Core Knowledge

Chemistry

Syllabus

Introduction

This syllabus describes a 13-week course that is designed to impart a fundamental understanding of chemical principles to elementary and middle school teachers. The course covers a range of topics similar to those in an introductory level college chemistry class, but with a strong emphasis on teaching the chemical basis for readily observable, everyday phenomena, in a manner that can eventually be conveyed to K-8 school children. The lectures, assigned problems, and examinations are predominantly descriptive in nature. Quantitative manipulation has been limited to a few essential topics such as dimensional analysis, mass-mole calculations, basic stoichiometry, and the pH scale. The week-to-week schedule has been divided by major concepts. The amount of relevant material may not divide precisely with the allotted lecture time, so some flexibility with the lecture schedule is encouraged.

In keeping with the spirit of the course, the laboratory portion consists of both traditional chemistry labs and "demonstration" type labs that can be carried out with household chemicals and other readily available materials. Thus, the students in the class can experience chemical principles first hand, and also carry over some of their activities as demonstrations in their own classrooms without prohibitive expenses or safety concerns. The topics of the labs run approximately 0-2 weeks behind the lecture so that important concepts can be reinforced while still fresh. Each lab period should conclude with a class discussion of the observations and the relevant chemical principles behind the results.

A midterm exam (25%), a final exam (50%), and a lab grade (15%) will make up the bulk of the course grade. There is also an assignment of a short paper (3-5 pages, 10%) that describes the life and scientific contributions of one of the eminent chemists listed in the syllabus. Depending on the class size, it may be possible to have the students give short presentations on their research papers. The homework problems from the text should be assigned for practice, and not be graded.

Overview:

Recommended texts:

Lecture:
Introductory Chemistry, 3rd edition
Charles H. Corwin
Prentice Hall (ISBN 0-13-087470-1)

Note: This text is widely used for introductory level college chemistry courses. It can often be found in university and public libraries. It was chosen because its accompanying laboratory manual seemed most suitable for the course. Some instructors, however, may prefer a different text, and there are many for which this syllabus can easily be adapted. A few possibilities are

Introductory Chemistry, 2nd Edition
Darrell D. Ebbing
R. A. D. Wentworth
ISBN: 0-395-87118-2
Houghton Mifflin

Basic Concepts of Chemistry, 6th Edition
Alan Sherman

Sharon J. Sherman
Leonard Russikoff
ISBN: 0-395-74038-X
Houghton Mifflin

Introduction to Chemistry, 8th Edition
T. R. Dickson
ISBN: 0-471-18042-4
Wiley

Foundations of College Chemistry, 10th Edition
Morris Hein
Susan Arena
ISBN: 0-534-35749-0
Brooks/Cole

Lab:
Laboratory Experiments in Basic Chemistry, 6th edition
Alan Sherman, Sharon J. Sherman & Leonard Russikoff
Houghton Mifflin Company (ISBN 0-395-74039-8)

Schedule

Week 1: Introduction to chemistry and chemical measurement

Week 2: Matter and energy

Week 3: Atomic theory and structure

Week 4: Electronic structure of atoms

Week 5: The periodic table and chemical reactions

Week 6: The mole concept and stoichiometry

Week 7: Chemical bonding

MIDTERM EXAM

Week 8: Gases

Week 9: Solids, liquids, and solutions

Week 10: Reaction rates and equilibrium

Week 11: Acid/base and redox reactions

Week 12: Nuclear chemistry

Week 13: Introduction to biochemistry

FINAL EXAM

Research assignment

Research and write a 3-5 page paper that provides a short biography and a summary of scientific accomplishments for one of the following scientists:

Svante August Arrhenius

Niels Bohr

Marie Curie

John Dalton

Michael Faraday

Dorothy Hodgkin

Antoine Lavoisier

Dmitri Mendeleev

Linus Pauling

Ernest Rutherford

Use of this Syllabus

This syllabus was created by Anthony Bishop, Damon Runyon-Walter Winchell Postdoctoral Research Fellow at The Scripps Research Institute in La Jolla, California, as part of *What Elementary Teachers Need to Know*, a teacher education initiative developed by the Core Knowledge Foundation. Although the syllabus is copyrighted by the foundation, and may not be marketed by third parties, anyone who wishes to use, reproduce, or adapt it for educational purposes is welcome to do so. However, we do ask individuals using this syllabus to notify us so we can assess the distribution and spread of the syllabi and serve as a repository of information about how they may be improved and more effectively used. Please contact Matthew Davis, Core Knowledge Foundation, 801 East High Street, Charlottesville, VA 22902. Phone: 434-977-7550, x. 224. E-mail: mdavis@coreknowledge.org

Syllabus:

<u>Week 1: Introduction to chemistry and chemical measurement</u>

Reading: Corwin Ch. 1-3.8

Homework: End-of-chapter exercises Corwin Ch. 2 & Ch. 3 (except #57-68)

Labs: Sherman Exp. #2: *Use of the Balance: Determining the Densities of Some Common Objects*

Demo Lab: *Oil and Water*

Lecture 1:

Introduction to chemistry, scientific measurement and notation: Students should understand that chemistry helps us understand and model many everyday processes and that it is a central science with direct applications in physics, geology, and biology. Students should learn the importance of scientific measurement, including the metric system, SI units, scientific notation, significant figures, and the difference between precision and accuracy.

Lecture 2:

The mathematical definitions of density, percent composition, and the Celsius and Kelvin temperature scales should be presented. The majority of the period should be spent performing unit conversion problems.

<u>Week 2: Matter and energy</u>

Reading: Corwin Ch. 3.9-4

Homework: End-of-chapter exercises Corwin Ch. 3 (#57-68) and Ch. 4

Labs: Sherman Exp. #3: *Separation of Solids from Liquids*

Demo Lab: *Chromatographic Separation of a Mixture*

Lecture 3:

Students should learn the definition of matter and qualitative descriptions of the solid, liquid, and gaseous states. The difference between physical and chemical changes should be presented, leading into a discussion of mixtures, solutions, and separation processes (filtration, distillation, chromatography). A brief discussion of the laws of conservation of mass and energy will help prepare the students for lecture #4.

Lecture 4:

This lecture should focus on the concept of energy. Kinetic and potential energy should be defined, along with the calorie and the food Calorie. The definition of specific heat should be presented, with particular attention to everyday applications (ex: cooling a metal pan with water). Students will learn that the term enthalpy is used to describe the heat given off (exothermic) or absorbed (endothermic) in a chemical process and that physical processes like melting or evaporation also have associated enthalpy changes.

<u>Week 3: Atomic theory and structure</u>

Reading: Corwin Ch. 5.1-5.5

Homework: End-of-chapter exercises Corwin Ch. 5 (#1-38)

Labs: Sherman Exp. #18: *The Personal Air Pollution Test: The Analysis of Solids in Cigarette Smoke*

Demo Lab: *Preparation of Oxygen*

Lecture 5:

This lecture should focus on atomic theory. A brief history of atomic theory with particular focus on John Dalton's contributions would be helpful. Students should know the difference between elements and compounds and become comfortable with the symbols of common elements as well as simple molecular formulas. Students should learn to relate familiar compounds such as water and NaCl with their molecular structure (molecular vs. ionic compounds). An introduction to electricity will help introduce the following lecture on atomic structure.

Lecture 6:

Students should learn the structure of the atom, including the charges and masses of protons, neutrons, and electrons. Atomic number and mass number should be defined and students should know the meaning of atomic mass. A brief introduction to ions (names, formulas & charges, including polyatomic ions) and isotopes should follow.

<u>Week 4: Electronic structure of atoms</u>

Reading: Corwin Ch. 5.6-5.11

Homework: End-of-chapter exercises Corwin Ch. 5 (#39-100)

Labs: Sherman Exp. #6: *Analyzing Unseen Things (Electrons in Atoms and Objects in "Black Boxes")*

Demo Lab: *Electrons and Magnetism*

Lecture 7:

Students should be introduced to the electromagnetic spectrum and understand how energy corresponds with wavelength. They should understand the Bohr hydrogen model and principle quantum numbers qualitatively. The concept of electron orbitals and the shapes of *s* and *p* orbitals should be introduced.

Lecture 8:

This lecture should focus on atomic orbital configurations. Students should practice writing electron configurations using the Pauli exclusion principle, the Aufbau principle, and Hund's rule. The relationship of atomic size to electronic configuration can be briefly introduced.

Week 5: The periodic table and chemical reactions

Reading: Corwin Ch. 6 & 8

Homework: End-of-chapter exercises Corwin Ch. 6 (#29-36, 71-92), Ch.8 (#19-26)

Labs: Sherman Exp. #8: *The Periodic Table: The Chemistry of Elements Within a Group*

Lecture 9:

This lecture should focus on the rationale and usefulness of the periodic table of the elements. Students should learn the history of the table (Mendeleev), and that elements are now organized by increasing atomic number. Students should learn to identify metals, non-metal, metalloids, and chemically important groups (alkali metals, alkaline earth elements, chalcogens, halogens, noble gases) on the table.

Lecture 10:

Students should learn the definitions of reactants and products and become familiar with writing and balancing chemical equations. A few specific reaction types should be briefly introduced: combustion, acid/base, redox, precipitation.

