products using the value of the equilibrium constant.	Unit 11	Lesson 11 ii
Statement C5.4	**Phase Change/Diagrams**		
C5.4A	Compare the energy required to raise the temperature of one gram of aluminum and one gram of water the same number of degrees.	Unit 9	
C5.4B	Measure, plot, and interpret the graph of the temperature versus time of an ice-water mixture, under slow heating, through melting and boiling.	Unit 9	
Statement C5.4x	**Changes of State**		
C5.4c	Explain why both the melting point and boiling points for water are significantly higher than other small molecules of comparable mass (e.g., ammonia and methane).	Unit 8	Lesson 8 v
C5.4d	Explain why freezing is an exothermic change of state.	Unit 8	
C5.4e	Compare the melting point of covalent compounds based on the strength of IMFs (intermolecular forces).	Unit 8	
Statement C5.5	**Chemical Bonds — Trends**		
C5.5A	Predict if the bonding between two atoms of different elements will be primarily ionic or covalent.	Unit 2	
C5.5B	Predict the formula for binary compounds of main group elements.	Unit 2	
Statement C5.5x	**Chemical Bonds**		
C5.5c	Draw Lewis structures for simple compounds.	Unit 2	
C5.5d	Compare the relative melting point, electrical and thermal conductivity, and hardness for ionic, metallic, and covalent compounds.	Unit 2	
C5.5e	Relate the melting point, hardness, and electrical and thermal conductivity of a substance to its structure.	Unit 9	Lesson 9 iv
Statement C5.6x	**Reduction/Oxidation Reactions**		
C5.6a	Balance half-reactions and describe them as oxidations or reductions.	Unit 11	Lesson 11 i Lesson 11 iv
C5.6b	Predict single replacement reactions.	Unit 6	
C5.6c	Explain oxidation occurring when two different metals are in contact.	Unit 11	Lesson 11 iv
C5.6d	Calculate the voltage for spontaneous redox reactions from the standard reduction potentials.	Unit 11	Lesson 11 iv Lesson 11 v
C5.6e	Identify the reactions occurring at the anode and cathode in an electrochemical cell.	Unit 11	Lesson 11 i
Statement C5.7	**Acids and Bases**		
C5.7A	Recognize formulas for common inorganic acids, carboxylic acids, and bases formed from families I and II.	Unit 10	Lesson 10 ii Lesson 10 iv
C5.7B	Predict products of an acid-based neutralization.	Unit 10	Lesson 10 iv
C5.7C	Describe tests that can be used to distinguish an acid from a base.	Unit 10	Lesson 10 i Lesson 10 iii
C5.7D	Classify various solutions as acidic or basic, given their pH.	Unit 10	Lesson 10 ii Lesson 10 v

Chemistry Cross Reference Guide

HSCE Code	Expectation	Clarification Unit	Instructional Example
C5.7E	Explain why lakes with limestone or calcium carbonate experience less adverse effects from acid rain than lakes with granite beds.	Unit 10	
Statement C5.7x	**Brønsted-Lowry**		
C5.7f	Write balanced chemical equations for reactions between acids and bases and perform calculations with balanced equations.	Unit 10	Lesson 10 iii Lesson 10 iv
C5.7g	Calculate the pH from the hydronium ion or hydroxide ion concentration.	Unit 10	Lesson 10 v
C5.7h	Explain why sulfur oxides and nitrogen oxides contribute to acid rain.	Unit 10	
C5.r7i	Identify the Brønsted-Lowry conjugate acid-base pairs in an equation. *(recommended)*	R	
Statement C5.8	**Carbon Chemistry**		
C5.8A	Draw structural formulas for up to ten carbon chains of simple hydrocarbons.	Unit 4	Lesson 4 iii
C5.8B	Draw isomers for simple hydrocarbons.	Unit 4	Lesson 4 iii
C5.8C	Recognize that proteins, starches, and other large biological molecules are polymers.	Unit 4	

Units by Content Expectation

CHEMISTRY

Unit 1: Atomic Theory

Code	Content Expectation
C2.5x	*Nuclear Stability* Nuclear stability is related to a decrease in potential energy when the nucleus forms from protons and neutrons. If the neutron/proton ratio is unstable, the element will undergo radioactive decay. The rate of decay is characteristic of each isotope; the time for half the parent nuclei to decay is called the half-life. Comparison of the parent/daughter nuclei can be used to determine the age of a sample. Heavier elements are formed from the fusion of lighter elements in the stars.
C2.5a	Determine the age of materials using the ratio of stable and unstable isotopes of a particular type.
C3.5x	*Mass Defect* Nuclear reactions involve energy changes many times the magnitude of chemical changes. In chemical reactions matter is conserved, but in nuclear reactions a small loss in mass (mass defect) will account for the tremendous release of energy. The energy released in nuclear reactions can be calculated from the mass defect using $E = mc^2$.
C3.5a	Explain why matter is not conserved in nuclear reactions.
C4.7x	*Solutions* The physical properties of a solution are determined by the concentration of solute.
C4.7b	Compare the density of pure water to that of a sugar solution.
C4.8	*Atomic Structure* Electrons, protons, and neutrons are parts of the atom and have measurable properties, including mass and, in the case of protons and electrons, charge. The nuclei of atoms are composed of protons and neutrons. A kind of force that is only evident at nuclear distances holds the particles of the nucleus together against the electrical repulsion between the protons.
C4.8A	Identify the location, relative mass, and charge for electrons, protons, and neutrons.
C4.8B	Describe the atom as mostly empty space with an extremely small, dense nucleus consisting of the protons and neutrons and an electron cloud surrounding the nucleus.
C4.8C	Recognize that protons repel each other and that a strong force needs to be present to keep the nucleus intact.
C4.8D	Give the number of electrons and protons present if the fluoride ion has a -1 charge.

C4.10	*Neutral Atoms, Ions, and Isotopes* A neutral atom of any element will contain the same number of protons and electrons. Ions are charged particles with an unequal number of protons and electrons. Isotopes are atoms of the same element with different numbers of neutrons and essentially the same chemical and physical properties.
C4.10A	List the number of protons, neutrons, and electrons for any given ion or isotope.
C4.10B	Recognize that an element always contains the same number of protons.
C4.10x	*Average Atomic Mass* The atomic mass listed on the periodic table is an average mass for all the different isotopes that exist, taking into account the percent and mass of each different isotope.
C4.10e	Write the symbol for an isotope, $^{A}X_{Z}$, where Z is the atomic number, A is the mass number, and X is the symbol for the element.
C5.2	*Chemical Changes* Chemical changes can occur when two substances, elements, or compounds interact and produce one or more different substances whose physical and chemical properties are different from the interacting substances. When substances undergo chemical change, the number of atoms in the reactants is the same as the number of atoms in the products. This can be shown through simple balancing of chemical equations. Mass is conserved when substances undergo chemical change. The total mass of the interacting substances (reactants) is the same as the total mass of the substances produced (products).
C5.2C	Draw pictures to distinguish the relationships between atoms in physical changes in terms of the properties of the reactants and products.

CHEMISTRY

Unit 1: Atomic Theory

Big Ideas (Core Concepts):
Order in the universe is exhibited through the location and function of subatomic particles and the likeness of atoms of individual elements

A strong force is needed to hold the nucleus together in all atoms.

Radioactive dating is the direct function of the timed decay of radioactive atoms.

Standard(s):
C2: Forms of Energy
C3: Energy Transfer and Conservation
C4: Properties of Matter
C5: Changes in Matter

Content Statement(s):
C2.5x: Nuclear Stability
C3.5x: Mass Defect
C4.7x: Solutions
C4.8: Atomic Structure
C4.10: Neutral Atoms, Ions, and Isotopes
C4.10x: Average Atomic Mass
C5.2: Chemical Change

Content Expectations: **(Content Statement Clarification)**
NOTE: C2.5a, C3.5a and C4.7b are considered to be engaging topics that set the stage for the unit topic of Atomic Theory.

C2.5a: Determine the age of materials using the ratio of stable and unstable isotopes of a particular type.

Clarification: Examples should be limited to the first 20 elements except for the long half life elements of uranium, iodine and cobalt.

C3.5a: Explain why matter is not conserved in nuclear reactions.

Clarification: Calculations are not necessary here except to illustrate $E=mc^2$.

C4.7b: Compare the density of pure water to that of a sugar solution.

Clarification: Compare properties that influence density. i.e. particle mass and packing of particles.

C4.8A: Identify the location, relative mass, and charge for electrons, protons, and neutrons.

Clarification: The relative mass of the proton is 1, the neutron is 1 and the electron is approximately zero. The relative charge of the electron is -1, the proton is +1 and the neutron is zero.

C4.8B: Describe the atom as mostly empty space with an extremely small, dense nucleus consisting of the protons and neutrons and an electron cloud surrounding the nucleus.

Clarification: It is not necessary to teach the electron orbital concept in detail a general discussion relating electron orbitals to a region of space (electron cloud) with higher probability regions that electrons are most likely to be found will suffice.

C4.8C: Recognize that protons repel each other and that a strong force needs to be present to keep the nucleus intact.

Clarification: Reinforce that the strong force is one of the four fundamental forces.

C4.8D: Give the number of electrons and protons present if the fluoride ion has a -1 charge.

Clarification: A modern periodic table must be made available.

C4.10A: List the number of protons, neutrons, and electrons for any given ion or isotope.

Clarification: Examples should be limited to the first 20 elements along with these other common elements: iron, gold, silver, mercury, iodine, chromium, and copper.

C4.10B: Recognize that an element always contains the same number of protons.

Clarification: None

C4.10e: Write the symbol for an isotope, $X^Z{}_A$, where Z is the atomic number, A is the mass number, and X is the symbol for the element.

Clarification: To teach this topic for conceptual understanding students should be given exercises with the location of the A and Z switched so students don't memorize the location as the key to the answer. Example: $X^Z{}_{A,}$ $X^A{}_Z$

C5.2C: Draw pictures to distinguish the relationships between atoms in physical and chemical changes.

Clarification: Use shapes of circles, triangles, squares, etc. to represent atoms for reactants and products to illustrate physical change and chemical change. Hands-on objects can be used also, example: nuts and bolts.

Vocabulary

Atomic mass
Atomic nucleus
Atomic number
Atomic theory
Atomic weight
Charged object
Decay rate
Electrically neutral
Electron
Electron cloud
Elementary particle
Ion
Isotope
Nuclear reaction
Neutron mass to energy conversion
Proton
Radioactive dating
Radioactive decay
Radioactive isotope
Relative mass
Stable
Strong force
Transforming matter and/or energy
Weight of subatomic particles

Real World Context:

Radioactive isotopes are used in the health fields to monitor internal bodily functions or to kill cancerous tissue.

Historical items may be placed in proper chronology using radioactive decay A process called radioactive dating compares quantities of an isotope present in the item with the same isotopes present in a contemporary item.

Half life of drugs in the body can be used in forensic science. Examples of half-life: caffeine, 4.9 hours; aspirin, 0.25 hours; nicotine, 2.0 hours; Bromide ion, 168 hours.

The large amount of energy available from nuclear reactions (fission in nuclear reactors, or fusion in stars) comes from the mass defect in atoms. Mass defect is the difference between the sums of the mass of individual particles in an atom (neglecting the electrons) compared to the actual mass of the same atom from the periodic table. The actual mass is always larger than the experimental mass whenever the nucleus contains more than one particle. The difference in mass (mass defect) is converted into energy that holds the nucleus together and can be released in nuclear reactions.

The chemical reactivity or stability of real world materials is based on the electron stability in atoms. Unstable or highly reactive elements are the result largely of outer electrons being lost or gained by neutral atoms. The noble gases for example are very stable and don't gain or lose electrons to other atoms under normal conditions and are used in light bulbs, deep sea diving, and between window panes.

Static electricity is the result of the outer electrons being pulled from or pulled to neutral atoms creating ions (the process that drives photocopying).

Ions are discussed in advertising about acid balance in living organisms, swimming pools, shampoos, etc.

Charged particles in a solution will allow current electricity to be conducted across or through the solution. Blood and other body fluids are able to transmit messages through electrical conductivity.

Common terminology in today's world is to refer to the relative comparison of facts (i.e.: a measure of one object relative to the same measure in another object)

Problems that are encountered in our daily lives are analyzed through the creation of models like scientists did with the atomic theory.

Observations of nuclear energy through observations of changes in systems containing radioactive substances, such as:

- Water used to cool down nuclear reactions in nuclear power plants: observable temperature increase in the water
- Radioactive isotopes of elements: emission of particles

Thermonuclear reactions: light emission

Instruments, Measurement, and Representations:
Use analogies to describe radioactive decay.

Models of atoms to represent the Bohr model

Historical reflection on Rutherford's Gold Foil experiment

Geiger counter

Instructional Examples:
i. Inquiry
CE: C1.1C, C1.1E, C1.1f, C1.1g

What is the location and shape of the object inside?

