Advanced Placement Chemistry

Course Description

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Welcome!
This course is designed to be the equivalent of a general chemistry course usually taken during the first college year. This is an academic, quantitative chemistry course. Chemistry is the study of atoms and molecules and how they interact according to physical laws. Such study is applicable to everyday life, and this will be demonstrated repeatedly throughout the year. Topics of study include matter, states of matter, reactions, descriptive chemistry, and chemical calculations. The class meets 5 days a week for 75 minutes.

The Philosophy
Learning is NOT a spectator sport. The ultimate responsibility for success in learning lies with the student. Learning is a process by which a person seeks to make sense out of the world. The view of the world, or any part of it, held by a person is as individual as fingerprints. The only way we are able to share knowledge is by verbal and written communication. The quality of the communication depends on the quality of the language used. A teacher is a facilitator for learning. The teacher structures experiences which provide the maximum probability that students can make sense out of the material presented.

Roles of the Teacher
Our responsibility is to present experiences that will assist you to make sense out of chemistry. These experiences can be altered, based on your input, to improve your chances of being successful. It is imperative that communications are two ways so that we can remain informed about how you view the concepts we are studying.
It is also our responsibility to design tests, quizzes and lab assessments that fairly evaluate the level of your success. You have the right to know where you stand at all times during the semester and to ask when you do not understand why you were evaluated in a particular manner.

Roles of the Students
You are in control of your success in chemistry. To some extent your success will be a function of your background, but the major factor in your success will be the quality and quantity of time and effort you put into your studies. You must keep in mind that 1) your principal job at this stage in your life is to be a student; and 2) chemistry is not the only course you are taking. You must balance your time such that you maximize success in all courses. We will provide several vehicles to assist you. You must elect to use them.

Study Groups
You are encouraged to form study groups. Much of the research on what helps students understand chemistry suggests that studying in groups, for many students, is more effective than studying alone. Be careful when forming a study group that it does not exceed four students. If six or more students like to work together, break into subgroups of three or four when studying. Do not allow "leaches" in your group. Some people think they can join a group and let others do the work. This practice harms the leach the most, but nobody likes to do another's work.

Help…Office Hours
One of the AP chem teachers is generally available before or after school. Please come with your materials, attempted work, and specific questions.

Classroom Rules and Etiquette
1. WEAR YOUR IDs!!
2. Respect yourselves, your classmates, and your teacher.
3. Students are to complete all assignments before arriving to class and are to bring all required materials (textbooks, pens/pencils, calculator, etc.) to class EVERY SINGLE DAY.
4. Students are to sit in their assigned seats when the bell rings at the start of class and should be in their seats when the bell rings at the end of class. The bell does not dismiss the class, the teacher dismisses the class.
5. Pay attention.
6. All backpacks, purses, bags, etc. must be stowed under the desk or on the floor.
7. Students are expected to NOT eat or drink, litter, sell candy, harm anyone, throw things, use profanity, deface property, play games, tamper with equipment, etc. in the classroom.
8. When the teacher calls for the class's attention, give it to her immediately.
9. Be safe in lab. Any action that compromises the safety of yourself or other students will result in removal from lab and a zero on your assignment.
10. Passes will be given on an as needed basis at the teacher's discretion and will only be written in your student planner in the designated area.

If there are violations of this policy, the consequences will be unpleasant but logical. In general, if students disrupt the teaching and learning process for any reason, the teacher has the right and obligation to eliminate the disruption. The teacher has the right and obligation to TEACH and EVERY student has the right to learn.

Required Materials
- Pen (blue or black) or pencils
- Binder
- Loose leaf paper or a spiral notebook
- Graphing calculator
- Lab notebook (or a binder with lab notebook paper)
- Textbook

Textbook and Other Resources
- Zuhmdahl, Chemistry, 8th Ed. 2010
- 5 Steps to a 5: AP Chemistry by John Moore and Richard Langley
- SPARKNOTES AP Chemistry Power Pack
- The Ultimate Chemical Equations Handbook

Tardies / Absences / Make-up Work
Tardies: *BE ON TIME*!!! You will receive 1 warning, and the next tardy will result in a detention
Absences: You are responsible for asking about what you missed and getting the materials that were handed out in class. These materials will be in a folder labeled with the day of the week it was handed out. If you are absent on the day an assignment is due, you are expected to submit the assignment the day you return. You have as many days as you are absent plus one to make up any additional work that was assigned. Any data collected during a lab will be provided for you. You can make up tests or quizzes in the Test Center during 0 or 10th period or with me after school.
Late Work: Late work will be accepted for a 20% reduction in points. You can submit late work through the end of a unit. If you do not turn in the assignment during that time, it will be a zero.

Structure of the Course

Homework: Homework will be assigned every night (and on weekends) and should take 30 minutes to an hour to complete. Homework will be checked at the beginning of class. You should show formulas, units, work, and circled answers for full credit. The quizzes will be related to homework assignments. If you do not feel that you understand the material after completing the assigned problems, feel free to do other problems in the text.

Lectures: Lectures will be somewhat formal, but with interchange between the instructor and student the norm. Questions are encouraged. Attendance is expected. You will find that failure to attend class will negatively affect your performance. Students are expected to pay attention in class. Reading the newspaper, chatting with another student, use of disruptive electronic devices, doing homework for another class, or sleeping shows disrespect to the instructor, to peers, and to the academic discipline. At the discretion of the instructor, students will be asked to leave the classroom for any of the aforementioned activities.

Daily Quizzes: There will be daily quizzes to check for understanding of the material taught the previous day and weekly nomenclature and equation writing quizzes. Each quiz is worth 10-30 points. *Quizzes are free response in format. If you are absent when the quiz is given, make arrangements with the instructor ASAP. Material for quizzes includes the homework and any lab. You will not be allowed extra time beyond the time given in class to complete a quiz.

Chapter Quizzes: There will be a quiz at the end of each chapter. Each quiz is worth 40-70 points. *see above for additional information. There are no retakes of quizzes; each grade is final.

Exams: There will be four 2+ hour exams during the semester. Each exam consists of multiple choice questions and free response problems. The exams are timed. If you have an acceptable reason for missing an exam, it must be made up as soon as possible. It is your responsibility to make these arrangements with the instructor. All exams are cumulative in nature. However, emphasis on a particular exam will be on the material covered since the previous one. Questions will be taken mostly from national AP Chemistry exams from previous years. While all students are invited to come in to go over each exam, it is recommended that any student with a C or lower grade on an exam come in to go over the exam during the week after the exam is returned. There are no retakes of tests; each grade is final.

Laboratory: The laboratory experience is an essential part of your understanding of chemistry. The experiments that you will perform have been chosen carefully to fulfill this purpose. Your complete participation is needed for you to take advantage of the experiences designed for you. To assess your understanding of the labs, you will submit complete lab reports for each experiment and answer questions on quizzes and hour exams that can only be answered if you complete the experiment and report.
For the lab, you will use a lab notebook. Detailed instructions for preparation of the notebook and writing of reports will be provided. All materials needed to prepare and conduct the experiments will also be provided. The grading rubric must be attached as the first page of the report.

Grading

Categories:

Tests	50%
Chapter Quizzes	20%
Labs	25%
Homework/Daily Quizzes	5%

Grading Scale Cutoffs: 90% = A, 80% = B, 70% = C, 60% = D, below 60% = F

Academic Integrity

All students deserve a healthy learning environment and evaluations that are based on their honest independent efforts. A clear sense of academic honesty and responsibility is fundamental to good scholarship and learning.

