ACT Course Standards
Chemistry

A set of empirically derived course standards is the heart of each QualityCore® science course. The ACT Course Standards represent a solid evidence-based foundation in science. They were developed from an intensive study of high-performing high schools with significant minority and low-income enrollments that produced many graduates who met or exceeded ACT College Readiness Benchmark Scores (See http://www.act.org/path/policy/reports/success.html).

This document contains a list of ACT Course Standards for a rigorous Chemistry course—what students should know and be able to do in the course—and a worksheet teachers can use to compare their course content to these standards. The ACT standards encompass the following overarching themes and/or foundational concepts:

A. Understanding Chemistry as Inquiry

B. Exploring the Physical World

C. Discovering the Language of Chemistry

D. Building Models of Matter

E. Integrating the Macroscopic, Microscopic, and Symbolic Worlds

ACT Course Standards—Chemistry

I. UNDERSTANDING CHEMISTRY AS INQUIRY (Note: Some of the process standards in this section are similar to those found in Biology and Physics.)
A. Foundations
1. Scientific Inquiry
a. Identify and clarify research questions and design experiments
b. Design experiments so that variables are controlled and appropriate numbers of trials are used
c. Collect, organize, and analyze data accurately and use techniques and equipment appropriately
d. Interpret results and draw conclusions, revising hypotheses as necessary and/or formulating additional questions or explanations
e. Write and speak effectively to present and explain scientific results, using appropriate terminology and graphics
f. Safely use laboratory equipment and techniques when conducting scientific investigations
g. Routinely make predictions and estimations
2. Mathematics and Measurement in Science
a. Distinguish between precision and accuracy with respect to experimental data
b. Use appropriate SI units for length, mass, time, temperature, quantity of matter, area, volume, and density; describe the relationships among SI unit prefixes (e.g., centi-, milli-, kilo-); recognize commonly used non-SI units
c. Use the correct number of significant figures in reporting measurements and the results of calculations
d. Use appropriate statistical methods to represent the results of investigations
e. Express numbers in scientific notation when appropriate
f. Solve for unknown quantities by manipulating variables
g. Use graphical, mathematical, and/or statistical models to express patterns and relationships inferred from sets of scientific data

© 2010 by ACT, Inc. All rights reserved.

3. Science in Practice
a. Explain and apply criteria that scientists use to evaluate the validity of scientific claims and theories
b. Explain why experimental replication and peer review are essential to eliminate as much error and bias as possible in scientific claims
c. Explain the criteria that explanations must meet to be considered scientific (e.g., be consistent with experimental/observational evidence about nature, be open to critique and modification, use ethical reporting methods and procedures)
d. Explain why all scientific knowledge is subject to change as new evidence becomes available to the scientific community
e. Use a variety of appropriate sources (e.g., Internet, scientific journals) to retrieve relevant information; cite references properly
f. Identify and analyze the advantages and disadvantages of widespread use of and reliance on technology
g. Compare the scientific definitions of fact, law, and theory, and give examples of each in chemistry
II. EXPLORING THE PHYSICAL WORLD
A. Introduction to Chemistry
1. Mass, Volume, and Density
a. Explain why mass is used as a quantity of matter and differentiate between mass and weight
b. Explain density qualitatively and solve density problems by applying an understanding of the concept of density
2. Elements, Atomic Mass, and Nomenclature
a. Use the IUPAC symbols of the most commonly referenced elements
b. Compare the characteristics of elements, compounds, and mixtures
c. Compare characteristics of isotopes of the same element
B. Properties of Matter and Gases
1. Phases of Matter, Phase Changes, and Physical Changes
a. Compare the definition of matter and energy and the laws of conservation of matter and energy
b. Describe how matter is classified by state of matter and by composition
c. Describe the phase and energy changes associated with boiling/condensing, melting/freezing, sublimation, and crystallization (deposition)
d. Explain the difference between chemical and physical changes and demonstrate how these changes can be used to separate mixtures and compounds into their components
e. Define chemical and physical properties and compare them by providing examples
2. The Nature of Gases
a. Define gas pressure and the various pressure units (e.g., torr, kilopascals, mm Hg, atmospheres)
b. Describe the use and operation of mercury barometers and manometers to find atmospheric pressure or relative gas pressures
c. Define the gas laws given by Boyle, Charles, Gay-Lussac, and Dalton and solve problems based on these laws
d. Predict boiling point changes based on changes in atmospheric pressure
e. Explain the basis for gaseous diffusion and effusion
f. Describe Avogadro's hypothesis and use it to solve stoichiometric problems
3. Ideal Gas Law
a. Explain the difference between an ideal and real gas, the assumptions made about an ideal gas, and what conditions favor ideal behavior for a real gas
b. Apply the mathematical relationships that exist among the volume, temperature, pressure, and number of particles in an ideal gas
c. Compute gas density when given molar mass, temperature, and pressure
d. Apply the ideal gas law to determine the molar mass of a volatile compound
e. Solve gas stoichiometry problems at standard and nonstandard conditions

© 2010 by ACT, Inc. All rights reserved.