Use a hat pin to probe a clay ball with a penny embedded inside. Students should collect data each time they probe into the clay. They should record position, hit or no miss and depth if the object hits something solid. Explain the analogy of the clay ball to our model of the atom.

Extend the inquiry by asking another related question and experimenting to find the answer.

Position on Clay	Hit or Miss	Depth if a Hit (cm)
North Center		
West Center		
Top Center		
N mid center and edge		
W mid center and edge		
Top mid center and edge		

ii. Reflection
CE: C1.2i

Review the human perspective on the atom beginning with the early times before the Greek philosophers. Include the early Greeks, Dalton, J.J. Thompson, Rutherford, and Bohr.
http://www.lancs.ac.uk/ug/cooked1/atomictheory .ppt#258,3,Slide 3,

iii. Enrichment
CE: C4.8A

Find the relative mass of several common objects, (ex. Various seeds, bean, pencils, pen, 15 cm ruler, etc).

Find the actual mass of each object and arrange the objects in a table with the lowest to highest mass. Add a column to the table listing the relative mass of each object if the lightest object has a mass of 1.00. Arrange the objects again with the second smallest object having the relative mass of 1.00.

iv. General
CE: C4.8D

Construct a two dimensional or a three dimensional model to represent the number and location of the three subatomic particles in a fluoride ion with a -1 charge and represent the path (toward or away from the model) that the extra particle took to change the neutral fluorine atom to the fluoride ion.

v. Intervention
CE: C4.10A, C4.10B
Using cut out shapes that represent protons, neutrons and electrons including mass and charge. Students should demonstrate their understanding of obtaining the number of protons, electrons and neutrons from the atomic number and the atomic mass.

protons + neutrons = the atomic mass
protons = electrons
protons = atomic number

Students should demonstrate their understanding of the following elements through use of manipulatives.

Element	Atomic Number	Atomic Mass
Hydrogen	1	1
Helium	2	4
Lithium	3	7
Beryllium	4	10
Boron	5	11

Units by Content Expectation

CHEMISTRY

Unit 2: Periodic table

Code	Content Expectation
C4.9	*Periodic Table* In the periodic table, elements are arranged in order of increasing number of protons (called the atomic number). Vertical groups in the periodic table (families) have similar physical and chemical properties due to the same outer electron structures.
C4.9A	Identify elements with similar chemical and physical properties using the periodic table.
C4.9x	*Electron Energy Levels* The rows in the periodic table represent the main electron energy levels of the atom. Within each main energy level are sublevels that represent an orbital shape and orientation.
C4.9b	Identify metals, non-metals, and metalloids using the periodic table.
C4.9c	Predict general trends in atomic radius, first ionization energy, and electronegativity of the elements using the periodic table.
C4.10x	*Average Atomic Mass* The atomic mass listed on the periodic table is an average mass for all the different isotopes that exist, taking into account the percent and mass of each different isotope.
C4.10c	Calculate the average atomic mass of an element given the percent abundance and mass of the individual isotopes.
C4.10d	Predict which isotope will have the greatest abundance given the possible isotopes for an element and the average atomic mass in the periodic table.
C5.2x	*Balancing Equations* A balanced chemical equation will allow one to predict the amount of product formed.
C5.2g	Calculate the number of atoms present in a given mass of element.
C5.5	*Chemical Bonds-Trends* An atom's electron configuration, particularly of the outermost electrons, determines how the atom can interact with other atoms. The interactions between atoms that hold them together in molecules or between oppositely charged ions are called chemical bonds.
C5.5A	Predict if the bonding between two atoms of different elements will be primarily ionic or covalent.
C5.5B	Predict the formula for binary compounds of main group elements.

C5.5x	*Chemical Bonds* Chemical bonds can be classified as ionic, covalent, and metallic. The properties of a compound depend on the types of bonds holding the atoms together.
C5.5c	Draw Lewis structures for simple compounds.
C5.5d	Compare the relative melting point, electrical and thermal conductivity, and hardness for ionic, metallic, and covalent compounds.

CHEMISTRY

Unit 2: Periodic Table

Big Idea (Core Concepts):
The periodic table organizes the known elements into periods and families with similar properties.

The periodic table is organized to display trends in the characteristics of elements.

The type of chemical bonding determines some characteristic properties of materials.

Standard(s):
C4: Properties of Matter
C5: Changes in Matter

Content Statement(s):
C4.9: Periodic Table
C4.9x: Electron Energy Levels
C4.10x: Average Atomic Mass
C5.2x: Balancing Equations
C5.5: Chemical Bonds-Trends
C5.5x: Chemical Bonds

Content Expectations: **(Content Statement Clarification)**
C4.9A: Identify elements with similar chemical and physical properties using the periodic table.

Clarification: None

C4.9b: Identify metals, non-metals, and metalloids using the periodic table.

Clarification: The "stair step" on the right side of the periodic table conveniently separates the elements with physical properties of metals from the nonmetals. The metalloids are approximately on the "stair step".

C4.9c: Predict general trends in atomic radius, first ionization energy, and electronegativity of the elements using the periodic table.

Clarification: Given the names of two or three elements either from the same family or from the same period, arrange them from greatest to least with respect to atomic radius, first ionization energy and electronegativity. Limit examples to elements 1 – 20.

C4.10c: Calculate the average atomic mass of an element given the percent abundance and mass of the individual isotopes.

Clarification: Atomic mass numbers of isotopes will be given.

C4.10d: Predict which isotope will have the greatest abundance given the possible isotopes for an element and the average atomic mass in the periodic table.

Clarification: No calculations are required here. This expectation should just require conceptualizing the isotope in greatest amount. Example: If B has only isotopes of B^{11} and B^{10} but the atomic mass is listed as $B^{10.81}$; atoms of isotope 11 must be more abundant than isotope 10.

C5.2g: Calculate the number of atoms present in a given mass of element.

Clarification: Avogadro's number, (6.02×10^{23} atoms/gram atomic mass), is a constant and a conversion factor. Examples should include only monatomic elements.

C5.5A: Predict if the bonding between two atoms of different elements will be primarily ionic or covalent.

Clarification: Electronegativity tables will not be provided. Bonds can be differentiated by looking at physical properties of the compound and/or by looking at whether the atoms are metallic or nonmetallic on the periodic table. Ionic compounds consist of a metal and a nonmetal, they are brittle, will conduct electricity if melted or dissolved in water, and they have high melting points. Ionic bonds will be favored when atoms from groups 1 and 2 in the periodic table, bond with atoms from groups 16 and 17. Ionic bonding can also be expected if a compound consists of a metal atom and one of the common anions listed in the C4.2c clarification.

Covalent bonding can be predicted when two nonmetal atoms bond or when a metalloid atom bonds with a nonmetal atom. Physical properties can also be used to predict covalent bonding. If physical properties do not indicate ionic bonding then the bond should be assumed to be covalent.

C5.5B: Predict the formula for binary compounds of main group elements.

Clarification: The main group elements are found in columns 1, 2, and 13-18 on modern periodic tables. Column 18 does not react under normal conditions and will not be used here.

C5.5c: Draw Lewis structures for simple compounds.

Clarification: Lewis structures can only be drawn for covalent compounds. Examples should be limited to nonmetal binary compounds with single center atoms, for example: H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2, H_2O, H_2S, HCl, HBr, HI, SF_2, SCl_2, SBr_2, SI_2, NCl_3, NBr_3, NI_3, PCl_3, PBr_3, PI_3, CH_4.
Exclusion: Resonance structures and expanded octets

C5.5d: Compare the relative melting point, electrical and thermal conductivity, and hardness for ionic, metallic, and covalent compounds.

Clarification: Comparing properties should lead to understanding trends.
Examples: Ionic, NaCl; metallic, Na; covalent, paraffin

<u>Vocabulary:</u>
Actual mass
Atomic bonding principles
Avogardo's hypothesis
Binary compound
Chemical bond
Chemical properties of elements
Covalent bond
Earth's elements
Electrical conductivity
Electronegativity
Electron sharing
Electron transfer
Element family
Elements of matter
Energy sublevels
Periodic table of the elements
Ionic bond
Ionization energy
Lewis structures
Main energy level
Main group elements
Metalloids
Metallic bond
Orbital shape
Outer electron
Thermal conductivity

Real World Context:
Ionic bonds form very strong bonds. They form salts like table salt, $NaCl$. They are brittle, and while they dissolve easily in water they have high melting points, they are nonconductors as solids and don't readily corrode (react with gases in the air).
Among the many covalently bonded compounds are: plastics ceramics/glasses, waxes, and common room temperature liquids and gases.

Plastic and glass are used as electrical insulators for power lines.

Glass can be made with special properties by adding different kinds of atoms to the glass. Adding cobalt makes glass blue; manganese makes glass purple, etc.

Corning Glass Company in 1912 found that by adding boron oxide to glass it became shock resistant to temperature changes (Pyrex).

Photochromic glasses (transition lenses in eyeglasses) are made by adding silver ions to the glass. The darkening is the result of the silver ions (Ag^{+}) converting to metallic silver (Ag) by picking up an electron. This color is lost again in the dark.

Glass that is very stable (doesn't react with other materials) is being developed to store nuclear waste material.

In physiology, the primary ions or electrolytes are sodium, (Na^{+}), potassium (K^{+}), calcium (Ca^{2+}), magnesium (Mg^{2+}), chloride (Cl^{-}), phosphate (PO_4^{3-}), and hydrogen carbonate (HCO_3^{-}).
Muscle contraction is dependent upon the presence of calcium ion (Ca^{2+}), sodium (Na^{+}), and potassium (K^{+}). Without sufficient levels of these key electrolytes, muscle weakness or severe muscle contractions may occur.

Today's sport drinks are packed with electrolytes (ions), potassium (K^{+}), magnesium (Mg^{2+}), calcium (Ca^{2+}), and sodium (Na^{+})

Instruments, Measurement, and Representations:
Graph of trends in periodic properties for elements in periods and families

Models of atoms or cross sections of atoms to highlight characteristics

Percentage occurrence of isotopes is used to predict average atomic mass

Avogadro's number is needed to calculate the number of atoms present in a given mass.

Write the formula for binary compounds given the two elements and the periodic table.

Draw Lewis structures for binary compounds.

<u>Instructional Examples:</u>
i. Inquiry
CE: C1.1A, C1.1D, C1.2i, C4.9A

Can everyday food be arranged into families with similar properties?

Students design a periodic table of everyday objects with 20 components such as food.
Extension: Use game equipment, hardware store parts, clothing, etc., as alternate objects to be arranged.

ii. Reflection
CE: C1.2C, C1.2h, C5.2g

a) Review the division line between ionic compounds and covalent compounds using the difference in electronegativity. Investigate several sources to determine the dividing line separating ionic compounds from covalent compounds.
b) Using the concept of the difference in electronegativity and values that are available in part a, explain the distinction between observations, hypotheses, laws and theories.
c) Investigate the, "bond triangle", for various compounds with covalent bonding, ionic bonding, and metallic bonding at the corners.

iii. Enrichment
CE: C1.1g. C1.1h, C1.2C, C4.9c

a) Graph the atomic number vs. the atomic radius for atoms in the 2nd period or row in the periodic table. Find the atomic radius values from resources. Emphasize in drawings the characteristics that help determine the trend observed.
b) Repeat part a using the 1st ionization energy.
c) Repeat part a using the electronegativity.
d) graph atomic radius versus electronegativity and atomic radius versus ionization energy.

iv. General
CE: C4.10c, C4.10d

a) Using paper cut-outs (shown below) of isotopes of Boron, (B^{10} and B^{11}). Fill in the subatomic particle inventory for each atom - using 5 atoms of B^{10} and 5 atoms of B^{11}
b) Find the total mass in atomic mass units of all the 10 atoms (sum of the protons and neutrons for all 10 atoms)
c) Find the hypothetical atomic mass or the average mass of 1 atom.
d) Write the symbol for B and write the new average mass to 2 significant figures beside the symbol.

B^{10}	B^{11}
___ Protons	___ Protons
___ Electrons	___ Electrons
___ Neutrons	___ Neutrons

e) Make up a new hypothetical percentage of B^{10} and B^{11} and repeat part a. For example $B^{10.5}$ or $B^{10.2}$.
http://www.ionsource.com/Card/Mass/mass.htm
http://www.carlton.srsd119.ca/chemical/molemass/isotopes.htm

f) Illustrate the relative abundance of isotopes by using familiar objects like M&Ms - plain and peanut.

v. Intervention
CE: C1.1g, C1.2D, C4.9c

a) Build or draw 3-D cross section models of atoms from common household materials. The models should show the comparison of neighbor atoms in the same period (example: Li and Be) emphasizing trends in atomic radius, 1st ionization energy, and electronegativity. Let students be creative in how to show the variation. They could use increasing numbers of pieces of string to show higher I.E. or lines extending out from the atom showing attractive forces for electronegativity.
b) Models can also be made with neighbors in the same family (example: Li and Na).
c) Students should explain their models to peer groups. Peer review should follow the explanations.

