Chemistry	Content	Skills
Unit 1 **Measuring, Calculating, Scientific Notation and Density** **Weeks 1-3**	1. Accuracy is how close a measurement is to the actual value. 2. Precision refers to the consistency of a group of measurements. 3. Significant digits are the digits that are measured or estimated when making a measurement. 4. The accuracy of a measurement is dependent upon the accuracy of the measuring device that was used to make the measurement. 5. When adding or subtracting numbers in science, the final answer can only be as accurate as the **least** accurate value in the data set. 6. When multiplying and dividing numbers in science, the final answer must have the **same** number of significant figures as the **least** number in the original data set. 7. When using scientific notation, the decimal point in a number is moved so that there is only one digit left of the decimal point. The one digit left of the decimal must be a nonzero digit; each time the decimal is moved is expressed as a power of 10. (Ex. 10^4) 8. Given two of the three variables in the density equation, the third can be calculated.	1. Express measurements with the proper level of accuracy based on the measuring device used to make the measurement. 2. Describe the accuracy of a measuring device that was used to make a measurement when a measurement is given. 3. Correctly apply the four rules of significant figures to identify how many significant digits are present in a given measurement. 4. Complete addition and subtraction problem and express the final answer with the correct level of accuracy. 5. Complete multiplication and division problems and express the final answer with the correct number of significant digits. 6. Be able to express a number in scientific notation or express a number that is given in scientific notation in standard form. 7. Distinguish the difference between accuracy and precision. 8. Manipulate the equation D=M/V and be able to calculate all three variables.
Unit 2 **Dimensional Analysis** **Week 4-6**	1. SI prefixes are used to form decimal multiples and submultiples of SI units and should be used to avoid very large or very small numeric values. 2. The factor-label method is a convenient way to convert from one SI base unit to another a well as any conversion problem.	1. Memorize the SI prefixes and multipliers for Mega to pico 2. Utilize the SI prefixes in combination with meters, liters, and grams 3. Convert any of the prefix units to a different unit

	3. Rules for significant figures must be applied when completing conversion problems.	4. Use the factor-label method to convert units in a logical, written format 5. Apply the dimensional analysis (conversion method) to other concepts, such as density or molecular mass 6. Applying significant figure knowledge from the previous unit in calculating and labeling properly
Unit 3 **Specific Heat Capacity** **Weeks 7-9**	1. Thermal energy of a substance is dependent on the amount of matter (mass), temperature and chemical composition of a substance. 2. The equations $q = (m)(C_p)(\Delta T)$ and $(m_1)(Cp_1)(\Delta T_1) = (m_2)(Cp_2)(\Delta T_2)$ can be used to solve various heat capacity problems. 3. Materials with low heat capacities change temperature with the addition of a minimal amount of energy. 4. Materials with high heat capacities change temperature with the addition of a relatively large amount of energy. 5. The Earth's temperature is influenced by the high heat capacity of water.	1. Determine the amount of heat lost or gained by various materials by way of a heat capacity lab 2. Identify individuals who have contributed to our current atomic theory
Unit 4 **Atomic Structure (Atomic Models/Theory)** **Weeks 10-11** .	1. Atoms are composed of protons, neutrons and electrons. 2. Various scientists have contributed to our current view of the atom. 3. Several experiments are associated with the development of the planetary atomic model. 4. Our view of the atom has evolved as new information has been obtained.	1. Identify the scientists that contributed to the development of the early atomic model as well as the experimental apparatus associated with each scientist. 2. Identify the properties of subatomic particles that make up an atom and state there location within an atom. 3. Describe the evolution of our current view of the atom based on the

		work of various scientists. 4. Create and explain a planetary model of the atom. 5. Analyze emission and absorption spectrums and explain how they correlate to our current view of the atom. 6. Apply the equations $c=\lambda\nu$ and $E = h\nu$ when working the electromagnetic spectrum.

Unit 5 **Atomic Theory (Electrons)** **Weeks 12-13**	1. An equation was developed treating an electron as a wave bringing about the Wave-particle Duality of Nature. 2. The Aufbau Principle, Hund's Rule and the Pauli Exclusion Principle led to the quantum model of the atom. 3. Electrons within atoms are located in energy levels. 4. Energy levels are composed sublevels.. 5. Sublevels are composed of orbitals. 6. Each orbital can hold up to two electrons. 7. The location of an electron within an atom can be shown with the use of electron configurations. 8. An element can be identified with its electron configuration. 9. Valence electrons in an atom can be represented with electron dot diagrams.	1. State how the quantum mechanical model of the atom describes electrons within an atom. 2. State how the Aufbau principle, Pauli exclusion principle and Hund's rule apply to the quantum model. 3. Describe the relationship between electrons and energy levels, sublevels and orbitals within an atom. 4. Describe the shape of the s,p,d and f orbitals. 5. Write the correct electron configuration for any element on the periodic table 6. Identify an element based on it's electron configuration. 7. Draw the correct electron dot structures for all of the elements on the periodic table.