Week 6: The mole concept and stoichiometry

Reading: Corwin Ch. 9 (except 9.5-9.6) & 10

Homework: End-of-chapter exercises Corwin Ch. 9 (#1-22, 39-82), Ch. 10 (#1-28)

Labs: Sherman Exp. #10: *Types of Chemical Reactions*

Lecture 11:

This lecture should focus on the concept of the mole. Particular attention should be paid to the usefulness of the mole in relating macroscopic and microscopic quantities, and students should learn Avogadro's number. Students should also learn to define and use molar mass and molecular mass.

Lecture 12:

This lecture should primarily focus on determining molar relationships using stoichiometry. Students should grasp that with a balanced equation they can predict the moles (and mass) of an expected product. The concepts of limiting reagent and actual vs. theoretical yield can be presented qualitatively.

Week 7: Chemical bonding

Reading: Corwin Ch. 12

Homework: End-of-chapter exercises Corwin Ch. 12

Labs: Sherman Exp. #5: *Empirical Formula of a Compound*

Demo Lab: *Vinegar and Baking Soda*

Lecture 13:

Building on the ability to generate electron configurations, students should learn that they can predict the reactivity of an atom based on the number of valence electrons. Lewis dot structures represent an atom with its valence electrons. Students should learn that many elements "give up" or "gain" electrons to generate compounds in which the atoms possess noble gas electronic configurations. Certain groups are worth memorizing: halogens form one bond (or gain one electron), chalcogens form two, etc.

Electronegativity should be introduced briefly and qualitatively, with a particular focus on predicting if a compound will be ionic, polar covalent, or non-polar covalent.

Lecture 14:

Most of this lecture should focus on drawing Lewis dot structures for compounds. Students should practice putting together atomic Lewis structures to form diatomics (and bigger molecules). It should be stressed that a correct Lewis dot structure allows us to predict a molecule's shape using VSEPR theory.

Week 8: Gases

Reading: Corwin Ch. 11

Homework: End-of-chapter exercises Corwin Ch. 11 (#1-62)

Labs: Sherman Exp. #16: *Charles's Law: A Look at One of the Gas Laws*

Lecture 15:

This lecture should focus on a qualitative description of the kinetic molecular theory of gases. This should lead to a discussion of the definition of pressure and the relationship between the temperature and kinetic energy of a gas. Students should learn that heavier gases diffuse more slowly at a given temperature and that partial pressures are additive.

***Lecture 16*:**

The majority of this lecture should focus on the gas laws for ideal gases. Boyle's law and Charles' law should be presented with an emphasis on their qualitative ramifications (ex: what happens when you heat an enclosed container?). The students should be made familiar with the terms of the ideal gas law, but proficiency in its quantitative use won't be expected.

Week 9: Solids, liquids and solutions

Reading: Corwin Ch. 13 & 14

Homework: End-of-chapter exercises Corwin Ch. 13 & 14

Labs: Sherman Exp. #11: *Determination of the Amount of Phosphate in Water*

Demo Lab: *Temperature and Solubility*

***Lecture 17*:**

This lecture should present a description of the intermolecular forces that hold together solids and liquids. Students should be made aware that polar molecules (through dipole-dipole interactions) and large molecules (through dispersion forces) often have higher boiling and melting points. Particular attention should be paid to hydrogen bonding and the unique properties that it imparts to water. Students should understand that liquids exert a given vapor pressure at a given temperature and that the liquid boils when the vapor pressure equals the atmospheric pressure. The concept of "like dissolves like" should be introduced with special emphasis on the water solubility and miscibility of common compounds.

***Lecture 18*:**

The definition of molarity should be presented here. Students should be able to perform simple molarity calculations. Colligative properties will be treated descriptively. Emphasis should placed on boiling point elevation and freezing point depression, with real-life examples.

Week 10: Reaction rates and equilibrium

Reading: Corwin Ch. 16.1-16.5

Homework: End-of-chapter exercises Corwin Ch. 16 (#1-34)

Labs: Sherman Exp. #25: *Chemical Kinetics: Rates of Reaction*

***Lecture 19*:**

A brief and qualitative description of collision theory should be presented. Students should learn that the rate of a chemical reaction is governed by several factors, including the concentration of reactants, the temperature, and the activation energy. Catalysts should be defined. The lecture should touch on the practical aspects of controlling reaction rates: cooling food to retard spoilage, catalytic converters, etc.

***Lecture 20*:**

It is important that the students learn about equilibrium at a conceptual level. This lecture should focus on imparting the concept that reactions do not proceed only from left to right, but may reach equilibrium. LeChâtlier's principle should be introduced qualitatively. Students should understand that reaction rate and equilibrium are separate concepts, and that highly favorable reactions can sometimes proceed slowly. Students should practice drawing reaction coordinate diagrams to help decouple the concepts of equilibrium and rate.

Week 11: Acid/base and redox reactions

Reading: Corwin Ch. 15 & 17 (except 17.3-17.4)

Homework: End-of-chapter exercises Corwin Ch. 15 & Ch. 17 (#1-18)

Labs: Sherman Exp. #14: *The Acetic Acid Content of Vinegar*

Demo Lab: *Acid-base Indicators from Red Cabbage*

***Lecture 21*:**

Students should learn the definition of a Brønsted-Lowry acid/base and be able to name a few strong and weak acids/bases. Students should be able to identify an acid or base in a written chemical equation and use the pH scale proficiently. This may require some review of logarithms. The lecture should also describe buffered solutions qualitatively.

***Lecture 22*:**

This lecture should focus on oxidation/reduction reactions. Students should be able to recognize oxidized and reduced ionic species, but oxidation states for neutral species needn't be covered. A fair portion of the lecture should cover batteries and students should be able to sketch basic voltaic and electrolytic cells.

<u>Week 12: Nuclear chemistry</u>

Reading: Corwin Ch. 18

Homework: End-of-chapter exercises Corwin Ch. 18 (#1-32)

Labs: Sherman Exp. #19: *Voltaic and Electrolytic Cells*

***Lecture 23*:**

This lecture should serve as a broad introduction to nuclear chemistry. Students should understand that nuclear reactions are fundamentally different from standard chemical reactions and that the energy changes in nuclear reactions can be much larger. Students should learn the basic types of radiation (alpha, beta, gamma) and that nuclear reactions follow the laws of mass and energy conservation. Students should have the opportunity to see a Geiger counter demo and cloud chamber.

***Lecture 24*:**

This lecture should focus on the applications of nuclear chemistry. Fission and fusion should be defined with an emphasis on their use in energy and weapon production. The concept of half-life should be introduced, using carbon dating as an illustration.

Week 13: Introduction to biochemistry

Reading: Corwin Ch. 20

Homework: No new problems; review old problem sets

Labs: Sherman Exp. #22: *Carbohydrates, Lipids, and Proteins*

Demo Lab: *Curds and Whey*

***Lecture 25*:**

Biochemistry should be introduced with an emphasis on the point that biology is governed by chemical principles. Students should learn that our health is controlled by the rates and equilibria of chemical reactions in our cells. This lecture should introduce the major biomolecules (proteins, DNA, carbohydrates) and their functions. It would be helpful for the instructor to present the chemical structures of a few amino acids, nucleotides, and carbohydrates.

***Lecture 26*:**

Building on the introduction to biochemistry, this lecture should focus on animal respiration and metabolism. Students should learn that animals chemically derive their energy from the O_2 and the carbohydrates that are synthesized by plant photosynthesis. The chemical similarity between respiration and fuel combustion should be pointed out. Also, hemoglobin, and its role in oxygen transportation should be described, so students know a concrete example of protein function.

Lecture notes:

Week 1: Introduction to chemistry and chemical measurement

Lecture 1:

A. Why study chemistry?; chemistry as the central science; relationship to physics and biology

- Chemistry addresses the nature and structure of the world around us.
- Chemistry provides a link between physics and biology. It provides many applications for physical principles, as well as a conceptual framework for understanding biology.
- Like all science, it helps us understand every day occurrences (from ice floating in water to the mechanism of action of penicillin) and makes life more interesting. This understanding enables us to rationally address problems beyond the everyday occurrence. For example, we can use chemistry to generate new antibiotics when bacteria become resistant to penicillin.

B. Measurement; precision and accuracy; significant figures; scientific notation

- Our ability to understand chemistry relies on the quality of our measurements.
- Measurements are subject to imperfections in accuracy (how close is the measured value to a true value?) and precision (how reproducible is the measurement?).

- The term “significant figures” refers to the number of digits that we can measure with confidence. We report measured values with the appropriate number of significant digits.
- Large and small numbers are expressed in scientific notation with the correct number of significant digits. For example, if we know that, to the closest billion, there are 6 billion people in the world, we express that as 6×10^9 people (one significant digit).
- The metric system and SI units are used in chemistry. Some important SI units to keep in mind are mass (kilogram, kg), length (meter, m), time (second, s), temperature (kelvin, K), amount of substance (mole, mol)

Lecture 2:

C. Temperature scales; density

- Kelvin is an absolute temperature scale that is used in chemical calculations. Absolute zero (0 K) is the lowest possible temperature. Celsius temperatures can be converted to kelvin by adding 273.15: $T(K) = [t°C + 273.15]$ K
- Density is the ratio of a substance’s mass to its volume, usually expressed in g/mL. Therefore, the mass and volume of a substance with a known density can always be interconverted by the formula: density = mass/volume. The density of water at 293 K is 1.00 g/mL. Gold has a density of 19.3 g/mL at the same temperature.

D. Unit conversion; dimensional analysis

- Problems in chemistry can often be simplified by interconverting units using dimensional analysis. This is accomplished by using unit factors (fractions that relate a quantity in a certain unit to 1; example: 10^3 m/ 1 km) to cancel units, and to convert to the units requested in a problem.