III. DISCOVERING THE LANGUAGE OF CHEMISTRY
A. Formulas and Equations
1. Empirical Formulas, Molecular Formulas, and Percentage Composition
a. Distinguish between chemical symbols, empirical formulas, molecular formulas, and structural formulas
b. Interpret the information conveyed by chemical formulas for numbers of atoms of each element represented
c Use the names, formulas, and charges of commonly referenced polyatomic ions
d. Provide the interconversion of molecular formulas, structural formulas, and names, including common binary and ternary acids
e. Calculate the percent composition of a substance, given its formula or masses of each component element in a sample
f. Determine the empirical formulas and molecular formulas of compounds, given percent composition data or mass composition data
g. Determine percent composition experimentally and derive empirical formulas from the data (e.g., for hydrates)
2. Mole Concept, Molar Mass, Gram Formula Mass, and Molecular Mass
a. Explain the meaning of mole and Avogadro's number
b. Interconvert between mass, moles, and number of particles
c. Distinguish between formula mass, empirical mass, molecular mass, gram molecular mass, and gram formula mass
3. Chemical Equations and Stoichiometry
a. Explain how conservation laws form the basis for balancing chemical reactions and know what quantities are conserved in physical, chemical, and nuclear changes
b. Write and balance chemical equations, given the names of reactants and products
c. Describe what is represented, on a molecular and molar level, by chemical equations
d. Use the appropriate symbols for state (i.e., solid, liquid, gaseous, aqueous) and reaction direction when writing chemical equations
e. Classify chemical reactions as being synthesis, decomposition, single replacement, or double replacement reactions
f. Predict the products of synthesis, combustion, and decomposition reactions and write balanced equations for these reactions
g. Predict products of single replacement reactions, using the activity series, and write balanced equations for these reactions
h. Predict the products of double replacement reactions, using solubility charts to identify precipitates, and write balanced equations for these reactions
i. Use chemical equations to perform basic mole-mole, mass-mass, and mass-mole computations for chemical reactions
j. Identify limiting reagents and use this information when solving reaction stoichiometry problems
k. Compute theoretical yield, actual (experimental) yield, and percent yield
l. Calculate percent error and analyze experimental errors that affect percent error
m. Write ionic equations, identifying spectator ions and the net ionic equation

IV. BUILDING MODELS OF MATTER
A. Microscopic Nature of Matter
1. Structure of Liquids and Solids
a. Describe differences between solids, liquids, and gases at the atomic and molecular levels
b. Describe and perform common separation techniques (e.g., filtration, distillation, chromatography)
2. Kinetic Molecular Theory of Gases
a. Use the kinetic molecular theory to explain the states and properties (i.e., microscopic and macroscopic) of matter and phase changes
b. Explain the basis and importance of the absolute temperature scale and convert between the Kelvin and Celsius scales
c. Use the kinetic-molecular theory as a basis for explaining gas pressure, Avogadro's hypothesis, and Boyle's/Charles's laws

© 2010 by ACT, Inc. All rights reserved.