Unit 6 **Atomic Structure** **(Atoms, Ions and Isotopes)** **Weeks 14-15**	1. The atomic number represents the number of protons in an atom. 2. An atom's mass number is the sum of protons and neutrons in the nucleus of the atom. 3. An atom has a neutral charge due to the fact that the #protons = # electrons. 4. An ion is a charged particle where the (# protons ≠ # electrons) 5. An Anion has a negative charge. (#protons < # electrons) 6. A cation has positive charge. (#protons > #electrons) 7. Isotopes have the same number of protons and different number of neutrons. 8. The average atomic mass of an element is an average of the masses of the isotopes of the element.	1. Identify an element when given the elements atomic number. 2. Determine the charge of an ion based on the number of protons and electrons present in the ion. 3. Determine the identity of an ion or the number of electrons present in the ion when given the charge of the ion and/or the mass number of the ion. 4. calculate the charge of an ion when given the atomic number and number of electrons present in the ion. 5. Calculate the mass number of an atom given the atomic number and number of neutrons present in an atom. 6. Calculate the relative atomic mass of an element based on the mass of each of the elements isotopes.

Unit 7 **Periodic Table/Semester Review** **Weeks 16-18**	1. The periodic table is organized based on the properties of the elements. 2. Elements in the same column have similar electron configurations. 3. There are trends on the periodic table in regards to atomic radii, electronegativities, ionization energies and electron affinities. 4. Elements on the periodic table can be classified as metals, nonmetals or metalloids.	1. State how the elements are arranged on the periodic table according to atomic number and similar properties 2. Explain the organization of periodic table according to groups, families, periods, metal, nonmetal and metalloids 3. Identify the names of the groups/families, series/periods on the periodic table 4. Show how the electron configurations and electron dot diagrams of elements in the same group are similar 5. Classify elements based on their electron configurations and properties 6. Demonstrate how the repeating pattern in the electron configurations of elements explains the many trends and properties of the elements 7. Relate how the atomic radii, ionic radii, first ionization energies, electronegativities and phases are supported by atomic theory and bonding

Unit 8 **Intramolecular Chemical Bonding** **(Ionic, Covalent and Metallic)** **Weeks 19-20**	1. Elements combine to form compounds with ionic, polar covalent and nonpolar covalent bonds. 2. Metal atoms combine with metallic bonds. 3. The difference in electronegativities determine the type of bond formed between two atoms. 4. The type of bond determines the properties of a compound. 5. Formulas of compounds formed between two atoms can be determined with the use electron dot diagrams.	1. Determine what type of bond will form between two atoms based on their electronegativities. 2. Distinguish the difference between ionic, polar covalent, pure covalent and metallic bonds and identify properties associated with each type of bond. 3. Recognize that covalent bonding occurs between two nonmetals 4. Recognize that ionic bonding will occur between a nonmetal and a metal 5. Predict how an atom interacts with other atoms based on the atom's electron configuration, electron dot diagrams and determine the chemical formula of the compound most likely to form between the two atoms. 6. Draw electron dot structures to support the visual representation of chemical bonding.
Unit 9 **Molecular Geometry** **(VSPER Theory)** **Weeks 21**	1. The shapes of molecules can be determined based on the total number of valence electrons available from each of the atoms involved in the compound. 2. There are six geometric shapes associated with molecules based on the number of lone pairs of electrons and bonding pair electrons involved in the molecule. 3. Electron dot diagrams are used to determine the number of lone pairs of electrons and bonding pairs of electrons in a compound. 4. Common reference charts are available to determine the shape of covalently bonded compounds.	1. Explain what the acronym VSPER represents. 2. Recognize that covalent bonding occurs between two nonmetals and do not form ions. 3. Draw the Lewis Structure of a compound to determine the number of lone pairs of electrons and bonding pairs of electrons present in a covalent compound. 4. Predict the geometric shape of the molecules based on the number of bonding and lone pair electrons and applying the information from a reference sheet..