Week 2: Matter and energy

Lecture 3:

A. Matter; states of matter; types of change, physical and chemical

- Matter has mass and occupies space.
- Matter can exist as solids, liquids, or gases. Solids have defined shapes and volumes, liquids have defined volumes. Gases can be readily compressed and have neither defined shapes nor volumes.
- Substances can undergo physical and chemical changes. Physical changes do not change the chemical composition of the substance [ex: freezing of water; $H_2O(l) \rightarrow H_2O(s)$]. Chemical changes alter the chemical composition of the substance (ex: splitting water; $2\ H_2O \rightarrow O_2 + 2\ H_2$).

B. Mixtures vs. solutions; distillation

- Mixtures of substances that are not uniform throughout and have boundaries between the phases are called heterogeneous. Homogenous mixtures contain one uniform phase. A heterogeneous mixture of a solid and a liquid can be separated by filtration.
- Solutions are a particular type of homogenous mixture. This term is usually used to represent a liquid (solvent) that contains a dissolved substance (solute).
- Mixtures of components with distinct boiling points can be separated by distillation. This is accomplished by heating the mixture to boil away the low

boiling point (volatile) component. The volatile component can then be re-condensed by cooling in a separate flask.

- Chromatography is a process used to separate components of a mixture based on their differing affinities for a solid support.

Lecture 4:

C. What is energy?; kinetic vs. potential

- Energy is the capacity to do work.
- Energy can be have various forms: light, heat, chemical, mechanical, electrical. All of these forms have chemical applications. Chemistry deals predominantly with chemical and heat energy.
- Energy can be stored as potential energy. An object in motion has kinetic energy.

D. 1^{st} law of thermodynamics

- Energy can never be created or destroyed. Much of chemistry deals with understanding how energy is converted from one form to another.

E. Units of measurement; definitions of Joule and calorie

- Energy can be quantified. It is measured often measured in calories (cal) or Joules (J). 1 cal = 4.18 Joules = the amount of energy needed to raise 1 g of $H_2O(l)$ from 14.5°C to 15.5°C; 1 Calorie (food) = 1000 cal

F. Heat capacity and specific heat; comparison of water with other substances

- Different substances require differing amounts of heat for a temperature change. A substance with a high specific heat © requires more heat energy (q) for a change in temperature (ΔT): $C= q/ (mass \times \Delta T)$
- Liquid water has a fairly high specific heat (4.184 J g^{-1} K^{-1}). Compare this to some common metals (ex: C (Fe) = 0.451 J g^{-1} K^{-1}). The high heat capacity of water allows it to temper climates. Bodies of water can absorb large amounts of heat causing only small temperature changes.

G. Definition of enthalpy; enthalpy changes in chemical reactions; endothermic vs. exothermic processes; heats of fusion and vaporization. Demo: $NaOH_{(s)}$ + H_2O, increase temperature, $NaC_2H_3O_{2(s)}$ + H_2O, decrease temperature

- Heat energy can be thought of as a reagent in a reaction. The heat change of a system at constant pressure (the condition of most chemical reactions) is called the change in enthalpy (ΔH). Reactions that release heat are called exothermic ($\Delta H < 0$). Reactions that absorb heat are called endothermic ($\Delta H > 0$).
- Exothermic reactions are generally more favorable (more likely to proceed spontaneously).
- Every compound also has a change in enthalpy associated with its melting and boiling points. Both melting and boiling are endothermic processes. When water condenses, it releases a large amount of heat (exothermic). This process makes steam burns particularly bad.

Week 3: Atomic theory and structure

Lecture 5:

A. History of atomic theory

- The Greek philosopher Democritus (460-370 BC) put forth the idea that matter consists of "uncuttable" fundamental pieces that he called atoms.
- Antoine Lavoisier (1743-194) found that no change in mass occurred when he carried out a chemical reaction. This was called the law of conservation of matter.
- Joseph Louis Proust (1754-1826) built on this principle by showing that a pure substance always contained the same ratio of components. This was called the law of constant composition.
- John Dalton (1766-1844) puts these ideas together and proposed his atomic theory. This states the following: all matter is composed of atoms; atoms are indivisible; all atoms of a given element are identical; atoms of different elements have different masses and properties; compounds are formed by the combination of two or more types of atoms with whole number ratios; atoms are rearranged in chemical reactions, but are not created, destroyed, or converted into other atom types.

B. What is an element?; names and symbols

- An element is a pure substance that consists only of atoms with the same number of protons. There are currently over 100 elements known.

- Each element is represented by a chemical symbol. It is helpful to be familiar with some common symbols, particularly non-obvious ones [ex: sodium (Na), iron (Fe), gold (Au), silver (Ag), lead (Pb)].
- Elements can consist of molecules of two or more atoms. Several elements occur naturally as diatomics (H_2, O_2, N_2).
- Allotropes are different forms of the same element that exist at the same temperature and pressure. Oxygen (ozone), carbon, phosphorous, and sulfur all have interesting allotropes.

C. Elements vs. compounds

- The "heart of chemistry" is understanding how elements combine to form new species.
- A compound is simply a chemical species that is composed of multiple elements in a defined proportion and arrangement.

D. Molecular compounds vs. ionic compounds

- Molecules are groupings of atoms held together by covalent bonds.
- Ionic compounds consist of electrically charged atoms (ions) held together in a lattice by electrostatic forces (positive ions being attracted to negative ions).

Lecture 6:

E. Atomic structure; protons, neutrons, electrons

- Fundamental experiments by Joseph John Thomson (1856-1940), Robert Andrews Millikan (1868-1953), and Ernest Rutherford (1871-1937) helped deduce the structure of the atom.
- Atoms are composed of protons, neutrons, and electrons.

- Protons are positively charged (+1) and have a mass of 1.007276 amu.
- Neutrons have no charge and have a mass almost identical to that of protons (1.008665 amu).
- Electrons are negatively charged (-1) and have a mass that is much smaller than that of protons or neutrons (0.0005485799 amu).
- Protons and neutrons occupy the nucleus (center) of the atom. The nucleus is surrounded by "clouds" of orbiting electrons. Although the nucleus contains the vast majority of an atom's mass, it is very small. The radius of an atom is generally on the order of 10^{-10} m. The nucleus is approximately 100,000 times smaller (10^{-15} m).

F. Atomic number; mass number and atomic mass

- The atomic number equals the number of protons in an atom's nucleus. This defines the element type of the atom.
- Mass number is the sum of the number of protons and neutrons.
- Atomic mass is distinct from mass number. For example a normal hydrogen atom (^{1}H) has a mass number of 1 (1 proton + 0 neutrons). However, this atom has an exact mass of 1.0078 amu.
- These values can be written with the chemical symbol as follows:

mass number → $^{35}_{17}Cl^{-}$ ← charge (if not neutral)

atomic number →

G. Ions; isotopes (radioactivity)

- Atoms with equal numbers of protons and electrons are neutral. If an atom loses electrons it becomes positively charged (cation), if it gains electrons it becomes negatively charged (anion). Ions can also be composed of multiple atoms (polyatomic ions).
- Some elements are found in nature with varying numbers of neutrons. These are called isotopes, and have differing masses.
- Often, certain isotopes of an element have unstable nuclei. These can decay to more stable nuclei, emitting radioactivity. This process is fundamentally different from most chemistry, which purely deals with the rearrangement of electrons and NOT of atomic nuclei.

Week 4: Electronic structure of atoms

Lecture 7:

A. The light spectrum; relationship between energy and wavelength; important wavelengths (UV, visible, γ-rays, etc.)

- Electromagnetic radiation is a form of energy. The energy of a photon (packet) of light is proportional to the frequency and inversely proportional to the wavelength of that light: $E = h\nu = hc/\lambda$
- There is a continuous spectrum of electromagnetic radiation ranging all the way from gamma rays ($\lambda \approx 10^{-12}$ m) to radio waves ($\lambda \approx 10^{0}$ m).
- Visible (white) light consists of a spectrum in the region from $\lambda \approx$ 400-800 nm.
- Surprisingly, when elements are excited by the input of some energy, they emit photons of light with discreet colors, not a continuous spectrum. These line

spectra suggest that only certainly energy levels are accessible for the excited state of an element.

B. The Bohr model for hydrogen; quantum numbers

- Niels Bohr (1885-1962) proposed a model for the simplest atom (hydrogen) that helped explain the observed line spectra. In this model the electron of hydrogen can only occupy discreet, "quantized" energy states.

C. Electron orbitals; shells and subshells

- Each principle quantum number represents a "shell" of orbitals. The first shell contains only a single spherical orbital called *1s*. The second shell contains an *s* orbital (*2s*) and three *p* orbitals ($2p_x$, $2p_y$, $2p_z$). The third shell contains an *s* orbital, three *p* orbitals, and five *d* orbitals. The fourth shell contains an *s* orbital, three *p* orbitals, five *d* orbitals, and seven *f* orbitals.
- An electron can have one of two spin states. The Pauli exclusion principle states that no two electrons in the same orbital can have the same spin. Therefore, each orbital can contain no more than two electrons and each shell can contain up to $2n^2$ electrons [n = 1 (2), n = 2 (8), n = 3 (18), n = 4 (32)].