B. Atomic Structure and Chemical Bonding
1. Atomic Theory (Dalton), Atomic Structure, and Quantum Theory
a. Describe the importance of models for the study of atomic structure
b. Describe the crucial contributions of scientists and the critical experiments that led to the development of the modern atomic model
c. Describe characteristics of a wave, such as wavelength, frequency, energy, and speed
d. Describe the role of probability in orbital theory
e. Describe atomic orbitals (s, p, d, f) and their basic shapes
f. Apply Hund's rule and the Aufbau process to specify the electron configurations of the elements
2. Periodic Table and Periodicity
a. Describe the historical development of the modern periodic table, including work by Mendeleev and then Moseley
b. Describe and explain the organization of elements into periods and groups in the periodic table
c. Use the periodic table to determine the atomic number; atomic mass; mass number; and number of protons, electrons, and neutrons in isotopes of elements
d. Calculate the weighted average atomic mass of an element from isotopic abundance, given the atomic mass of each contributor
e. Identify regions (e.g., groups, families, series) of the periodic table and describe the chemical characteristics of each
f. Compare the periodic properties of the elements (e.g., metal/nonmetal/metalloid behavior, electrical/heat conductivity, electronegativity and electron affinity, ionization energy, atomic/covalent/ionic radius) and how they relate to position in the periodic table
g. Use the periodic table to predict and explain the valence electron configurations of the elements, to identify members of configuration families, and to predict the common valences of the elements
3. Intermolecular Forces and Types of Bonds
a. Describe the characteristics of ionic and covalent bonding
b. Explain ionic stability, recognize typical ionic configurations, and predict ionic configurations for elements (e.g., electron configurations, Lewis dot models)
c. Describe the nature of the chemical bond with respect to valence electrons in bonding atoms
d. Explain how ionic and covalent compounds differ
e. Describe the unique features of bonding in carbon compounds
f. Compare the different types of intermolecular forces (e.g., van der Waals, dispersion)
g. Explain and provide examples for dipole moments, bond polarity, and hydrogen bonding
h. Describe the unique physical and chemical properties of water resulting from hydrogen bonding
i. Explain the relationship between evaporation, vapor pressure, molecular kinetic energy, and boiling point for a single pure substance
j. Explain the relationship between intermolecular forces, boiling points, and vapor pressure when comparing differences in the properties of pure substances
k. Classify solids as ionic, molecular, metallic, or network
4. Orbital Theory Applied to Bonding
a. Use Lewis dot diagrams to represent bonding in ionic and covalent compounds
b. Draw Lewis structures for molecules and polyatomic ions, including those that must be represented by a set of resonance structures
c. Use VSEPR theory to explain geometries of molecules and polyatomic ions
d. Describe how orbital hybridization models relate to molecular geometry
e. Describe the molecular orbital models for double bonds, triple bonds, and delocalized pi electrons
f. Describe the relationship between molecular polarity and bond polarity

© 2010 by ACT, Inc. All rights reserved.

V. INTEGRATING THE MACROSCOPIC, MICROSCOPIC, AND SYMBOLIC WORLDS
A. Solutions
1. Types of Solutions, Concentration, and Solubility
a. Define solution, solute, and solvent
b. Compare properties of suspensions, colloids, and true solutions
c. Define the terms *saturated*, *unsaturated*, *supersaturated*, *dilute*, and *concentrated* as they pertain to solutions
d. Give examples of solid, liquid, or gas medium solutions
e. Define and calculate the molarity of a solution
f. Define and calculate the percent composition of a solution
g. Describe the preparation and properties of solutions
h. Solve stoichiometry calculations based on reactions involving aqueous solutions
i. Describe the relationship between temperature or pressure and the solubility of gases in liquids
j. Describe the relationship between solvent character and solute character and explain miscibility
k. Apply the general rules of solubility to aqueous salt solutions
l. Describe the factors affecting the solubility of a solute in a given solvent and its rate of solution
2. Colligative Properties
a. Describe qualitatively the effect of adding solute on freezing point, boiling point, and vapor pressure of a solvent
b. Define molality and mole fraction
c. Calculate changes in the boiling point and freezing point when nonvolatile, nonelectrolyte solutes are added to solvents
B. Kinetics, Equilibrium, and Thermodynamics
1. Chemical Equilibrium and Factors Affecting Reaction Rates; Le Châtelier's Principle
a. Explain the collision theory of reactions
b. Analyze factors (e.g., temperature, nature of reactants) affecting reaction rates in relation to the kinetic theory
c. Relate reaction mechanism, rate-determining step, activated complex, heat of reaction, and activation energy to reaction kinetics
d. Interpret potential energy diagrams for chemical reactions
e. Describe the conditions that define equilibrium systems on a dynamic molecular level and on a static macroscopic scale
f. Apply Le Châtelier's principle to explain a variety of changes in physical and chemical equilibria
g. Define K_{sp} and manipulate K_{sp} to predict solubility
h. Explain the law of concentration (mass) action and write equilibrium law expressions for chemical equilibria
i. Determine solubility product constants from solubilities (and vice versa) for a given solubility equilibrium system
2. Mechanism, Rate-Determining Step, Activation Energy, and Catalysts
a. Relate the rate of a chemical reaction to the appearance of products and the disappearance of reactants
b. Describe the meaning of reaction mechanism and rate-determining step
c. Relate collision theory to the factors that affect the rate of reaction
d. Describe the meaning of activation energy and activated complex
e. Interpret and label a plot of energy versus reaction coordinate
f. Explain the effects of catalysts on reaction rates (e.g., mechanism, activation energy/activated complex)
3. Chemical Processes and Heat; Calorimetry
a. Explain the law of conservation of energy in chemical reactions
b. Describe the concept of heat, and explain the difference between heat energy and temperature