You are encouraged to form study groups and to problem solve together. The normal expectation is that the work on quizzes and exams is your own and that homework and lab reports, while discussed with other students, is of your own creation. Assignments submitted by pairs or groups of students will not be accepted. Academic dishonesty will not be tolerated. Please refer to the student handbook.

The AP Exam

All students enrolled in AP Chemistry are expected to take the AP exam in May. This year the AP Chemistry exam is on May ___. Students who take the AP exam are exempt from taking the final at the end of second semester. To help you to be successful on the AP exam (and in class), you should follow some guidelines:

- Take careful notes, including notes about demos and examples presented in class
- Turn in all assignments on time
- Come to class every day
- Spend time on chemistry every night
- Be open to a variety of learning styles and activities
- Find a study partner or a study group
- Get one-on-one help before or after school
- Become an independent learner – Take the initiative to ask questions and do extra work when necessary. Be reflective about your learning – ask yourself, "do I really understand this?"
- As soon as you feel behind or are having trouble, TALK TO ME!

AP Chemistry Primary Objectives

AP chemistry students should be able to:

- think analytically to reduce problems to identifiable, answerable questions;
- attain a depth of understanding of fundamentals and a reasonable competence in dealing with chemical problems;
- enhance abilities to think clearly and to express their ideas, orally and in writing, with clarity and logic;
- obtain college credit on the AP chemistry exam;
- understand problems expressed as experimental questions;
- design and carry out experiments that answer questions;
- manipulate data acquired during an experiment;
- make conclusions and evaluate the quality and validity of such conclusions;
- propose further questions for study; and
- communicate accurately and meaningfully about observations and conclusions.

AP Chemistry Lab Philosophy

Chemistry lab is NOT a spectator sport! Each student must be actively involved throughout each experiment. In some experiments, you will work by yourself. In other experiments, you will work with one or more lab partners. Regardless of the size of the working group, each member of the group must significantly contribute to the group activity in the laboratory. Trade tasks throughout the experiment, so all members have the opportunity to become proficient in the task at hand. You will be required to collect, process, and manipulate data taken from physical observations, both measured and unmeasured, and then develop and formally report your conclusions. Laboratory reports will consist of the following sections: purpose, procedure, data analysis, error analysis, and conclusions. You may also be called on to present your laboratory results to the class.

Educational research suggests that each individual constructs his/her own knowledge. Thus, it is considered important in AP Chemistry lab that students have the opportunity to contribute, as much as they can, to the procedure they will carry out to accomplish the stated purpose of a particular experiment. To that end, some experiments are presented with a minimum of instructions. It is the responsibility of the student to devise his/her own procedure, using the equipment and supplies provided. Of course, it is permissible, and encouraged, for students to consult and discuss their ideas with their peers. For the first few experiments, students typically feel very unsure and uncomfortable with this process. Please persevere, for as the course proceeds, you will find yourself gaining in confidence and ability. You should find this skill valuable as you participate in laboratory activities in other courses and when you get the opportunity to conduct independent research. In other experiments, detailed instructions will be provided. In any case, it is expected that all students will have the procedure (as well as any pre-lab questions) completed when they come to lab.

Lab Handouts and Experiments
All lab handouts will be taken from the Advanced Chemistry with Vernier Lab Book. Students will be required to complete their lab reports in a chemistry lab notebook and attach the appropriate grading rubric. The following table lists the experiments that should be completed during the school year.

Experiment Number	Name of Experiment
1	Determination of a Chemical Formula – Tin(?)Oxide
2	Determination of the Percentage of Water in a Compound
3	The Molar Mass of a Volatile Liquid
4	Using Freezing Point Depression to Determine Molecular Weight
5	The Molar Volume of a Gas
6	Standardizing a Solution of Sodium Hydroxide
7	Acid-Base Titration
8	Determining the Mole Ratios in a Chemical Reaction
9	The Determination of an Equilibrium Constant
10	Investigating Indicators
11	The Decomposition of Hydrogen Peroxide
12	Determining the Enthalpy of a Chemical Reaction
13	Separation and Qualitative Analysis of Cations and Anions
14	The Synthesis and Analysis of Alum
15	Conductimetric Titration and Gravimetric Determination of a Precipitate
16	Determining the Concentration of a Solution: Beer's Law
17	Liquid Chromatography
18	Buffers
19	Electrochemistry: Voltaic Cells
20	Electroplating
21	The Synthesis and Analysis of Aspirin
22	Determining the Ksp of Calcium Hydroxide
23	Determining Ka by the half titration of Weak Acid
24	Determining the Rate and Order of a Chemical Reaction
25	Exploring the Properties of Gases
26	Determining Avogadro's Number
27	Vapor Pressure and Heat of Vaporization
28	Rate Determination and Activation Energy

AP Chemistry Course Outline

Summer Packet Objectives:

1. Count the number of significant figures in a measurement.
2. Add, subtract, multiply, and divide with the correct number of significant figures.
3. Use dimensional analysis to solve problems.
4. Name and form basic ionic compounds.
5. Write and balance equations.
6. Do conversions involving moles, molecules, and molar mass.
7. Calculate percent by mass.
8. Calculate the average atomic mass of an isotope using percent abundance.
9. Do stoichiometry problems, including those that use limiting reactants and excess.
10. Determine empirical and molecular formulas through calculation.
11. Calculate the empirical formula of an unknown hydrocarbon through a combustion reaction and calculation.

<u>Semester One</u>

Chapter 1-3: Chemical Foundations (2 weeks)
Atoms, Molecules, and Ions
Stoichiometry

Topics:

I. Measurement topics
II. Atomic theory
III. Symbols and formulas
IV. Periodic table
V. Ionic and covalent bonds
VI. Nomenclature
VII. Reactions
VIII. Stoichiometry
  - A. Percent composition
  - B. Empirical formulas
  - C. Solutions
  - D. Mole relationships
    - 1. percent (%) yield
    - 2. Limiting reagents
  - E. Titrations and other analyses

Objectives:

1. Recall a definition of chemistry.
2. Understand the process and stages of scientific (logical) problem solving.
3. Recall the three states of matter, their general properties and the methods for their interconversion.
4. Understand and recall definitions for physical and chemical change.
5. Know the difference between elements, mixtures and compounds including the difference between heterogeneous and homogeneous mixtures.
6. Understand and be able to use scientific notation (standard form).
7. Recall and use SI units and prefixes.
8. Be able to convert between units.
9. Understand the concept of derived units and use relationships relating to density.
10. Recall the meaning of uncertainty and understand and be able to use the rules for determining significant figures and rounding off.
11. Understand the differences between, and be able to apply, the concepts of accuracy and precision.
12. Learn, and be able to use, formulae for the conversion of the three different temperature units studied.
13. Learn and be able to apply the formula for percentage error