Units by Content Expectation

CHEMISTRY

Unit 3: Quantum Mechanics

Code	Content Expectation
C2.4x	*Electron Movement* For each element, the arrangement of electrons surrounding the nucleus is unique. These electrons are found in different energy levels and can only move from a lower energy level (closer to nucleus) to a higher energy level (farther from nucleus) by absorbing energy in discrete packets. The energy content of the packets is directly proportional to the frequency of the radiation. These electron transitions will produce unique absorption spectra for each element. When the electron returns from an excited (high energy state) to a lower energy state, energy is emitted in only certain wavelengths of light, producing an emission spectra.
C2.4a	Describe energy changes in flame tests of common elements in terms of the (characteristic) electron transitions.
C2.4b	Contrast the mechanism of energy changes and the appearance of absorption and emission spectra.
C2.4c	Explain why an atom can absorb only certain wavelengths of light.
C2.4d	Compare various wavelengths of light (visible and nonvisible) in terms of frequency and relative energy.
C4.8x	*Electron Configuration* Electrons are arranged in main energy levels with sublevels that specify particular shapes and geometry. Orbitals represent a region of space in which an electron may be found with a high level of probability. Each defined orbital can hold two electrons, each with a specific spin orientation. The specific assignment of an electron to an orbital is determined by a set of 4 quantum numbers. Each element and, therefore, each position in the periodic table is defined by a unique set of quantum numbers.
C4.8e	Write the complete electron configuration of elements in the first four rows of the periodic table.
C4.8f	Write kernel structures for main group elements.
C4.8g	Predict oxidation states and bonding capacity for main group elements using their electron structure.
C4.8h	Describe the shape and orientation of *s* and *p* orbitals.
C4.8i	Describe the fact that the electron location cannot be exactly determined at any given time.

CHEMISTRY

Unit 3: Quantum Mechanics

Big Idea (Core Concepts)
The emission spectrum of individual elements is always identical and can be used to identify the elements.

Electron transition within energy levels can account for a specific energy emission or absorption within atoms.

Standard(s):
C2: Forms of Energy
C4: Properties of Matter

Content Statements:
C2.4x: Electron Movement
C4.8x: Electron Configuration

Content Expectations: (Content Statement Clarification)

C2.4a: Describe energy changes in flame tests of common elements in terms of the (characteristic) electron transitions.

Clarification: Limit the salts (nitrates or sulfates) to the following elements: potassium, calcium, sodium, lithium and copper for flame tests. No calculations are needed.

C2.4b: Contrast the mechanism of energy changes and the appearance of absorption and emission spectra.

Clarification: No calculations are necessary, conceptual understanding is sufficient.

C2.4c: Explain why an atom can absorb only certain wavelengths of light.

Clarification: None

C2.4d: Compare various wavelengths of light (visible and nonvisible) in terms of frequency and relative energy.

Clarification: None

C4.8e: Write the complete electron configuration of elements in the first four rows of the periodic table.

Clarification: Included in the first four rows are two exceptions to filling in order of increasing energy, the Aufbau principle, (Cr and Cu). Students should see the exceptions and understand the idea of stability over lowest energy.

C4.8f: Write kernel structures for main group elements.

Clarification: Introduce the kernel to simplify electron configurations. The kernel is a structure used to shorten an electron configuration. A kernel is an inert gas symbol in brackets that stands in place of all of the filled orbitals contained in the inert gas. It is also called the base unit or shortened version.
Example: [Ne] is a kernel, it represents an electron configuration of $1s^2 2s^2 2p^6$; Na= [Ne],$3s^1$). Limit to elements 1-20.

C4.8g: Predict oxidation states and bonding capacity for main group elements using their electron structure.

Clarification: Main group elements are those in columns 1 – 2 and 13-18. (Transition elements are not included in the main groups.)

C4.8h: Describe the shape and orientation of s and p orbitals.

Clarification: Emphasize the idea that orbitals are three dimensional not two and that the orbitals represent space with high probability of where electrons would be located.

C4.8i: Describe the fact that the electron location cannot be exactly determined at any given time.

Clarification: None

<u>Vocabulary</u>
Absorbance spectrum
Atomic motion
Bright line spectrum
Chemical bond
Electromagnetic field
Electromagnetic radiation
Electromagnetic spectra
Electromagnetic wave
Electron
Electron configuration
Emission spectra
Energy level
Excited state

Kernel
Ground state
Orbitals
Probability
Quantum energy
Quantum numbers
Release of energy
Sublevel
Valence electrons
Wave amplitude
Wavelength

Real World Context
Fireworks produce specific colors because of the compounds used and the energy released when they burn.

Lighting, both commercial (neon lights) and highway or backyard lighting (mercury vapor or sodium) are a result of excited state electrons.

A rainbow is an example of a continuous spectrum being broken down into its different wavelengths as a result of rain droplets in the air.

Scientists can learn what stars are made of by observing the spectrum they emit.

The use of UV blockers in suntan lotions

Gas discharge tubes are used in UPC scanners

Photoelectric panels on solar houses, cars, and calculators

Aurora borealis (northern lights) or aurora australis (southern lights)

Instruments, Measurement, and Representations
Formulas can be used to calculate energy changes and then related to specific wavelengths and type of radiation.

Electron configurations can be written for elements and ions, both with and without a kernel (noble gas base) in the first four periods. Given a configuration of a main group element, determine the oxidation state (i.e. ns^2np^3 – will have a -3 oxidation state).

Spectroscopes can be used to observe different light sources. Light sources might include the following: sunlight; lights in classroom; gas tubes containing hydrogen, neon, or other.

Models which represent s and p orbitals can be drawn.

Instructional Examples:
i. Inquiry
CE: C1.1D, C2.4a, C2.4b

Can you identify the composition of an unknown light source?
Using a hand held spectroscope, examine a variety of light sources. (Light sources might include the following: sunlight; lights in classroom; gas tubes containing hydrogen, neon, or other.)
Also observe the resulting spectrum of white light that is passed through a colored solution. Using colored pencils, draw what is observed in each case. Explain why they are not all the same. Classify them as line spectra, absorption spectra, or continuous spectra.

ii. Reflection
CE: C1.2i, C2.4d, C2.4c

Review the concepts of the atomic theory and how they have changed as new knowledge has become available. Including the information advanced in the field of quantum mechanics by Heisenberg and Schrödinger.

iii. Enrichment
CE: C4.8h

This is a mini-probability exercise. This exercise can be accomplished by having them drop small ball bearings onto a target which consists of ten concentric rings, each one centimeter wide. Balls should be dropped from a height of about six feet, at arm's length while aiming at the bullseye. By attaching a second target to the first and placing a piece of carbon paper between them, the hits will be recorded on the bottom target. Use 100 drops into the rings to make probability of a given area easier. After counting the number of hits in rings in each ring, the hit density (hits/ring area) can be calculated for each concentric ring. This will generally show that the likelihood of hitting a given ring decreases with the distance from the bullseye. This can then be related to the likelihood of where electrons would be found in the hydrogen atom and the probable shape of the s orbital. The electron charge density is greatest at the nucleus. (Graphing hit density vs. distance from center of target can help support the idea that the electrons will be close to the nucleus but not generally in it.) Caution should be used since this exercise will have a directional effect to it which electron probability does not. (This is only representative of an s orbital.)

iv. General
CE: C2.4d, C4.8i

After observing the hydrogen spectra, draw what has been observed. The spectra should have four lines showing up in difference colors. Next, draw two diagrams which represent a hydrogen atom. In one, have the electron in n=1. In the second, have the electron in n=5. Which of the drawings represents a ground state configuration? If the electron in the second diagram was to fall to n=2, would a continuous or line spectra be produced? What color light would be admitted, based on what you observed earlier?

Extension: Determinate the actual wavelength of the light that was produced. This is possible using the information provided below.
$\Delta E = E_{\text{higher orbit}} - E_{\text{lower orbit}} = E_{\text{photon}}$
$E_n = -2.178 \times 10^{-18}\ J / n^2$
$E_{\text{photon}} = hv$
$\lambda v = c$
h=Planck's constant ($6.626 \times 10^{-34} J \cdot s$)
v = frequency
λ = wavelength; c = speed of light (2.998×10^{8} m/s)

v. Intervention
CE: C4.8e, C4.8f, C4.8g

Complete the table below. Either a condensed or full electron configuration is acceptable.

Element	Elect. Conf.	Valence electrons	Oxidation state
Li			
	$1s^22s^22p^63s^23p^1$		
		$2s^22p^5$	
O			-2
		$3s^2$	
	$[Ar]4s^23d^{10}4p^3$		

Units by Content Expectation

CHEMISTRY

Unit 4: Introduction to Bonding

Code	Content Expectation
C2.1x	*Chemical Potential Energy* Potential energy is stored whenever work must be done to change the distance between two objects. The attraction between the two objects may be gravitational, electrostatic, magnetic, or strong force. Chemical potential energy is the result of electrostatic attractions between atoms.
C2.1a	Explain the changes in potential energy (due to electrostatic interactions) as a chemical bond forms and use this to explain why bond breaking always requires energy.
C2.1b	Describe energy changes associated with chemical reactions in terms of bonds broken and formed (including intermolecular forces).
C3.2x	*Enthalpy* Chemical reactions involve breaking bonds in reactants (endothermic) and forming new bonds in the products (exothermic). The enthalpy change for a chemical reaction will depend on the relative strengths of the bonds in the reactants and products.
C3.2b	Describe the relative strength of single, double, and triple covalent bonds between nitrogen atoms.
C3.3x	*Bond Energy* Chemical bonds possess potential (vibrational and rotational) energy.
C3.3c	Explain why it is necessary for a molecule to absorb energy in order to break a chemical bond.
C4.4x	*Molecular Polarity* The forces between molecules depend on the net polarity of the molecule as determined by shape of the molecule and the polarity of the bonds.
C4.4a	Explain why at room temperature different compounds can exist in different phases.
C4.4b	Identify if a molecule is polar or nonpolar given a structural formula for the compound.
C5.8	*Carbon Chemistry* The chemistry of carbon is important. Carbon atoms can bond to one another in chains, rings, and branching networks to form a variety of structures, including synthetic polymers, oils, and the large molecules essential to life.
C5.8A	Draw structural formulas for up to ten carbon chains of simple hydrocarbons.
C5.8B	Draw isomers for simple hydrocarbons.
C5.8C	Recognize that proteins, starches, and other large biological molecules are polymers.

CHEMISTRY

Unit 4: Introduction to Bonding

Big Idea (Core Concepts)
Chemical bonds form either by the attraction of a positive nucleus and negative electrons or the attraction between a positive ion and a negative ion.

The strength of chemical bonds can be measured by the changes in energy that occur during a chemical reaction.

Standard(s):
C2: Forms of Energy
C3: Energy Transfer and Conservation
C4: Properties of Matter
C5: Changes in Matter

Content Statement(s):
C2.1x: Chemical Potential Energy
C3.2x: Enthalpy
C4.4x: Molecular Polarity
C5.8: Carbon Chemistry

Content Expectations: (Content Statement Clarification)
C2.1a: Explain the changes in potential energy (due to electrostatic interactions) as a chemical bond forms and use this to explain why bond breaking always requires energy.

Clarification: None

C2.1b: Describe energy changes associated with chemical reactions in terms of bonds broken and formed (including intermolecular forces).

Clarification: None

C3.2b: Describe the relative strength of single, double, and triple covalent bonds between nitrogen atoms.

Clarification: The three bond examples in increasing order of strength are: single < double < triple.

C3.3c: Explain why it is necessary for a molecule to absorb energy in order to break a chemical bond.

Clarification: None

C4.4a: Explain why at room temperature different compounds can exist in different phases.

Clarification: None

C4.4b: Identify if a molecule is polar or nonpolar given a structural formula for the compound.

Clarification: The polarity of a molecule is based on two ideas. One is the bonding itself, whether it is polar or nonpolar. The second part is the geometry or shape of the molecule and whether or not the polar bonds cancel out. Symmetric molecules are always nonpolar. Polar molecules will align themselves a set way within an electric field because they have a greater electron density on one side then another. CH_2Cl_2 is polar molecule whereas CCl_4 is nonpolar molecule. They both have the same geometry but one is symmetrical and the polar bonds cancel out.

C5.8A: Draw structural formulas for up to ten carbon chains of simple hydrocarbons.

Clarification: Simple hydrocarbons should include alkanes, alkenes, and alkynes to take into account the versatility of carbon and the fact that multiple bonds were introduced in C3.2b.