Lecture 8:

D. Predicting electron configurations: Aufbau, Hund

- In the ground state electrons occupy the orbitals of lowest energy. This order is generally: *1s*, *2s*, *2p*, *3s*, *3p*, *4s*, *3d*, *4p*, *5s*, *4d*, *5p*, *6s*, *4f*, *5d*, *6p*, *7s*, *5f*, *6d*

- Hund’s rule states that electrons prefer to occupy different orbitals in the same subshell with parallel spins. Thus, the ground state of carbon is *1s* [↑↓] *2s* [↑↓] *2p* [↑][↑] and NOT *1s* [↑↓] *2s* [↑↓] *2p* [↑↓]

Week 5: The periodic table and chemical reactions

Lecture 9:

A. Grouping elements

- Different elements can have very different physical and chemical properties. It is important to be able to classify elements to understand their similarities and differences.
- Most metals are solid at room temperature and conduct electricity and heat. They are often ductile and malleable.
- Nonmetals can be solids, liquids, or gases at room temperature. They are generally not good conductors.

B. The periodic table; Mendeleev; utility of periodic table

- Dmitri Ivanovitch Mendeleev (1834-1907) first ordered the elements in a table based on their chemical properties. Astonishingly, he left spaces for elements that were not yet discovered, showing the predictive power of the table.

C. Layout of the table; chemical periodicity; metals, non-metals, metalloids

- Mendeleev arranged his table in order of increasing atomic mass, the modern table is arranged in order of increasing atomic number.

- Metals are found on the left side of the periodic table, non-metals on the right. Metalloids (B, Si, Ge, etc.) form the "border" between the metals and non-metals.

D. Important groups: Group 1A (alkali metals), Group 2A (alkaline earth elements), transition elements, Group 6A (chalcogens), Group 7A (halogens), Group 8A (noble gases).

- Groups (elements in a vertical column) often have very similar chemical reactivities.
- Alkali metals (Li, Na, K, etc.) all react with water to form H_2 and alkaline solutions. They also react with oxygen to give compounds with the formula X_2O, similarly to the non-metal hydrogen.
- The alkaline earth elements have some characteristics in common with the alkali metals. However, they form oxides of the formula XO.
- The halogens are diatomics in their elemental form and they react vigorously with metals to form salts (ionic compounds). NaCl is the major component of table salt.
- All of the Group 8A elements are very unreactive. They are called noble gases because they are found as monatomic gases and do not easily react with "common" elements. The noble gases are not found in great abundance on earth, but He is the second most abundant element in the universe (after H).

E. Relationship of electron configurations to the periodic table

- Elements within a group of the periodic table have very similar outer electron configurations. These configurations of the elements help explain their periodic reactivities.

- Noble gases have filled electron shells and generally do not share, take on, or give up electrons. The alkali metals all have one electron beyond a noble gas configuration ([NG] ns^1). These elements all readily form +1 ions. Likewise, the halogens readily react to form –1 ions. The chalcogens are two electrons short of a noble gas configuration and these elements form compounds in a 1:2 ratio with Group 1A elements (ex: H_2O, H_2S, Li_2O).
- Atomic radius generally decreases proceeding left to right across a row of the periodic table. This is due to the fact that the electrons are subjected to the attraction of an increased effective nuclear charge.

Lecture 10:

A. Chemical equations; balancing equations

- Chemical reactions are written in the form: $2\ A + B_2 \rightarrow 2\ AB$. A and B_2 are reactants. AB is the product.
- Atoms do not change identity in a chemical reaction. Therefore the number of all atom types must be balanced for a reaction to be balanced.

B. Types of reactions: combustion, combination/decomposition, acid/base, oxidation/reduction, precipitation

- All chemical reactions are governed by the same principles, but some types are very common and have been given special names.
- Combustion refers to reacting a species with O_2, generally releasing heat energy. Combination and decomposition reactions simply refer to putting elements

together to make compounds (combination) and taking compounds apart to make elements (decomposition).

- Acidic compounds donate H^+ ions in water. Basic compounds accept H^+ ions in water, forming OH^-. Soluble metal and nonmetal oxides form basic and acidic solutions, respectively.

a. Oxidation/reduction (redox) reactions occur when one reactant loses electrons (is oxidized) and one reactant gains electrons (is reduced).

When ionic compounds are dissolved in water, the ions are surrounded by polar water molecules, and the ionic lattice is pulled apart. Different salts have very different solubilities in water. These tendencies can be used to selectively precipitate a desired ion from solution.

Week 6: The mole concept and stoichiometry

Lecture 11:

E. Intro to the mole; Avogadro's number; molar mass; molecular mass; % composition by mass and mol %

- The concept of a "mole" is used to connect the macroscopic property of mass to the microscopic property of number of atoms.
- A mole is an amount of substance that has been defined arbitrarily. It is defined as the number of atoms in exactly 12 g of ^{12}C.
- 1 mole = 6.022×10^{23} particles. This number is called Avogadro's number after Amedeo Avogadro (1776-1856) who put forth the idea that equal volumes of

gases at the same conditions contained equal number of molecules. This concept will be elucidated further in the section on gas laws.

- Molar mass (M), which is expressed g/mol, is numerically equal to atomic mass (in amu). This relationship allows us to directly convert an element's mass to it's number of atoms, or (more practically) number of moles.
- Molecular mass can be calculated by summing the molar masses of each atom in a compound.

Lecture 12:

C. Stoichiometry; limiting reagent

- Using a balanced equation we can calculate the amount of expected product.
- If one or more reactants are in excess, the reaction will be limited by the reactant that is consumed first. For example, if we burn 1 mole of methane (CH_4) in air from the atmosphere: $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$, we'll only get one mole of CO_2, even though there is lots of O_2 leftover.

D. Theoretical and actual yield

- Theoretical yield is the amount of products we would retrieve if both the reaction and our retrieval methods are perfect. In the above example, the theoretical yield of CO_2 is 1 mole (44 g). Actual yield is the amount we actually retrieve.

- Dividing the actual yield by the theoretical yield gives the % yield, a useful measure for the efficiency of our chemical process.

Week 7: Chemical bonding

Lecture 13:

A. Valence electrons and the octet rule

- The *s* and *p* electrons in an atom's outer shell are known as the valence electrons
- The octet rule states that atoms in the main group elements tend to form bonds to fill the "octet" of their outer *s* and *p* orbitals.
- Practically, this means that metals may lose electrons to form noble gas configuration cations, non-metals may gain electrons to form noble gas configuration anions, or atoms may share electrons in a covalent bond to fill their valence shell.

B. Use of the octet rule to predict molecular and ionic formulas

- Group 1A elements will thus tend to form +1 cations, 2A will form +2 cations.
- Group 7A elements will tend to form -1 anions, 6A will form -2 anions.
- In covalent interactions, the same principle holds; C tends to form 4 bonds, N forms 3 bonds, O forms 2 bonds, F forms 1 bond to fill their respective octets.

C. How do we predict if an interaction is ionic or covalent?; electronegativity; polar and non-polar covalent bonds

- Each element has an intrinsic ability to attract electron density to itself in a covalent bond. This property is called electronegativity and it loosely increases as one proceeds up and to the right of the periodic table.
- Several scales have been developed to quantify electronegativity, but the term only has meaning in comparing two types of atoms. There is no absolute scale.
- Bonds between atoms with very different electronegativities have high ionic character (ex: NaCl). Bonds between atoms of the same electronegativity are

non-polar (ex: Cl-Cl). Bonds between atoms with somewhat different electronegativities are referred to polar covalent (ex: HCl).

- A molecule with polar bonds can be non-polar if the polarities cancel one another (ex: CCl_4).

Lecture 14:

D. Lewis dot structures

- Lewis structures are a useful tool for representing the electronic distribution in simple molecules.
- These structures are generated by rearranging the valence electrons in a molecule until all of the atoms appear to be in their most stable possible state, with filled octets and minimized formal charges. Covalent bonds are represented as dashes, lone electron pairs are represented as a pair of dots.

E. Valence shell electron pair repulsion theory (VSEPR)

- VSEPR is a simple model that allows us to predict the actual shapes of molecules using Lewis dot structures.
- VSEPR assumes that electron pairs (bonded or non-bonded) will move as far away from one another in space as possible due to electrostatic repulsion. Molecular geometry can be predicted based on the number of electron groups around a central atom.
- 2 electron groups $\Rightarrow$ 180°, linear; 3 electron groups $\Rightarrow$ 120°, trigonal planar; 4 electron groups $\Rightarrow$ 109.5°, tetrahedral; 5 electron groups $\Rightarrow$ trigonal bipyramidal; 6 electron groups $\Rightarrow$ octahedral

- Tetrahedral carbon is very important in all organic compounds (pharmaceuticals, etc) as well as all life forms (protein, DNA structure).

Week 8: Gases

Lecture 15:

A. What's a gas?; kinetic molecular theory; Graham's law

- Gases are forms of matter in which the atoms or molecules interact with one another only very weakly and transiently. They are compressible and have low densities. Gases fill containers and exert pressure on the container uniformly.
- The kinetic molecular theory of gases described these characteristics more precisely by making the following approximations: (i.) the molecules of a gas are widely spaced and occupy a negligible amount of volume (ii.) gas molecules are in rapid motion, and collide with each other and the container walls elastically (iii.) the intermolecular forces of a gas are negligible (iv.) different gases have the same kinetic energy at the same temperature.
- The diffusion (or effusion) rates of two gases are inversely proportional to the square root of their molecular masses.

Lecture 16:

B. How gases are affected by pressure and temperature; gas laws

- The pressure of a gas is defined as the force it exerts per unit area. Pressure is usually expressed in the units atm or torr (1 atm = 760 torr).