Unit 10 **Chemical Formulas** **Weeks 22-23**	1. The algebraic sum of the charges of the ions in a compound must equal zero. 2. All binary compounds end with "ide". 3. Parenthesis must be used when writing chemical formulas for compounds that involve more than one of a particular polyatomic ion. 4. Roman numerals are required when writing chemical formulas of compounds containing multiple oxidation state elements. 5. Greek prefixes must be used when a nonmetal is bonded to a nonmental.	1. Name and write chemical formulas for binary ionic compounds formed between a metal and a nonmetal. 2. Name and write chemical formulas for compounds involving polyatomic ions. 3. Name and write chemical formulas for compounds containing multiple oxidation state elements. 4. Name and write chemical formulas for binary compounds formed between two nonmetals.
Unit 11 **The Mole** **Weeks 24-26**	1. Avogadro's number (constant) is useful for comparing amounts in a more uniform manner. 2. The Periodic Table is in grams/mole and useful when determining masses or moles from masses 3. Two steps are needed to mathematically convert atoms or molecules to grams and visa versa. 4. The empirical formula represents the simplest whole number ratio. 5. The molecular formula represents the actual conmposition of a compound. 6. The percent composition accurately represents the amount of each element present in a compound. 7. Hydrates are substances that include water in their formula. The water is not actually part of the chemical substance and this is reflected in the way the formula is written.	1. Recognize the significance of using the "mole" as a universal unit in order to compare elements having different masses 2. Determine the significance of Avogadro's number, 6.02×10^{23} particles, in comparing particle amounts 3. Understand that particles can be formula units (ionic compounds), molecules (covalent compounds), ions, or particles. 4. Convert molecules to moles and moles to molecules using Avogadro's number 5. Recognize that the Periodic Table can be labeled as grams/mole and that each individual element has a different mass 6. Convert grams to moles and moles to grams using the Periodic Table atomic masses 7. Convert grams to moles and moles to grams for compounds, using molecular masses (added up) of various compounds 8. Convert atoms or molecules to grams in a two step process involving Avogadro's number and the Periodic Table (and visa versa

Unit 12 **Chemical Reactions** **Weeks 27-29**	1. There are five basic types of chemical reactions (synthesis, decomposition, single and double replacement and combustion reactions) 2. Chemical reactions are represented with chemical equations. 3. In a chemical equation, reactants are on the left side of the equation and products re on the right side of the equation. 4. Chemical equations must be balanced in order to satisfy the Law of Conservation of Mass. 5. To balance a chemical equation, only the coefficients may be changed.	9. Differentiate between and locate products and reactants in a chemical equation. 10. Correctly balance a given chemical equation. 11. Identify and classify the five main types of chemical reactions. 12. Predict products and write a complete balanced equation when given the reactants in word form,
Unit 13 **Stoichiometry** **Weeks 30-32**	1. Coefficients in a balanced equation represent the moles of 2. each substance involved in a chemical reaction. 3. Given amounts of substances in chemical reactions must be converted to moles. 4. Molarity is defined as moles of solute per 1000ml (1L) of solution 5. The limiting reactant controls the amount of product that be produced in a chemical reaction. 6. Percent yield is equal to experimental yield divided by the theoretical yield times 100.	1. Calculate the amount of one substance required to react with the amount of another substance in a chemical reaction. 2. Calculate the amount of products produced in a chemical reaction based on the amount of reactants available.. 3. Convert grams, moles, volume or particles of a known substance to grams, moles, volume or particles of an unknown substance. 4. Determine the limiting reactant in a chemical reaction 5. Calculate the percent yield in a chemical reaction by comparing the experimental yield to the theoretical yield.
Unit 14 **Gas Laws/Semester Review** **Weeks 33-36**	1. The relationship between the pressure, volume and temperature of a gas sample can be represented with gas laws. 2. Boyle's Law deals with the pressure and volume of a gas. 3. Charles' Law deals with the volume and temperature of a gas.	1. Define Boyle's, Charles', Gay-Lussac's and the combined gas laws and represent them in equation form. 2. Convert all temperatures to Kelvin when applying the gas laws 3. Recognize that motion theoretically stops at absolute zero 4. Apply the ideal gas law and represent it in equation form

	4. Gay-Lussac's Law deals with the pressure and temperature of a gas. 5. The combined gas law $P_1V_1/T_1 = P_2V_2/T_2$ can be used when there are multiple variable changes occur. 6. The ideal gas law is represented with the equation PV =nRT 7. The ideal gas law can be used to complete stoichiometric calculations	recognizing "R" as the ideal gas constant. 5. Comprehend that the value of "R" varies depending on the units of measure for pressure 6. Apply the gas law equations to solve various gas law problems. 7. Relate the pressure, volume, and temperature to each other. 8. Maintain one variable as constant and quantify the correlation between the remaining two variables. 9. Explain P, V, and T relationships via the kinetic molecular theory (KMT). 10. Relate real world phenomenon examples of gas laws (hot air balloons, automobile tires, basketballs)