© 2010 by ACT, Inc. All rights reserved.

c.	Explain physical and chemical changes as endothermic or exothermic energy changes
d.	Solve heat capacity and heat transfer problems involving specific heat, heat of fusion, and heat of vaporization
e.	Calculate the heat of reaction for a given chemical reaction when given calorimetric data
	4. Enthalpy and Entropy
a.	Define enthalpy and explain how changes in enthalpy determine whether a reaction is endothermic or exothermic
b.	Compute ΔH_{rxn} from $\Delta H_f°$ values and explain why the $\Delta H_f°$ values for elements are zero
c.	Explain and apply, mathematically, the relationship between $\Delta H_{rxn}°$ (forward) and $\Delta H_{rxn}°$ (reverse)
d.	Define entropy and explain the role of entropy in chemical and physical changes, and explain the changes that favor increases in entropy
	C. Salts, Acids, and Bases
	1. Acid/Base Theories
a.	Describe the nature and interactions of acids and bases
b.	Describe the hydronium ion and the concept of amphoterism
c.	Describe Arrhenius and Brønsted-Lowry acids and bases; identify conjugate acids and bases in reactions
d.	Relate solvent interaction to the formation of acidic and basic solutions
e.	Define the water constant, K_w, and the pH scale
f.	Describe characteristics of strong and weak acids and bases, and identify common examples of both
	2. Acid/Base Constants and pH; Titration; Buffers
a.	Write and balance a simple equation for a neutralization reaction
b.	Calculate hydrogen ion concentration, hydroxide ion concentration, pH, and pOH for acidic or basic solutions
c.	Explain how the acid-base indicators work
d.	Define percent ionization, K_a, and K_b and explain how they relate to acid/base strength
e.	Conduct a titration experiment in order to determine the concentration of an acid or base solution
f.	Qualitatively understand the behavior of a buffer and explain why buffer solutions maintain pH upon dilution
	D. REDOX Reactions and Electrochemistry
a.	Define REDOX reaction, oxidation, reduction, oxidizing agent, and reducing agent
b.	Assign oxidation numbers (states) to reaction species; identify the species oxidized and reduced, and the oxidizing agent and reducing agent, in a REDOX reaction
c.	Balance REDOX equations by the ion-electron and half-reaction methods
d.	Diagram and explain the operation of a voltaic cell
e.	Determine the net voltage obtained when standard half-cells are paired to form a voltaic cell, and use this voltage to predict reaction spontaneity
	E. Nuclear Chemistry
a.	Describe alpha, beta, and gamma decay, half-life, and fission and fusion
b.	Write appropriate equations for nuclear decay reactions, using particle balance; describe how the nucleus changes during these reactions and compare the resulting radiation with regard to penetrating ability

© 2010 by ACT, Inc. All rights reserved.

ACT Course Standards Worksheet—Chemistry

This worksheet gives teachers an opportunity to compare their course content to ACT's QualityCore® program. Completing the worksheet also allows teachers who teach the same course to ensure their courses have similar outcomes.

Gap Analysis 1—Individual Teacher Review

This analysis allows individual teachers to identify "gaps" between ACT Course Standards and their course content. They should review the ACT standards on the following worksheet, then determine whether the ACT standard **is** or **is not** included in the course as it is currently taught. "Included" means the standard is taught and students are expected to demonstrate proficiency by the end of the course. "Not Included" means the standard is not taught in the course, is taught in another course, or is already mastered. In the "Gap 1" column on the worksheet, place an "I" for "Included" or an "NI" for "Not Included." Analyze any gaps between the current course standards and the ACT Course Standards. Identify reasons the standards receiving a "Not Included" designation are not included in the course.