- Boyle's law states that (at constant temperature and moles) the pressure of a gas and its volume are inversely proportional: $P_1V_1 = P_2V_2$
- Charles's law states that (at constant pressure and moles) the volume of a gas and its temperature are proportional: $T_2V_1 = T_1V_2$
- Also, Gay-Lussac's law states that (at constant volume and moles) the temperature and pressure of a gas are proportional: $T_2P_1 = T_1P_2$
- Combining these laws yields: $P_1V_1/T_1 = P_2V_2/T_2$
- Avogadro discovered that (at constant pressure and temperature) the volume of a gas is proportional to the number of moles of the gas (n): $V_1n_2 = V_2n_1$
- All of these relationships can be combined into the ideal gas law: $PV = nRT$, where R is a constant. This relationship is true for gases that fit the kinetic molecular theory approximations well.
- It follows from our assumptions that in a mixture of gases the total pressure is simply the sum of the individual pressures of the gases: $P_{tot} = P_1 + P_2 + \ldots P_N$

Week 9: Solids, liquids and solutions

Lecture 17:

A. Properties of solids and liquids; special properties of water

- Substances with substantial intermolecular forces at a given temperature form solids or liquids.
- Most compounds contract upon freezing. Water expands upon freezing due to a complex network of hydrogen bonds.

B. Intermolecular forces

- Attractive forces that hold discreet molecules together in a solid or liquid are called intermolecular forces. Generally, the stronger the intermolecular forces, the higher the melting and boiling points of a species.
- Dipole-dipole interactions: polar molecules can arrange themselves such that the partial positive charge of one molecule is close in space to the partial negative charge of another molecule.
- van der Waal's (London) forces: even non-polar molecules are attracted to one another through instantaneous dipole-induced dipole interactions. Because electrons are moving, every molecule has instantaneous dipoles that can induce dipoles in adjacent molecules. These forces generally are more significant in larger molecules, because the electron clouds are more polarizable.
- Hydrogen bonds: a special type of electrostatic intermolecular force between a hydrogen atom that is bonded to an electronegative atom, and a non-bonded electron pair on another electronegative atom. This interaction is extremely important in life. It helps give water a very high boiling point (for its molecular weight) and it helps define the structure of proteins and DNA.

C. Phase changes; vapor pressure; phase diagrams; heats of fusion, vaporization

- Liquids exert a vapor pressure that increases with increasing temperature. The boiling point is reached when the vapor pressure equals the pressure of the atmosphere.
- Every compound has a characteristic phase diagram that represents its phase behavior at varying pressures and temperatures. The phase diagram of water stands out because increasing pressure lowers the freezing temperature of water.

D. Solutions; miscibility and solubility; temperature effects on solubility

- Solutions consist of a solute dissolved in a solvent.
- "Like dissolves like": a polar solvent such as water is a good solvent for polar compounds, such as glucose and many salts. Less polar organic solvents (ex: hexane, $CHCl_3$) are generally not miscible with water and they are not good solvents for polar compounds. They cannot overcome the ionic or dipole-dipole interactions that hold the solid lattice together. However, organic solvents are generally good solvents for non-polar organic compounds.
- The solubility of solids and liquids usually increases with increasing temperature.

Lecture 18:

E. Measuring concentration; molarity

- The concentrations of solutions are generally defined in % by mass or molarity.
- Molarity (M) = mols solute/ liter of solution

F. Conductivity of solutions; colligative properties of solutions

- Solutions have different physical properties than pure liquids.
- Electrolytes: ionic compounds (and a few non-ionic, ex: HCl) dissociate to yield ions in water. Such solutions conduct electricity much better than pure water.
- Boiling point, freezing point, and osmotic pressure are colligative properties of solutions, meaning that they depend on the relative amounts of solute and solvent, NOT the chemical characteristics of the solute.
- Increased solute concentration lowers the vapor pressure (raises the boiling point) of a solution by reducing the number of volatile molecules at the

solution/atmosphere interface. Thus, spaghetti cooks faster in boiling salt water (higher boiling point) than in boiling tap water.

- Freezing points are depressed with increasing solute concentration because the solute competes with the solvent molecules at the solid/liquid interface. Thus, spreading salt on the sidewalk keeps ice from forming.

Week 10: Reaction rates and equilibrium

Lecture 19:

A. Collision theory; effects of temperature and concentration on rates of reactions

- For a reaction between two molecules to take place, they must collide with sufficient energy and the proper orientation.
- Raising the temperature of a reaction mixture, increases the average kinetic energy of the molecules, increasing the proportion of collisions with sufficient energy. This speeds up the reaction.
- Increasing the concentration of the reactants increases the frequency of the collisions, also increasing the rate of the reaction.

B. Activation energy; catalysts

- In a bimolecular, one step reaction ($A_2 + B_2 \rightarrow 2\ AB$), the A_2 and B_2 molecules collide. Energy is required to start the breaking of the A-A and B-B bonds. This leads to high energy transition state (or activated complex). As the formation of the of the A-B bonds proceeds, the complex is stabilized, leading to the products.
- The difference in energy between the reactants and transition state is called the activation energy. The higher the activation energy, the slower the reaction.

- A catalyst is a substance that speeds up a reaction, without being used up in the reaction. Catalysts work by stabilizing the transition state complex, thereby lowering the activation energy.
- Catalysts play very important roles in atmospheric chemistry (Cl atoms catalytically destroy ozone) and biochemistry (enzymes).

Lecture 20:

C. Reversible reactions; equilibrium; LeChâtlier's principle

- Until now, we have been showing reactions in one direction. In fact, reactions proceed both forwards and backwards:

$$A_2 + B_2 \rightleftharpoons 2\ AB$$

- A reaction has reached equilibrium when the rates of the forward and backward reactions are equal. Thus, the concentrations of the reactants and products are not changing.
- LeChâtlier's principle states that when a stress (chemical, temperature, pressure) is applied to a system at equilibrium, the system will react to partially counteract the stress. In other words, if the reactant concentration is artificially increased, more product will be formed as part of the added reactant is used up.

D. Reaction coordinate diagrams

- The relative stabilities of reactants, products, and transition states can be shown clearly through reaction coordinate diagrams. Favorable reactions are overall "downhill." Unfavorable reactions are overall "uphill." However, the activation energy (thus, the rate) is independent of the free energy change of the reaction. Reaction coordinate diagrams show that a reaction can be like crossing a mountain range. Walking from Yosemite Valley to San Francisco, is "downhill" (favorable), but it will be very slow because of the activation energy of climbing over the Sierras.

Week 11: Acid/base and redox reactions

Lecture 21:

A. Definition of Brønsted-Lowry acid; strong vs. weak acids; polyprotic acids and bases

- Acids are defined as compounds that dissociate to release H^+ when dissolved in water. Bases accept an H^+ and form OH^- in aqueous solution. In every acid reaction there is also something acting as a base (usually water) and vice versa.
- Acids and bases neutralize one another to yield water and dissolved salts.
- $HCl + H_2O \rightarrow H_3O^+ + Cl^-$: HCl is the acid. H_2O is the base. H_3O^+ is the conjugate acid. Cl^- is the conjugate base.
- Strong acids (ex: HCl, HI, HNO_3) essentially dissociate completely. Weak acids dissociate partially.
- Polyprotic acids (H_3PO_4, H_2SO_4) can donate more than one H^+ in solution.
- Water can act as an acid and a base and can autoionize. The ion product ($[H_3O^+][OH^-]$) is a constant, 1.0×10^{-14}.

B. Simple acid/base equilibrium calculations; pH scale

- The pH is a useful measure for the acidity or basicity of a solution: pH < 7 is considered an acidic solution: $pH = -\log[H_3O^+]$, $pOH = -\log[OH^-]$.
- For any aqueous solution: $pH + pOH = 14.00$

C. Qualitative understanding of buffers

- Buffers are solutions of weak acids and their conjugate bases. These are useful because they resist significant pH changes, even when a strong acid or base is added.

Lecture 22:

D. Definitions of oxidation and reduction

- Every oxidation reaction has an associated reduction.
- The oxidized species loses electrons; it is the reducing agent (reductant). The reduced species gains electron; it is the oxidizing agent (oxidant). Molecular oxygen (O_2) is a common oxidant.
- The oxidations and reduction can be written separately as half-reactions.

E. Oxidation state; recognizing redox reactions

- Redox reactions can be identified by finding changes in oxidation states.
- Oxidation numbers are simply a bookkeeping method for keeping track of electrons. An oxidation number of +3, does NOT signify a +3 ion.
- Elements have the oxidation number 0.
- Monatomic ions have the oxidation number of their ionic charge.
- Redox reactions are slightly more complicated to balance because we must make sure that charge (electrons) are conserved, as well as atom types.

F. Qualitative understanding of batteries; voltaic and electrolytic cells

- Redox reactions are very commercially important because they are used in batteries.

- Voltaic cells use a favorable redox reaction to generate an electrical current. The oxidation vessel contains the anode and it is separated from the reduction vessel which contains the cathode. A salt bridge allows movement of ions between the compartments.
- Electrolytic cells use electrical energy to drive a non-favorable chemical reaction. Such cells can be used to “split” water into hydrogen and oxygen, or plate precious metals onto a less expensive metal.