C5.8B: Draw isomers for simple hydrocarbons.

Clarification: Isomers should be limited to structure only at this point (no geometric isomers). Most likely limit examples and work to six carbon compounds for alkanes and either four or five for alkenes and alkynes.

C5.8C: Recognize that proteins, starches, and other large biological molecules are polymers.

Clarification: Limit other large biological molecules to nucleic acids and cellulose.

<u>Vocabulary</u>
Bond energy
Carbon atom
Charged object
Chemical bond
Crystalline solid
Double bond
Electric force
Electron
Electron sharing
Electron transfer
Endothermic process

Enthalpy
Exothermic process
Hydrocarbon
Intermolecular force
Ion
Isomers
Monomer
Moving electric charge
Polarity
Potential energy
Protein
Release of energy
Single bond
Synthetic polymer

Real World Context

In addition to NaCl, many minerals exist as ionic solids, such as pyrite (FeS_2), cinnabar (HgS), hematite (Fe_2O_3), fluorite (CaF_2), beryl ($Be_3Al_2Si_6O_{18}$), and barite ($BaSO_4$).

N_2 is an extremely stable and thus nonreactive substance. Fertilizers generally contain nitrogen in the form of ammonia or ammonium compounds because most plants cannot use the nitrogen out of the air (it exists as a stable N_2 molecule). The legumes and a few other plants are considered very important because they "fix" the atmospheric nitrogen into a usable form.

N_2 is used in the food industry. For example, many manufacturers use nitrogen to fill the space in the potato chip bags and to reduce or prevent oxidation from occurring before the bag is opened.

Water drops that form on plant blossoms from the early morning's dew is based on strong attractive forces between the highly polar water molecules.

Water striders are able to stay on top of the water, rather than sink, because of the water tension or attractive forces of the molecules for one another.

Salts dissolved in the oceans and most of the substances that comprise the earth's crust are held together by ionic bonds. Most seashells are made of the ionic compound calcium carbonate, but they are insoluble in sea water.

Instruments, Measurement, and Representations
The periodic table can be used to make predictions about the type of bonds that will be formed.

Structural formulas of simple hydrocarbon can be drawn along with any isomers that exist.

Changes in energies that result from bonds breaking and being made can be calculated using bond energy charts.

Lewis structures can be used to show how single, double, or triple bonds are produced.

Diagrams of phase changes

Instructional Examples:
i. Inquiry
CE: C1.1g, C4.4b

Given the following scenario, review your knowledge about bonding and try to answer the question.
When a nylon comb is run through your hair, electrons are transferred to the comb. (The comb becomes negatively charged because of an excess of electrons. This situation is very evident in wintertime when your home environment tends to have low relative humidity.) When the comb is then brought near to a small stream of running water from a faucet, the stream will bend and move toward the comb. Explain what is happening and why based on your knowledge of molecules and bonding.

ii. Reflection
CE: C1.2C, C4.4b

Access the internet and learn how a microwave oven works. Predict what would happen in the following cases and justify your answers.
a) Would the microwaves have the same effect on a piece of ice as it has on liquid water?
b) If a sample of liquid carbon dioxide is placed in a microwave oven, would it heat the sample like it would a sample of liquid water?

iii. Enrichment
CE: C4.4b, C5.8A, C5.8B

Build models and draw structural representations for the following substances: HCN, O_2, CO_2, $CHCl_3$, PH_3, and H_2S. (If isomers also exist, construct them as well.) After the model is built, then pretend to place it in an electric field and decide if the molecule will be polar or nonpolar. Identify the bonds in the molecule as being polar or nonpolar covalent.

iv. General
CE: C2.1a, C2.1b, C3.3c

Are Chemical reactions endothermic or exothermic?
Using average bond energy charts, calculate the change in energy for the following process and identify if the process will be exothermic or endothermic.

a) $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{2(g)}$
b) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

v. Intervention
CE: C4.4a

Calcium chloride is an ionic substance and is often a produced when you ingest an antacid tablet. Elemental chlorine is used in treatment of water. Based on your knowledge of bonding, answer the following questions:

a) What is the most probable state of matter for the calcium chloride at room temperature?
b) Which of the two mentioned materials would have the highest melting point?
c) Which would have the lowest boiling point?
d) Analyze the bonding pattern for a single chlorine atom in each of the two substances and use this information to explain the differences noted above in their properties.

Units by Content Expectation

CHEMISTRY

Unit 5: Nomenclature and Formula Stoichiometry

Code	Content Expectation
C4.1x	*Molecular and Empirical Formulae* Compounds have a fixed percent elemental composition. For a compound, the empirical formula can be calculated from the percent composition or the mass of each element. To determine the molecular formula from the empirical formula, the molar mass of the substance must also be known.
C4.1a	Calculate the percent by weight of each element in a compound based on the compound formula.
C4.1b	Calculate the empirical formula of a compound based on the percent by weight of each element in the compound.
C4.1c	Use the empirical formula and molecular weight of a compound to determine the molecular formula.
C4.2	*Nomenclature* All compounds have unique names that are determined systematically.
C4.2A	Name simple binary compounds using their formulae.
C4.2B	Given the name, write the formula of simple binary compounds.
C4.2x	*Nomenclature* All molecular and ionic compounds have unique names that are determined systematically.
C4.2c	Given a formula, name the compound.
C4.2d	Given the name, write the formula of ionic and molecular compounds.
C4.2e	Given the formula for a simple hydrocarbon, draw and name the isomers.
C4.6x	*Moles* The mole is the standard unit for counting atomic and molecular particles in terms of common mass units.
C4.6a	Calculate the number of moles of any compound or element given the mass of the substance.
C4.6b	Calculate the number of particles of any compound or element given the mass of the substance.

CHEMISTRY

Unit 5: Nomenclature and Formula Stoichiometry

Big Idea (Core Concepts):
Chemical compounds always have the same formula and the same composition.

The formal charge on ions determines the ratio of the ions in an ionic compound, just as the apparent charge on atoms determines the ratio of the atoms in a covalent compound.

Standard(s):
C4: Properties of Matter

Content Statement(s):
C4.1x: Molecular and Empirical Formulae
C4.2: Nomenclature
C4.2x: Nomenclature
C4.6x: Moles

Content Expectations: (Content Statement Clarification)
C4.1a: Calculate the percent by weight of each element in a compound based on the compound formula.

Clarification: Compounds should include hydrates and compounds containing two or three different elements. A modern periodic table must be made available.

C4.1b: Calculate the empirical formula of a compound based on the percent by weight of each element in the compound.

Clarification: Compounds should include hydrates and compounds containing two or three different elements. A modern periodic table must be made available.

C4.1c: Use the empirical formula and molecular weight of a compound to determine the molecular formula.

Clarification: Compounds should include hydrates and compounds containing two or three different elements. A modern periodic table must be made available.

C4.2A: Name simple binary compounds using their formulae.

Clarification: Use the first 20 elements from the periodic table plus copper, iron, lead and mercury. Limit problems to metals combined with nonmetals. (Molecular compounds with two nonmetals are found in C4.2c and C4.2d.)

C4.2B: Given the name, write the formula of simple binary compounds.

Clarification: Same as C4.2A

C4.2c: Given a formula, name the compound.

Clarification: Use the first 20 elements from the periodic table plus copper, iron, lead and mercury. Problems should include molecular compounds (two nonmetals) three element compounds with common ions. Common ions should be limited to: acetate, hydroxide, sulfate, sulfite, nitrate, nitrite, carbonate and ammonium.

C4.2d: Given the name, write the formula of ionic and molecular compounds.

Clarification: Same as C4.2c.

C4.2e: Given the formula for a simple hydrocarbon, draw and name the isomers.

Clarification: Limit hydrocarbons to 6 carbon compounds with all single bonds. Isomer names should be limited to IUPAC naming rules.

C4.6a: Calculate the number of moles of any compound or element given the mass of the substance.

Clarification: Notice these calculations should include compounds as well as elements. A modern periodic table must be made available.

C4.6b: Calculate the number of particles of any compound or element given the mass of the substance.

Clarification: A modern periodic table must be made available.

Vocabulary

Binary
Carbon atom
Carbon dioxide
Empirical formula
Fossil fuel
Hydrocarbons
Isomers
Mole
Molecular formula
Organic matter

Real World Context:
Pharmacists make some special solutions and ointments using percent composition.

Examples of formulas in the work world are usually expressed in proportions of various compounds mixed together. An example is concrete which changes strength when the volume ratio of cement: sand : gravel is changed (1:2:4 is stronger than 1:3:6, a 1:1:2 mixture is used when concrete is used under water.) Another example is steel which changes properties when the formula of percent carbon is changed (carbon steel, 1% carbon; cast iron, 4% carbon; cementite, 6.7% carbon).

Minerals are everyday examples of empirical formulas (galena, PbS; magnetite, Fe_3O_4; pyrite, FeS_2; quartz, SiO_2; cinnabar, HgS).

Formula nomenclature is helpful when chemical formulas and/or chemical compounds are mentioned in news reports or in medical information.

Mole calculations are examples of packaging objects in larger units for ease of description or understanding. Terms are used to convey numbers in astronomy, the light year; or in the computer world, bytes, etc.

Instruments, Measurement, and Representations
Use a laboratory balance to find masses and then calculate percent composition.

Make calculations of percent composition.

Illustrate in a drawing the percent by weight of each element in a compound.

Use atomic or molecular weight in grams = 1 mole = 6.02×10^{23} atoms

Alkanes are compounds of hydrogen and carbon that have all single bonds.

Molecular formulas are whole number ratios of empirical formulas.

Instructional Examples:
i. Inquiry
CE: C1.1B, C1.1f, C1.1h, C4.1a

a) Design a laboratory investigation to determine the percent composition of a hydrate.
b) This lab can be extended by comparing class data for the same hydrate and/or comparing data for different hydrates.

ii. Reflection
CE: C1.1h, C4.2d

a) Using a list of names and formulas of compounds containing suffixes of –ate and –ite, suggest an hypothesis for the naming system.
b) Research to find out the systematic reason for the -ate and -ite suffixes on polyatomic ions with the same root name, such as sulfate and sulfite.

iii. Enrichment
CE: C4.1b, C4.1c

Complete the following table:

Empirical Formula	Molecular Weight	Molecular Formula
C_2H_5		C_4H_{10}
	1670	$H_{180}O_{90}$
HO	34	

iv. General
CE: C4.6a

Calculate the number of moles of a material in large quantities. Examples: If 2 billion people eat one egg each day how many dozen eggs are eaten in 100 days? How many moles of eggs are eaten in 100 days? If there are 20 drops in one milliliter, how many moles of drops are in a 100,000 liter swimming pool?

v. Intervention
CE: C4.1a

a) Make a compound out of common items such as marshmallows and jelly beans. Take the compound apart to find the mass of each component and determine the percent by weight of each component.
b) Trade your compound with someone else in class and find the percent composition by weight of the new compound.

Units by Content Expectation

CHEMISTRY

Unit 6: Equations and Stoichiometry

Code	Content Expectation
C3.4	*Endothermic and Exothermic Reactions* Chemical interactions either release energy to the environment (exothermic) or absorb energy from the environment (endothermic).
C3.4A	Use the terms endothermic and exothermic correctly to describe chemical reactions in the laboratory.
C3.4x	*Enthalpy and Entropy* All chemical reactions involve rearrangement of the atoms. In an exothermic reaction, the products have less energy than the reactants. There are two natural driving forces: (1) toward minimum energy (enthalpy) and (2) toward maximum disorder (entropy).
C3.4c	Write chemical equations including the heat term as a part of equation or using ΔH notation.
C5.2	*Chemical Changes* Chemical changes can occur when two substances, elements, or compounds interact and produce one or more different substances whose physical and chemical properties are different from the interacting substances. When substances undergo chemical change, the number of atoms in the reactants is the same as the number of atoms in the products. This can be shown through simple balancing of chemical equations. Mass is conserved when substances undergo chemical change. The total mass of the interacting substances (reactants) is the same as the total mass of the substances produced (products).
C5.2A	Balance simple chemical equations applying the conservation of matter.
C5.2B	Distinguish between chemical and physical changes in terms of the properties of the reactants and products.
C5.2x	*Balancing Equations* A balanced chemical equation will allow one to predict the amount of product formed.
C5.2d	Calculate the mass of a particular compound formed from the masses of starting materials.
C5.2e	Identify the limiting reagent when given the masses of more than one reactant.
C5.2f	Predict volumes of product gases using initial volumes of gases at the same temperature and pressure.
C5.6x	*Reduction/Oxidation Reactions* Chemical reactions are classified according to the fundamental molecular or submolecular changes that occur. Reactions that involve electron transfer are known as oxidation/reduction (or “redox”).
C5.6b	Predict single replacement reactions.