14. Recall a very brief history of Atomic Theory
15. Know and understand the five main aspects of Dalton's Atomic Theory
16. Recall some of the experiments that led to the identification of sub-atomic particles
17. Know the three particles that make up the atom and their relative charges, masses and positions in the atom
18. Be able to use the Atomic # and Mass # of an isotope to calculate the numbers of protons, neutrons and electrons present
19. Know what the term isotope means and be able to perform simple calculations relating to isotopic data
20. Know the approximate locations of metals, non-metals and metalloids on the periodic table
21. Learn the lists of common anions and cations (including polyatomic ions) studied.
22. Know how to combine those anions and cations in the correct proportions to form ionic compounds with no net charge
23. Be able to name binary ionic compounds of a metal and a non-metal
24. Be able to name binary molecular compounds of two non-metals
25. Be able to name simple binary acids
26. Be able to name ionic compounds containing polyatomic anions
27. Be able to name oxoacids and compounds containing oxyanions
28. Be able to name hydrated salts
30. Be able to write chemical equations in words
31. Be able to write chemical equations using chemical formulae and chemical symbols (this requires knowledge, and correct use of, chemical nomenclature)
32. Understand, and be able to use, state symbols as part of chemical equation writing
33. Be able to balance chemical equations
34. Understand why balancing chemical equations is important
35. Understand the concept of percentage by mass
36. Be able to calculate empirical formulae from percentage by mass data
37. Be able to convert empirical formulae to molecular formulae by using Molar Mass data
38. Understand and be able to apply the concept of the mole in chemical calculations (including the application of Avogadro's number)
39. Be able to use combustion data to calculate empirical formulae of compounds
40. Understand the importance of, and be able to apply, the concept of stoichiometric coefficients relating to reacting ratios
41. Know how to calculate the number of moles of a solid substance present in a reaction from data
42. Be able to perform calculations relating to molarity
43. Be able to perform calculations relating to dilution
44. Be able to perform calculations relating to molality
45. Be able to calculate the formulae of hydrated salts from experimental data
46. Understand, and be able to apply, the concept of a limiting reactant
47. Understand, and be able to apply, the concept of percentage yield

Laboratory:

- Intro to lab: volume and mass
- Determination of a chemical formula
- Determination of the percentage of water in a compound
- Determining Avogadro’s number

Chapter 4: Types of Chemical Reactions and Solution Stoichiometry (3 weeks)

Topics:

I. Reaction types
    A. Acid base reactions
        1. Concepts of
            a) Arrhenius
            b) Lowry-Bronsted
            c) Lewis
    B. Precipitation reactions
    C. Oxidation reduction reactions
        1. Oxidation number
        2. Electron transport
        3. Electrochemistry

II. Stoichiometry

III. Net ionic equations

IV. Balancing equations including redox

V. Mass-volume relationships with emphasis on the mole

VI. Chemical Equations According to AP

Objectives:

1. Define a solution and explain the components that make up a solution.
2. Identify the factors that affect the formation of a solution and the amount of solute that will dissolve in a solvent.
3. Identify major properties of water and explain why it is known as a “universal” solvent.
4. Explain what it means to be a strong electrolyte, weak electrolyte, or nonelectrolyte.
5. Explain the difference between ionization and solubility (ionizing vs. dissolving)
6. Calculate the molarity of a solution given grams or moles of solute and L or mL of solvent.
7. Utilize the dilution equation to determine the new molarity of a solution.
8. Recognize double replacement reactions and predict the products of a double replacement reaction.
9. Use the solubility rules to predict what substances will precipitate in a double replacement or acid-base reaction.
10. Write molecular equations.
11. Write complete ionic equations.
12. Write net ionic equations.
13. Define and recognize spectator ions.
14. Recall the three major steps necessary to solve a stoichiometry problem
15. Solve stoichiometry problems that utilize a net ionic equation.
16. Solve stoichiometry problems with limiting reactants.
17. Solve stoichiometry problems where they must solve for the concentration of ALL species at the end of a reaction.
18. Give the Arrhenius definition of an acid and a base.
19. Give the Bronsted-Lowry definition of an acid and base.
20. Define what it means to be a “strong” acid or a base.
21. Identify strong acids and bases by name and formula.
22. Write molecular, complete ionic, and net ionic equations for a reaction between a strong acid and strong base, weak acid and strong base, or strong acid and weak base.
23. Identify what is equal when doing an acid-base titration and solve stoichiometry problems using this equality.
24. Assign oxidation numbers for elements, molecules, compounds, and ions.
25. Define oxidation and reduction and state what happens to the oxidation number for each.
26. Define oxidation agent and reducing agent.
27. Recall the steps for balancing an oxidation-reduction reaction.
28. Balance redox reactions in acidic solution and basic solution. (Ch.18.1)

Laboratory:
- Determination mole ratios in a chemical reaction
- Standardization of a solution of sodium hydroxide

Chapter 6: Thermochemistry (3 weeks)

Topics:

I. Thermal energy, heat, and temperature
II. Calorimetry
III. Enthalpy changes
IV. Hess's Law

Objectives:
1. Learn the meaning of the following thermodynamic terms: enthalpy, ΔH, exothermic, endothermic, system, surroundings, universe, heat of formation, heat of reaction, calorimetry, heat, calorie, joule, standard molar enthalpy of formation, molar heat of combustion.
2. Explain the first law of thermodynamics.
3. Define heat and explain how the definition can be applied in a handshake.
4. Distinguish between system and surroundings.
5. Identify an exothermic reaction in terms of drawing a potential energy diagram, strength of reactant/product bonds, placing heat on the correct side of the equation, and the correct sign for ΔH.
6. Identify an endothermic reaction in terms of drawing a potential energy diagram, strength of reactant/product bonds, placing heat on the correct side of the equation, and the correct sign for ΔH.
7. Define internal energy and solve for internal energy of a system.
8. Determine the sign of heat, work, and internal energy in a typical exothermic or endothermic reaction.
9. Solve for work done on the system given pressure and volume.
10. Recall the conversion from L·atm to J, and solve for ΔE of a system.
11. Recall the equation used to solve for change in enthalpy.
12. Solve calorimetry problems involving $q = mC\Delta T$.
13. Explain what it means for an object to have low or high specific heat.
14. Calculate the specific heat capacity or molar heat capacity of an object.
15. Solver for the combustion per gram or per mole of a substance.
16. Apply the ΔH to the correct side of the equation and solve stoichiometry problems with heat.
17. Use and rearrange the reaction mechanism to solve for the overall ΔH of a reaction.
18. Use Hess's Law to solve for heat of reaction.
19. Use stoichiometric principles to solve heat problems.
20. Explain how to determine the heat of formation of a compound from its elements.
21. Solve for the overall ΔH of a compound using heats of formation of each element or compound in a reaction.
22. Solve for the standard heat of formation of a compound in a reaction given the overall change in enthalpy and the heats of formation for all other parts of the reaction.

Laboratory:
- Determination of enthalpy of a chemical reaction

Chapter 7, 19: Atomic Structure and Periodicity (3 weeks)
The Nucleus: A Chemist's View

Topics:

I. Electronic Structure
   A. Evidence for the atomic theory
   B. Atomic masses

C. Atomic number and mass number
D. Electron energy levels: atomic spectra, quantum numbers, atomic orbitals
E. Periodic relationships

II. Nuclear structure
A. Nuclear equations
B. Half-lives
C. Radioactivity
D. Chemical application

Objectives:

