Detailed Description of Content		
Goals and objectives	**Content – What are students expected to know and be able to do?**	**A = Recommended Activities** **O = Optional Enrichment**
COMPETENCY GOAL 1: The learner will develop abilities necessary to do and understand scientific inquiry.		
1.01 Design, conduct and analyze investigations to answer questions related to chemistry. • Identify questions and suggest hypotheses. • Identify variables. • Use a control when appropriate. • Select and use appropriate measurement tools. • Collect and organize data in tables, charts and graphs. • Analyze and interpret data. • Explain observations. • Make inferences and predictions. • Explain the relationship between evidence and explanation. • Identify how scientists share findings.	This goal and these objectives are an *integral* part of *each of the other goals*. In order to measure and investigate scientific phenomena, students must be given the opportunity to design and conduct their own investigations in a safe laboratory. The students should use questions and models to formulate the relationship identified in their investigations and then report and share those finding with others **Students will be able to:** • Identify questions and suggest hypotheses. • Identify variables. • Use a control when appropriate. • Select and use appropriate measurement tools. • Collect and organize data in tables, charts and graphs. • Analyze and interpret data. • Explain observations. • Make inferences and predictions. • Use questions and models to determine the relationships between variables in investigations. • Identify how scientists share findings.	Activities and Labs listed below are **suggested** resources only. This list is not inclusive and only reflects labs and activities recommended by the Chemistry Curriculum Committee. Teachers should substitute other labs appropriate to the equipment available in their school. Effort should be made in each course to provide students opportunities to practice the skills of a chemist. **A:** Students should be given numerous opportunities to design and conduct experiments within the context of the entire course.
1.02 Evaluate reports of scientific investigations from an informed scientifically-literate viewpoint including considerations of: • Appropriate sample. • Adequacy of experimental controls.	**Students will be able to:** Analyze reports of scientific investigations from an informed scientifically-literate viewpoint including considerations of: appropriate sample, adequacy of experimental controls, replication of findings,	**A:** Integrate follow-up questions for lab activities and experiment summaries which focus on experiment design and appropriate conclusions.

• Replication of findings. • Alternative interpretations of the data	alternative interpretations of the data.	
1.03 Evaluate experimental designs with regard to safety and use safe procedures in laboratory investigations: • Identify and avoid potential safety hazards given a scenario. • Differentiate between safe and unsafe procedures. • Use information from the MSDS (Material Safety Data Sheets) to assess chemical hazards.	**Students will be able to:** • Identify and avoid potential safety hazards given a scenario. • Differentiate between safe and unsafe procedures. • Use information from the MSDS (Material Safety Data Sheets) to assess chemical hazards.	**A:** Data sets and experimental outcomes should be presented to students for analysis within the context of the entire course. Each lab activity should include safety hazards and MSDS data as appropriate.
COMPETENCY GOAL 2: The learner will build an understanding of the structure and properties of matter. (34%)		
2.01 Analyze the historical development of the current atomic theory. • Early contributions: Democritus and Dalton. • The discovery of the electron: Thomson and Millikan. • The discovery of the nucleus, proton and neutron: Rutherford and Chadwick. • The Bohr model. • The quantum mechanical model.	**Students should be able to:** • Describe the composition of the atom and the experiments that led to that knowledge. • Describe how Rutherford predicted the nucleus. • Understand the inverse relationship between wavelength and frequency, and the direct relationship between energy and frequency. • Analyze diagrams related to the Bohr model of the hydrogen atom in terms of allowed, discrete energy levels in the emission spectrum. • Describe the electron cloud of the atom in terms of a probability model.	**A: Animations from www.dlt.ncssm.edu/TIGER** (Click on chemistry by topic and go to atomic structure.) **O:** Calculations of energies and wavelengths in Bohr atom using $c=f\lambda$, $E=hc/\lambda$, and $E=-R_h(1/n^2)$.
2.02 Examine the nature of atomic structure • Subatomic particles: protons, neutrons, and electrons. • Mass number. • Atomic number.	**Students should be able to:** • Characterize the protons, neutrons, electrons: location, relative charge, relative mass (p=1, n=1, e=1/2000). • Use symbols: A= mass number, Z=atomic number	**A: The Atom Activity** "What is an Atom?" curriculum support activity **O:** Calculating average atomic mass of atoms from relative abundance (%) and