CHEMISTRY

Unit 6: Equations and Stoichiometry

<u>Big Idea</u> (Core Concepts):
Balanced chemical equations always exhibit conservation of mass and conservation of heat.

The same number of all gaseous molecules will occupy the same volume under the same conditions.

Chemical reactions carried out in the same fashion will always produce the same products.

Breaking of chemical bonds consumes energy while formation of bonds releases energy.

<u>Standard(s)</u>:
C3: Energy Transfer and Conservation
C5: Changes in Matter

<u>Content Statement(s)</u>:
C3.4: Endothermic and Exothermic Reactions
C3.4x: Enthalpy and Entropy
C5.2: Chemical Changes
C5.2x: Balancing Equations
C5.6x: Reduction/Oxidation Reactions

<u>Content Expectations:</u> (Content Statement Clarification)
C3.4A: Use the terms endothermic and exothermic correctly to describe chemical reactions in the laboratory.

Clarification: Examples: exothermic, steel wool plus vinegar; endothermic, vinegar plus sodium bicarbonate.

C3.4c: Write chemical equations including the heat term as a part of equation or using ΔH notation.

Clarification: Do not calculate ΔH from heat of formation tables.

C5.2A: Balance simple chemical equations applying the conservation of matter.

Clarification: Photosynthesis and respiration reactions are good examples. Other reaction examples are oxidation-reduction and acid base reactions.

C5.2B: Distinguish between chemical and physical changes in terms of the properties of the reactants and products.

Clarification: None

C5.2d: Calculate the mass of a particular compound formed from the masses of starting materials.

Clarification: Expected product mass can be calculated given a balanced chemical equation, the formula masses of reactants, and starting masses of reactants.

C5.2e: Identify the limiting reagent when given the masses of more than one reactant.

Clarification: Limiting reagents can be predicted given a balanced chemical equation, the formula masses of reactants, and starting masses of reactants.
Simple calculations should suffice here and can be determined using mental math (simple multiplication of atomic weights).
Example: 200 grams of Fe are reacted with 100 grams of O_2.

$$4Fe\ (s) \quad + \quad 3O_2\ (g) \quad \rightarrow \quad 2Fe_2O_3(s)$$

Notice in this example the masses are close to the stoichiometric value for iron and oxygen, but iron is slightly less and oxygen is slightly greater. If oxygen is greater then the iron would need to be greater also, therefore the iron is limited. Further examples should follow this pattern. Keep the mass of at least one of the reagents nearly a multiple of the stoichiometric value.

C5.2f: Predict volumes of product gases using initial volumes of gases at the same temperature and pressure.

Clarification: The product gases should be at the same temperature and pressure as the reactant gases. Simple whole numbers for coefficients and volumes given with two significant figures will work here. Calculations should require mental math only. Limiting reagents can be added to these problems.
Example: 50 liters of hydrogen and 25 liters of oxygen as starting gases or,
50 liters of hydrogen and 10 liters of oxygen as starting gases

$$2H_2\ (g) \quad + \quad O_2\ (g) \quad \rightarrow \quad 2H_2O\ (g)$$

C5.6b: Predict single replacement reactions.

Clarification: Students should learn to predict hypothetical products from single replacement reactions and then predict if the reaction will actually form products indicated using the appropriate activity series table.

$$3CuCl_2\ (aq) \quad + \quad 2Al\ (s) \quad \rightarrow \quad 3Cu\ (s) \quad + \quad 2AlCl_3\ (aq)$$

Vocabulary
Delta (meaning change)
Endothermic reaction
Exothermic reaction
Limiting reagent
Molar Volume
Oxidation-Reduction reactions
Pressure
Product
Properties of reactants
Reactant
Reagent
Release of energy

Real World Context
Chemical reactions in everyday situations produce or absorb heat. Examples: Hot and Cold Packs, Toilet bowl cleaners, Burning (paper, fuel, food), Photosynthesis, Respiration, etc.

The mass of materials produced or consumed in chemical reactions can be used to help understand natural phenomenon such as: global warming (CO_2 produced or consumed), oxygen needs of organisms, etc.

The amount of reagents that are added to a chemical reaction will determine the amount of product produced. Examples are: smoky fires (not getting enough oxygen), Mentos and diet coke reaction too small (add more Mentos), fizzing too little in vinegar in baking soda (add more vinegar or baking soda).

Proper volumes of gases are required in certain situations such as air bag deployment, carbon dioxide produced by ingredients in rising bread, production of ammonia gas for industry, etc.

Instruments, Measurement, and Representations
Make models to demonstrate balanced chemical reactions.

Use thermometers to determine how much heat is involved in a chemical reaction.

Write stories involving exothermic and endothermic reactions in the world.

Use models to demonstrate balanced chemical equations including adding the heat term to the reactant or product side of the equation.

Complete tables to show conservation of mass in chemical equations.

Use laboratory balances to determine reactant and/or product masses.
From simple chemical reactions involving gases, draw pictures of the volumes of reactants and products that demonstrate the understanding of molar volume. Any standard symbols can be used for a molar volume.
Example: $3H_2\,(g) \;+\; N_2\,(g) \;\rightarrow\; 2NH_3\,(g)$

Instructional Examples:
i. Inquiry
CE: C1.1C, C1.1f, C1.1h, C5.2e

Can the mass of reactants needed in a chemical reaction be determined by laboratory investigation?

By adding incremental additions of one reagent to a fixed amount of a second reagent determine the point where the mass of both reagents is adequate (without a limiting reagent).
Sample reactions might be aluminum with copper II chloride, or magnesium with hydrochloric acid.

ii. Reflection
CE: C1.1C, C1.2c. C5.2d

Compare the mass of products formed through investigation with the amount of product determined by the stoichiometry of a reaction.

a) (CAUTION: Sodium hydroxide is a caustic substance. Observe proper precautions in handling and cleanup of any spills. Be sure none gets on the skin.)
   React a given mass of sodium hydroxide in solution with a given mass of hydrochloric acid in solution and then recover the salt formed through evaporation.
b) Calculate the mass of salt that would be formed through the balanced chemical equation and mass/mass stoichiometry.
c) Determine your percentage yield for the laboratory exercise. Compare this with data from the entire class.

iii. Enrichment
CE: C1.1B, C.1C, C5.2f

a) Investigate a chemical reaction in a baggy. After determining the volume of the bag, decide how much of two reactants must be added to produce a full bag of gas.
$NaHCO_3$ (s) + $HC_2H_3O_2$ (aq) $\rightarrow$ CO_2 (g) + $NaC_2H_3O_2$ (aq) + H_2O (l)

b) Determine the amount of reactants needed to produce enough gas to fill a blimp with hydrogen gas using the following reaction:
Zn (s) + 2HCl (aq) $\rightarrow$ H_2 (g) + $ZnCl_2$ (aq)

Note: Assume the volume of one mole of hydrogen will be 22.4 dm^3
Estimate the volume of the blimp and/or research acceptable volumes.

iv. General
CE: C3.4A, C3.4c

Make a list of situations in a day that are encountered that could be improved if either heat was added or heat was absorbed. Suggest some ways that this might be done chemically.
Make a table to label the ideas as either exothermic or endothermic and include whether ΔH would be positive or negative.

v. Intervention
CE: C1.2D, C5.2e

Given routine activities write scenarios that will describe a limiting reagent in action.

a) Making Smores around a camp fire (grahams, marshmallows, chocolate, roasting sticks.)
b) Building bicycles from parts in a factory (frames, tires, chains).
c) Ordering at a fast food drive-thru window.
d) Have students peer review the scenarios with a partner.

Units by Content Expectation

CHEMISTRY

Unit 7: States of Matter

Code	Content Expectation
C2.2	*Molecules in Motion* Molecules that compose matter are in constant motion (translational, rotational, vibrational). Energy may be transferred from one object to another during collisions between molecules.
C2.2A	Describe conduction in terms of molecules bumping into each other to transfer energy. Explain why there is better conduction in solids and liquids than gases.
C2.2B	Describe the various states of matter in terms of the motion and arrangement of the molecules (atoms) making up the substance.
C2.2x	*Molecular Entropy* As temperature increases, the average kinetic energy and the entropy of the molecules in a sample increases.
C2.2c	Explain changes in pressure, volume, and temperature for gases using the kinetic molecular model.
C2.2f	Compare the average kinetic energy of the molecules in a metal object and a wood object at room temperature.
C3.3	*Heating Impacts* Heating increases the kinetic (translational, rotational, and vibrational) energy of the atoms composing elements and the molecules or ions composing compounds. As the kinetic (translational) energy of the atoms, molecules, or ions increases, the temperature of the matter increases. Heating a sample of a crystalline solid increases the kinetic (vibrational) energy of the atoms, molecules, or ions. When the kinetic (vibrational) energy becomes great enough, the crystalline structure breaks down, and the solid melts.
C3.3A	Describe how heat is conducted in a solid.
C3.3B	Describe melting on a molecular level.
C4.3	*Properties of Substances* Differences in the physical and chemical properties of substances are explained by the arrangement of the atoms, ions, or molecules of the substances and by the strength of the forces of attraction between the atoms, ions, or molecules.
C4.3A	Recognize that substances that are solid at room temperature have stronger attractive forces than liquids at room temperature, which have stronger attractive forces than gases at room temperature.

C4.3B	Recognize that solids have a more ordered, regular arrangement of their particles than liquids and that liquids are more ordered than gases.
C4.5x	*Ideal Gas Law* The forces in gases are explained by the ideal gas law.
C4.5a	Provide macroscopic examples, atomic and molecular explanations, and mathematical representations (graphs and equations) for the pressure-volume relationship in gases.
C4.5b	Provide macroscopic examples, atomic and molecular explanations, and mathematical representations (graphs and equations) for the pressure-temperature relationship in gases.
C4.5c	Provide macroscopic examples, atomic and molecular explanations, and mathematical representations (graphs and equations) for the temperature-volume relationship in gases.

CHEMISTRY

Unit 7: States of Matter

Big Idea (Core Concepts):
Particles in all matter are in constant motion until the temperature reaches absolute zero.

The order and organization in the universe is illustrated in the pressure, volume and temperature relationships which can be predicted by models, mathematical equations and graphs.

Standard(s):
C2: Forms of Energy
C3: Energy Transfer and Conservation
C4: Properties of Matter

Content Statement(s):
C2.2: Molecules in Motion
C2.2x: Molecular Entropy
C3.3: Heating Impacts
C4.3: Properties of Substances
C4.5x: Ideal Gas Law

Content Expectations: (Content Statement Clarification)
C2.2A: Describe conduction in terms of molecules bumping into each other to transfer energy. Explain why there is better conduction in solids and liquids than gases.

Clarification: None

C2.2B: Describe the various states of matter in terms of the motion and arrangement of the molecules (atoms) making up the substance.

Clarification: None

C2.2c: Explain changes in pressure, volume, and temperature for gases using the kinetic molecular model.

Clarification: Emphasize the understanding of the kinetic model. No calculation or relationship of measurable quantities is needed.

C2.2f: Compare the average kinetic energy of the molecules in a metal object and a wood object at room temperature.

Clarification: Note both objects are at the same temperature.

C3.3A: Describe how heat is conducted in a solid.

Clarification: None

C3.3B: Describe melting on a molecular level.

Clarification: None

C4.3A: Recognize that substances that are solid at room temperature have stronger attractive forces than liquids at room temperature, which have stronger attractive forces than gases at room temperature.

Clarification: None

C4.3B: Recognize that solids have a more ordered, regular arrangement of their particles than liquids and that liquids are more ordered than gases.

Clarification: None

C4.5a: Provide macroscopic examples, atomic and molecular explanations, and mathematical representations (graphs and equations) for the pressure-volume relationship in gases.

Clarification: None

C4.5b: Provide macroscopic examples, atomic and molecular explanations, and mathematical representations (graphs and equations) for the pressure-temperature relationship in gases.

Clarification: None

C4.5c: Provide macroscopic examples, atomic and molecular explanations, and mathematical representations (graphs and equations) for the temperature-volume relationship in gases.

Clarification: None

Vocabulary
Conduction
Kinetic molecular model
Kelvin temperature
Order
Pressure-temperature relationship
Pressure-volume relationship
Rotational motion
Temperature-volume relationship
Translational motion
Vibrational motion

Real World Context
Hot liquids can make the handle of a metal spoon hot through conduction.

Air pressure in automobile tires increases while driving due to friction within the tire and friction between the road and the tire. Recommended tire pressure is based on cold pressure (before driving).

Weather balloons are never filled to capacity because they continue to inflate as they rise due to changes in the air pressure.

Cooking pans get hot because of conduction of heat.

Perfume and smoke spread out in a room or area because of the motion of the particles. Through diffusion from an area of high concentration to areas of less concentration smoke or perfume spreads throughout a room.

Pressure relief values are used on hot water boilers and in pressure cookers as safety devices.