1. Understand the Bohr model of the atom
2. Understand how line emission spectra are formed
3. Appreciate that the electron can be considered to have wave like properties as well as particle type properties
4. Explain what it means when energy is "quantized" and explain how energy can have particle qualities.
5. Define wavelength and frequency and give the units associated with each.
6. Differentiate between different types of electromagnetic radiation in terms of wavelength, frequency, and energy.
7. Recall the relationship between energy with frequency and calculate its value.
8. Recall the relationship between energy and wavelength and calculate its value.
9. Relate Einstein's equation and solve for the mass of light given the wavelength.
10. Understand and use equations that relate the Energy, frequency, speed and wavelength of waves
11. Determine the wavelength and type of light found given an energy transition for the hydrogen atom in the Bohr model.
12. Understand the concept of electrons in shells and the use of quantum numbers
13. Understand the use of the terms s, p, d and f and their use in orbital notation
14. Recall and understand the rules for filling orbitals and determining electronic configuration, including the Pauli exclusion principle, Hund's rule of maximum multiplicity and notable exceptions
15. Be able to construct the electronic configuration of the elements using the s, p and d and f notation
16. Be able to construct the electronic configuration of the elements using the noble gas core and s, p, d and f notation
17. Be able to construct the electronic configuration of simple ions (including d block ions)
18. Recall the shapes of the s, p and d orbitals
19. Recall that orbitals are electron probability maps
20. Be able to describe electronic configurations using the electrons in boxes notation
21. Define what each quantum number (n, l, m_l, m_s) stand for and give the possible values for each.
22. Assign quantum numbers for the valence electron of compounds.
23. Determine the maximum number or electrons for a given quantum condition.
24. Determine the number of valence electrons in an atom.
25. Understand that regular, repeatable patterns occur across periods and within groups on the periodic table
26. Appreciate that these patterns sometimes have notable exceptions
27. Recall and understand that the noble gases have full outer shells that represent stable electronic configurations
28. Recall how, and understand why, group I, II, VI and VII elements achieve pseudo noble gas electronic configurations
29. Recall the definition of ionization energy and electron affinity.
30. Recall and understand the variation in ionization energy and electron affinity when moving about the periodic table
31. Be able to predict the group an element is in from ionization energy data

32. Recall how and why atomic and ionic size vary when moving about the periodic table
33. Understand how many physical properties change gradually when moving about the periodic table, including densities, boiling points, and melting points.
34. Explain what happens to the size of an ion when it becomes a cation and when it becomes an anion and explain the trend in ionic radius down a group.
35. Use the number of paired and unpaired electrons to determine the color or transition metals and the magnetism of an element.
36. List which transition metals are brightest in color.
37. Define diamagnetism, paramagnetism, and ferromagnetism, give examples of each, and determine which gives the greatest magnetic field.
38. Understand the phenomenon of radioactivity and the properties of radioactive particles
39. List the types of radioactive emissions.
40. Be able to write and balance nuclear equations
41. Understand the concept of half-life and be able to perform calculations related to it
42. Recall some uses of radioactivity
43. Understand the term mass deficit
44. Be able to use neutron:proton ratio to make predictions about stability
45. Understand the terms nuclear fission and fusion
46. Understand, that in very general terms, radioactivity involves the rearrangement of the nucleus and chemical reactions involve the rearrangement of electrons

Laboratory:

- A & B Synthesis of Alum and Its Chemical Analysis

Chapter 8, 9: Bonding – General Concepts (3 weeks)
Covalent Bonding – Orbitals

Topics:

I. Binding forces
- A. ionic
- B. covalent
- C. organic
- D. metallic
- E. hydrogen bonding
- F. Van der Waals

II. Relationships to states, structure, and properties of matter

III. Polarity of bonds, Electronegativities

IV. Molecular models
- A. Lewis structures
- B. Valence bond: Hybridization of orbitals, resonance, sigma and pi bonds

V. VSEPR
- A. Geometry of molecules and ions
- B. Structural, geometric, optical, and conformational isomerism of:
    1. Organic molecules
    2. Coordination complexes

Objectives:

1. Explain the two reasons that atoms or molecules tend to form bonds.
2. Understand the essential difference between intra and inter bonding
3. Differentiate between ionic, polar covalent, and nonpolar covalent bonds and explain the location of electrons in each case.
4. Understand that ionic bonding and covalent bonding are at two ends of a sliding scale of bond type
5. Understand the concept of electronegativity
6. Understand that polarization caused by small highly charged cations leads to ionic compounds exhibiting some covalent character

7. Understand that differences in electronegativity in covalent molecules causes dipoles and some ionic character in covalent compounds
8. Use electronegativity values to determine the type of bond formed and its polarity.
9. Draw Lewis dot diagrams for atoms
10. Draw Lewis structures for ionic compounds
11. Calculate the energy of attraction or lattice energy of a compound using Coulomb's Law.
12. Tell what a negative and positive result for Coulomb's Law represents.
13. Tell what energy is directly and inversely related to in Coulomb's Law and which is more significant.
14. Determine the number of valence electrons in a given atom.
15. Identify atoms and situations where the octet rule is violated.
16. Draw Lewis structures of covalent compounds.
17. Understand the concept of resonance related to Lewis structures
18. Determine when a molecule will have lone pairs or multiple bonds.
19. Determine the Lewis geometry and the VSEPR shape of molecules.
20. Determine the bond angles associated with Lewis structures of molecules.
21. Determine the polarity of a molecule given its structure and shape.
22. Determine the hybridization around the central atom.
23. Recall the formula for determining the enthalpy of a reaction using bond energies.
24. Calculate the enthalpy of a reaction by drawing Lewis structures and calculating bond energies.
25. Calculate ΔH of a bond given a reaction and the overall enthalpy.
26. Solve for the formal charge of atoms to determine the most likely Lewis structure.
27. Understand the concept of the dative (co-ordinate) bond related to Lewis structures
28. Understand the nature of sigma and pi bonds
29. Understand and be able to identify different types of orbital hybridization

Laboratory:
- Separation and qualitative analysis of cations and anions
- The molar mass of a volatile liquid
- Conductimetric titration and gravimetric determination of a precipitate

Chapter 5, 10: Gases (3 weeks)
Liquids and Solids

Topics:

I. Gas Laws
- A. Ideal gases
- B. Boyle's law
- C. Charles' law
- D. Dalton's law of partial pressure
- E. Graham's law
- F. Henry's law
- G. Van der Waal's equation of state

II. Kinetic-Molecular theory
- A. Avogadro's hypothesis and the mole concept
- B. Kinetic energy of molecules
- C. Deviations from Ideality

III. Liquids and Solids
- A. Intermolecular Forces
- B. Liquids and solids from the K-M viewpoint
- C. Phase diagrams of one-component systems
- D. Changes of state
- E. Structure of solids including lattice energies

Objectives:
1. Be able to convert between different units of pressure
2. Be able to convert between different units of temperature
3. Recall and be able to use Boyle's law in calculations
4. Recall and be able to use Charles' law in calculations
5. Recall and be able to use Gay-Lussac's law in calculations
6. Recall and be able to use Avogadro's law in calculations
7. Recall and be able to use the Combined gas law and the General gas law in calculations
8. Recall and be able to use the Ideal gas law in calculations
9. Understand and be able to use the van der Waals equation (modified ideal gas law) in calculations
10. Recall and be able to use Dalton's law of partial pressures in calculations
11. Recall the conditions that are used as standard in calculations
12. Be able to use molar gas volume in calculations
13. State and discuss the major tenants of the kinetic-molecular theory.
14. Apply the kinetic-molecular theory to liquids and solids, as well as gases.
15. Understand the concept of, and be able to perform calculations involving, the root-mean-square-speed of gases
16. Understand the terms effusion and diffusion and be able to perform calculations relating to those concepts
17. Interpret heating curves as to melting point, boiling point, and specific heat.
18. Interpret phase diagrams and correctly define terms such as triple point, critical temperature, and critical pressure.
19. Discuss the phenomena of boiling, and be able to relate it to pressure.
20. Carry out a distillation to separate substances with differing boiling points.
21. Distinguish between crystalline and amorphous solids.
22. Use the unit cell information for calculation of ionic radii.
23. Use X-ray diffraction data to calculate unit cells.
24. Discuss intermolecular forces and relate them to physical properties such as boiling point.