Gap Analysis 2—Group Consensus

This analysis allows groups of teachers who teach the same course and who have completed Gap Analysis 1 individually to identify differences in how they evaluated the gaps between ACT Course Standards and current course standards. In the "Gap 2" column of the worksheet, place an "X" where members of the group differed in their assessment of whether a particular ACT standard is included in the course as it is currently taught.

The following questions can guide discussion of the gaps:

Overarching Questions

1. What should students know and be able to do before going to the next course?
2. Do all teachers teaching this course have a shared understanding of the intent or meaning of each course standard and topic area?

Gap Analysis 1 Questions

1. Which ACT Course Standards were identified as not included in the course?
2. What is the level of agreement among the group of teachers about the skills and knowledge that is or is not taught in the course?
3. Are there sound pedagogical reasons for not including specific ACT standards in the course?
4. What implications will any decisions have on students' future learning and academic achievement?

Gap Analysis 2 Questions

1. Which of the ACT Course Standards elicited differences of opinion?
2. What are the possible reasons for different opinions about the standards that are or are not included in the course?
3. Are there sound pedagogical reasons for including or not including these disputed standards in the course?
4. What implications will any decisions have on students' future learning and academic achievement?

© 2010 by ACT, Inc. All rights reserved.

Finally, document the necessary steps to address the outcomes of the discussion. Be sure to note whether course standards will be added, deleted, or modified; identify who will be responsible for communicating any changes to other teachers; and note any other decisions. Document responsibilities and establish a timetable for continuing the discussion and implementing the decisions.

NOTE: This course content review is most effective as a continuous process that generates feedback throughout the year. ACT recommends, at minimum, monthly status update meetings for teachers and departments involved in the review.