• Isotopes.	• Use notation for writing isotope symbols: $^{235}_{92}U$ or U-235 • Identify isotope using mass number and atomic number and relate to number of protons, neutrons and electrons. • Have a conceptual awareness of the nature of average atomic mass. (Relative abundance of each isotope determines the average- no calculations).	actual isotopic mass. Mass defect and $E=mc^2$
2.03 Apply the language and symbols of chemistry. • Name compounds using the IUPAC conventions. • Write formulas of simple compounds from their names.	**Students should be able to:** • Use the state of matter symbols: (s), (l), (g), (aq) • Write binary compounds of two nonmetals: use Greek prefixes (di-, tri-, tetra-, …) • Write binary compounds of metal/nonmetal* • Write ternary compounds (polyatomic ions)* • Write, with charges, these polyatomic ions: nitrate, sulfate, carbonate, acetate, and ammonium. • Know names and formulas for these common laboratory acids: HCl, HNO_3, H_2SO_4, $HC_2H_3O_2$, (CH_3COOH) **The Stock system is the correct IUPAC convention for inorganic nomenclature.*	**O:** "-ic/-ous" method for copper, iron, manganese, mercury, tin, etc. "-ite/-ate", "hypo-/per-" with names of polyatomic ions Names and formulas for other acids Organic nomenclature
2.04 Identify substances using their physical properties: • Melting points. • Boiling points. • Density. • Solubility.	**Students should be able to:** • Apply information (BP, MP, density) from the reference tables to identify an unknown. • Calculate density. (D=m/V) • Apply the solubility rules. • Use graph of solubility vs. temperature to identify a substance based on solubility at a particular temperature. Use graph to relate the degree of saturation of solutions to temperature. Use graph	**A**: **Density Lab** curriculum support activity. Density Activity Unknown Liquid Lab

	to make simple calculations about solutions.	

2.05 Analyze the basic assumptions of kinetic molecular theory and its applications:
• Ideal Gas Equation.
• Combined Gas Law.
• Dalton's Law of Partial Pressures.
	Students should be able to:
	• Know characteristics of ideal gases
	• Apply general gas solubility characteristics
	Students should be able to use the following formulas and concepts of kinetic molecular theory.
	• 1 mole of any gas at STP=22.4 L
	• Ideal gas equation (PV=nRT),
	• Combined gas law ($P_1V_1/T_1 = P_2V_2/T_2$) and applications holding one variable constant
	  • (PV=k), $P_1V_1 = P_2V_2$ *Boyle's Law*
	  • (V/T=k), $V_1/T_1= V_2/T_2$ *Charles' Law*
	  • (P/T=k), $P_1/T_1 = P_2/T_2$ *Gay-Lussac's Law*
	(Note: Students should be able to derive and use these gas laws, but are not necessarily expected to memorize their names.)
	• Avogadro's Law (n/V=k), $n_1/V_1 = n_1/V_2$
	• Dalton's Law (P_t=P_1+P_2+P_3 …)
	• Vapor pressure of water as a function of temperature (conceptually)
		A: **Boyle's Law Lab** uses CBL/LabPro Technology
		Dalton's Law Demonstration
		Mg-HCl - Drying a Gas
		
		O: Graham's Law
		Calculate MW from effusion of gases
		Calculation of KE of gas molecules
		Distribution of speeds as a function of temperature
		Real gases and the van der Waals equation
		Differentiate between ideal gas and real gas (conceptually, no calculations)
		Density/MW variation of Ideal Gas equation calculations

2.06 Assess bonding in metals and ionic compounds as related to chemical and physical properties.	**Students should be able to:** • Describe how ions are formed and which arrangements are stable (filled d-level, or half-filled d-level). • Appropriately use the term cation as a positively charged ion and anion as negatively charged ion. • Predict ionic charges for representative elements based on valence electrons. • Describe ionic bond's intermolecular attraction as electrostatic attraction. • Determine that a bond is predominately ionic by the location of the atoms on the Periodic Table (metals combined with nonmetals) or when ΔEN	**A: Bonding- Type Triangle** **O:** Coulomb's Law ($F=kq_1q_2/r^2$) Crystal shapes Lattice energies