Laboratory:
- Exploring the properties of gases
- The molar volume of a gas
- Triple Point of Dry Ice

<u>Semester Two</u>
Chapter 11: Properties of Solutions (2.5 weeks)
Topics:
I. Types of solutions
II. Factors affecting solubility
III. Concentration issues
IV. Raoult’s law and Colligative properties
V. Nonideality

Objectives:
1. Identify the two components that make up a solution
2. Identify what type of a mixture a solution is and its properties
3. Explain how to create a 50/50 mixture of two components in several different ways, and explain why 50 mL of ethanol and 50 mL of water do not give 100 mL total
4. Explain conceptually the difference between a dilute and a concentrated solution
5. Calculate molarity and concentration of ions
6. Calculate molality or solve for moles or grams of solute using molality
7. Calculate % composition by mass
8. Calculate the mole fraction of either the solute or the solvent

9. Find the normality of both an acid-base and a redox solution using the molarity and the number of equivalents
10. Give three examples based on structure of a solute that would allow it to mix with a solvent such as water
11. Recall the three thermodynamic steps associated with solution making and explain why they are either endothermic or exothermic
12. Find the overall enthalpy of a solution from its steps
13. Explain what a change in temperature will do, in general, to the solubility of solids and to the solubility of gases
14. Show how a change in pressure can affect the solubility of a gas using Henry's Law, and calculate the pressure or concentration using Henry's Law.
15. Explain what colligative properties are always based upon and explain how this term is used in colligative properties calculations
16. Explain what happens to the boiling point of a solution when a solute is added
17. Explain what happens to the freezing point of a solution when a solute is added
18. Calculate the new boiling/freezing point of a substance
19. Calculate the mass or the molar mass of a solute from freezing/boiling information
20. Explain what happens to the vapor pressure of a solution when a solute is added
21. Calculate the new vapor pressure of a solution using Raoult's Law
22. Explain when a solute/solvent pair are "ideal" using Raoult's Law
23. Explain and calculate the osmotic pressure of a solution

Laboratory:
- Buffers
- Using freezing point to find molecular weight

Chapter 12: Chemical Kinetics (2.5 weeks)
Topics:
I. Rate of reaction
II. Order of the reaction
III. Factors that change the rate of the reaction
    A. Temperature
    B. Concentration
    C. Surface Area
    D. Nature of substance
    E. Catalysts
IV. Relationship between the rate-determining step and the reaction mechanism

Objectives:
1. Define kinetics and explain what it means
2. Explain what happens to the rate of reaction over time at constant temperature
3. Define "rate" and explain how to calculate the average rate of reaction
4. Differentiate between average and instantaneous rate, and calculate the instantaneous rate of reaction
5. Use stoichiometry to find rates of other products or reactants given one.
6. Identify and explain the three major parts of collision theory that are needed for a reaction to occur.
7. Identify and explain how the five factors affect the rate of a reaction in terms of collision theory.
8. Identify how to find the orders of the reactants and explain the effect on rate is the concentration is changed fin a zero order, first order, or second order reaction.
9. Identify when to use a differential or integrated rate law by looking at the data given
10. Use experimental data to calculate orders of reactants in a differential rate law
11. Use experimental data to calculate the orders of reactants when one reactant is not held constant in any two trials.

12. Calculate the rate constant, k, using experimental data and determine the correct units for the rate constant.
13. Determine the order of the reaction by analyzing graph data, and identify the variable that should be on the y-axis for zero order, first order, and second order reaction
14. Solve for the rate constant in an integrated rate law using the slope of the line or by using the experimental data (UNITS are really important again)
15. Recall and utilize the equation used in a first order integrated rate law and solve for either time or concentration
16. Recall the equation to find the half-life of a first order reaction and calculate it
17. Recall and utilize the equation used in a second order integrated rate law and solve for either time or concentration
18. Recall the equation to find the half-life of a second order reaction and calculate it
19. Recall and utilize the equation used in a zero order integrated rate law and solve for either time or concentration
20. Recall the equation to find the half-life of a zero order reaction and calculate it
21. Define “mechanism” and explain how it is used to verify a rate law
22. Explain what a rate determining step is and identify it given a rate law and a reaction mechanism
23. Define an intermediate and identify one in a reaction mechanism
24. Explain how to identify a catalyst in a reaction mechanism
25. Simplify a reaction mechanism to verify that the balanced equation for the mechanism is the same as the overall reaction
26. Define “molecularity” and give examples of balanced equations that would be unimolecular, bimolecular, and termolecular.
27. Analyze the experimentally determined rate law and the rate law provided by the mechanism to verify the rate law.
28. Draw a potential energy diagram and label the activation energy of the reactants, the activated complex, the heat of reaction, and the same quantities for the reverse reaction
29. Draw a potential energy diagram and explain what is changed and what is unchanged when a catalyst is added.
30. Explain what changing the rate of concentration does to the rate constant, k
31. Identify and explain 3 things that can be done to change the rate constant, k, of a reaction
32. Recall the Arrehenius equation and use it to solve for the rate constant, temperature, or activation energy.

Laboratory:
- The rate & order of a chemical reaction
- Rate determination & activation energy

Chapter 13: Chemical Equilibrium (2 weeks)
Topics:
I. Concept of dynamic equilibrium including Le Chatelier’s principle
II. Equilibrium constants and the law of mass action

Objectives:
1. Identify what is equal at equilibrium
2. Explain why equilibrium is considered dynamic
3. Draw a graph and identify when a reaction has reached equilibrium and whether is favors the reactants or products at equilibrium
4. Identify the three factors that determine whether the reaction at equilibrium will favor the reactants or the products
5. Explain the relationship between Gibbs free energy and equilibrium.
6. Write equilibrium expressions for reactions at equilibrium.
7. Calculate the equilibrium constant, K, for a reaction at equilibrium

8. Write and calculate K_p expressions using partial pressures instead of concentrations.
9. Identify the two things that will change the equilibrium constant, K.
10. Predict and calculate K based on the original equation and given a new manipulated equation and the original equilibrium constant.
11. Recall and utilize the equation needed to convert K (K_c) to K_p
12. Use the value calculated or given for K to determine whether the equilibrium lies towards products or the reactants
13. Calculate where a reaction is at any point by solving for the reaction quotient, Q.
14. Compare Q and K and interpret which side of the reaction is favored at that point and the shift that the reaction will make to reach equilibrium.
15. Calculate the equilibrium constant given initial concentrations and the concentration of a reactant at equilibrium (ICE)
16. Calculate the equilibrium concentrations of ALL species given initial concentrations and the equilibrium constant (ICE)
17. Correctly use the 5% rule when applicable to prove that it is acceptable through calculation (ICE)
18. Define and explain Le Chatelier's principle
19. Explain how the equilibrium shifts with a change in concentration and explain the effect on the concentration/amount of a given reactant/product and the constant, K.
20. Explain how the equilibrium shifts with a catalyst and explain the effect on the concentration/amount of a given reactant/product and the constant, K.
21. Explain how the equilibrium shifts with a change in temperature and explain the effect on the concentration/amount of a given reactant/product and the constant, K.
22. Explain how the equilibrium shifts with a change in pressure and explain the effect on the concentration/amount of a given reactant/product and the constant, K.
23. Explain how the equilibrium shifts with a change in volume and explain the effect on the concentration/amount of a given reactant/product and the constant, K.