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
I. UNDERSTANDING CHEMISTRY AS INQUIRY (Note: Some of the process standards in this section are similar to those found in Biology and Physics.)			
A. Foundations			
1. Scientific Inquiry			
a. Identify and clarify research questions and design experiments			
b. Design experiments so that variables are controlled and appropriate numbers of trials are used			
c. Collect, organize, and analyze data accurately and use techniques and equipment appropriately			
d. Interpret results and draw conclusions, revising hypotheses as necessary and/or formulating additional questions or explanations			
e. Write and speak effectively to present and explain scientific results, using appropriate terminology and graphics			
f. Safely use laboratory equipment and techniques when conducting scientific investigations			
g. Routinely make predictions and estimations			
2. Mathematics and Measurement in Science			
a. Distinguish between precision and accuracy with respect to experimental data			
b. Use appropriate SI units for length, mass, time, temperature, quantity of matter, area, volume, and density; describe the relationships among SI unit prefixes (e.g., centi-, milli-, kilo-); recognize commonly used non-SI units			
c. Use the correct number of significant figures in reporting measurements and the results of calculations			
d. Use appropriate statistical methods to represent the results of investigations			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
e. Express numbers in scientific notation when appropriate			
f. Solve for unknown quantities by manipulating variables			
g. Use graphical, mathematical, and/or statistical models to express patterns and relationships inferred from sets of scientific data			
3. Science in Practice			
a. Explain and apply criteria that scientists use to evaluate the validity of scientific claims and theories			
b. Explain why experimental replication and peer review are essential to eliminate as much error and bias as possible in scientific claims			
c. Explain the criteria that explanations must meet to be considered scientific (e.g., be consistent with experimental/observational evidence about nature, be open to critique and modification, use ethical reporting methods and procedures)			
d. Explain why all scientific knowledge is subject to change as new evidence becomes available to the scientific community			
e. Use a variety of appropriate sources (e.g., Internet, scientific journals) to retrieve relevant information; cite references properly			
f. Identify and analyze the advantages and disadvantages of widespread use of and reliance on technology			
g. Compare the scientific definitions of fact, law, and theory, and give examples of each in chemistry			
II. EXPLORING THE PHYSICAL WORLD			
A. Introduction to Chemistry			
1. Mass, Volume, and Density			
a. Explain why mass is used as a quantity of matter and differentiate between mass and weight			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
b. Explain density qualitatively and solve density problems by applying an understanding of the concept of density			
2. Elements, Atomic Mass, and Nomenclature			
a. Use the IUPAC symbols of the most commonly referenced elements			
b. Compare the characteristics of elements, compounds, and mixtures			
c. Compare characteristics of isotopes of the same element			
B. Properties of Matter and Gases			
1. Phases of Matter, Phase Changes, and Physical Changes			
a. Compare the definition of matter and energy and the laws of conservation of matter and energy			
b. Describe how matter is classified by state of matter and by composition			
c. Describe the phase and energy changes associated with boiling/condensing, melting/freezing, sublimation, and crystallization (deposition)			
d. Explain the difference between chemical and physical changes and demonstrate how these changes can be used to separate mixtures and compounds into their components			
e. Define chemical and physical properties and compare them by providing examples			
2. The Nature of Gases			
a. Define gas pressure and the various pressure units (e.g., torr, kilopascals, mm Hg, atmospheres)			
b. Describe the use and operation of mercury barometers and manometers to find atmospheric pressure or relative gas pressures			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
c. Define the gas laws given by Boyle, Charles, Gay-Lussac, and Dalton and solve problems based on these laws			
d. Predict boiling point changes based on changes in atmospheric pressure			
e. Explain the basis for gaseous diffusion and effusion			
f. Describe Avogadro's hypothesis and use it to solve stoichiometric problems			
3. Ideal Gas Law			
a. Explain the difference between an ideal and real gas, the assumptions made about an ideal gas, and what conditions favor ideal behavior for a real gas			
b. Apply the mathematical relationships that exist among the volume, temperature, pressure, and number of particles in an ideal gas			
c. Compute gas density when given molar mass, temperature, and pressure			
d. Apply the ideal gas law to determine the molar mass of a volatile compound			
e. Solve gas stoichiometry problems at standard and nonstandard conditions			
III. DISCOVERING THE LANGUAGE OF CHEMISTRY			
A. Formulas and Equations			
1. Empirical Formulas, Molecular Formulas, and Percentage Composition			
a. Distinguish between chemical symbols, empirical formulas, molecular formulas, and structural formulas			
b. Interpret the information conveyed by chemical formulas for numbers of atoms of each element represented			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
c. Use the names, formulas, and charges of commonly referenced polyatomic ions			
d. Provide the interconversion of molecular formulas, structural formulas, and names, including common binary and ternary acids			
e. Calculate the percent composition of a substance, given its formula or masses of each component element in a sample			
f. Determine the empirical formulas and molecular formulas of compounds, given percent composition data or mass composition data			
g. Determine percent composition experimentally and derive empirical formulas from the data (e.g., for hydrates)			
2. Mole Concept, Molar Mass, Gram Formula Mass, and Molecular Mass			