Week 12: Nuclear chemistry

Lecture 23:

A. Introduction to nuclear chemistry; types of radiation

- Normal chemical reactions deal only with the rearrangement of electrons. Nuclear reactions deal with the rearrangement of atomic nuclei. Generally, the amount of energy absorbed and/or released in nuclear reactions is much higher than in chemical reactions.
- There are several types of emitted radiation in common nuclear reactions. α-particles are helium nuclei. β-particles are electrons. An emitted electron often results from the conversion of a neutron into a proton. Positrons (positively charged electrons) can also be emitted, converting a proton to a neutron. γ-rays are high energy photons that can be emitted by themselves or in combination with other emission events.
- A proton can also be converted into a neutron through electron capture

B. Balancing nuclear reactions

- The considerations in balancing nuclear reactions are different from chemical reactions. We still have to balance mass and charge, but the elements are changing in the process. The mass numbers and atomic numbers must be balanced, a point we take for granted in chemical reactions:

^{14}C decay: $^{14}_{6}\mathrm{C} \longrightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}\beta$

fission: $^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \longrightarrow {}^{141}_{56}\mathrm{Ba} + {}^{92}_{36}\mathrm{Kr} + 3\ {}^{1}_{0}\mathrm{n}$

fusion: $4\ {}^{1}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + 2\ {}^{0}_{+1}\beta$

Lecture 24:

C. Half-life; carbon dating

- Unstable isotopes have characteristic half-lives ($t_{1/2}$), in which half the material decays.
- ^{14}C has a $t_{1/2}$ of 5,730 years and is normally present at about 14 dpm per gram of carbon in living organisms. After an organism dies, it no longer takes in ^{14}C, and the ^{14}C activity of its biomass decreases. This allows researchers to use the ^{14}C activity to date materials that derived from living things.

D. Nuclear fission and fusion

- Two fundamentally different types of nuclear reactions are important in generating energy.
- Nuclear fission refers to the splitting of heavy nuclei into lighter ones. ^{235}U, which was used in the nuclear bombing of Hiroshima, can undergo a nuclear chain reaction. The fission is started by neutron bombardment. The fission of

one ^{235}U nucleus leads to the emission of 3 more neutrons which can cause the splitting of adjacent nuclei.

- Fusion involves bringing light nuclei together to make heavy ones. The fusion reaction of hydrogen atoms to helium is fundamental to life, as it is the source of most of the sun's energy. It is very difficult to harness the power of fusion for our energy needs, because extremely high temperatures are needed.

Week 13: Introduction to biochemistry

Lecture 25:

A. What is biochemistry?; is it really chemistry?

- The chemistry going on in living cells is governed by the exact same principles (energy change, rates, equilibrium) as reactions in test tubes.
- The fundamental questions are: what are these reactions and how are they controlled?

B. Important biochemicals: proteins, DNA, RNA, carbohydrates, lipids

- Proteins, DNA, and RNA are all biopolymers.
- Proteins are composed of twenty different amino acids. The standard amino acids contain C, H, N, O, and S. These amino acid chains fold to give shaped proteins with precise cellular functions. Proteins are the workhorses of the cell, controlling the majority of biochemical processes. Many cellular proteins are

enzymes, which are simply protein catalysts. These function just like other chemical catalysts, stabilizing the transition state of a particular reaction.

- DNA and RNA are each made up of four nucleotide building blocks. Chemically, there are very small differences between DNA and RNA, but these give rise to large differences in the biological function. DNA's primary function is to hold the genetic information of the cell. This information encodes the sequences of the functional proteins. RNA has several different roles. Messenger RNA (mRNA) relays the DNA sequence information to the site of protein synthesis (the ribosome). Transfer RNAs (tRNA) physically bring the amino acid building blocks together on the ribosome.
- Carbohydrates are molecules composed of C, H, and O which are crucial for cellular energy storage. The oxidation of carbohydrates is the primary energy-generating process in metabolism. Carbohydrates also play critical structural roles. Cellulose (a polymer of glucose) is the major carbohydrate in woody plants and is the most abundant polymer in the biosphere.

Lecture 26:

C. Animal respiration and metabolism

- Animals chemically derive their energy from the O_2 and carbohydrates that are synthesized by plant photosynthesis.
- Air is about 21 mol % oxygen. When animals breath, oxygen enters the lungs and is absorbed into the blood. Red blood cells contain a protein called hemoglobin that binds O_2 by surrounding it with amino acids and an iron-containing organic

molecule called heme. In this way, oxygen can be transported to cells throughout the body.

- The chemistry of animal respiration is essentially a combustion reaction. Carbohydrates are oxidized with O_2 to yield CO_2, H_2O, and energy. Unlike burning gasoline, the process is broken down into many individual chemical steps that are catalyzed by enzymes.
- Respiration can be defined as the oxidative breakdown and energy release from nutrients. The overall balanced reaction for oxidation of glucose is:

  $C_6H_{12}O_6 + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$

- This reaction yields 2880 kJ/mol of energy. This energy is used to convert adenosine diphosphate (ADP) into adenosine triphosphate (ATP). ATP is the energy "currency" that cells use in many reactions. Thus, it is more complete to write the respiration reaction as follows:

$C_6H_{12}O_6 + 36\ ADP + 36\ H_3PO_4 + 6\ O_2 \rightarrow 6\ CO_2 + 36\ ATP + 42\ H_2O$

Chemistry Lab Syllabus

Week #1
Quantitative Lab: Sherman Exp. #2: *Use of the Balance: Determining the Densities of Some Common Object*
Demo Lab (see below): *Oil and Water*

Necessary items:
- two small glasses
- water
- vegetable oil
- ice

Procedure:
1.) Pour about 150 mL of water in one glass and 150 mL of vegetable oil in another.
2.) Place an ice cube into each glass. Observe.
3.) Remove the ice and pour the oil into the glass of water. Observe.

Discussion questions:
1.) What can you conclude about the relative densities of water, ice, and vegetable oil?
2.) Is there anything surprising about your results?
3.) The density of vegetable oil is roughly 0.8 g/mL. Provide a rough estimate for the density of ice.

Week #2
Quantitative Lab: Sherman Exp. #3: *Separation of Solids from Liquid*
Demo Lab (see below): *Chromatographic Separation of a Mixture*

Necessary items:
- drinking glass
- water
- paper towels
- pencil
- water-soluble markers

Procedure:
1.) Cut a paper towel into a circle. Fold it in half twice to form a shape resembling 1/4 of a pie.
2.) Make a mark with a water-soluble marker about 2 cm from the point of the "pie slice".
3.) Place about 2.5 cm of water into a low, wide juice glass.

4.) Stick a pencil through the marked paper towel close to the curved portion. Place the pencil on the rim of the glass so that the paper towel is suspended with the point immersed in the water, but with the marked area ABOVE the water level.
5.) Observe what happens as the water moves up the paper towel. When the water is almost to the top of the paper towel, remove the paper towel from the water to stop the experiment. What do you see?
6.) Repeat the experiment with various color pens.

Discussion questions:
1.) Is ink a pure compound? How many components could you detect?
2.) Are there attractive forces between the water and the paper? Between the ink and the paper? Between the water and the ink?
3.) What are some possible applications for this type of "chromatographic" separation?

Week #3

Quantitative Lab: Sherman Exp. #18: *The Personal Air Pollution Test: The Analysis of Solids in Cigarette Smoke*
Demo Lab (see below): *Preparation of Oxygen*

Necessary items:
- household liquid bleach (USE WITH CAUTION IN A VENTILATED AREA)
- household hydrogen peroxide (3% solution, USE WITH CAUTION)
- several small glasses
- plastic wrap
- matches
- toothpicks

Procedure:
1.) Pour about 1 cm of bleach into the glass.
2.) Wrap the top of the glass with plastic wrap.
3.) Add two tablespoons of the hydrogen peroxide solution to another glass.
4.) Quickly fold back a piece of the plastic wrap over the bleach and add the hydrogen peroxide solution. Replace the plastic wrap.
5.) Light a toothpick on fire. Blow out the flame so that the toothpick is glowing hot.
6.) Quickly fold back a piece of the plastic wrap and insert the toothpick. Observe. Repeat.

Discussion questions:
1.) What are the molecular formulas for hydrogen peroxide and oxygen?
2.) Formulate a hypothesis that addresses your observations.
3.) In this experiment you generated oxygen gas. List a few properties of oxygen gas that you knew already or that you learned from this experiment.

Week #4

Quantitative Lab: Sherman Exp. #6: *Analyzing Unseen Things (Electrons in Atoms and Objects in "Black Boxes")*
Demo Lab (see below): *Electrons and Magnetism*

Necessary items:
- 9 volt alkaline battery
- twist ties
- scotch tape
- iron paper clips, cut into small (3cm and 0.5 cm) pieces

Procedure:
1.) Peel the paper portion back from the edges of a twist tie. Coil the covered portion around one of the 3 cm paper clip lengths.
2.) Twist the exposed ends of the twist tie around the contacts of the 9 volt battery and tape them in place.
3.) See if you can pick up small the small pieces of paper clips with your electromagnet.
4.) After about 1-2 minutes of connection, remove the 3 cm paper clip length from the twist tie.
5.) Test the paper clip length to see if it can behave as a magnet. Compare the magnetized length of paper clip with a non-magnetized length of paper clip.

Discussion questions:
1.) Find iron on the periodic table. Make some guesses concerning its electron configuration. Does it have unpaired electrons?
2.) What is the electronic basis for magnetic compounds? Why do we think of metals having stronger magnetic properties than non-metals?
3.) What was the purpose of the battery?