Laboratory:
- Determination of the equilibrium

Chapter 14, 15: Acids and Bases (3 weeks)
Acid-Base Equilibria

Topics:

I. Arrhenius theory
- A. Properties of acids and bases
- B. Acid base neutralization

II. Lowry-Bronsted theory
- A. Amphiprotic species
- B. Relative strengths of acids and bases
- C. Polyprotic acids

III. Lewis acids and bases. Comparison of all three definitions.

IV. Weak acids and bases
- A. pH
- B. pOH
- C. Common Ion Effect
- D. Buffer systems
- E. Indicators
- F. Hydrolysis

V. Solubility Product
- A. Factors involving dissolution
- B. Molar solubility

Objectives:

1. Identify many properties of acids and bases
2. Differentiate between strong and weak acids and bases
3. Identify all of the strong acids and strong bases
4. Recall the Arrhenius definition for acids and bases
5. Recall the Bronsted-Lowry definition for acids and bases
6. Identify conjugate acid-base pairs in a Bronsted-Lowry reaction, and predict the products of a reaction with an acid and a base
7. Explain the relative strength of a strong acid/base and its conjugate
8. Explain the relative strength of a weak acid/base and its conjugate
9. Recall the lewis definition of an acid and base and identify Lewis acids/bases
10. Compare the strength of acids/bases by looking at their formulas and K_a values
11. Calculate the K_a or K_b or a substance given the other (K_a or K_b)
12. Calculate the concentration of $[H^+]$ or $[OH^-]$ given the other
13. Calculate pH or pOH given the concentration of $[H^+]$ or $[OH^-]$
14. Calculate pH or pOH given the other
15. Calculate $[H^+]$ or $[OH^-]$ given the pH or pOH
16. Calculate the pH of a strong acid solution
17. Calculate the pH of a weak acid solution
18. Calculate the percent dissociation of an acid (same calc as 5% rule)
19. Calculate the pH and concentration of species given a mixture of acids
20. Calculate the pH and concentration of all species for a polyprotic acid
21. Calculate the pH of a solution containing an acid and its conjugate base (common ion)
22. Calculate the pH of a strong base solution and explain how it differs in reactivity from a strong acid
23. Calculate the pH of a weak base solution
24. Predict whether a salt will be acidic, basic, or neutral in an aqueous solution
25. Explain why some cations and anions yield a neutral solution
26. Explain what makes a salt solution basic, and calculate the pH of the solution given a concentration of the anion and the K_a value for its conjugate acid
27. Explain why a conjugate acid of a weak base in a salt will make the solution acidic, and calculate the pH given a concentration and the K_b of the weak base
28. Explain why highly charged cations will yield an acidic solution and write an equation for hydrolysis hat occurs
29. Calculate the pH and the percent dissociation of a solution that contains both a weak acid and its conjugate base (or a weak base and its conjugate acid)
30. Explain what the presence of common ion does to the percent dissociation of an acid/base
31. Write equations for the reaction of a acid or a base with a buffer and explain the expected change in pH
32. Identify the components of an ideal buffer
33. Recall the Henderson-Hasselbach equation for pH and pOH
34. Calculate the new pH of a buffer solution after the addition of either acid or base using stoichiometry and the Henderson-Hasselbach equation (or ICE)
35. Draw a curve for a strong acid-strong base, strong base-strong acid, weak acid-strong base, and weak base-strong acid titration and on it: give an estimated initial pH, identify the species present at each point and the ideal buffer region, the estimate equivalence point, and explain why the equivalence point occurs at this pH.
36. Calculate the initial pH in all four types of titration
37. Use stoichiometry and the Henderson-Hasselbach equation (or ICE) to find the pH at any point prior to the equivalence point.
38. Calculate the volume needed to reach the halfway point and the equivalence point of a titration using a balanced equation and solution stoichiometry
39. Use stoichiometry and the concentration of excess acid or base to find the pH at any point past the equivalence point.

40. Recall what type of chemical an indicator is and use Le Chatelier's principle to explain its reaction with acids or bases and the associated color change
41. Calculate the range of an indicator using the Henderson-Hasselbach equation and the 1/10 rule
42. Select the appropriate indicator for a solution that has a given pH, and use calculations to prove this indicator is most effective
43. Calculate the pH of a solution and then determine what color the indicator will appear at this pH.
44. Explain what is occurring at equilibrium in a Ksp problem
45. Write Ksp expressions for ionic compounds
46. Solve for Ksp given the molar solubility of a substance
47. Solve for the molarity solubility of a substance given Ksp
48. Compare two solutes that have the same number of ions and determine the relative solubility
49. Compare two solutes that have different numbers of ions and determine the relative solubility
50. Explain what the common ion effect does to solubility and calculate a Ksp problem that involves common ions
51. Calculate Q for a dissolution process and determine whether a precipitate is made by comparing the value to Ksp, explaining why this is the case.

Laboratory
- Acid base titration
- Determination of the KSP of calcium hydroxide
- Investigating indicators
- Determining Ka by the half titration of a weak acid

Chapter 16: Solubility and Complex Ion Equilibria (1 week)
Topics:
I. Names and structures of complexions
II. Bonding in coordination systems
III. Formation of complex ions (reactions).
IV. Practical applications

Objectives:
1. Define the following: central ion or atom, coordination sphere, coordination number, polydentate ligand, ligand, chelating agent, cis and trans isomers, t2g and e.g. orbitals, ligand field splitting, 10 Dq, LFSE, low spin complex, high spin complex, paramagnetic, diamagnetic, % transmittance, Absorbance, Beer's law, spectrometer.
2. Name coordination complexes.
3. Write net ionic equations involving complex ions.

Laboratory:
- Determination of concentration of a solution: Beer's Law

Chapter 17: Spontaneity, Entropy, and Free Energy (2 weeks)
Topics:
I. State functions
II. Laws of thermodynamics
III. Relationship of change of free energy to equilibrium constants

Objectives:
1. Explain what it means for a reaction to be spontaneous
2. Explain how spontaneous reactions depend on temperature
3. Explain how a catalyst would affect spontaneity

4. Explain how some reactions that are endothermic are also spontaneous
5. Give the driving force in a spontaneous exothermic reaction
6. Define entropy in terms of disorder and the number of possible arrangements
7. Explain 4 ways to increase entropy in terms of state of matter, number of molecules, type of reaction, and temperature
8. Recall the second law of thermodynamics
9. Explain the effect on the entropy of the system and the surroundings in both an endothermic and an exothermic reaction.
10. Solve for entropy of the system or surroundings given its enthalpy and temperature
11. Solve for the entropy change of a reaction from standard values
12. Recall the third law of thermodynamics
13. Define Gibbs free energy and explain what values of ΔG are spontaneous
14. List the three factors that influence the value of ΔG
15. Calculate Gibbs free energy from enthalpy, temperature, and entropy
16. Calculate ΔG given values for each part of a reaction
17. Predict at what conditions reactions will be spontaneous based on signs of ΔH and ΔS
18. Give the signs for ΔH, ΔS, and ΔG for a nonspontaneous exothermic process, freezing, dissolving of calcium oxide, and evaporating
19. Explain why a solution may dissolve spontaneously even though it has a slightly positive ΔH in terms of entropy and Gibbs free energy.