a. Explain the meaning of mole and Avogadro's number			
b. Interconvert between mass, moles, and number of particles			
c. Distinguish between formula mass, empirical mass, molecular mass, gram molecular mass, and gram formula mass			
3. Chemical Equations and Stoichiometry			
a. Explain how conservation laws form the basis for balancing chemical reactions and know what quantities are conserved in physical, chemical, and nuclear changes			
b. Write and balance chemical equations, given the names of reactants and products			
c. Describe what is represented, on a molecular and molar level, by chemical equations			
d. Use the appropriate symbols for state (i.e., solid, liquid, gaseous, aqueous) and reaction direction when writing chemical equations			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
e. Classify chemical reactions as being synthesis, decomposition, single replacement, or double replacement reactions			
f. Predict the products of synthesis, combustion, and decomposition reactions and write balanced equations for these reactions			
g. Predict products of single replacement reactions, using the activity series, and write balanced equations for these reactions			
h. Predict the products of double replacement reactions, using solubility charts to identify precipitates, and write balanced equations for these reactions			
i. Use chemical equations to perform basic mole-mole, mass-mass, and mass-mole computations for chemical reactions			
j. Identify limiting reagents and use this information when solving reaction stoichiometry problems			
k. Compute theoretical yield, actual (experimental) yield, and percent yield			
l. Calculate percent error and analyze experimental errors that affect percent error			
m. Write ionic equations, identifying spectator ions and the net ionic equation			
IV. BUILDING MODELS OF MATTER			
A. Microscopic Nature of Matter			
1. Structure of Liquids and Solids			
a. Describe differences between solids, liquids, and gases at the atomic and molecular levels			
b. Describe and perform common separation techniques (e.g., filtration, distillation, chromatography)			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
2. Kinetic Molecular Theory of Gases			
a. Use the kinetic molecular theory to explain the states and properties (i.e., microscopic and macroscopic) of matter and phase changes			
b. Explain the basis and importance of the absolute temperature scale and convert between the Kelvin and Celsius scales			
c. Use the kinetic-molecular theory as a basis for explaining gas pressure, Avogadro's hypothesis, and Boyle's/Charles's laws			
B. Atomic Structure and Chemical Bonding			
1. Atomic Theory (Dalton), Atomic Structure, and Quantum Theory			
a. Describe the importance of models for the study of atomic structure			
b. Describe the crucial contributions of scientists and the critical experiments that led to the development of the modern atomic model			
c. Describe characteristics of a wave, such as wavelength, frequency, energy, and speed			
d. Describe the role of probability in orbital theory			
e. Describe atomic orbitals (s, p, d, f) and their basic shapes			
f. Apply Hund's rule and the Aufbau process to specify the electron configurations of the elements			
2. Periodic Table and Periodicity			
a. Describe the historical development of the modern periodic table, including work by Mendeleev and then Moseley			
b. Describe and explain the organization of elements into periods and groups in the periodic table			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
c. Use the periodic table to determine the atomic number; atomic mass; mass number; and number of protons, electrons, and neutrons in isotopes of elements			
d. Calculate the weighted average atomic mass of an element from isotopic abundance, given the atomic mass of each contributor			
e. Identify regions (e.g., groups, families, series) of the periodic table and describe the chemical characteristics of each			
f. Compare the periodic properties of the elements (e.g., metal/nonmetal/metalloid behavior, electrical/heat conductivity, electronegativity and electron affinity, ionization energy, atomic/covalent/ionic radius) and how they relate to position in the periodic table			
g. Use the periodic table to predict and explain the valence electron configurations of the elements, to identify members of configuration families, and to predict the common valences of the elements			
3. Intermolecular Forces and Types of Bonds			
a. Describe the characteristics of ionic and covalent bonding			
b. Explain ionic stability, recognize typical ionic configurations, and predict ionic configurations for elements (e.g., electron configurations, Lewis dot models)			
c. Describe the nature of the chemical bond with respect to valence electrons in bonding atoms			
d. Explain how ionic and covalent compounds differ			
e. Describe the unique features of bonding in carbon compounds			
f. Compare the different types of intermolecular forces (e.g., van der Waals, dispersion)			
g. Explain and provide examples for dipole moments, bond polarity, and hydrogen bonding			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
h. Describe the unique physical and chemical properties of water resulting from hydrogen bonding			
i. Explain the relationship between evaporation, vapor pressure, molecular kinetic energy, and boiling point for a single pure substance			
j. Explain the relationship between intermolecular forces, boiling points, and vapor pressure when comparing differences in the properties of pure substances			
k. Classify solids as ionic, molecular, metallic, or network			
4. Orbital Theory Applied to Bonding			
a. Use Lewis dot diagrams to represent bonding in ionic and covalent compounds			
b. Draw Lewis structures for molecules and polyatomic ions, including those that must be represented by a set of resonance structures			
c. Use VSEPR theory to explain geometries of molecules and polyatomic ions			
d. Describe how orbital hybridization models relate to molecular geometry			
e. Describe the molecular orbital models for double bonds, triple bonds, and delocalized pi electrons			
f. Describe the relationship between molecular polarity and bond polarity			
V. INTEGRATING THE MACROSCOPIC, MICROSCOPIC, AND SYMBOLIC WORLDS			
A. Solutions			