Week #5
Quantitative Lab: Sherman Exp. #8: *The Periodic Table: The Chemistry of Elements Within a Group*

Week #6
Quantitative Lab: Sherman Exp. #10: *Types of Chemical Reaction*

Week #7
Quantitative Lab: Sherman Exp. #5: *Empirical Formula of a Compound*
Demo Lab (see below): *Vinegar and Baking Soda*

Necessary items:
- vinegar
- baking soda
- a narrow-necked bottle (50-100 mL capacity)
- latex balloon

Procedure:

1.) Pour about 5 teaspoons of vinegar into the bottle.
2.) Carefully spoon 1 teaspoon of baking soda into the latex balloon.
3.) Slip the open end of the balloon over the neck of the bottle.
4.) Invert the balloon and lightly shake the bottle to make sure that all of the baking soda falls into the vinegar. Observe.
5.) Repeat the same experiment with the same quantity of baking soda and varying quantities of vinegar. Observe.
6.) Repeat with the same quantity of vinegar and varying quantities of baking soda. Observe.

Discussion questions:

1.) What did you observe when you added the baking soda? What reaction is taking place?
2.) What happened when you changed the amounts of vinegar and baking soda? Did the balloon inflate more or less?
3.) What was the limiting reagent under the conditions of the initial experiment?

Week #8

Quantitative Lab: Sherman Exp. #16: *Charles's Law: A Look at One of the Gas Laws*

Week #9

Quantitative Lab: Sherman Exp. #11: *Determination of the Amount of Phosphate in Water*

Demo Lab (see below): *Temperature and Solubility*

Necessary items:

- aspirin tablets
- several small glasses
- coffee filters
- cold water
- boiling hot water
- ice

Procedure:

1.) Crush two aspirin tablets and place the powder in a glass. Repeat with two more aspirin tablets in a second glass.
2.) Add 100 mL of cold water to one of the glasses. Stir vigorously. Record your observations.
3.) Add 100 mL of hot water to the other glass. Stir vigorously. Record your observations.

4.) While the solution in glass #2 is still hot, pour it through a coffee filter into a third glass.
5.) Let the solution in glass #3 cool for 30-60 minutes. Record your observations.
6.) If you don't observe anything besides cooling, add a tiny crumb of crushed aspirin to glass #3 and observe. If you still do not observe anything, immerse the glass in a bowl of ice water.

Discussion questions:
1.) Describe the crystals you observed. What do you think they are composed of?
2.) What can you conclude about the solubility of aspirin in hot and cold water? Do you think your conclusions hold true for other solids? Liquids? Gases? Why or why not?
3.) What was the purpose of the tiny crumb of aspirin? Of the ice water?
4.) Recrystallization is often used to purify organic compounds like aspirin. How do you think this works?

Week #10
Quantitative Lab: Sherman Exp. #25: *Chemical Kinetics: Rates of Reaction*

Week #11
Quantitative Lab: Sherman Exp. #14: *The Acetic Acid Content of Vinegar*
Demo Lab (see below): *Acid-base Indicators from Red Cabbage*

Necessary items:
- chopped red cabbage
- a pan and stove
- small glasses
- water
- saturated solution of baking soda
- milk of magnesia
- white vinegar
- laundry detergent
- seltzer water
- aspirin tablets
- household ammonia (USE WITH CAUTION IN A VENTILATED AREA)
- saliva
- pH paper

Procedure:
1.) Place 1 cup of chopped red cabbage in a pan with 250 mL of water.
2.) Bring the water to a boil and simmer the cabbage for 5 minutes.

3.) Let the pan cool. While the pan is cooling, you can start setting up the solutions in step #6.
4.) Decant the liquid into a clean glass, leaving the cabbage behind. Use the cabbage juice as soon as possible.
5.) Prepare a data table so you can record the color of each test.
6.) Set up nine different glasses containing the following items:
    (i.) 100 mL plain water
    (ii.)100 mL vinegar
    (iii.) 100 mL seltzer water
    (iv.) 100 mL saturated solution of baking soda
    (v.) 100 mL of household ammonia solution
    (vi.) 1 teaspoon laundry detergent dissolved in 100 mL of water
    (vii.) a few drops of milk of magnesia dissolved in a teaspoon of water
    (viii.) four crushed aspirin tablets stirred in 100 mL warm water
    (ix.) a small amount of saliva
7.) To each of your glasses (one glass at a time) add 1 teaspoon of your cabbage indicator solution Stir and closely observe each color change.
8.) Perform an acid-base titration. Place 100 mL vinegar and two teaspoons of the indicator solution in a glass. Add one teaspoon of milk of magnesia and stir well. Repeat, carefully observing the color after each milk of magnesia addition. Stop when you no longer can detect a color change.

Discussion questions:
1.) Make a chart of your colors for each indicator solution. Which solutions do you think were acidic and which were basic?
2.) The pH for these solutions increase in the following order: vinegar (pH 2.5), seltzer water (pH 4), saliva (pH 7), saturated baking soda (pH 8), milk of magnesia (pH 10), laundry detergent (pH 11). Add pH values to your chart from question #1. Estimate the pH for your aspirin and ammonia solutions.
3.) Calculate the concentration of H^+ ions in saturated baking soda.

Week #12
Quantitative Lab: Sherman Exp. #19: *Voltaic and Electrolytic Cells*

Week #13
Quantitative Lab: Sherman Exp. #22: *Carbohydrates, Lipids, and Proteins*
Demo Lab (see below): *Curds and Whey*

Procedure:
1.) Pour about 150 mL of milk into a glass.
2.) Add two tablespoons of vinegar to the milk. Stir. Let it set for 5 minutes. Observe.
3.) Filter the milk to remove any solid material. Pat the solid material with a paper towel to dry it. Observe.

Discussion questions:

1.) Is milk an element? A compound? A solution? A heterogeneous mixture?
2.) Did the status change after the addition of vinegar?
3.) Formulate a hypothesis concerning the composition of the filtered solid material.

Chemistry

Midterm Exam

This exam is designed for a 90 minute examination period. Time pressure should not be a factor. A periodic table of the elements is attached to the back of the exam. Please pay attention to significant digits when reporting numerical answers.

Constants and relationships that may or may not be helpful:
1 pound = 16 ounces = 453.59 g
Avogadro's number: N = 6.02×10^{23} mol^{-1}
Gas constant: R = 8.31 J K^{-1} mol^{-1} = 0.821 L atm K^{-1} mol^{-1}
Planck's constant: h = 6.63×10^{-34} J sec
Speed of light (in a vacuum): c = 3.00×10^{8} m sec^{-1}
1 cal = 4.184 J
1 atm = 760 torr
K = °C + 273.15 °C
Density of water: 1.0 g mL^{-1}
specific heat of water: 4.18 J g^{-1} K^{-1}

Question 1: 12 pts
Question 2: 9 pts
Question 3: 9 pts
Question 4: 9 pts
Question 5: 7 pts
Question 6: 10 pts
Question 7: 9 pts
Question 8: 10 pts
Question 9: 9 pts
Question 10: 16 pts

Total possible: 100 pts

Question 1 (12 points, 3 points each):

List the number of protons, neutrons, and electrons in each of the following:

a. ^{241}Am

b. $^{71}Ga^{3+}$

c. ^{13}C

d. $^{131}Xe^{4+}$

Question 2 (9 points):

What was incorrect about Mendeleev's original concept of the periodic table? Explain how the modern periodic table differs from Mendeleev's.

Question 3 (9 points):

Explain why people often feel cold when they get out of a swimming pool or shower. Be sure that the word "enthalpy" is in your answer.

Question 4 (9 points):

You are shipwrecked on a Pacific island and you miraculously discover an old, deserted chemistry lab. Explain how you will use the lab to make the ocean salt water drinkable.

Question 5 (7 points):

Two crystalline forms of white phosphorous are known Both forms contain P_4 molecules, but the molecules pack in different arrangements. The α form is obtained when the liquid freezes at 44.1 °C. However, below –76.9 °C the α form spontaneously converts to the β form. For the reaction as written, which of the following are true (circle all that apply)?

$$P_4\,(\alpha \text{ form}) \rightleftharpoons P_4\,(\beta \text{ form})$$

a. $\Delta H > 0$

b. $\Delta H < 0$

c. $\Delta H = 0$

d. all of the above

e. none of the above

Question 6 (10 points)

a. In your own words, define the term "photon". What is the relationship between a photon's energy and its wavelength?

b. List three types of electromagnetic radiation in order of **INCREASING** energy.

Question 7 (9 points, 3 points each):

Balance the following chemical equations:

a. $(CH_3)_2N_2H_2(l) \quad + \quad N_2O_4(l) \quad \rightarrow \quad N_2(g) \quad + \quad H_2O(g) \quad + \quad CO_2(g)$

b. $CaNCN(s) \quad + \quad H_2O(l) \quad \rightarrow \quad CaCO_3(s) \quad + \quad NH_3(g)$

c. $C_{16}H_{34}\,(l) \quad + \quad O_2(g) \quad \rightarrow \quad CO_2(g) \quad + \quad H_2O(g)$

Question 8 (10 points):

a. Name (not just chemical symbols) two transition elements, two halogens, two metalloids, and two alkali metals.

b. For the halogens that you listed in part a, write a molecular formula for a stable, neutral compound that can be formed by the halogen (one molecular formula for each halogen).

Question 9 (9 points):

An average baseball bat weighs 30 ounces. How many kilograms does a mole of baseball bats weigh?