Laboratory:

- Vapor pressure & heat of vaporization

Chapter 18: Electrochemistry (2 weeks)

Topics:

I. Galvanic cells and cell potentials
II. Electrolytic cells
III. Redox equations

Objectives:

1. Use the half reaction method to balance redox equations
2. Recall the definition of oxidation and reduction in terms of electrons
3. Distinguish between an electrolytic cell and a voltaic cell in terms of function and ΔG
4. Solve problems using Faraday's law.
5. Predict reaction products for both electrolytic and voltaic cells.
6. Discuss the importance of and draw a diagram of a standard hydrogen electrode.
7. Understand and recall the definition of standard electrode potential
8. Understand and recall how to construct a cell diagram (line notation) and draw a diagram (picture) of the apparatus needed
9. Recall the conditions that standard electrode potentials are measured under
10. Understand the nature and purpose of a salt bridge
11. Be able to predict the likelihood or otherwise of chemical reactions using standard electrode potentials and understand how those predictions may not prove to be accurate
12. Understand and use the Nernst equation
13. Understand the relationship between Gibbs free energy, equilibrium constants and Ecell, and be able to perform related calculations
14. Understand electrolysis and be able to perform quantitative calculations relating to it
15. Discuss and give examples of primary cells, secondary cells, and fuel cells.

Laboratory

- An oxidation-reduction titration the reaction of Fe^{2+} and Ce^{4+}
- Electrochemistry: Voltaic cells
- Electroplating

Review for Exam (1 week)

Topics:

I.test taking strategies

II. 3 complete practice AP exams

III. review of test performance

Welcome to AP Chemistry!

I hope you will find next year's chemistry class to be interesting and enjoyable. I know that you will find it challenging. I have enclosed a packet of materials that I would like you to work on over the summer. This is the material I expect you to know when you walk into class. The first section consists of review material. The second part is the assessment for this review material. I would prefer that you would work in small groups and discuss any problems you have with each other.

You should work hard to make sure you are up to speed on this material. Understanding of this material will be expected but is by no means a prerequisite to be in the class. It will, however, make the start of the year much easier and relieve much of the stress and anxiety that can come at the beginning of an Advanced Placement class. The material in this packet will be briefly reviewed, but it is expected that you feel comfortable with most of it. All you'll need for this packet is a calculator and a periodic table.

If you need some help:
You can email me – my email is akraus@grantbulldogs.org.

I'm looking forward an exciting year in chemistry as well as successful scores on the Advanced Placement Test.

Enjoy the rest of the summer and have fun doing a little chemistry.

Ms. Amber Kraus

AP Chemistry Summer Review Packet

1. How many significant figures are in each of the following:
    a. 2.71 g
    b. 0.00047 kg
    c. 7.0×10^5 m
    d. 1030 L
    e. 150 pencils
    f. 37500 g
    g. 0.1010 cm

2. Express each of the following in proper scientific notation (pay attention to sig figs and units)
    a. 0.000125 m
    b. 155.0 mL
    c. 123,030,000 kg
    d. 481.9×10^{-9} cm

3. Calculate the correct answer with proper units and sig figs for each of the following:
    a. 12 g + 0.677 g + 86.33 g =
    b. (355.78 g) / (0.056 g) =
    c. 97.34 mL – 34.1 mL =
    d. 14.68 x 5 =

4. Dimensional Analysis Practice
    a. A large railroad car is filled with 1745 gallons of milk. The car springs a leak in the bottom, and milk starts dripping out at a rate of 204.84 mL/sec. If the train is traveling at a speed of 65.4 miles per hour, calculate how many miles it will travel before all the milk has drained out of the car.
    (1 gal = 3.78 L, 1 mi = 5280 ft, 1 in = 2.54 cm)
    b. A student can eat 4.0 M&Ms every 1.00 seconds. If an M&M has a mass of 63 mg, determine how many kilograms of M&M's can be eaten by a class of 20 students in 3.75 hours.

5. Name or formula for the following compounds:
    a. Sodium fluoride
    b. K_2O
    c. Calcium phosphate
    d. $FeCl_3$
    e. Iron(II)chloride
    f. Hg_2O
    g. Sodium sulfate
    h. $CaCO_3$
    i. Lithium phosphate
    j. SO_2
    k. Calcium hydroxide
    l. H_2SO_4
    m. Cupric chloride

6. Write and balance the following equations:
    a. Iron metal reacts with oxygen to form solid rust, iron(III)oxide.

    b. Calcium metal reacts with water to produce aqueous calcium hydroxide and hydrogen gas.

    c. Aqueous barium hydroxide reacts with aqueous sulfuric acid to produce solid barium sulfate and water.

7. Problems:
   a. Calculate the mass of 500. atoms of iron (Fe).

   b. Aspartame is an artificial sweetener that is 160 times sweeter than sucrose (table sugar) when dissolved in water. It is marketed as Nutra-Sweet. The molecular formula of aspartame is $C_{14}H_{18}N_2O_5$.
      i. Calculate the molar mass of aspartame.

      ii. Find the % by mass of each element in this compound.

      iii. Calculate the mass in grams of 1.56 moles of aspartame.

      iv. How many molecules are in 5.0 mg of aspartame?

      v. How many atoms of nitrogen are in 1.2 grams of aspartame?

      vi. What is the mass of one molecule of aspartame?

   c. The reusable booster rockets of the U.S. space shuttle employs a mixture of aluminum and ammonium perchlorate for fuel. A possible reaction for this is:

$$\_\_\_Al_{(s)} + \_\_\_NH_4ClO_{4\,(s)} \rightarrow \_\_\_Al_2O_{3\,(s)} + \_\_\_AlCl_{3\,(s)} + \_\_\_NO_{(g)} + \_\_\_H_2O_{(g)}$$

      i. Balance the following reaction.
      ii. If 4.0 g of aluminum reacted with 15.0 g of ammonium perchlorate, what would be the limiting reactant? How much excess reactant would you have?

      iii. Using the above information, how much aluminum chloride would be produced in grams?

      iv. If you actually collected 4.18 g of aluminum chloride what would be your percent yield?

d. You add aluminum to a solution of copper(II)chloride and it reacts exothermically.
   i. Write and balance the equation.

   ii. What does the word exothermic mean. Come up with at least 4 facts about an exothermic reaction.

   iii. You reacted 1.25 g of Al, how much copper(II)chloride do you need to add for the Al to fully react?

   iv. How much of <u>each</u> product would you collect?

e. Phenol is a compound that contains 76.57% carbon, 6.43% hydrogen, and 17.0% oxygen.
   i. Calculate the empirical formula.

   ii. If its molecular weight is 188 g/mol, what would be its molecular formula?

f. When 125.0 g of ethylene (C_2H_4) burns in 60.0 g of oxygen to give carbon dioxide and water, how many grams of CO_2 are formed. (Hint: balance first!)

AP Chemistry Summer Review Packet

Assessment of Review Material

Answer all questions completely. If work is needed show all work.

1. How many significant figures are in the following measurement? 2.40 cm _____
2. How many significant figures are in the following measurement? 0.1020 cm _____

Perform the following computations using the measurements in 1 and 2.

3. Multiply
4. Divide (1÷2)
5. Add
6. Subtract

7. Burning paper undergoes what kind of change. Explain.

8. Italian salad dressing is what kind of mixture, homogeneous or heterogeneous? Explain.

Identify the following as a compound, element, or mixture. Completely answer all questions asked in the problem.