	> 1.7. • Explain how ionic bonding in compounds determines their characteristics: high MP, high BP, brittle, and high electrical conductivity either in molten state or in aqueous solution. • Explain how covalent bonding in compounds determines their characteristics: low MP, low BP, poor electrical conductivity, polar nature, etc. • Describe metallic bonds: "metal ions plus 'sea' of mobile electrons". • Explain how metallic bonding determines the characteristics of metals: high MP, high BP, high conductivity, malleability, ductility, and luster.	
2.07 Assess covalent bonding in molecular compounds as related to molecular geometry and chemical and physical properties. • Molecular. • Macromolecular. • Hydrogen bonding and other intermolecular forces (dipole/dipole interaction, dispersion). • VSEPR theory.	**Students should be able to:** • Apply the concept that sharing electrons form a covalent compound that is a stable (inert gas) arrangement. • Determine that a bond is predominately covalent by the location of the atoms on the Periodic Table (nonmetals combined with nonmetals) or when $\Delta EN < 1.7$. • Know that the diatomic elements have single, double, or triple bonds (For example: F_2, O_2, N_2). • Describe carbon bonds as either single, double or triple bonds. • Apply the relationship between bond energy and length of single, double, and triple bonds (conceptual, no numbers). • Draw Lewis (dot diagram) structures for simple compounds with one central atom. • Apply Valence Shell Electron Pair Repulsion Theory (VSEPR) for these electron pair	A: Molecular and Intermolecular Bonds Lab **Bond-Type Triangle Activity** **O:** Molecular geometries (VSEPR) expanded octet Valence bond theory: hybrid orbitals sigma and pi bonds Molecular orbital theory Formal charge calculations Additional macromolecules Dipole moment – conceptual

	geometries and molecular geometries, and bond angles. • Electron pair - Molecular (bond angle) • Linear framework – linear • Trigonal planar framework– trigonal planar, bent • Tetrahedral framework– tetrahedral, trigonal pyramidal, bent • Bond angles (include distorting effect of lone pair electrons – no specific angles, conceptually only) • Describe bond polarity. Polar/nonpolar molecules (relate to symmetry) ; relate polarity to solubility—"like dissolves like" • Describe macromolecules and network solids: water (ice), graphite/diamond, polymers (PVC, nylon), proteins (hair, DNA) intermolecular structure as a class of molecules with unique properties. • Describe intermolecular forces for molecular compounds. • H-bond as attraction between molecules when H is bonded to O, N, or F. Dipole-dipole attractions between polar molecules. • London dispersion forces (electrons of one molecule attracted to nucleus of another molecule) – i.e. liquefied inert gases. • Relative strengths (H>dipole>London/van der Waals).	
2.08 Assess the dynamics of physical equilibria. • Interpret phase diagrams. • Factors that affect phase changes.	**Students should be able to:** • Describe physical equilibrium: liquid water-water vapor. Vapor pressure depends on temperature and concentration of particles in	A: **Heat of Fusion** "Ice Cream Lab" **O:** Calculations with Raoult's law

	solution. (conceptual only – no calculations) • Draw phase diagrams of water and carbon dioxide (shows how sublimation occurs). Identify regions, phases and phase changes using a phase diagram. • Know that phase changes occur with changes in temperature and/or pressure. Relate change of phase to heating and cooling curves.	
COMPETENCY GOAL 3: The learner will build an understanding of regularities in chemistry.		
3.01 Analyze periodic trends in chemical properties and use the periodic table to predict properties of elements. • Groups (families). • Periods. • Representative elements (main group) and transition elements. • Electron configuration and energy levels. • Ionization energy. • Atomic and ionic radii. • Electronegativity.	**Using the Periodic Table, students should be able to:** **Groups (families)** • Identify groups as vertical columns on the periodic table. • Know that main group elements in the same group have similar properties, the same number of valence electrons, and the same oxidation number. • Understand that reactivity increases as you go down within a group for metals and decreases for nonmetals. **Periods** • Identify periods as horizontal rows on the periodic table. **Representative elements (main group) and transition elements** • Identify representative (main group) elements as A groups <u>or</u> as groups 1, 2, 13-18. • Identify alkali metals, alkaline earth metals,	**A: "Periodic Table Mystery"** **"Building and Using a 3-D Periodic Table"** **Photoelectron Spectroscopy activity** **O:** Electron affinity Examine shapes of orbitals

	halogens, and noble gases based on location on periodic table. • Identify transition elements as B groups <u>or</u> as groups 3-12. **Electron configuration and energy levels** • Write electron configurations, including noble gas abbreviations (no exceptions to the general rules). Included here are extended arrangements showing electrons in orbitals. • Identify s, p, d, and f blocks on Periodic Table. • Identify an element based on its electron configuration. (Students should be able to identify elements which follow the general rules, not necessarily those which are exceptions.) • Determine the number of valence electrons from electron configurations. • Predict the number of electrons lost or gained and the oxidation number based on the electron configuration of an atom. **PERIODIC TRENDS, including…** **Ionization energy** • Define ionization energy. • Know group and period general trends for ionization energy. • Apply trends to arrange elements in order of increasing or decreasing ionization energy. • Explain the reasoning behind the trend. **Atomic and ionic radii** • Define atomic radius and ionic radius. • Know group and period general trends for atomic radius.	