Regulators are used in SCUBA diving to match water pressure with the air pressure going into the lungs.

Aerosol cans work because of the pressure (propellant) in the can.

Instruments, Measurement, and Representations
Draw diagrams and pictures to illustrate heat conduction.

Create models to demonstrate molecules in motion (translational, rotational, and vibrational).

Act out the motion and arrangement particles in a substance.

Construct model or use people to demonstrate order and disorder.

Graph relationships between pressure and volume (P & V), pressure and temperature (P & T), and volume and temperature (V & T).

These expectations (C4.5a, C4.5b, C4.5c) can be calculated using the relationships between the variables changing in each situation (all other variables remain constant).
Pressure varies inversely with the volume, $P_1V_1 = P_2V_2$
Pressure varies directly with the Kelvin temperature, $P_1/T_1 = P_2/T_2$
Volume varies directly with the Kelvin temperature, $V_1/T_1 = V_2/T_2$

Exclusion: There is no need to make calculations using the Ideal Gas Law.

Instructional Examples:

i. Inquiry
CE: C1.1D, C4.5a

How exactly does volume change with changes in pressure? Using the provided equipment, a rubber-plugged or capped syringe, textbooks, ring stand, utility clamp, and graph paper, conduct an experiment to collect the necessary data. Graphically and mathematically present your results. Set a syringe volume at the maximum, seal the tip with a solid rubber stopper and support the syringe with a clamp attached to a ring stand. Read the volume of air in the syringe before adding any weights (books work best) and then read the volume after adding each weight. NOTE: The books represent added pressure on the gas. Each book is additional pressure and can be graphed as books or can be changed to mass.

ii. Reflection
CE: C1.2C, C4.5a, C4.5b, C4.5c

Investigate a hobby, sport, or activity that involves changes in gas pressure, volume, or temperature. (Some Possible choices are hot air ballooning, SCUBA diving, mountain climbing.) Report your results by writing a paper, making a poster presentation or small group presentation.

iii. Enrichment
CE: C4.3B

Create models that show changes in disorder for the following conditions: water at -5^{o}C, water at room temperature, water at 75^{o}C, and above 100^{o}C

iv. General
CE: C4.3B

Act out or create flipbooks to show the motion of particles in a solid, liquid, or a gas.

v. Intervention
CE: C4.5c

Investigate the relationship between volume and temperature by using a water plug in a glass tube. Insert a small diameter glass tube, approximately 30-40 cm in length into a one-hole rubber stopper which fits tightly into a 250 mL Erlenmeyer flask. Draw a color plug of water, about 1 cm high, into the glass tube. The plug of water should be about the middle of the tube. Cover the end of the tube with your finger and insert stopper (the tube is inside the stopper) into the flask. The flask can then be heated or cooled to various temperatures to investigate the relationship. (Alternate method, use a partially inflated balloon and measure the diameter with a piece of string.)

Units by Content Expectation

CHEMISTRY

Unit 8: Advanced Bonding Concepts

Code	Content Expectation
C4.3x	*Solids* Solids can be classified as metallic, ionic, covalent, or network covalent. These different types of solids have different properties that depend on the particles and forces found in the solid.
C4.3c	Compare the relative strengths of forces between molecules based on the melting point and boiling point of the substances.
C4.3d	Compare the strength of the forces of attraction between molecules of different elements. (For example, at room temperature, chlorine is a gas and iodine is a solid.)
C4.3e	Predict whether the forces of attraction in a solid are primarily metallic, covalent, network covalent, or ionic based upon the elements' location on the periodic table.
C4.3f	Identify the elements necessary for hydrogen bonding (N, O, F).
C4.3g	Given the structural formula of a compound, indicate all the intermolecular forces present (dispersion, dipolar, hydrogen bonding).
C4.3h	Explain properties of various solids such as malleability, conductivity, and melting point in terms of the solid's structure and bonding.
C4.3i	Explain why ionic solids have higher melting points than covalent solids. (For example, NaF has a melting point of 995°C while water has a melting point of 0° C.)
C5.4x	*Changes of State* All changes of state require energy. Changes in state that require energy involve breaking forces holding the particles together. The amount of energy will depend on the type of forces.
C5.4c	Explain why both the melting point and boiling points for water are significantly higher than other small molecules of comparable mass (e.g., ammonia and methane).
C5.4d	Explain why freezing is an exothermic change of state.
C5.4e	Compare the melting point of covalent compounds based on the strength of IMFs (intermolecular forces).

CHEMISTRY

Unit 8: Advanced Bonding Concepts

Big Idea (Core Concepts)
Many physical properties of substances can be determined by knowing the type of bond structure that exists within the substance.
Forces that exist between atoms can be classified into specific categories.

Standard(s):
C4: Properties of Matter
C5: Changes in Matter

Content Statement(s):
C4.3x: Solids
C5.4x: Changes of State

Content Expectations: (Content Statement Clarification)
C4.3c: Compare the relative strengths of forces between molecules based on the melting point and boiling point of the substances.

Clarification: None

C4.3d: Compare the strength of the forces of attraction between molecules of different elements. (For example, at room temperature, chlorine is a gas and iodine is a solid.)

Clarification: Compare the elements within one family only at a time. (i.e. alkali metals, alkaline earth metals, halogens, noble gases)

C4.3e: Predict whether the forces of attraction in a solid are primarily metallic, covalent, network covalent, or ionic based upon the elements' location on the periodic table.

Clarification: None

C4.3f: Identify the elements necessary for hydrogen bonding (N, O, F).

Clarification: None

C4.3g: Given the structural formula of a compound, indicate all the intermolecular forces present (dispersion, dipolar, hydrogen bonding).

Clarification: None

C4.3h: Explain properties of various solids such as malleability, conductivity, and melting point in terms of the solid's structure and bonding.

Clarification: None

C4.3i: Explain why ionic solids have higher melting points than covalent solids. (For example, NaF has a melting point of 995°C while water has a melting point of 0° C.)

Clarification: None

C5.4c: Explain why both the melting point and boiling points for water are significantly higher than other small molecules of comparable mass (e.g., ammonia and methane).

Clarification: None

C5.4d: Explain why freezing is an exothermic change of state.

Clarification: None

C5.4e: Compare the melting point of covalent compounds based on the strength of IMFs (intermolecular forces).

Clarification: Within a family or a group of compounds with similar formulas, the dispersion forces increase as molecular mass increases. The larger the molecule, the greater the number of electrons available to create a temporary dipole are, and thus results in a stronger force of attraction.

Vocabulary

Atomic weight
Boiling point
Chemical bond
Dipole-dipole bond
Dispersion forces
Endothermic process
Exothermic process
Hydrogen bonding
Ion
Ionic solid (crystal)
Melting point
Metal
Network solid
Relative mass
Release of energy
Temporary dipole

Real World Context

Wiring in homes is mostly done with Cu. However, depending on cost factors, sometimes Al has been used. In computers, gold is used in some connections because of its better conductivity.

The properties of water, which we depend on greatly, are the result of its special bonding.

Carbon dioxide, dry ice, is used to keep food products cold during shipment.

Diamond, a covalent network, is used in manufacturing processes (cutting and drilling) because of its hardness and in jewelry because of its lasting ability and beauty.

Carbon is covalently bonded when used to form polymers. These polymers may be used in various types of plastics, such as garbage bags, milk cartons, shrink wrap, automobile parts and toys.

Ionic bonded compounds are used in fertilizers, K_2CO_3, NH_4NO_3, and $Ca(H_2PO_4)_2$ because of their ability to dissolve in water.

Some ionic compounds, such as anhydrous compounds, are used as drying agents and are packaged with electronic equipment and many other substances, to remove moisture after manufacturing and before consumer use.

Water drops that form on plant blossoms from the early morning's dew are based on strong attractive forces between the highly polar water molecules.

Water striders are able to stay on top of the water, rather than sink, because of the water tension or attractive forces of the water molecules for one another.

Instruments, Measurement, and Representations

The periodic table can be used to make predictions about the type of bonds that will form.

Instructional Examples:

i. Inquiry
CE: C1.1f, C4.3d, C4.3h

How can the physical properties of a metal be changed?
Find out about the differences that exist between spring, annealed, and hardened steel and then using a bobby pin, design an experiment that prepares and tests each of three types. Observe and record the differences.

ii. Reflection
CE: C1.2C, C4.3h

Research the following materials which are both ionic compounds and used in over-the-counter drugs: magnesium hydroxide and magnesium sulfate. What are they used in and what purpose(s) do they serve. Evaluate your findings and make a presentation of the results.

iii. Enrichment
CE: C4.3d, C4.3f. C5.4c

If hydrogen bonding did not exist, especially with oxygen, what changes would exist on earth?

iv. General
CE: C4.3d, C4.3g

Chromatography is a way to separate components of a mixture based on differences in polarity. How might chromatography be used to separate and analyze the composition of the dyes in a marker?

v. Intervention
CE: C4.3e, C4.3g, C4.3i

Place the following compounds in increased order of melting point: potassium chloride, paraffin and ice. Explain your ordering system using what you know about bonding structure.

Units by Content Expectation

CHEMISTRY

Unit 9: Thermochemistry and Solutions

Code	Content Expectation
C2.1x	*Chemical Potential Energy* Potential energy is stored whenever work must be done to change the distance between two objects. The attraction between the two objects may be gravitational, electrostatic, magnetic, or strong force. Chemical potential energy is the result of electrostatic attractions between atoms.
C2.1c	Compare qualitatively the energy changes associated with melting various types of solids in terms of the types of forces between the particles in the solid.
C2.2x	*Molecular Entropy* As temperature increases, the average kinetic energy and the entropy of the molecules in a sample increases.
C2.2d	Explain convection and the difference in transfer of thermal energy for solids, liquids, and gases using evidence that molecules are in constant motion.
C3.1x	*Hess's Law* For chemical reactions where the state and amounts of reactants and products are known, the amount of energy transferred will be the same regardless of the chemical pathway. This relationship is called Hess's law.
C3.1c	Calculate the ΔH for a chemical reaction using simple coffee cup calorimetry.
C3.1d	Calculate the amount of heat produced for a given mass of reactant from a balanced chemical equation.
C3.4x	*Enthalpy and Entropy* All chemical reactions involve rearrangement of the atoms. In an exothermic reaction, the products have less energy than the reactants. There are two natural driving forces: (1) toward minimum energy (enthalpy) and (2) toward maximum disorder (entropy).
C3.4g	Explain why gases are less soluble in warm water than cold water.
C4.7x	*Solutions* The physical properties of a solution are determined by the concentration of solute.
C4.7a	Investigate the difference in the boiling point or freezing point of pure water and a salt solution.
C5.4	*Phase/Change Diagrams* Changes of state require a transfer of energy. Water has unusually high-energy changes associated with its changes of state.
C5.4A	Compare the energy required to raise the temperature of one gram of aluminum and one gram of water the same number of degrees.

C5.4B	Measure, plot, and interpret the graph of the temperature versus time of an ice-water mixture, under slow heating, through melting and boiling.
C5.5x	*Chemical Bonds* Chemical bonds can be classified as ionic, covalent, and metallic. The properties of a compound depend on the types of bonds holding the atoms together.
C5.5e	Relate the melting point, hardness, and electrical and thermal conductivity of a substance to its structure.

CHEMISTRY

Unit 9: Thermochemistry and Solutions

Big Idea (Core Concepts)
Heat released or absorbed in chemical reactions is proportional to the amounts of reactants consumed.

When a reversible process occurs, the same amount of energy is involved no matter which way the reaction proceeds. The difference will be if the energy is released or absorbed.

Standard(s):
C2: Forms of Energy
C3: Energy Transfer and Conservation
C4: Properties of Matter
C5: Changes in Matter

Content Statement(s):
C2.1x: Chemical Potential Energy
C2.2x: Molecular Entropy
C3.1x: Hess's Law
C3.4x: Enthalpy and Entropy
C4.7x: Solutions
C5.4: Phase/Change Diagrams
C5.5x: Chemical Bonds

Content Expectations: (Content Statement Clarification)
C2.1c: Compare qualitatively the energy changes associated with melting various types of solids in terms of the types of forces between the particles in the solid.

Clarification: Compare a variety of substances, free elements (monatomic and/or diatomic), ionic compounds, molecular compounds, and something with hydrogen bonding. You might consider looking at melting points of common materials, such as Na, O_2, CH_4, H_2O, and NaCl.

C2.2d: Explain convection and the difference in transfer of thermal energy for solids, liquids, and gases using evidence that molecules are in constant motion.

Clarification: None

C3.1c: Calculate the ΔH for a chemical reaction using simple coffee cup calorimeter.

Clarification: None

C3.1d: Calculate the amount of heat produced for a given mass of reactant from a balanced chemical equation.

Clarification: The heat of reaction for the balanced equation must be given.