9. $MgCO_3$
   a. shows covalent or ionic bonding? ____________
   b. label and name the polyatomic ion. ___________
   c. write the name. __________________________

10. Ag–107
   a. atomic number. ________
   b. number of protons. _________
   c. number of neutrons. ________
   d. number of electrons. ________
   e. what is the name of the group where it is found? _________

11. $CuCl_2$
   a. what is one way we can write the name of this substance? _________________________

12. Aluminum has a density of 2.70 g/mL. If the mass of the block is 24.60 g, find the volume of the substance. Show all work and use sig figs.

13. This material has two isotopes 63 and 65. The relative abundance of each of 69.09% and 30.91% respectively. Calculate the atomic mass to 5 significant figures. Show all work.

AP Chemistry Summer Review Packet

Multiple Choice Questions

1. [(2.30 g – 0.055 g) / 2.30 g] x 100 (non-measurement value) =
    a. 98%
    b. 100%
    c. 97.8%
    d. 97.4%
    e. 99.1%
2. Silver has 2 naturally occurring isotopes, one with a mass of 107 and the other with a mass of 109. The percentage of Ag-107 is 51.82%. What is the atomic mass of all known silver atoms?
    a. 107.9636 amu
    b. 108 amu
    c. 108.2342 amu
    d. 107.5 amu
    e. Can't determine
3. Which one of the following is the molar mass of aluminum hydroxide?
    a. 43.988
    b. 78.004
    c. 97.948
    d. 35.226
    e. 101.348
4. Which of the following is NOT a correct way to interpret the following equation:

    $$2\ H_2O_{(l)} \rightarrow 2\ H_{2\ (g)} + O_{2\ (g)}$$

    a. Liquid water decomposes into hydrogen gas and oxygen gas
    b. 2 molecules of water breaks down into 2 molecules of hydrogen and 1 molecule of oxygen
    c. 2 moles of water decomposes into 2 moles of hydrogen and 1 mole of oxygen
    d. 2 grams of water will give 2 grams of hydrogen and 1 gram of oxygen
    e. All of them are correct
5. 0.45 g of NaCl contain how many moles of chlorine ions?
    a. 0.0038 moles
    b. 0.0039 moles
    c. 0.0077 moles
    d. 130 moles
    e. 0.004 moles
6. Which of these alkaline earth metal oxides has the greatest percent by mass of oxygen?
    a. Barium oxide
    b. Calcium oxide
    c. Beryllium oxide
    d. Magnesium oxide
    e. Strontium oxide
7. Which oxides of manganese, Mn, have percent by mass of manganese that is greater than 50%?

    I. MnO   II. MnO_2 III. Mn_2O_3

    a. II only
    b. III only
    c. I and III only
    d. II and III only
    e. I, II, and III

8. Which pair of samples contains the same number of oxygen atoms in each compound?
    a. 0.10 mol Al_2O_3 and 0.50 mol BaO
    b. 0.20 mol Cl_2O and 0.10 mol HClO
    c. 0.20 mol SnO and 0.20 mol SnO_2
    d. 0.10 mol Na_2O and 0.10 mol Na_2SO_4
    e. 0.20 mol $Ca(OH)_2$ and 0.10 mol $H_2C_2O_4$
9. If 18 grams of X completely reacts with 14 grams of Y to produce 10 grams of M and 2.0 moles of Z, the molar mass of Z is:
    a. 11
    b. 16
    c. 22
    d. 32
    e. 42
10. $2\ Al_{(s)} + 3\ S_{(s)} \rightarrow Al_2S_{3\ (s)}$
    What mass of Al_2S_3 is produced when 3.00 moles of aluminum reacts with excess sulfur according to the equation above?
    a. 81.0 g
    b. 96.0 g
    c. 123 g
    d. 150 g
    e. 225 g
11. $2\ KClO_{3\ (s)} \rightarrow 2\ KCl_{(s)} + 3\ O_{2\ (g)}$
    Which expression gives the mass of O_2 produced when 15 g $KClO_3$ is heated, according to the equation above, in an open vessel until no further weight loss is observed?
    a. $15.0 \times \frac{122.5}{1} \times \frac{2}{3} \times \frac{32}{1}$
    b. $15.0 \times \frac{1}{122.5} \times \frac{3}{2} \times \frac{32}{1}$
    c. $15.0 \times \frac{1}{122.5} \times \frac{3}{2} \times \frac{1}{32}$
    d. $15.0 \times \frac{1}{122.5} \times \frac{2}{3} \times \frac{1}{32}$
    e. $15.0 \times \frac{122.5}{1} \times \frac{3}{2} \times \frac{32}{1}$
12. $Cu_{(s)} + 4\ HNO_{3\ (aq)} \rightarrow 2\ NO_{2\ (g)} + Cu(NO_3)_{2\ (aq)} + 2\ H_2O_{(l)}$
    What volume of $NO_{2\ (g)}$ measured at 1 atm and 273 K can be produced by the reaction of 48 grams of copper with excess concentrated nitric acid according to the equation above?
    a. 11.2 liters
    b. 22.4 liters
    c. 33.8 liters
    d. 44.8 liters
    e. 67.2 liters
13. $2\ C_2H_{6\ (g)} + 7\ O_{2\ (g)} \rightarrow 4\ CO_{2\ (g)} + 6\ H_2O_{(g)}$
    What quantity of reactant remains after ignition of a mixture that contains 0.40 moles of C_2H_6 mixed with 1.60 moles of O_2?
    a. 0.20 mol O_2
    b. 0.80 mol O_2
    c. 1.20 mol O_2
    d. 0.20 mol C_2H_6
    e. 0.30 mol C_2H_6

14. ____$Fe_{(s)}$ + ____$H_2O_{(g)}$ → ____$Fe_3O_{4\ (s)}$ + ____$H_{2\ (g)}$
A mixture in a reaction vessel contains 12 moles of iron and 12 moles of steam. A reaction occurs according to the equation above. After the reaction, all gas escapes from the reaction vessel. Which describes the solid phase remaining after heating has occurred?
   a. 4.0 moles Fe_3O_4 and no excess Fe
   b. 12 moles of Fe_3O_4 and no excess Fe
   c. 2.0 moles of Fe_3O_4 and 6.0 moles excess Fe
   d. 3.0 moles of Fe_3O_4 and 3.0 moles excess Fe
   e. 4.0 moles of Fe_3O_4 and 4.0 moles excess Fe

Free Response

1. A compound is reported to contain 26.6% potassium, 35.3% chromium, and 38.1% oxygen by mass.
   a. Calculate the empirical formula.
   b. If the molecular weight is reported to be 294.2, calculate the molecular formula.
   c. If there is 3.40 moles of this compound, determine how many atoms would be found.

2. ____$ScCl_{3\ (aq)}$ + ____$KOH_{(aq)}$ → ____$Sc(OH)_{3\ (s)}$ + ____$KCl_{(aq)}$
   a. Balance the following.
   b. If 0.60 grams of $ScCl_3$ and 0.15 grams of KOH are mixed together, identify the limiting reactant.
   c. Using the above information, how many grams of $Sc(OH)_3$ will be produced?
   d. Performing the experiment in lab, you collected 0.059 g of $Sc(OH)_3$. Calculate the percent yield.
   e. How many grams of excess reactant will you have?

AP Chemistry Summer Review Packet

One Killer of a Problem: Gold Star if you can get this!

3. Menthol, the substance we can smell in mentholated cough drops, is composed of C, H, and O. A 0.1005 g sample of menthol is combusted producing 0.2829 g of CO_2 and 0.1159 g H_2O. What is the empirical formula for menthol. Show work.