	• Apply trends to arrange elements in order of increasing or decreasing atomic radius. • Compare cation radius to neutral atom. Compare anion radius to neutral atom. • Explain the reasoning behind the trend. **Electronegativity** • Define electronegativity. • Know group and period general trends for electronegativity. • Apply trends to arrange elements in order of increasing or decreasing electronegativity. • Explain the reasoning behind the trend. • Use differences in electronegativity to predict bond type (ionic, polar covalent, and nonpolar covalent).	
3.02 Apply the mole concept, Avogadro's number and conversion factors to chemical calculations: • Particles to moles. • Mass to moles. • Volume of a gas to moles. • Molarity of solutions. • Empirical and molecular formula. • Percent composition.	**Students should be able to:** **Calculate formula mass.** **Particles to moles** • Convert representative particles to moles and moles to representative particles. (Representative particles are atoms, molecules, formula units, and ions.) **Mass to moles** • Convert mass of atoms, molecules, and compounds to moles and moles of atoms, molecules, and compounds to mass. • Convert representative particles to mass and mass to representative particles. **Volume of a gas to moles**	**A: Observations Copper Chloride and Aluminum Lab** **O:** Molality Write formulas for and name hydrates.

	• Convert moles to volume and volume to moles at STP. **Molarity of a solution** • Calculate molarity given mass of solute and volume of solution. • Calculate mass of solute needed to create a solution of a given molarity and volume. • Solve dilution problems: $M_1V_1 = M_2V_2$. **Empirical and molecular formula** • Calculate empirical formula from mass or percent using experimental data. • Calculate molecular formula from empirical formula given molecular weight. **Percent composition** • Determine percentage composition by mass of a given compound. • Perform calculations based on percent composition. • Calculate using hydrates.	
3.03 Calculate quantitative relationships in chemical reactions (stoichiometry): • Moles of each species in a reaction. • Mass of each species in a reaction. • Volumes of gaseous species in a reaction.	**Students should be able to:** • Interpret coefficients of a balanced equation as mole ratios. • Use mole ratios from the balanced equation to calculate the quantity of one substance in a reaction given the quantity of another substance in the reaction. (given moles, particles, mass, or volume and ending with moles, particles, mass, or volume of the desired substance)	**A:** Investigation involving quantitative mole relationships **Construction of Air Bags** **Observations: Copper Chloride and Aluminum** **O:** Limiting reactant problems Percent yield

COMPETENCY GOAL 4: The learner will build an understanding of energy changes in chemistry.		

4.01 Analyze the Bohr model in terms of electron energies in the hydrogen atom.
• The spectrum of electromagnetic energy.
• Emission and absorption of electromagnetic energy as electrons change energy levels.
	A student should be able to:
	• Understand that energy exists in discrete units called quanta.
	• Describe the concepts of
	  o an atom being "excited" above its ground state by the addition of energy, resulting in the electron(s) moving to a higher energy level.
	  o when the atom returns to its ground state, the electron(s) releases that energy gained as electromagnetic radiation (emissions spectrum).
	• Articulate that this electromagnetic radiation is given off as a photon(s). This photon represents the physical difference between ground state and excited state.
	• Use the "Bohr Model for Hydrogen Atom" and "Electromagnetic Spectrum" diagrams from the Reference Tables to relate color, frequency, and wavelength of the light emitted to the energy of the photon.
		A: Spectrum Lab
		
		O: Calculations of wavelengths and energies in Bohr atom using $c=f\lambda$, $E=hc/\lambda$, and $E=-R_h(1/n^2)$

	• Explain that Niels Bohr produced a model of the hydrogen atom based on experimental observations. This model indicated that:
	  o an electron circles the nucleus only in fixed energy ranges called orbits;
	  o an electron can neither gain or lose energy inside this orbit, but could move up or down to another orbit;
	  o and that the lowest energy orbit is closest to the nucleus.
	• Recognize the historical contribution that this model gave to our modern theory of the structure of the atom; however, also realize the limitations of this model (applicable only to the hydrogen atom).
	• Describe the wave/particle duality of electrons.	