C3.4g: Explain why gases are less soluble in warm water than cold water.

Clarification: As temperature increases the more disordered state is favored. Dissolved gases have less entropy than undissolved gases so as temperature increases the change is forced toward the gaseous phase.

C4.7a: Investigate the difference in the boiling point or freezing point of pure water and a salt solution.

Clarification: None

C5.4A: Compare the energy required to raise the temperature of one gram of aluminum and one gram of water the same number of degrees.

Clarification: Specific heat values must be given.

C5.4B: Measure, plot, and interpret the graph of the temperature versus time of an ice-water mixture, under slow heating, through melting and boiling.

Clarification: None

C5.5e: Relate the melting point, hardness, and electrical and thermal conductivity of a substance to its structure.

Clarification: Physical properties of a substance depend on the strength and types of bonding holding it together. Use examples: common covalent network (diamond or silicon dioxide), a metal (copper or gold), and ionic substance (sodium chloride).

<u>Vocabulary</u>

Boiling point elevation
Calorie
Change of state
Chemical bond
Concentration
Convection current
Convection heating
Crystalline solid
Electrostatic attractions
Enthalpy
Entropy
Equilibrium
Exothermic reaction

Freezing point depression
Hess's Law
Ionic motion
Joules
Kinetic energy
Mass to energy conversion
Potential energy
Release of energy
Solute
Specific heat
Transforming matter and/or energy

Real World Context

The decreasing solubility of gases with increasing temperatures is also responsible for the formation of boiler scale. (At higher temperatures, the amount of $CO_{2\,(g)}$ decreases which in turn causes the following reaction to occur: $HCO_3^{1-}{}_{(aq)} \leftrightarrow H_2O_{(l)} + CO_{2(g)} + CO_3^{2-}{}_{(aq)}$ If Ca^{2+} ions are present, calcium carbonate, which has low solubility in water will form, which is boiler scale.)

Thermal pollution in rivers and lakes causes a decrease in the amounts of dissolved oxygen.

Putting salt on the icy roads in winter to melt ice, lowers the freezing point of water.

The differences in cooking pans are due to differences in specific heat capacities. An iron pan will not heat up as quickly as an aluminum or copper pan.

In order to maintain body temperature, part of the cooling process is done by convection. Heat is lost by virtue of heating air that is in contact with the body. The heated air rises and is replaced by cooler air and the process continues.

Instruments, Measurement, and Representations

Phase change diagrams can be drawn, read, and/or interpreted.

Calculations of heat, based on thermochemical equations can be done. (Stoichiometric relationships)

Calorimeter calculation using: $Heat_{(surr)} = q_{(surr)} = m \times C \times \Delta T$ [m= mass of substance, C = specific heat capacity, and ΔT is the change in temperature]

$Heat_{(system)} = q_{(system)} = -q_{(surr)}$ [(surr) = surroundings = liquid in the cup which is absorbing the heat; (system) = reactants – products of reaction]

Instructional Examples:

i Inquiry
CE: C1.1D, C4.7a

How does the amount of solute in a solution affect its boiling temperature? Using equal amounts of water dissolve different amounts of the same solute in the water and determine the boiling point of the resulting solution. What conclusions can you reach regarding the solute used? What conclusions can you reach about amount of solute?

ii Reflection
CE: C1.2C, C3.4g

Since most gases are only slightly soluble in water, try to explain the following situation based on your current and prior knowledge of chemistry. A liter of water saturated with oxygen gas at room temperature contains about 0.05 grams of oxygen gas. On the other hand, approximately 255 grams of ammonia gas, enough to make approximately a 15 mole/liter solution can be dissolved in the same amount of water at the same temperature. What can possibly account for the huge difference?

iii Enrichment
CE: C3.1c; C3.1d

(CAUTION: Sodium hydroxide is a caustic substance. Observe proper precautions in handling and cleanup of any spills. Be sure none gets on the skin.)
How much energy is involved in the dissolving of sodium hydroxide in water?

Using a simple coffee cup calorimeter, design and carry out a laboratory experiment to find the answer.

Write a balanced thermo chemical equation showing your results.

If the accepted value of reaction is 44.51 kJ/mol, what is your percentage error?

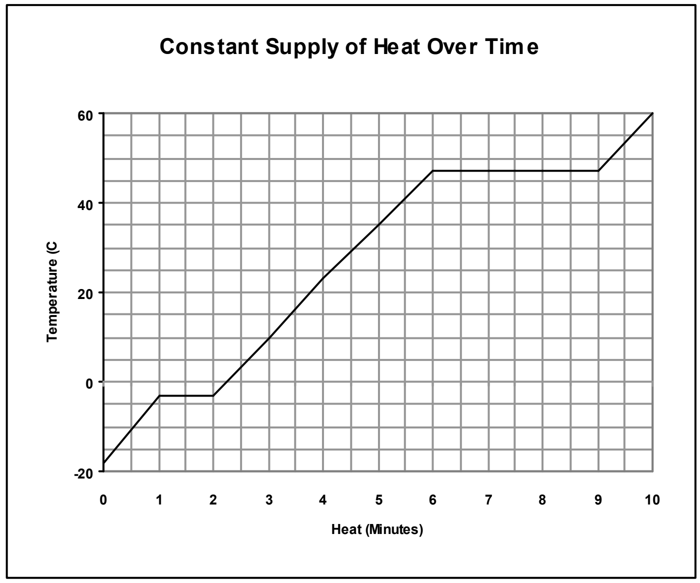
How might you revise the experiment in order to achieve better results?

iv. General
CE: C2.1c; C5.4B: C5.5e

Using the diagram to the right, answer the following questions:

A. What is the boiling point of the substance?
B. What is the freezing temperature of the substance?
C. What type of bonding does the substance most likely contain?
D. Under normal conditions, what phase would it exist in at room temperature?

Speculate as to the identity of the element.



v. Intervention
CE: C2.2d

Using a beaker of cold water, add a drop of food coloring. Watch what happens and record your observations. Next, starting with a fresh sample of the same amount of water, only heat it about 15°C warmer than the first time. Add a drop of food coloring and again observe what happens. What differences if any did you observe? Repeat the same process a third time but this time make the water even warmer. What happens this time? Explain what you observed in terms of convection.

Units by Content Expectation

CHEMISTRY

Unit 10: Acid/Base

Code	Content Expectation
C5.7	*Acids and Bases* Acids and bases are important classes of chemicals that are recognized by easily observed properties in the laboratory. Acids and bases will neutralize each other. Acid formulas usually begin with hydrogen, and base formulas are a metal with a hydroxide ion. As the pH decreases, a solution becomes more acidic. A difference of one pH unit is a factor of 10 in hydrogen ion concentration.
C5.7A	Recognize formulas for common inorganic acids, carboxylic acids, and bases formed from families I and II.
C5.7B	Predict products of an acid-based neutralization.
C5.7C	Describe tests that can be used to distinguish an acid from a base.
C5.7D	Classify various solutions as acidic or basic, given their pH.
C5.7E	Explain why lakes with limestone or calcium carbonate experience less adverse effects from acid rain than lakes with granite beds.
C5.7x	*Bronsted-Lowry* Chemical reactions are classified according to the fundamental molecular or submolecular changes that occur. Reactions that involve proton transfer are known as acid/base reactions.
C5.7f	Write balanced chemical equations for reactions between acids and bases and perform calculations with balanced equations.
C5.7g	Calculate the pH from the hydronium ion or hydroxide ion concentration.
C5.7h	Explain why sulfur oxides and nitrogen oxides contribute to acid rain.

CHEMISTRY

Unit 10: Acid/Base

Big Idea (Core Concepts):
The environment is impacted by chemical reactions on earth.

Acids, bases and pH are systems developed by man to help describe natural systems.

Standard(s):
C5: Changes in Matter

Content Statement(s):
C5.7: Acids and Bases
C5.7x: Brønsted-Lowry Chemical Reactions

Content Expectations: (Content Statement Clarification)
C5.7A: Recognize formulas for common inorganic acids, carboxylic acids, and bases formed from families I and II.

Clarification:
Limit examples of inorganic acids to: HCl, HBr, and HI
Limit common oxy-acids to: H_2SO_4 and HNO_3
Limit carboxylic acids to: H_2CO_3 and $HC_2H_3O_2$.
Limit bases to: hydroxides of alkali and alkaline earth metals.

C5.7B: Predict products of an acid-based neutralization.

Clarification: Use examples to illustrate that a salt and water are products of the reaction.

Example: $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

C5.7C: Describe tests that can be used to distinguish an acid from a base.

Clarification: Limit indicators to litmus, phenolphthalein and universal indicator both in the aqueous form and treated paper form. For universal indicator, color changes or a color chart will be given. Other properties could also be used such as acidic foods taste sour and bases taste bitter and feel slippery. Acids react with most metals to produce hydrogen gas.

C5.7D: Classify various solutions as acidic or basic, given their pH.

Clarification: Solutions may also be neutral and therefore, would not be classified as acidic or basic.

C5.7E: Explain why lakes with limestone or calcium carbonate experience less adverse effects from acid rain than lakes with granite beds.

Clarification: Focus on the neutralization reaction.

$CaCO_3 + H_2SO_4 \rightarrow CaSO_4 + CO_2 + H_2O$

Granite + $H_2SO_4 \rightarrow$ No Reaction

C5.7f: Write balanced chemical equations for reactions between acids and bases and perform calculations with balanced equations.

Clarification: Use only strong acids (HCl, H_2SO_4 and HNO_3) and strong bases, (the hydroxides of alkali and alkaline earth metals). Calculations would review concepts presented in Unit 6, "Equations and Stoichiometry."

C5.7g: Calculate the pH from the hydronium ion or hydroxide ion concentration.

Clarification: Calculate the pH from hydronium ion concentration, $-\log[H_3O^+]$, or using Kw, 1×10^{-14}.
Exclusion: Molarity is not a vocabulary word.
$[H^+O]$ = Hydronium ion concentration and the units are mole/liter. Molarity is not required

C5.7h: Explain why sulfur oxides and nitrogen oxides contribute to acid rain

Clarification: Sulfuric and nitric acids can be formed when sulfur oxides and nitrogen oxides mix with rain water.

Vocabulary:
Acid rain
Acid/base reaction
Acidic
Alkaline
Basic
Bronsted-Lowry
Carboxyl group
Hydrogen ion
Hydronium ion
Hydroxide
Ion
K_w,
Neutral
Neutralize
pH

Real World Context:

Household cleaners are acidic or basic. Examples: soaps, shampoos, window cleaners, toilet bowl cleaners, vinegar and drain cleaners, etc.

Foods and medicines are acidic or basic. Examples are: soda pop, antacids, vinegar, salad dressing etc.

Food processing requires adherence to strict pH ranges: canning, meat tenderizer, etc.

Indicators are used to test pH of soil and swimming pools.

Red cabbage juice and grape juice are common substances that can be used as a pH indicator.

Some plants may change flower color due to the pH of the soil. Some hydrangeas bloom blue in acid soil and pink in alkaline soil.

Acid rain can have economic and aesthetic effects on lakes and structures (limestone, marble and metals)

Nitrous oxides are produced from nitrogen in the air reacting at high temperatures with the oxygen in the air. Examples: internal combustion engines and lightning

Instruments, Measurement, and Representations:

Use formulas to calculate pH.

$pH = -\log[H_3O^+]$

$1 \times 10^{-14} = [H_3O^+][OH^-]$

pH meters or electronic probes

Concentration = mole/liter

Exclusion: Molarity

Instructional Examples:

i. Inquiry

CE: C1.1A, 1.1C, C5.7C

Can we make our own pH indicator?

Using red cabbage; design an experiment to extract the juice from the cabbage.
Make a table showing the color changes associated with red cabbage juice.
Extension: Grape juice can also be used.

ii. Reflection
CE: C1.2C, C5.7A, C5.7D

Research antacids used in health care. Make a poster presentation describing how they work and some of the dangerous and beneficial side effects.
Extension: Determine the "best" buy based on cost.

iii. Enrichment
CE: C1.1B, C1.1h, C5.7f, C5.7C

Determine the unknown concentration of an acid using a known concentration of base given the acid and base formulas and phenolphthalein indicator.

iv. General
CE: C5.7A, C5.7B, C5.7f

Taste various foods and classify them as acidic or basic based on taste alone. Make a table to display your results and add a column for verification using litmus paper. Add another column if you can find the acidity of the foods from references.
Food Examples: vinegar, lemon, Brussels sprouts, broccoli, tomato, milk of magnesia, etc.

v. Intervention
CE: C5.7D, C5.7g

Substance	$[H_3O^+]$	$[OH^-]$	pH	Acidic or Basic
Apple	1.00×10^{-3}			
Ginger ale		3.16×10^{-12}		
Human blood	3.98×10^{-8}			
Maple syrup	1.99×10^{-7}			
Milk of magnesia		3.33×10^{-4}		
Tomato	6.31×10^{-5}			
Window cleaner		1.11×10^{-12}		

Units by Content Expectation

CHEMISTRY

Unit 11: Redox/Equilibrium

Code	Content Expectation
C5.3x	*Equilibrium* Most chemical reactions reach a state of dynamic equilibrium where the rates of the forward and reverse reactions are equal.
C5.3a	Describe equilibrium shifts in a chemical system caused by changing conditions (Le Chatelier's Principle).
C5.3b	Predict shifts in a chemical system caused by changing conditions (Le Chatelier's Principle).
C5.3c	Predict the extent reactants are converted to products using the value of the equilibrium constant.
C5.6x	*Reduction/Oxidation Reactions* Chemical reactions are classified according to the fundamental molecular or submolecular changes that occur. Reactions that involve electron transfer are known as oxidation/reduction (or "redox").
C5.6a	Balance half-reactions and describe them as oxidations or reductions.
C5.6c	Explain oxidation occurring when two different metals are in contact.
C5.6d	Calculate the voltage for spontaneous redox reactions from the standard reduction potentials.
C5.6e	Identify the reactions occurring at the anode and cathode in an electrochemical cell.