Question 10 (16 points):

a. Predict the ground state electron configuration of a neutral silicon atom.

b. Predict the electron configuration of the LOWEST ENERGY excited state of a neutral silicon atom.

c. Predict the value of x for a stable, neutral compound: $SiCl_x$. In 25 words or less, describe why you picked your value of x.

Chemistry

Final Exam

This exam is designed for a 3 hour examination period. Time pressure should not be a factor. A periodic table of the elements is attached to the back of the exam. Please pay attention to significant digits when reporting numerical answers.

Constants and relationships that may or may not be helpful:
Avogadro's number: N = 6.02×10^{23} mol^{-1}
Gas constant: R = 8.31 J K^{-1} mol^{-1} = 0.821 L atm K^{-1} mol^{-1}
Planck's constant: h = 6.63×10^{-34} J sec
Speed of light (in a vacuum): c = 3.00×10^{8} m sec^{-1}
1 cal = 4.184 J
1 atm = 760 torr
K = °C + 273.15 °C
Density of water: 1.0 g mL^{-1}
specific heat of water: 4.18 J g^{-1} K^{-1}

Question 1: 8 pts
Question 2: 10 pts
Question 3: 16 pts
Question 4: 12 pts
Question 5: 10 pts
Question 6: 8 pts
Question 7: 8 pts
Question 8: 8 pts
Question 9: 10 pts
Question 10: 14 pts
Question 11: 14 pts
Question 12: 10 pts
Question 13: 10 pts
Question 14: 8 pts
Question 15: 8 pts
Question 16: 14 pts
Question 17: 16 pts
Question 18: 16 pts

Total possible: 200 pts

Question 1 (8 points):

Assume that the water is already boiling. Which option (a or b) takes longer to cook? Briefly explain your answer

a. Hard-boiling an egg on top of Mount Everest

b. Hard-boiling an egg at the bottom of Death Valley

Question 2 (10 points):

Briefly describe how Carbon-14 is used to date ancient organic materials.

Question 3 (16 points):

Mercuric oxide (HgO) decomposes upon heating to afford elemental mercury and elemental oxygen.

a. Write a balanced chemical equation for this reaction.

b. If I want to produce 25 g of mercury using this process, how many grams of mercuric oxide do I need?

c. I discover that I can collect the oxygen and recycle it in another process in my chemical plant. How grams of oxygen do I get when I make my 25 g of mercury?

Question 4 (12 points):
Balance the following NUCLEAR reactions by filling in the question marks. Be sure to include the mass number, atomic number, and chemical symbol for each entry.

a. $${}^{96}_{42}\mathrm{Mo} + {}^{2}_{1}\mathrm{H} \longrightarrow {}^{1}_{0}\mathrm{n} + \ ?$$

b. $${}^{246}_{96}\mathrm{Cm} + {}^{12}_{6}\mathrm{C} \longrightarrow 4\ {}^{1}_{0}\mathrm{n} + \ ?$$

c. $$? \longrightarrow {}^{24}_{12}\mathrm{Mg} + {}^{0}_{-1}\beta$$

Question 5 (10 points):

Water is a very unique compound. As you probably know by now, its molecular formula is H_2O and its molecular weight is 18 g/mol.

a. Draw a Lewis dot structure of water. Are the chemical bonds of water ionic, non-polar covalent, or polar covalent?

b. Discuss the melting and boiling points of water. Are they unusual (high or low) for an 18 g/mol molecule? If so, what intermolecular force(s) governs the unique melting and boiling characteristics of water ?

Question 6 (8 points):

a. What volume of 0.123 M NaOH contains 25.0 g NaOH?

Question 7 (8 points)

If you need 300 mL of 0.500 M K_2CrO_7, which method would you use to prepare the solution?

i. Dilute 250 mL of 0.600 M K_2CrO_7 to 300 mL with water

ii. Add 50.0 mL of water to 250 mL of 0.250 M K_2CrO_7

iii. Dilute 125 mL of 1.00 M K_2CrO_7 to 300 mL

iv. Add 30.0 mL of 1.50 M K_2CrO_7 to 270 mL of water

Question 8 (8 points):

Four metals (A, B, C, D) exhibit the following properties:

i.) only A and C react with 1.0 M HCl to give $H_2(g)$
ii.) when C is added to solutions of ions of the other metals, metallic B, D and A are formed
iii.) metal D reduces B^{+n} ions to give metallic B and D^{+n} ions

From this information, arrange the four metals in order of increasing reducing strength.

Question 9 (10 points):

For an ideal gas, draw a graph that shows the relationship between pressure and volume at a constant temperature. Graph volume on the x-axis and pressure on the y-axis.

Question 10 (14 points):

Nickel carbonyl, $Ni(CO)_4$, can be prepared by the reaction of elemental nickel metal with gaseous CO. This is the basis for purifying nickel on an industrial scale.

a. Write a balanced equation for this process.

b. You are the manager of a chemical plant that utilizes this process. You find that the plant is generating $Ni(CO)_4$ at an 85% yield. Define what is meant by "percent yield". List two factors that may help account for a percent yield of less than 100 in any chemical process.

Question 11 (14 points):

In a VERY SHORT paragraph (3-4 sentences), explain what is meant by the biochemical term "respiration". Make sure you include:
i.) a description of what chemicals are involved in the respiration process
ii.) a description of why respiration is important to animal life
iii.) a chemical reaction describing respiration (it doesn't have to be balanced)

Question 12 (10 points):

A popular soft drink has a pH of 3.3.

a. What is the concentration of hydronium ions in the beverage?

b. Is the soft drink acidic or basic?

Question 13 (10 points)

a. Name a compound that can act both as an acid and as a base.

b. Write balanced chemical equations for your chosen compound acting as an acid and as a base (one equation for each).

Question 14 (8 points):

Using only three pure liquids, give an example of a miscible pair and an immiscible pair.

Question 15 (8 points):

Circle the aqueous solution with the lowest freezing point. Briefly explain your answer.

a. 1.0 m NaCl

b. 1.0 m sucrose

c. 1.0 m $CaCl_2$

Question 16 (14 points):

One thing you can be sure about in life is that every rule has exceptions. This is definitely true with the octet rule. Consider the molecule SF_6. If sulfur has 6 bonds to it, it must violate the octet rule, right?

a. Write out the ground state electron configuration of a sulfur atom.

b. Write out a modified electron configuration, showing how sulfur could make six bonds.

Question 17 (16 points):

One of the most common reactions in organic chemistry involves using $H_2(g)$ to hydrogenate a an unsaturated compound. An example of such a reaction is shown below:

$$C_5H_{10} + H_2 \rightleftharpoons C_5H_{12} \qquad K_{eq}\,(298\text{ K}) >> 1$$

$$\Delta H << 0$$

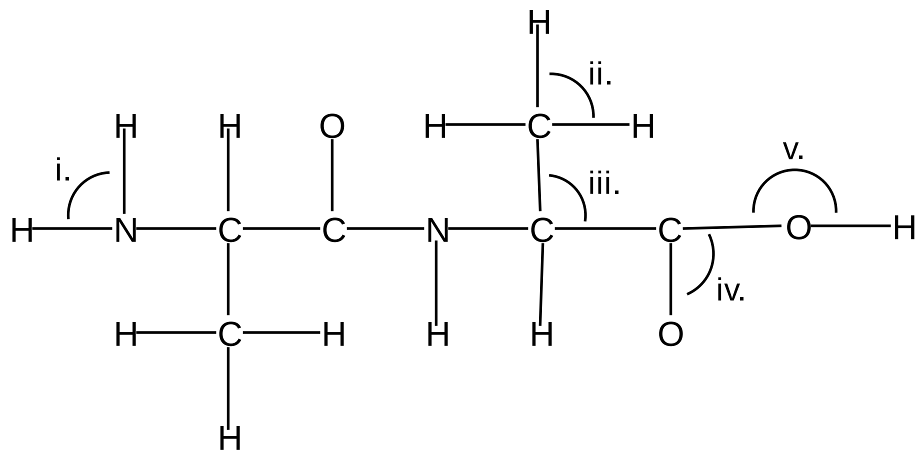
Hydrogenations are generally product-favored at room temperature, and the one shown above is no exception.

a. Assume this reaction proceeds in one step. Draw a qualitative reaction coordinate diagram for the reaction. On your diagram, show the reactants and products. Indicate the position of the transition state. Clearly label the activation energy and the enthalpy change for the reaction. Make sure that it is very clear from your diagram whether the free energy change is positive or negative.

b. This reaction can by catalyzed by the addition of palladium on carbon (Pd/C). Redraw (without all the labels) your diagram from part a. On the same axes (using another color pen or dashed lines) draw the new reaction coordinate diagram for the reaction in the presence of Pd/C.

Question 18 (16 points):

Vancomycin is an antibiotic that is often used clinically as the antibiotic of "last resort". Vancomycin kills bacteria by binding to the dipeptide D-Ala-D-Ala, which is used in the construction of the bacterial cell wall. Shown below is a "skeleton" structure of D-Ala-D-Ala (don't worry about what "D-Ala" means). Only the connectivity of the atoms is shown.



a. Redraw the structure, filling in the necessary bonds and lone pairs of electrons to generate a complete Lewis dot structure for the dipeptide. This is a neutral molecule and all of the atoms have a formal charge of 0 in the best Lewis structure.

b. Use VSEPR theory to predict the approximate size of the indicated bond angles (i.-v.).

d. Draw a Lewis dot structure of your compound from part c.