4.02 Analyze the law of conservation of energy, energy transformation, and various forms of energy involved in chemical and physical processes.
• Differentiate between heat and temperature.
• Analyze heating and cooling curves.
• Calorimetry, heat of fusion and heat of vaporization calculations.
• Endothermic and exothermic processes including interpretation of potential energy.
• Diagrams (energy vs reaction pathway), enthalpy and activation energy.
	A student should be able to:
	• Recognize that, for a closed system, energy is neither lost nor gained during normal chemical activity.
	• Explain that the total useful energy of an open system is constantly declining due to entropy.
	• Define and use the terms and/or symbols for: enthalpy, entropy, specific heat capacity, temperature, joule, endothermic reactions, exothermic reactions, and catalyst.
	• Interpret the following:
	  o heating and cooling curves (noting both significance of plateaus and the physical states of each segment
	  o Phase diagrams for H_2O and CO_2,
		A: **Heat of Solution Lab**
		
		O: Hess's Law calculations (multistep)
		Heats of formation
		Stoichiometric calculations with heat

	- Energy vs reaction pathway diagrams for both positive and negative values of ΔH (including activation energy). - Complete calculations of: $q=mC_p\Delta T$, $q = mH_f$, $q = mH_v$, and $q_{lost}=(-q_{gain})$ in water, including phase changes. - Contrast heat and temperature, including temperature as a measure of average kinetic energy, and appropriately use the units Joule, Celsius, and Kelvin.	
4.03 Analyze the relationship between entropy and disorder in the universe.	**A student should:** - Understand entropy as a measure of disorder. - Recognize that the entropy of the universe is increasing. - Explain that, along with a tendency for systems to proceed toward the lowest energy level, they also move in the direction of the greatest entropy. (Increasing Entropy: solid → liquid → gas; Ionic compounds → ions in solution)	**A: Phase Change Demonstration** **O:** Understand Gibbs free energy and how it is used to predict spontaneity. calculations with $\Delta G=\Delta H-T\Delta s$
4.04 Analyze nuclear energy. - Radioactivity: characteristics of alpha, beta and gamma radiation. - Decay equations for alpha and beta emission. - Half-life. - Fission and fusion.	**A student should be able to:** - Use the symbols for and distinguish between alpha (${}_2^4He$), and beta (${}_{-1}^0e$) nuclear particles, and gamma (γ) radiation include relative mass). - Use shorthand notation of particles involved in nuclear equations to balance and solve for unknowns. Example: The neutron is represented as (${}_0^1n$) . - Discuss the penetrating ability of alpha, beta, and gamma radiation. - Conceptually describe nuclear decay,	**A: "Penny's Lab or M & M Lab"** **O:** Calculations with $A=A_oe^{-kt}$ Complex calculations with half-life O: Transmutation

	including: ○ Decay as a random event, independent of other energy influences ○ Using symbols to represent simple balanced decay equations ○ Half-life (including simple calculations) • Contrast fission and fusion. • Cite illustrations of the uses of nuclear energy, including, but not limited to: electricity, Carbon-14 dating, and radioisotopes for medicine (tracers, ionizing radiation, gamma sterilization, etc).	
COMPETENCY GOAL 5: The learner will develop an understanding of chemical reactions.		
5.01 Evaluate various types of chemical reactions • Analyze reactions by types: single replacement, double replacement (including acid-base neutralization), decomposition, synthesis, and combustion of simple hydrocarbons. • Predict products.	**Students should be able to:** • Identify a reaction by type. • Predict product(s) in a reaction using the reference tables. • Identify acid-base neutralization as double replacement. • Write and balance ionic equations. • Write and balance net ionic equations for double replacement reactions. • Recognize that hydrocarbons (C,H molecule) and other molecules containing C, H, and O burn completely in oxygen to produce CO_2 and water vapor. • Use reference table rules to predict products for all types of reactions to show the conservation of mass.	**A: "Reaction Types Demonstration"** **" Solubility Rule Activity"**