CHEMISTRY

Unit 11: Redox/Equilibrium

<u>Big Idea</u> (Core Concepts):
Many redox (oxidation-reduction) reactions are a source of energy.

Redox reactions significantly impact humans in both positive and negative ways.

In a closed system, many reactions will reach equilibrium. Changes to the equilibrium can be predicted by using Le Châtelier's Principle.

<u>Standard(s)</u>:
C5: Changes in Matter

<u>Content Statement(s)</u>:
C5.3x: Equilibrium
C5.6x: Reduction/Oxidation Reactions

<u>Content Expectations:</u> (Content Statement Clarification)
C5.3a: Describe equilibrium shifts in a chemical system caused by changing conditions (Le Châtelier's Principle).

Clarification: Conditional changes are limited to changing the concentration of reactants or products as well as heat and pressure. Limit discussion to changing only one variable at a time.
Exclusion: Common ion effect

C5.3b: Predict shifts in a chemical system caused by changing conditions (Le Châtelier's Principle).

Clarification: None

C5.3c: Predict the extent reactants are converted to products using the value of the equilibrium constant.

Clarification: No calculations are necessary. If the K_{eq} is greater than 1 the products are favored when equilibrium is reached. If K_{eq} is less than 1 the reactants are favored.

C5.6a: Balance half-reactions and describe them as oxidations or reductions.

Clarification: Limit these reactions to balancing electrons on reactant or product side of the equation.

Example: $Mg \rightarrow Mg^{2+} + 2e^-$

$Cl_2 + 2e^- \rightarrow 2Cl^-$

Exclusion: Reactions in acidic or basic conditions

C5.6c: Explain oxidation occurring when two different metals are in contact.

Clarification: None

C5.6d: Calculate the voltage for spontaneous redox reactions from the standard reduction potentials.

Clarification: None

C5.6e: Identify the reactions occurring at the anode and cathode in an electrochemical cell.

Clarification: Oxidation occurs at the anode and reduction occurs at the cathode

Vocabulary:

Anode
Cathode
Electrochemical Cell
Equilibrium
K_{eq}
Le Châtelier
Oxidation
Oxidation-reduction reactions
Reduction

Real World Context:
Unprotected iron on automobiles or other steel structures will rust.

Batteries are electrochemical cells.

Hydrogen fuel cells produce water and energy using hydrogen and oxygen.

Outdoor grilling uses combustion, a redox reaction.

Commercially available hot and cold packs

Electroplating

Sacrificial anodes (made of magnesium or zinc generally) are used on ships, in water heaters, and on the Alaskan pipeline to prevent corrosion of the primary metal.

Instruments, Measurement, and Representations:
$E_{cell} = E_{red} - E_{oxid}$, spontaneous if > 0

Standard Reduction Potential table

OIL RIG (Oxidation Is Loss, Reduction Is Gain) in regards to electrons.

Models that demonstrate one metal protecting another metal as a sacrificial anode.

Instructional Examples:
i. Inquiry
CE: C1.1D, C5.6a, C5.6e

How was the Standard Reduction Potential table determined? Using six metals and their nitrate solutions, a twelve-cell well plate, small strips of filter paper soaked in potassium nitrate, and a voltmeter, design an experiment to create a reduction potential series.

ii. Reflection
CE: C1.2j, C5.3c

Investigate the pros and cons of hydrogen fuel cell energy vs. hydrocarbon fuels.

iii. Enrichment
CE: C5.3a, C5.3b

Given the following equilibrium reaction: $2SO_{3\ (g)} \rightarrow 2SO_{2\ (g)} + O_{2\ (g)}$
$\Delta H = 197$ kJ

What effect will each of the following have on the amount of SO_3 in equilibrium?

A. Oxygen gas is added.
B. The pressure is increased by decreasing the volume.
C. The temperature is decreased.
D. Gaseous sulfur dioxide is removed.

iv. General
CE: C5.6a, C5.6c, C5.6d

Conduct research on dry cell and wet cell batteries. Explain how the batteries are similar and different. Why is one used over the other for specific applications?

v. Intervention
CE: C5.6d

Design an experiment using copper pennies, aluminum foil, and wet (saltwater) paper towels that will demonstrate the electric potential difference. Investigate other metals.

Units by Content Expectation

CHEMISTRY

Unit 12: Thermodynamics

Code	Content Expectation
C2.2x	*Molecular Entropy* As temperature increases, the average kinetic energy and the entropy of the molecules in a sample increases.
C2.2e	Compare the entropy of solids, liquids, and gases.
C2.3x	*Breaking Chemical Bonds* For molecules to react, they must collide with enough energy (activation energy) to break old chemical bonds before their atoms can be rearranged to form new substances.
C2.3a	Explain how the rate of a given chemical reaction is dependent on the temperature and the activation energy.
C2.3b	Draw and analyze a diagram to show the activation energy for an exothermic reaction that is very slow at room temperature.
C3.1x	*Hess's Law* For chemical reactions where the state and amounts of reactants and products are known, the amount of energy transferred will be the same regardless of the chemical pathway. This relationship is called Hess's law.
C3.1a	Calculate the ΔH for a given reaction using Hess's Law.
C3.1b	Draw enthalpy diagrams for exothermic and endothermic reactions.
C3.2x	*Enthalpy* Chemical reactions involve breaking bonds in reactants (endothermic) and forming new bonds in the products (exothermic). The enthalpy change for a chemical reaction will depend on the relative strengths of the bonds in the reactants and products.
C3.2a	Describe the energy changes in photosynthesis and in the combustion of sugar in terms of bond breaking and bond making.
C3.4	*Endothermic and Exothermic Reactions* Chemical interactions either release energy to the environment (exothermic) or absorb energy from the environment (endothermic).
C3.4B	Explain why chemical reactions will either release or absorb energy.
C3.4x	*Enthalpy and Entropy* All chemical reactions involve rearrangement of the atoms. In an exothermic reaction, the products have less energy than the reactants. There are two natural driving forces: (1) toward minimum energy (enthalpy) and (2) toward maximum disorder (entropy).
C3.4d	Draw enthalpy diagrams for reactants and products in endothermic and exothermic reactions.

C3.4e	Predict if a chemical reaction is spontaneous given the enthalpy (ΔH) and entropy (ΔS) changes for the reaction using Gibb's Free Energy, $\Delta G = \Delta H - T\Delta S$ (Note: mathematical computation of ΔG is not required.)
C3.4f	Explain why some endothermic reactions are spontaneous at room temperature.

CHEMISTRY

Unit 12: Thermodynamics

<u>Big Idea</u> (Core Concepts):
Chemical compounds and chemical reactions strive toward states of highest disorder as does every thing in the universe.

Bond formation releases energy to the system.

<u>Standard(s)</u>:
C2: Forms of Energy
C3: Energy Transfer and conservation

<u>Content Statement(s)</u>:
C2.2x: Molecular Entropy
C2.3x: Breaking Chemical Bonds
C3.1x: Hess's Law
3.2x: Enthalpy
C3.4: Endothermic and Exothermic Reactions
C3.4x: Enthalpy and Entropy

<u>Content Expectations:</u> (Content Statement Clarification)
C2.2e: Compare the entropy of solids, liquids, and gases.

Clarification: None

C2.3a: Explain how the rate of a given chemical reaction is dependent on the temperature and the activation energy.

Clarification: None

C2.3b: Draw and analyze a diagram to show the activation energy for an exothermic reaction that is very slow at room temperature.

Clarification: The diagram to show a very slow exothermic reaction at room temperature is one in which the energy of activation is very large.

C3.1a: Calculate the ΔH for a given reaction using Hess's Law.

Clarification: Use reactions involving only a two step process when the overall reaction and the heats of formation are given.

C3.1b: Draw enthalpy diagrams for exothermic and endothermic reactions.

Clarification: Activation energies need to be included in all diagrams.

C3.2a: Describe the energy changes in photosynthesis and in the combustion of sugar in terms of bond breaking and bond making.

Clarification: None

C3.4B: Explain why chemical reactions will either release or absorb energy.

Clarification: None

C3.4d: Draw enthalpy diagrams for reactants and products in endothermic and exothermic reactions.

Clarification: (see C3.1b)

C3.4e: Predict if a chemical reaction is spontaneous given the enthalpy (ΔH) and entropy (ΔS) changes for the reaction using Gibb's Free Energy, $\Delta G = \Delta H - T\Delta S$ (Note: mathematical computation of ΔG is not required.)

Clarification: There are two driving forces for all reactions, (1) decreasing energy (ΔH = -), and (2) increasing entropy (ΔS = +). If both forces are favorable (ΔH = (-), ΔS = (+)) the reaction is always spontaneous. If both forces are unfavorable (ΔH = (+), ΔS = (-)) the reaction cannot be spontaneous. If one force is favorable and the other unfavorable the spontaneity will depend on the temperature. If ΔG is negative then the reaction is spontaneous. If ΔG is zero then the reaction is at equilibrium.

C3.4f: Explain why some endothermic reactions are spontaneous at room temperature.

Clarification: None

Vocabulary

Activation energy
Disorder
Endothermic reaction
Enthalpy
Entropy
Exothermic reaction
Gibb's Free Energy
Hess's Law
Reaction rate
Release of energy
Spontaneous

Real World Context
Ice packs and hot packs chemically react and free energy is put to work.

Fuels involve a tremendous output of energy

Food—digestion is the slow release of chemical energy

Plants—photosynthesis is the accumulation of energy from a chemical reaction.

The major difference between the formation of diamond versus graphite is due to the large change of entropy

Instruments, Measurement, and Representations
Enthalpy graphs of exothermic and endothermic reactions

Hess's Law problems

$\Delta G = \Delta H - T\Delta S$

Instructional Examples:
i. Inquiry
CE: C1.1A, C1.1C, C2.3A

Design an experiment using Alka-Seltzer tablets to determine the effect temperature has on the reaction rate. After conducting the experiment construct a table and draw conclusions. Generate questions for further investigations.

ii. Reflection
CE: C1.2E

Look into firefighting as a career through research. Plan a report either written or oral to discuss the training and the incidents that relate to thermodynamics.

iii. Enrichment
CE: C3.4e

Stretch a rubber band against your forehead or lips (note the relative temperature).
Stretch the rubber band and hold it tight. Touch it back to your skin again (note the temperature change).
Release the rubber band allowing it to return to its original shape. Touch it to your skin again (note the temperature change).

Questions:
1) Is this process of stretching the rubber band exothermic or endothermic?
2) If there is no change in enthalpy because there is no reaction, what do you expect to be the order for the entropy (positive or negative)?
3) Is there more order or more disorder?
4) What would account for the change in entropy?

iv. General
CE: C3.1a

a) Use Hess's Law to calculate the enthalpy for the reaction
$Mg(s) + ½O_2(g) \rightarrow MgO(s)$

using the following information:

Reaction			ΔH
$Mg(s) + 2HCl(aq)$	$\rightarrow$	$MgCl_2(aq) + H_2(g)$	-142.82 kJ/mole
$MgO(s) + 2HCl(aq)$	$\rightarrow$	$MgCl_2(aq) + H_2O(l)$	-218.17 kJ/mole
$H_2(g) + ½O_2(g)$	$\rightarrow$	$H_2O(l)$	-286 kJ/mole

v. Intervention
CE:2.2e

a) Make a list of activities that are encountered everyday that exhibit high or low entropy. Make two columns in a table to show the highest state of entropy and the lowest state of entropy. Examples: deck of cards, clothes, room
b) Make a list of chemical reactions that are encountered everyday that exhibit endothermic or exothermic properties. Examples: photosynthesis, rusting, food digestion, etc.

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