1. Types of Solutions, Concentration, and Solubility			
a. Define solution, solute, and solvent			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
b. Compare properties of suspensions, colloids, and true solutions			
c. Define the terms *saturated*, *unsaturated*, *supersaturated*, *dilute*, and *concentrated* as they pertain to solutions			
d. Give examples of solid, liquid, or gas medium solutions			
e. Define and calculate the molarity of a solution			
f. Define and calculate the percent composition of a solution			
g. Describe the preparation and properties of solutions			
h. Solve stoichiometry calculations based on reactions involving aqueous solutions			
i. Describe the relationship between temperature or pressure and the solubility of gases in liquids			
j. Describe the relationship between solvent character and solute character and explain miscibility			
k. Apply the general rules of solubility to aqueous salt solutions			
l. Describe the factors affecting the solubility of a solute in a given solvent and its rate of solution			
2. Colligative Properties			
a. Describe qualitatively the effect of adding solute on freezing point, boiling point, and vapor pressure of a solvent			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
b. Define molality and mole fraction			
c. Calculate changes in the boiling point and freezing point when nonvolatile, nonelectrolyte solutes are added to solvents			
B. Kinetics, Equilibrium, and Thermodynamics			
1. Chemical Equilibrium and Factors Affecting Reaction Rates; Le Châtelier’s Principle			
a. Explain the collision theory of reactions			
b. Analyze factors (e.g., temperature, nature of reactants) affecting reaction rates in relation to the kinetic theory			
c. Relate reaction mechanism, rate-determining step, activated complex, heat of reaction, and activation energy to reaction kinetics			
d. Interpret potential energy diagrams for chemical reactions			
e. Describe the conditions that define equilibrium systems on a dynamic molecular level and on a static macroscopic scale			
f. Apply Le Châtelier’s principle to explain a variety of changes in physical and chemical equilibria			
g. Define K_{sp} and manipulate K_{sp} to predict solubility			
h. Explain the law of concentration (mass) action and write equilibrium law expressions for chemical equilibria			
i. Determine solubility product constants from solubilities (and vice versa) for a given solubility equilibrium system			
2. Mechanism, Rate-Determining Step, Activation Energy, and Catalysts			
a. Relate the rate of a chemical reaction to the appearance of products and the disappearance of reactants			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
b. Describe the meaning of reaction mechanism and rate-determining step			
c. Relate collision theory to the factors that affect the rate of reaction			
d. Describe the meaning of activation energy and activated complex			
e. Interpret and label a plot of energy versus reaction coordinate			
f. Explain the effects of catalysts on reaction rates (e.g., mechanism, activation energy/activated complex)			
3. Chemical Processes and Heat; Calorimetry			
a. Explain the law of conservation of energy in chemical reactions			
b. Describe the concept of heat, and explain the difference between heat energy and temperature			
c. Explain physical and chemical changes as endothermic or exothermic energy changes			
d. Solve heat capacity and heat transfer problems involving specific heat, heat of fusion, and heat of vaporization			
e. Calculate the heat of reaction for a given chemical reaction when given calorimetric data			
4. Enthalpy and Entropy			
a. Define enthalpy and explain how changes in enthalpy determine whether a reaction is endothermic or exothermic			
b. Compute ΔH_{rxn} from ΔH_f^o values and explain why the ΔH_f^o values for elements are zero			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
c. Explain and apply, mathematically, the relationship between $\Delta H_{rxn}°$ (forward) and $\Delta H_{rxn}°$ (reverse)			
d. Define entropy and explain the role of entropy in chemical and physical changes, and explain the changes that favor increases in entropy			
C. Salts, Acids, and Bases			
1. Acid/Base Theories			
a. Describe the nature and interactions of acids and bases			
b. Describe the hydronium ion and the concept of amphoterism			
c. Describe Arrhenius and Brønsted-Lowry acids and bases; identify conjugate acids and bases in reactions			
d. Relate solvent interaction to the formation of acidic and basic solutions			
e. Define the water constant, K_w, and the pH scale			
f. Describe characteristics of strong and weak acids and bases, and identify common examples of both			
2. Acid/Base Constants and pH; Titration; Buffers			
a. Write and balance a simple equation for a neutralization reaction			
b. Calculate hydrogen ion concentration, hydroxide ion concentration, pH, and pOH for acidic or basic solutions			
c. Explain how the acid-base indicators work			
d. Define percent ionization, K_a, and K_b and explain how they relate to acid/base strength			

© 2010 by ACT, Inc. All rights reserved.

Chemistry Course Standards	Gap 1	Gap 2	Comments
e. Conduct a titration experiment in order to determine the concentration of an acid or base solution			
f. Qualitatively understand the behavior of a buffer and explain why buffer solutions maintain pH upon dilution			
D. REDOX Reactions and Electrochemistry			
a. Define REDOX reaction, oxidation, reduction, oxidizing agent, and reducing agent			
b. Assign oxidation numbers (states) to reaction species; identify the species oxidized and reduced, and the oxidizing agent and reducing agent, in a REDOX reaction			
c. Balance REDOX equations by the ion-electron and half-reaction methods			
d. Diagram and explain the operation of a voltaic cell			
e. Determine the net voltage obtained when standard half-cells are paired to form a voltaic cell, and use this voltage to predict reaction spontaneity			
E. Nuclear Chemistry			
a. Describe alpha, beta, and gamma decay, half-life, and fission and fusion			
b. Write appropriate equations for nuclear decay reactions, using particle balance; describe how the nucleus changes during these reactions and compare the resulting radiation with regard to penetrating ability			

© 2010 by ACT, Inc. All rights reserved.