	• Use activity series to predict whether a single replacement reaction will take place. • Use the solubility rules to determine the precipitate in a double replacement reaction if a reaction occurs.	
5.02 Evaluate the law of conservation of matter to the balancing of chemical equations. • Write and balance formulas and equations • Write net ionic equations**.**	**Students should be able to:** • Write and balance chemical equations. • Write net ionic reactions. • Predict and write formulas using the reference tables.	**O:** Balancing REDOX reactions by half-reaction method or electron transfer method. **A: "Solubility Rule Activity" write equations for reactions observed**
5.03 Identify and predict the indicators of chemical change: • Formation of a precipitate. • Evolution of a gas. • Color change. • Absorption or release of heat.	**Students should be able to determine if a chemical reaction has occurred based on the following criteria:** • Precipitate Ttie to solubility rules (Goals 2.04 and 5.01). • Product testing - Know the tests for some common products such as oxygen, water, hydrogen and carbon dioxide. (tests to know: burning splint for Oxygen, Hydrogen and Carbon Dioxide (include knowledge of safety precautions) lime water for Carbon Dioxide). • Color Change – Distinguish between color change as a result of chemical reaction, and a change in color intensity as a result of dilution. • Temperature change – Tie to endothermic/exothermic reaction. Express ΔH as (+) for endothermic and (–) for exothermic.	**O:** Calculate heat of reaction - ΔH **A: Decomposition of $NaHCO_3$** **Test for reaction products**
5.04 Assess the physical and chemical properties of acids and bases.	**Students should be able to** • Distinguish between acids and bases based on	**O:** Acid-base equilibria; Ka, Kb, Kw Lewis theory.

• General properties of acids and bases. • Concentration and dilution of acids and bases. • Ionization and the degree of dissociation (strengths) of acids and bases. • Indicators. • Acid-base titration. • pH and pOH.	formula and chemical properties. • Distinguish between Arrhenius acids and bases and Bronsted-Lowry acids and bases. • Compute concentration (molarity) of acids and bases in moles per liter (3.02). • Solve dilution problems: $M_1V_1 = M_2V_2$. • Differentiate between concentration (molarity) and strength (degree of dissociation). No calculation involved. • Use pH scale to identify acids and bases. • Interpret pH scale in terms of the exponential nature of pH values in terms of concentrations. • Relate the color of indicator to pH using pH ranges provided in a table. Range should involve various values of pH (for example: 3.3 or 10.8). • Determine the concentration of an acid or base using titration. Interpret titration curve for strong acid/strong base. • Compute pH, pOH, $[H^+]$, and $[OH^-]$. Calculations will involve only whole number values (for example: pH or pOH values such as 3, 5, 8. and $[H^+]$ and $[OH^-]$ values such as $1x10^{-4}$ or $1x10^{-10}$).	Weak acids and weak bases in titrations. Buffer systems Henderson-Hasselbalch equation **A:** an Acid Base Titration Lab
5.05 Analyze oxidation/reduction reactions with regard to the transfer of electrons. • Assign oxidation numbers to elements in REDOX reactions • Identify the elements oxidized and reduced. • Write simple half reactions. • Assess the practical applications of oxidation and reduction reactions.	**Students should be able to:** • Determine oxidation number of each element in a REDOX reaction, including peroxides. • Determine elements oxidized and reduced. • Write half reactions indicating gain or loss of electrons and identify the reaction as either reduction or oxidation. • Students should be aware of some practical applications of oxidation/reduction reactions.	**O:** Oxidizing agents, reducing agents Balancing REDOX reactions by half-reaction method or electron transfer method Determine anode and cathode Cell potential calculations **A:** “Electrochemistry Law’

	Some examples include: simple wet cell, dry cell, bleaching, and electroplating.	
5.06 Analyze the factors that affect the rates of chemical reactions. • The nature of the reactants. • Temperature. • Concentration. • Surface area. • Catalyst.	**Students should be able to:** • Explain collision theory – molecules must collide in order to react, and they must collide in the correct or appropriate orientation and with sufficient energy to equal or exceed the activation energy. (Goal 4.02) • Understand qualitatively that reaction rate is proportional to number of effective collisions. • Explain that nature of reactants can refer to their complexity and the number of bonds that must be broken and reformed in the course of reaction. • Interpret potential energy diagrams. • Explain how temperature (kinetic energy), concentration, and/or pressure affects number of collisions. • Explain how increased surface area increases number of collisions. • Explain how a catalyst lowers the activation energy, so that at a given temperature, more molecules will have energy equal to or greater than the activation energy.	**O:** Rate from concentration vs. time graphs Order of reaction and graphs used to determine order: [x] vs. time, ln[x] vs. time, 1/[x] vs. time Rate law, rate constant, overall order Determine mechanisms from rate law Rate determining step **A:** “Reaction Rates Lab – Mr. Potato Inquires”